Lesson 1: Measurement

text: appendices A, B, & C (pages A-1 through A-28), page 193 for temperature

what to know:

- -scientific notation, section §A-1 (A-1)
- -precision and accuracy, §A-2 (omit standard deviations)
- -significant figures, §A-3
- -basic SI units of measurements and derived SI units of area, volume, density, force and pressure, §B-1
- -commonly-used prefixes, page §A-13 (Know Table B-3) (omit pect and exa)
- -commonly-used temperature scales, §A-15, p193
- -unit factors and the unit-factor method in solving problems (making conversions), §B-2
- -using graphs, solving algebraic equations, using logarithms and ratios, §C-1, §C-2, §C-3 and §C-4

questions:

1. What unit factor or combination of unit factors would convert kg to mg? nm to km? ps to ìs?

2. Make the following conversions expressing all answers in scientific notation.
   a. 25.0 centiliters to L
   b. 0.0149 g to ìg
   c. 22,300 nm to km
   d. 8.0 x 10^7 watts to megawatts
   e. 1,135 ns to fs
   f. 0.059 L to ìL

3. The density of mercury is 13.6 g/cm^3.
   a. What is the density expressed in kg/L? (1.00 L = 1.00 dm^3, 1.00 dm = 10^{-1} m)?
   b. How many g of mercury will occupy 98.5 mL?

4. A child has a temperature of 103 ^oF. What is her temperature in ^oC? K?

5. A sample of water at 20.0 ^oC was heated to 65.0 ^oC, an increase of 45.0 Celsius degrees. What was the increase expressed in Fahrenheit degrees?

6. How many minutes does it take for the light from the sun to reach the earth? (the distance from the sun to the earth is 93 million miles; the speed of light is 3.00 x 10^8 m/s.)

7. If a cube is 3.3 cm per side, which is the best answer for the volume? 36 cm^3 35.9 cm^3 35.94 cm^3 35.937 cm^3 Which is the best answer if the cube is 3.30 cm per side?

8. A computerized instrument is rigged up to count atoms at a rate of 1.0 million(10^6) atoms per second. How many millions of years would it take to count 6.0 x 10^{23} atoms?

9. The logarithm of 3.78 is 0.577. Without the use of a calculator, determine the logarithm of 3.78 x 10^{23} and 3.78 x 10^{-12}.

10. If the ln of a/b is -15.69 and b is 5.9 x 10^{17}, what is the value for a?

11. Additional practice problems: #s 1, 3, 13, 19, page A-7 & 8, #s 13, 19, page A-17, and #s 13, 15 page A-28.

Lesson 2: Atomic Theory

text: 2-30

what to know:

- -names and symbols of selected elements, supplement-2 (those marked)
- -what chemistry is, 1-1
- -structure and classification of matter
- -historical development of the concept of the atom (Dalton's theory), law of conservation of mass, law of definite proportions, 1-3
- -chemical formulas and law of multiple proportions, 1-4 (omit law of combining volumes and Avogadro's hypothesis)
- -contributions of Millikan, Thomson, and Rutherford to an understanding of the nuclear atom, §1-5
- -nature and properties of subatomic particles, 1-5
- -atomic number, mass number, isotopes, mass spectrometer, fractional abundance,
- -molecular & formula masses, mole, molar mass, atomic mass unit, 1-7
- -how the number of grams, the number of moles and the number of atoms(or molecules or formula units) of a substance are related, 1-7

questions:

1. Using a specific example, explain how laws and theories are related and how the term "model" fits into this discussion.
2. Which explains <u>why</u> gases are compressible, Boyle's Law or the kinetic theory?
3. When do theories become laws?
4. Describe the experiment which indicated that the nucleus occupies a very small fraction of the volume of the atom. Who first did the experiment?
5. Are compounds considered to be mixtures. Explain your answer.
6. Describe the relative charges and masses of protons, electrons and neutrons.
7. According to the Law of Multiple Proportions, if 8 grams of oxygen combines with 8 grams of sulfur to form SO_2, how many grams of oxygen would combine with 8 grams of sulfur to form SO_3?
8. The identity of potassium is determined by the number of (electrons, neutrons, protons) in the atom.
9. Dalton's theory states that all atoms of a given element are identical. Is this true? Explain.
10. Given the nuclide, $^{131}_{53}I$.
    - a. How many protons are in its nucleus?
    - b. How many neutrons are in its nucleus?
    - c. How many electrons are in an iodine atom?
    - d. How many electrons are in its nucleus?
    - e. Give the symbol for an isotope of $^{131}_{53}I$?
    - f. What is the atomic number of $^{131}_{53}I$?
    - g. What is the mass number of $^{131}_{53}I$?
11. Which symbol represents an isotope of Co-60? $^{58}_{27}Co$, $^{60}_{28}Ni$, $^{87}_{38}Sr$, $^{58}_{28}Co$
12. What is the mass of 1.50 moles of calcium atoms(40) in grams?
13. How many atoms are there in 3.94 grams of gold(197)?
14. What is the mass (in grams) of one lead atom(207)?
15. How many atoms are there in 1.80 g of magnesium(24.0)?
16. How many g of calcium(40.0) have the same number of atoms as 8.0 g of sulfur(32.0)?
17. Does every atom of carbon have a mass of exactly 12.000 amu? Explain your answer.
18. Gallium is 60% Ga-69 and 40% Ga-71, what is the approximate average molar mass of gallium?
19. What is the molar mass of a metal if 21.9 g of the metal combines exactly with 40.0 g of Br(80)? Assume that one atom of the metal combines with 2 atoms of bromine.
20. What is the molar mass of $CaCO_3$? of $Al_2(SO_4)_3$? of $(NH_4)_3PO_4$?
21. Discuss the term "purity" as it relates to substances. 4-8
22. Additional practice problems: #s 19, 21, 29, 39, 49, 53, 61 pages 34-37

Lesson 3: Chemical Formulas

text: 42-49

what to know:

- difference between molecules and compounds, molecular formulas and formula units, empirical formulas and molecular formulas, §2-1
- how to determine per cent composition from formulas and empirical formulas from per cent composition or combining masses, §2-2

questions:

1. What is the % Al in aluminum oxide?

2. How many g of Ca(40) are in 26.0 g of calcium hydroxide?

3. Verify the following data for the compound $Al_2(HPO_4)_3$ (342 g/mole).

Element	Moles per mole compound	Grams per mole compound	Grams per 1.00 g compound	Moles per 1.00 g compound	Moles per one mole Al	Empirical formula
Al	2	54	0.16	0.0059	1.0	2
H	3	3	0.0088	0.0088	1.5	3
P	3	93	0.27	0.0088	1.5	3
O	12	192	0.56	0.035	6.0	12

4. Ethylene glycol, commonly used as an antifreeze, has the empirical formula, CH_3O. Its molar mass is around 60. What is its molecular formula?

5. A 50.0 g sample contains 20.0 g of calcium, 6.00 g of carbon, and oxygen. What is the empirical formula for the compound?

6. If 0.983 g of a compound composed of phosphorus and sulfur contains 0.386 g of phosphorus, what is the simplest formula for the compound? (atomic weights, P = 31.0, S = 32.0)

7. A certain metal oxide has a formula, MO_2. A 39.46 g sample of the compound is heated in an atmosphere of hydrogen to remove the oxygen as water molecules. At the end, 24.94 g of the metal M is left over. If oxygen has an atomic mass of 16.00 amu, calculate the atomic mass of metal M and identify the element.

8. Suppose you were buying fertilizer and found that the cost of urea, $(NH_2)_2CO$, and ammonium nitrate, NH_4NO_3, were the same cost per ton. Show that you be getting the most nitrogen for your $ with the urea.

9. Additional practice problems: #s 5,27,29 pages 66,67

Lesson 4: Equations and Stoichiometry

text: 49-61
what to know:
- what a chemical equation is and says and how to balance by trial and error, §2-3
- the following terms and symbols as they relate to chemical equations:
  reactants, products, balanced, coefficient, (g), (l), (s), (aq),+, =>, §2-3
- weight relationships using equations, §2-4 (omit volume relationships of gases, 55-56)
- concepts of limiting reagents, theoretical yields and percent yields, §2-5

questions:

1. Does a chemical equation give a complete description of what happens during a chemical reaction? Explain.
2. Why are reactants always shown on the left in chemical equations?
3. If we assume that 2 molecules of O_2 react with 1 molecule of N_2, a balanced equation can be written in two ways. Are both correct? How can one determine which is actually correct?
   $N_2(g) + 2O_2(g) \Longrightarrow 2NO_2(g)$ $N_2(g) + 2O_2(g) \Longrightarrow N_2O_4(g)$
4. Balance the following equations by trial and error.
   a. $C + O_2 \Longrightarrow CO_2$
   b. $Fe + O_2 \Longrightarrow Fe_2O_3$
   c. $K + H_2O \Longrightarrow KOH + H_2$
   d. $NaOH + H_3PO_4 \Longrightarrow Na_3PO_4 + H_2O$
   e. $Fe_2O_3 + CO \Longrightarrow Fe + CO_2$
   f. $KClO_3 \Longrightarrow KCl + O_2$
   g. $P_4O_{10} + H_2O \Longrightarrow H_3PO_4$
   h. $C_6H_6 + O_2 \Longrightarrow CO_2 + H_2O$
   i. $PH_3 + O_2 \Longrightarrow P_4O_{10} + H_2O$
5. Given the equation, $4Al(27) + 3O_2(32) \Longrightarrow 2Al_2O_3(102)$
   a. How many grams of Al would react with 3.20 g of oxygen?
   b. How many g of aluminum oxide could be formed from 10.0 g of Al and 10.0 g of oxygen?
6. Given the balanced equation: $4Ga(70) + 3O_2(32) \Longrightarrow 2Ga_2O_3(188)$
   a. How many moles of O_2 are required to make 1.88 g of Ga_2O_3?
   b. How many g of Ga are required to react with 1.60 g of O_2?
   c. How many moles of Ga_2O_3 can be formed from 7.0 g of Ga and 3.2 g of O_2?
7. Given the equation, $4NH_3(17) + 5O_2(32) \Longrightarrow 4NO(30) + 6H_2O(18)$. If 28.0 g of NH_3 react with excess O_2 to form 43.0 g of water, what is the percent yield?
8. Given the balanced equation, $3H_2(2) + N_2(28) \Longrightarrow 2NH_3(17)$
   a. How many moles of ammonia could be formed from 1.00 moles of hydrogen and excess nitrogen?
   b. How many grams of N_2 would react with 1.00 moles of H_2?
   c. How many grams of H_2 would be required to produce 0.85 g of ammonia?
   d. How many moles of ammonia could be formed from 14.0 g of N_2 and 6.00 g of H_2?
9. Consider the reaction: $MnO_2(87) + 4HCl(36.5) \Longrightarrow MnCl_2(126) + Cl_2(71) + 2H_2O(18)$
   If 0.86 moles of MnO_2 and 48.2 g of HCl are allowed to react,
   a. show which reagent would be used up first.
   b. how many grams of H_2O would be formed.
10. Additional practice problems: #s 39,49,55,65,84 pages 67-70

Lesson 5: Groups and Periodicity

text: 75-85, 689-696

what to know:

- general characteristics of metals, semimetals (metalloids) and nonmetals and of the following groups: alkali metals, alkaline earth metals, halogens and noble gases, §3-1 & §3-2
- allotropes, §3-1
- periodic law, periodic table, groups (families), periods, representative elements, transition elements, lanthanides, rare earth elements, actinides, §3-2
- periodic trends of atomic and ionic radii, §17-3
- concept of ions and isoelectronic series such as Al^{3+}, Mg^{2+}, Na^{+}, Ne, F^{-}, O^{2-}, N^{3-}
- periodicity and trends of ionization energies, p-689-690

questions:

1. Describe the general distinguishing characteristics of metals, nonmetals and metalloids.

2. Describe the general distinguishing characteristics of the following groups: alkali metals, alkaline earth metals, halogens and the noble gases.

3. Which has the largest atomic radius? K Ca Rb Sr

4. Which has the largest atomic radius? S^{2-} Cl^{-} K^{+} Ca^{2+}

5. Which has the smallest atomic radius? Na Na^{+} Mg^{2+} Al^{3+}

6. Which has the smallest atomic radius? N^{3-} O^{2-} F^{-} F

7. Which has the lowest first ionization energy? Li Be Na Mg

8. Which has the largest first ionization energy? F Ne Cl Ar

9. Which has the lowest second ionization energy? Na Mg Al

10. Predict the formula for lanthanum chloride.

11. Why is the second ionization energy always greater that the first ionization energy regardless of the element being considered?

12. Explain why negative ions are always larger than their corresponding atoms.

13. Explain why atomic radii decrease going across a period from left to right.

14. What is the relationship of the size of an atom to its ionization energy and electronegativity?

15. Name the elements which are gases at room conditions. Where are they found on the periodic table?

Lesson 6: Bonding

text: 86-92, 95-103
handouts on web: Common cations, Common anions
what to know:

- valence shell, valence electrons, Lewis dot symbol, cations, anions, ionic bonds, ionic compounds salts, octet, oxoanions, §3-3
- writing Lewis dot symbols for elements and predicting the charge on binary ions, §3-3
- naming ions and writing formulas for ionic compounds, §3-3
- covalent bonding, covalent compounds (molecules), Lewis structures, octet rule, lone pairs, double bonds triple bonds, resonance structures, §3-4 & §3-5, (omit formal charges)
- how to write Lewis structures for molecules and ions, §3-5
- exceptions to the octet rule

questions:

1. Write Lewis dot symbols for each of the following elements and ions.
   C, H, F, Cl, N, S, Al, Na^+, Br^-, S^{2-}

2. Write chemical formulas for the ionic compounds formed from:

calcium and fluorine	hydrogen and nitrogen
potassium and sulfur	lithium and phosphorus
aluminum and oxygen	sodium and nitrate
magnesium and bromine	barium and phosphate
cesium and sulfate	hydrogen and carbonate

3. Write Lewis structures for: H_2S, CH_4, C_2H_6, N_2H_4, N_2, C_2H_2, NH_3, NH_4^+, CS_2, SO_3, SO_2, CO_2

4. What is the theoretical basis for the octet rule?

5. How many unshared pairs of electrons are in a molecule of: water? ammonia? methyl alcohol (CH_3OH)?, carbon dioxide?

6. Write the Lewis structures for the resonance forms of CO_3^{2-}.

Lesson 7: Molecular Geometry and Polarity

text: 85-86, 103-110, 113-114, 700-706

supplement: Geometry of electron pairs and molecules –Handout – see web page

what to know:

- -electronegativity and percent ionic character, p-85, 86, §17-5 (omit related energy concepts and calculations)
- -VSEPR (valence shell electron-pair repulsion) theory, steric number, lone pairs, bonding pairs, linear, trigonal planar, tetrahedral, trigonal pyramidal, octahedral, bent, trigonal planar, bond angles, polarity, dipole, dipole moment, §3-6(omit steric number 5 items as on page 106)
- -using the VSEPR model to predict both the distribution of electrons and the shape of common simple molecules and ions, §3-6
- -predict the relative polarities of bonds and molecules, §3-6
- -coordination complexes and the characteristics of complex ions (solubility and color) and ligands (electron donors), §3-6
- -dry ice versus wet ice, (p-108)

questions:

1.Which has the highest electronegativity value? C N Si P

2.Using the VSEPR model, predict:

a. the orientation of electron pairs in and
b. the geometry(shape) of and
c. the bond angles for each of the following: H_2S, NH_3, NH_4^+, BF_3, SF_6, CO_2, SO_2, CH_4, C_2H_4

3.Consider the methyl alcohol molecule, CH_3OH. According to the VSEPR model, what:

a. is the valence shell electron orientation on the C?
b. is the valence shell electron orientation on the O?
c. is the bond angle on the C?
d. is the bond angle on the O?

4.Using relative electronegativities from the positions of the elements on the periodic table, predict which bond is the most polar in each of the following groups.

a. H-F H-Cl H-Br H-I b. H-N H-O H-S H-Br
c. C-H C-Cl C-S C-P d. C-H C-N C-O C-F

5. Given the molecules, H_2O, H_2S, H_2Se, CO_2.
Which has the greatest polarity? Which has the smallest polarity?

6.Why is SO_2 a polar molecule and CO_2 is not?

7.In each group, which is the most polar molecule?

a. H_2O H_2S H_2Se b. SiH_4 PH_3 H_2S HCl c. H_2O H_3N H_4C H_2 d. CH_4 CH_3Cl CO_2

8. What characteristics do ligands have?

Lesson 8: Nuclear Chemistry I

text: 613-630

what to know:

-mas§energy relationships in nuclei, §15-1
-concept of nuclear stability (belt of stability), §15-1
-binding energies (mass changes during nuclear reactions), §15-1
-nature of nuclear reactions and related terms, §15-2
-how to write balanced equations for nuclear reactions, §15-2
-nature of the naturally radioactive series, §15-2
-kinetics of radioactive decay and dating procedures, §15-3

questions:

1.What are the differences between nuclides, nucleons and isotopes?
2.Describe the nature of the emissions from natural radioactive sources.
3.Consider H, Fe and U nuclei.
a. Which has the greatest mass defect per nucleon?
b. Which has the greatest nuclear binding energy per nucleon?
c. Which one loses mass when it undergoes fission?
4.Which has the least binding energy per nucleon, a He or Cu nucleus?
5.What does the symbol, a) $^{0}_{-1}e$, b) $^{0}_{-1}â$ c) $^{0}_{1}e$ represent?
6.How do nuclear reactions differ from ordinary chemical reactions?
7.Identify X in the following nuclear equations.
a. $^{135}_{53}I$ ===> $^{135}_{54}Xe$ + X
b. $^{53}_{24}Cr$ + $^{4}_{2}He$ ===> $^{1}_{0}n$ + X
8. Given the nuclear reaction, $^{20}_{11}Na$ ====> $^{20}_{10}Ne$ + X.
a. What is the mass number of X?
b. What is the atomic number of X?
c. What is the name of X?
d. How does X originate in a nucleus?
9.If a nucleus is unstable because its neutron to proton ratio is too high, what type of particle emission

10.What nuclide is formed when $^{242}_{94}Pu$ undergoes alpha particle emission?
11.What is meant by "induced radioactivity"? K-electron capture?
12.Why are neutrons never used in nuclear bombardment reactions in particle accelerators?
13.Strontium-90 has a half-life of 28.8 years. How many grams of strontium-90 was present initially, if after 144 years 10.0 g remain?
14.The radioactive decay of Tl-206 to Pb-206 has a half-life of 4.40 minutes. Starting with 5.00×10^{22}

15.Why is C-14 dating not applicable to objects over 50,000 years old?

16.Estimates show that the total energy output of the sun is 5×10^{26} J/s. What is the corresponding

Lesson 9&10: Nuclear Chemistry II

text: 630-640

what to know:

- -radiation in biology and medicine, §15-4
- -nuclear binding energy, mass defect and nuclear binding energy per nucleon, §15-2
- -nuclear fission processes and related terms, §15-5
- -how various nuclear reactors work, §15-5
- -hazards of "nuclear energy"
- -concept of nuclear fusion and status of fusion reactors, §15-6
- -radioimmunoassay, p-633
- -how radiation affects life
- -problems related with nuclear waste disposal, §15-5
- -process of synthetic nuclear transmutation, §15-6

questions:

1.Explain why much heat is released during fission and fusion.

2.Describe the role of moderators and control rods in nuclear reactors.

3.What is meant by "chain reaction" and "critical mass"?

4.What is heavy water?

5.What are the hazards associated with nuclear reactor?

6.List at least 2 advantages a fusion process would have over a fission process for energy purposes.

7.If nuclear fusion occurs in the sun, what are the problems associated with using nuclear fusion

8.Which type of radiation from radioactive sources requires the most shielding, alpha, beta or gamma?

9.Why is Strontium-90 a particularly dangerous isotope for humans?

10.Radiation from radioactive sources is considered harmful to life. Exactly what does it do in a living

11.Describe the uses of radioactive isotopes in medicine, chemistry and commercial applications.

Lesson 11: Molarity, Titrations and Dilutions

text: 159-169

what to know:
- -concept of molarity and how to use it, §4-5
- -how to prepare and dilute standard solutions, §4-5
- -stoichiometry of reactions of aqueous solutions, §4-5
- -concept of titrations and related terms, §4-5

questions:

1. How many g of NaOH(40) are required to prepare 500 mL of 0.250 M solution?

2. How many grams of solute are in 35.0 mL of 0.125 M NaCl(58.5)?

3. If I dissolve 684 g of sugar(342 g/mole) in 1.00 L of water, would I have a 2.0 M solution? Explain your answer.

4. What volume of 0.90 M H_2SO_4 solution can be prepared from 10.0 g of sulfuric acid and all the water you need?

5. How many mL of a 0.550 M HCl solution are required to prepare 1.00 L of 0.200 M HCl by dilution?

6. How many mL of 0.120 M NaOH would be required to prepare 250 mL of a 0.030 M solution by dilution?

7. Assuming that volumes are additive, what is the molarity of sodium ion and chloride ion in a solution prepared by mixing:
   a. 100 mL of 0.20 M NaCl with 200 mL of 0.10 M HCl?
   b. 100 mL of 0.20 M NaCl with 50.0 mL of 0.40 M $AgNO_3$? (AgCl precipitates)
   c. 100 mL of 0.20 M Na_2SO_4 with 100 mL of 0.20 M $BaCl_2$? ($BaSO_4$ precipitates)

8. What is the molarity of a HCl solution if excess $AgNO_3$ is added to 250 mL of the HCl solution and the resulting dry AgCl(143.5) weighs 1.435 g?

9. What volume of 0.105 M $Ba(OH)_2$ is required to exactly neutralize 25.0 mL of 0.315 M HNO_3?

10. What volume of 0.20 M NaOH would react with 100 mL of 0.30 M H_2SO_4?

11. An acidic solution of potassium permanganate reacts with oxalate ions, $C_2O_4^{2-}$, to form carbon dioxide and manganese (II) ions. If 38.4 mL of 0.150 M potassium permanganate is required to titrate 25.2 mL of sodium oxalate solution, what is the molarity of the oxalate ion? It requires 2 permanganate ions to oxidize 5 oxalate ions.

12. Excess sodium sulfate is added to a 50.0 mL of a $Ba(OH)_2$ solution and the resulting $BaSO_4$ (233 g/mole) weighs 2.963 g. What is the molarity of the barium hydroxide solution.

Lesson 12: Dissolution and Precipitation Reactions

text: 125-134

handout or webout (haha, nice new word I coined huh?):
- Solubility of ionic compounds in water (10), Reactions and equations (11)

what to know:
- difference between chemical equations and chemical reactions, sup-11
- indicators that a reaction has occurred, make a list and ask in class
- process of dissolving including terms: dissolution, solute, solvent, aqueous, aquation, hydration, dissociation, ionize, strong (weak, non) electrolytes, saturated, solubility, miscible "like dissolves like", §4-1
- generalities regarding solubility, p-130
- precipitation reactions, ionic equations, spectator ions, net ionic equations, §4-2
- solubility of ionic compounds (solubility rules), ask for the simple way!!, p-130
- how to write net ionic dissolution and precipitation reactions, practice sheet

questions:

1. Write net ionic equations to illustrate the dissolving process for those which are water soluble compounds: $AgCl$, $Ba(NO_3)_2$, $BaSO_4$, $ZnSO_4$, NH_4Br, HgS, Na_3PO_4, $CaCO_3$, $Mn_3(PO_4)_2$, $Cu(OH)_2$, $(NH_4)_2CrO_4$, MgS

2. State whether a precipitate will form when aqueous solutions of the following are mixed and write net
   a. aluminum sulfate and sodium chloride
   b. lead(II) nitrate and ammonium chloride
   c. barium nitrate and chromium(II) sulfate
   d. potassium nitrate and sodium hydroxide
   e. nickel(II) chloride and sodium sulfide
   f. silver nitrate and copper(II) chloride
   g. titanium((IV) iodide and cadmium(II) nitrate
   h. sodium carbonate, barium hydroxide and potassium chloride
   I. ammonium phosphate, calcium sulfate and sodium nitrate
   j. magnesium acetate and tin(IV) chlorate
   k. iron(III) chloride and sodium hydroxide
   l. cadmium(II) sulfate and barium sulfide
   m. copper(II) sulfate and potassium sulfide
   n. lead(II) nitrate and potassium iodide

Lesson 12: Acid-base Reactions

text: 135-138, 140-145

what to know:

- -acids, bases, alkalis, indicators, neutralization, amphoteric, §4-3
- -Arrhenius theory of acids and bases, concept of strong and weak acids and bases. §4-3
- -list of strong acids and bases, §4-3, sup-10
- -acid and base anhydrides, §4-3
- -how to write net ionic equations for reactions involving acids and bases, §4-3, sup 11

questions:

1. What makes aqueous acid-base reactions go? What drives them?
2. For acid-base reactions in aqueous solution, what is the reacting species for each of the following acids or bases?
   nitric acid, hydrofluoric acid, barium hydroxide, nitrous acid, sulfuric acid,
   acetic acid, hydrogen sulfide, potassium hydroxide, hydroiodic acid, perchloric acid
   lactic acid($HC_3H_5O_3$), hypochlorous acid, ammonia
3. Write an equation to represent the <u>net</u> change occurring when the following are mixed and allowed to react. All are aqueous solutions unless otherwise specified. (i.e. the Net Ionic Equations)
   a. $AgNO_3$ and HBr
   b. $Ba(OH)_2$ and $(NH_4)_2CO_3$
   c. HCl and H_2SO_4
   d. NH_4Cl and NaOH
   e. $BaCl_2$ and NaOH
   f. $CuSO_4$ and KOH
   g. barium hydroxide and sulfuric acid
   h. perchloric acid and potassium hydroxide
   i. phosphoric acid and ammonia
   j. HF and barium hydroxide
   k. nitric acid and methylamine (CH_3NH_2)
   l. acetic acid and lithium hydroxide
   m. NH_4OH and H_2S
   n. HNO_3 and $NaC_2H_3O_2$
   o. Na_2CO_3 and HNO_2
   p. $CaCO_3$(s) and HBr
   q. ZnS(s) and H_2SO_4
   r. Na_2S and HF
   s. NH_4NO_3 and KI
   t. Na_2SO_4, $Pb(NO_3)_2$ and $Zn(NO_3)_2$

Lesson 13: Redox Reactions

text: 146-159

what to know:

- -concept of oxidation-reduction (redox), §4-4
- -how to determine oxidation numbers, use them in identifying redox reactions and balancing them, §4-4
- -how to recognize oxidizing and reducing agents, §4-4
- -examples of combination, decomposition, oxygenation, hydrogenation, displacement, and disproportionation reactions, §4-4
- -concept and use of the activity series, §4-4

questions:

1.What is the oxidation number of:

H in HCl?	O in $S_2O_3^{2-}$?	Cl in $NaClO_4$?	I in I_2?	Fe in Fe_2O_3?
H in H_2?	S in $S_2O_8^{2-}$?	Mn in Mn_2O_3?	I in IO_4^-?	N in NH_4^+?
Cl in Cl^-	Fe in $Fe_2(SO_4)_3$?	Cr in $Cr_2O_7^{2-}$?	O in O_2?	N in NO_3^-?

2.When Sn^{4+} is converted to Sn^{2+}, is it being oxidized or reduced?

3.Consider the reaction of Mg metal with N_2 gas to form Mg_3N_2.

a. What substance is being reduced?
b. What substance is the oxidizing agent?
c. Write the balanced reaction.

4. Classify each change as oxidation or reduction and classify each first named substance as either a reducing or oxidizing agent.

a. NO to NO_3^-
b. SO_3^{2-} to SO_4^{2-}
c. MnO_4^- to Mn^{2+}
d. MnO_4^- to MnO_2
e. aluminum metal to aluminum ion
f. bromide ion to Br_2
g. BrO_3^- to Br_2
h. $Cr_2O_7^{2-}$ to Cr^{3+}

5.Consider the reaction, $Mg(s) + H^+(aq) ====> Mg^{2+}(aq) + H_2(g)$.

a. What substance is being oxidized? b. gaining electrons?
c. What substance is serving as the oxidizing agent?
d. Balance the reaction.

6 For the following reactions give the oxidizing and reducing agents and give the number of electrons lost or gained per mole of each . Assume aqueous solutions.

a. Sn^{2+} reacts with IO_4^- to form Sn^{4+} and I^- in acid.
b. $Na_2Cr_2O_7$ plus $Fe(NO_3)_2$ forms $Cr(NO_3)_3$ and $Fe(NO_3)_3$ in acid.
c. Cu + HNO_3 forms cupric nitrate and NO.
d. $Bi(OH)_3$ + SnO_2^{2-} forms SnO_3^{2-} and Bi in base.
e. Br_2 forms BrO_3^- and Br^- in base
f. Zn(s) + $AgNO_3$ forms $Zn(NO_3)_2$ and Ag(s)

Lesson 14: Nomenclature of Inorganic Compounds

text: 89-90, 110-113, 138-140

know: Common cations and anions -handouts

what to know:

-classification of inorganic compounds, sup-12

-how to write formulas from names and names from formulas for inorganic compounds

Nomenclature Worksheet

Name the following:

Na_2O	FeS	$Bi(OH)_3$
N_2O_5	H_3PO_4	KClO
$Sn(NO_3)_4$	Al_2S_3	$ZnSO_3$
Hg_2Cl_2	$AgC_2H_3O_2$	$(NH_4)_2CO_3$
$HBrO_3$	$SiCl_4$	H_2S
Al_4C_3	$Fe(OH)_2$	$CsHCO_3$
$Co_2(Cr_2O_7)_3$	Mg_3N_2	$BaHPO_4$
P_2O_3	CrO_3	Cl_2O_7
$Hg(MnO_4)_2$	H_2SO_4	CaH_2
$Cd(CN)_2$	MnO_2	CuI_2
H_2SO_3	CH_4	$Sn(ClO)_4$
$NiCrO_4$	$CoCO_3$	HI

Write formulas for the following:

lithium nitrate	nitrogen trichloride	gold (III) fluoride
silicon dioxide	aluminum carbonate	ammonium acetate
cobaltic sulfate	calcium chlorate	antimony (V) hydrogen sulfate
periodic acid	cadmium hydroxide	ferric chloride
ammonium hydrogen sulfite	iodine bromide	magnesium hydroxide
silver bromite	zinc dichromate	stannous oxide
sulfur hexafluoride	nickel nitrite	lead (IV) permanganate
nitrous acid	cuprous sulfide	potassium hydroxide
sodium dihydrogen phosphate	ammonium bicarbonate	carbon disulfide

Lesson 15: Gas Laws

text: 179-200

what to know:

-nature of the gaseous state and examples of common gases, §5-1
-concept of pressure and methods of measuring and expressing it, §5-2
-relationships of the variables, P, T, V and n, s5-2, §5-3
-derivation and use of the ideal gas law, §5-4
-concept of standard T and P, §5-2
-relationship of the density of a gas to P and T, §5-4
-stoichiometry of gaseous reactions, §5-5
-how an air bag works, p-181

questions:

1.Which elements exist as gases under normal atmospheric conditions? Li, B, Cl, S, Kr, Br, H, C, U, He,

2.If a balloon rises in the air, would its volume increase, decrease or remain the same if the T remained constant?

3.Does the pressure inside a basketball increase, decrease or remain the same when its temperature is increased?

4.Is the density of SO_2 greater than, smaller than or the same as that of SO_3 at the same T, P and moles of gas?

5.Explain why the density of HBr gas will be less at 733 mm Hg and 46 ^oC than at STP.

6.Which of the following statements are true?

a. When the pressure of a gas is doubled at constant T and n, the volume is doubled.
b. When the pressure of a gas is doubled at constant V and n, the temperature is doubled.
c. When the pressure of a gas is doubled at constant V and T, the number of moles of gas is doubled.
d. When the temperature of a gas is halved at constant P and n, the volume of a gas is halved.
e. When the T of a gas is halved at constant P and V, the number of moles of gas is halved.
f. When the volume of a gas is tripled at constant P and T, the number of moles of gas is tripled.
g. Pressure is properly expressed in lb/in^3.
h. The density of a gas is inversely related to its molecular weight at any given P and T.
i. 16 g of O_2(32) in a 10.0 L container will have the same pressure as 14 g of N_2(28) in a 10.0 L
ii. container at constant T.
j. When a gas sample is heated from 100 ^oC to 200 ^oC, its volume will double at constant P.
k. 0.125 moles of a gas will occupy 2.80 L at a pressure of 1520 atm and 273 ^oC.
l. If 20.0 g of a gas occupies 5.60 L at STP, its molar mass is 80.
m. When a balloon filled with air is heated, the density of the air is increased.
n. According to the balanced equation, $N_2 + 3H_2 \Longrightarrow 2NH_3$, 6.0 L of N_2 would react with 18.0 L of H_2 if both gases are at the same T and P.
o. 2.0 moles of O_2(32) and 2.0 moles of H_2(2) will each occupy the same V at constant T and P.

7.If the density of a gas is 4.87 g/L at 1.50 atm and 27.0 ^oC, what is the molar mass of the gas?

8.What volume would 2.20 g of CO_2 (44.0 g/mole) occupy at 380 torr and 27.0 ^oC?

9.Given the equation, $N_2(g) + 3H_2(g) \Longrightarrow 2NH_3(g)$

a. How many L of N_2 at STP would be required to produce 10.0 L of ammonia under the same
b. conditions of P and T?
c. How many grams of H_2(2.0) would be required to react with 2.00 L of N_2(g) at a pressure of 152
d. mm Hg and a T of 27 ^oC?

10.If a balloon contains 0.50 moles of a gas with a volume of V L, a pressure of P atm, and 27.0 ^oC, what would the temperature have to be so that 0.60 moles of the same gas would have the same V and P?

11.What is the pressure of the air in a tire if its gauge pressure is 28.0 psi?

Lesson 16: Kinetic Theory, Dalton's and Graham's Laws

text: 200-210

what to know:

-concept of partial pressure and its applications (Dalton's law), §5-6
-mole fraction concept, §5-6
-kinetic theory, its underlying assumptions and how it explains the properties of gases, §5-7
-relationship of mass to the rate of diffusion and effusion (Graham's law), §5-7
-the relationship of real and ideal gases, §5-8
-Van der Waals equation of state and when you'd use it

questions:

1. A 20.0 L container is filled with a H_2-He mixture. The total pressure of the gases is 1.60 atm. The container has 0.30 moles of H_2 and the partial pressure of the H_2 is 0.40 atm.
   a. What is then partial pressure of the He?
   b. What is the mole fraction of the H_2?
   c. How many moles of He atoms are in the container?
2. The vapor pressure of water in a sample of air at temperature T is 29 mm Hg. The total pressure of the gas is 0.93 atm. What is the mole fraction of the water vapor in the air sample?
3. Cylinder A contains He(4) gas and cylinder B contains N_2(28) gas. Both gases are at the same temperature and have a pressure of 10.0 atm. Both cylinders have identical valves for gas effusion.
   a. What is the ratio of the average translational rate of the He atom to the that of the N_2 molecule?
   b. If both valves are opened at the same time and it takes exactly 8.0 minutes for the pressure of the He to go from 10.0 atm to 1.0 atm, how many minutes would it take for the pressure of the N_2 to go from 10.0 atm to 1.0 atm?
   c. What effect would an increase in the T have on the effusion times?
4. Using the kinetic molecular theory, explain the relationship of P and V under conditions of constant T and n.
5. A He-filled balloon deflates faster when opened than an identical air-filled balloon of comparable initial size and T. Explain.
6. When the temperature of a gas is increased at constant volume, the molecules hit the wall (more often, less often, at the same rate) (with more force, with less force, with the same force).
7. When the volume of a gas is increased at constant T, the gas molecules hit the wall (more often, less often, at the same rate) (with more force, with less force, with the same force).
8. Which of the following statements are true?
   a. One third(1/3) of the molecules of a mixture of N_2 and O_2 molecules are nitrogen, the molar masses are 28 and 32 g/mole respectively and the total pressure of the mixture is 2.10 atm. The partial pressure of nitrogen in this mixture is 2.10 x 28/32 x 1/3.
   b. On the average, a H_2(2) molecule hits the wall of a container with 2 of the force of a He(4) atom at the same temperature.
   b. According to the kinetic theory of gases, a H_2(2) molecule in a container at 0 ^oC, will hit the wall of the container with less force on the average than a H_2 molecule at 100 ^oC.
   d. The Kinetic Theory of gases explains Boyle's Law (PV = k at constant T), in that as the volume increases the gas particles hit the wall less often thus exerting less pressure on the wall.
   e. Real gases deviate more from Boyles Law (PV = k at constant T) at low pressures than at high pressures.

Lesson 17: Condensed Phases, Phase Changes, Water

text: 222-240 handout: Types of interactions

what to know:

-condensed phases, types of intermolecular forces and examples, §6-1, sup-14

-how the kinetic theory and intermolecular forces of liquids and solids relate to phase changes and the solubilities of various substances, §6-2, §6-3, §6-4

-vapor pressure and related terms, §6-3

-phase changes and related terms, §6-4

-structure of water and it determines its properties such as density, boiling point, melting point, surface tension, solubilizing ability, capillary action, conductance, specific heat, heat of vaporization, §6-4

questions:

1.Which of the following statements are true?

a. Molecular substances have relatively higher melting points than ionic substances.

b. As a rule, intermolecular forces are weaker than intramolecular (covalent) forces.

c. The fact that He can be liquefied is evidence that dispersion forces exist.

d. Dispersion forces are due to the attraction between temporary dipoles.

e. ICl has a higher boiling point than Br_2 because ICl is more polar.

f. The dispersion forces in N_2 are stronger than those in O_2.

g. The unusually high boiling point of HF results from the unusually strong type of dipole force called a "hydrogen bond".

h. The boiling point of ethyl alcohol(CH_3CH_2OH) is higher than that of dimethyl ether (CH_3OCH_3).

i. Methane (CH_4) is more soluble in benzene (C_6H_6) than in water.

j. The relative densities of liquid water and ice are explained by the way they hydrogen bond.

k. The fact that the O-H bond energy is much greater than the hydrogen bonding in water explains why it is much easier to melt ice than to decompose water into its elements.

l. The polar nature of water explains the water solubility of many ionic compounds.

2.When evaporation occurs from a liquid, what happens to the average kinetic energy of the remaining liquid? What effect does this have on the temperature of the remaining liquid?

3.What type of intermolecular forces are involved in: liquid helium? ice? $Na^+(aq)$, $O_2(aq)$

4.In each group, rank the compounds in order of those which would exhibit the greatest hydrogen bonding to the least.

a. HF, HI, HBr b. H_2S, H_2O, NH_3 c. ether (CH_3OCH_3), ethyl alcohol (CH_3CH_2OH)

5.Explain in terms of forces between structural units why:

a. ICl has a higher melting point than Br_2 and Br_2 has a higher boiling point than Cl_2.

b. C_2H_6 has a higher boiling point than CH_4.

c. NaCl is more soluble in water than in benzene(C_6H_6).

d. water has such a high surface tension and forms droplets.

e. ice is less dense than liquid water.

f. water will form a meniscus on glass but will not wet Teflon.

6.Vapor pressure of a liquid in a closed container depends on which of the following? How about the rate of evaporation? Explain your answers.

a. what the liquid is b. the volume above the liquid c. the temperature

d. the amount of liquid present e. the strength of the intermolecular forces

f. the density of the liquid g. the surface area of the liquid

7.Give two reasons why the vapor pressure of a given liquid increases as the temperature is increased.

` 8.Given air under two conditions, at 35 ^oC and 15 ^oC. If the relative humidity is 50% in each case, in which is the vapor pressure the highest? Explain.

9.Explain why cooking food in boiling water takes a lot longer on top of Pikes Peak and how a pressure cooker makes a difference.

Lesson 18: Phase Diagrams, Raoult's and Henry's laws

text: 240-252, 262-267

what to know:

-phase diagrams (water and carbon dioxide) and related terms, §6-5

-concept of critical temperature and pressure, §6-5

-mass percentage, mole fraction and molality and interconversions with molarity,p-245-250

-ideal solutions, Raoult's law, p-250-251,

-Henry's law, distillation, p-262-267

questions:

1.When solid carbon dioxide (dry ice) is left open at room temperature it disappears without melting. Explain what happens and why it happens. How could you observe liquid CO_2?

2.Compare and contrast the phase diagrams for water and carbon dioxide.

3. Under what conditions do the solid , liquid and vapor phases of water coexist in equilibrium?

4.Lead is a poisonous metal that especially affects children. Lead levels of 0.250 ppm in a child cause delayed cognitive development. How many moles of lead are present in 1.00 g of the child's blood would 0.250 ppm represent?

5.A 100 g sample of water from a lake is found to contain 1.36×10^{-3} mg of mercury. What is the concentration of the mercury in ppm?

6.Complete the following table for water solutions of oxalic acid, $H_2C_2O_4$.

	Mass Solute	Moles Solute	Volume Solution	Molarity
a.	12.5 g	_____	456 mL	_____
b.	_____	0.0375	_____	0.138
c.	_____	_____	1.75 L	0.496

7.The density of a 60.0% aqueous solution of sugar is 1.290 g/mL. Molar masses are 342 for sugar and 18.0 for water.

a. What is the molarity of the sugar solution?

b. What is the mole fraction of water in the solution?

c. What is the molality of the sugar solution?

d. Which of the above expressions of concentration are temperature dependent, molarity, mole fraction or molality?

8.Explain why the solubility of a gas in a liquid always decreases with an increase in temperature.

9.At 25 ^{o}C the vapor pressure of pure benzene is 0.125 atm. What is the mole fraction of any nonvolatile solute which would result in a solution vapor pressure of 0.100 atm?

10.A mixture of hexane (C_6H_{14}) and heptane (C_7H_{16}) is 0.50 mole fraction in hexane. What can you say about the composition of the vapor above the liquid? Describe how these can be separated by distillation.

11. Relate Henry's law to a can of soda.

12.What is an ideal solution?

Lesson 19: Colligative Properties

text: 245, 251-262

what to know:

-colligative properties, what they are, how they relate to the concentration of solution, and their application to life around us, §6-6

-osmosis, dialysis and your kidneys, p-263

questions:

1.Which if the following statements are true?

a. 1.0 M aqueous HCl would have a lower freezing point than a 1.0 M acetic acid solution because HCl is a strong acid and the concentration of dissolved particles is higher in the HCl.

b. The vapor pressure of pure benzene at 26.1 ^{o}C is 100 mm Hg. The mole fraction of a nonvolatile solute dissolved in the benzene is 0.20.

1) The vapor pressure of the solution is 0.2 x 100 mm Hg.

2) The boiling point of the solution would be lower than the boiling point of pure benzene.

c. In an ideal solution of two volatile components, the total vapor pressure at temperature T is equal to the sum of the vapor pressures of each of the pure substances at temperature T.

d. The freezing point of 0.200 M NaCl would be lowered twice as much as that of 0.100 M NaCl when compared to the freezing point of pure water.

e. If red blood cells are placed into a hypotonic solution, they will swell and burst.

e. If one knew the g/L of a solution of an unknown substance and determined the molarity of the solution by measuring the osmotic pressure, the MW of the substance could be easily calc'd

g. The following aqueous solutions are arranged according to decreasing freezing points:
0.35 M NaCl, 0.10 M Na_3PO_4, 0.20 M $MgCl_2$.

h. Water is forced up to the top of tall trees by a process of reverse osmosis.

i. Salt helps melt ice on sidewalks by supplying the energy required for the melting process.

j. A 5.0 M aqueous solution of methyl alcohol will protect a radiator from freezing just as well as a 5.0 M solution of ethylene glycol (normal antifreeze).

k. A red blood cell with dissolved particles of 0.25 M will shrink when placed into pure water.

l. 0.25 M solutions of NaCl and KNO_3 in water will have the same osmotic pressure.

2.Draw a phase diagram for water labeling the axes and various areas on the graph. Also show what the graph would look like for a solution of sugar in water.

3.Which is not a colligative property? osmotic pressure, solubility, boiling point elevation

4.What is reverse osmosis and how does it work?

5.As compared to the vapor pressure of 0.10 M aqueous glucose ($C_6H_{12}O_6$), the vapor pressure of 0.10 M aqueous sucrose ($C_{12}H_{22}O_{11}$) will be (-almost twice as high, -the same, -1/2 as high).

6.Given an aqueous unknown liquid with a freezing point of -0.50 ^{o}C. The boiling point of the liquid under the same conditions will be (-higher than pure water, -lower than pure water, -the same as pure water, -either lower or higher than pure water depending on the atmospheric pressure).

7.The temperature of a 6 molal aqueous solution of sodium chloride in equilibrium with ice will be:
-Zero degrees Celsius, -lower than zero degrees C, -dependent on the atmospheric pressure).

8.Why are mole fraction and molality used instead of molarity when dealing with colligative properties?

9.Why does an ice-water bath get colder when salt is dissolved in the water?

10.How does salt keep a Asalt-cured@ ham from spoiling?

Lesson 20: Gaseous Equilibrium

text: 280-297

what to know:

-concept of equilibrium and equilibrium constants, §7-1
-how to calculate equilibrium constants, §7-2
-reaction quotients are and how they are used, §7-3
- how to use equilibrium constants to predict results of chemical reactions, §7-4

questions:

1.Write equilibrium constant expressions (K_p) for the reaction, $P_4H_{10}(g) + 6PCl_5(g) <=> 10POCl_3(g)$

2.Consider the following reaction: $N_2O_4(g) <===> 2NO_2(g)$ $K_p = 11$ at 100 °C. & endothermic
a. Write the equilibrium constant expression, K_p.
b. What is the value for the K_p of the reverse reaction?
c. If the partial pressures of N_2O_4 and NO_2 in a 4.0 L container are 2.0 atm and 4.0 atm respectively, which direction will the reaction go?

3.Given the equation: $2SO_3(g) <===> 2SO_2(g) + O_2(g)$, $K_p = 1.30$
a. What is the K_p for the reaction, $2SO_2(g) + O_2(g) <===> 2SO_3(g)$
b. What is the K_p for the reaction, $4SO_3(g) <===> 4SO_2(g) + 2O_2(g)$?

4.Consider the reaction, $2A_2B_3(g) ====> 4A(g) + 3B_2(g)$. 12.0 atm of A_2B_3 is placed into a 10.0 L container and allowed to dissociate at constant T. Upon achieving equilibrium, it was found that 33.0% of the A_2B_3 was dissociated. Calculate the equilibrium constant, K_p, for the reaction.

5.Consider the reaction, $H_2(g) + I_2(g) <===> 2HI(g)$ with a K_p of 62.5 at 800 K. The initial partial pressure of H_2 was 0.20 atm I_2 was 0.20 atm and HI was 2.00 atm.
a. Determine which direction the reaction will go.
b. Calculate the equilibrium partial pressures of all three species under these conditions.

6.Consider the reaction, $Cl_2(g) <===> 2Cl(g)$. $K_p = 1.4 \times 10^{-4}$ at 1100 °C.
a. If one starts with 0.5 atm Cl_2 at 1100 °C, will the equilibrium system contain mostly Cl_2, mostly Cl or appreciable amounts of both species?
b. At another temperature the K_p for the reaction is 3.6×10^{-1}. At this temperature will the equilibrium system contain mostly Cl_2, mostly Cl or appreciable amounts of both species?

7.At 650 °C, the reaction, $H_2(g) + CO_2)g) <===> CO(g) + H_2O(g)$ has a K_p value of 0.771. If hydrogen and carbon dioxide, each at a partial pressure of 0.500 atm, are placed in a 4.00 L container and allowed to react, what will be the partial pressures of all four gases at equilibrium?

8.Carbonyl bromide, $COBr_2$, can be formed by reacting carbon monoxide with bromine gas $\{CO(g) + Br_2(g) <===> COBr_2(g)\}$.
A mixture of carbon monoxide, bromine and carbonyl bromide with partial pressures of 0.0800 atm, 0.0600 atm and 0.0400 atm, respectively, is sealed in a flask. Assuming the K_p to be 5.26 at this temperature, calculate the equilibrium pressures for all gases.

Lesson 21: LeChatelier's Principle

text: 298-313

what to know:

- how changes in concentrations of reactants and products affect equilibrium and concentrations of other reactants and products at equilibrium, §7-5
- how changes in volume, pressure and temperature affect equilibrium and concentrations of reactants and products at equilibrium, §7-5
- heterogeneous equilibrium, §7-6
- how a catalyst affects equilibrium
- concepts of extraction, partition coefficients and chromatography, §7-7
- gas synthesis, p-302, & hemoglobin and oxygen transport p-306-307

questions:

1. Write equilibrium constant expressions (K_p) for the reactions,
   a. $2PbS(s) + 3O_2(g) <===> 2PbO(s) + 2SO_2(g)$

   b. $2HgO(s) <===> 2Hg(l) + O_2(g)$

2. Consider the reaction, $N_2O_4(g) <===> 2NO_2(g)$ [K_p = 11 at 100 ^oC. &endothermic] at equilibrium in a closed container, and assume that V and T are constant unless otherwise specified.

   a. What would happen to the $P(NO_2)$ if more N_2O_4 was introduced?
   b. What would happen to the $P(NO_2)$ if the volume of the container was decreased?
   c. What would happen to the $P(N_2O_4)$ if the volume of the container was decreased?
   c. What would happen to the actual number of moles of NO_2 if the volume of the container was decreased?
   e. What would happen to the $P(NO_2)$ if the temperature was increased?
   f. What would happen to the K_p if the temperature was increased?
   g. What would happen to the K_p if the volume of the container was decreased?
   h . What would happen to the K_p if a catalyst was introduced?
   i. What would happen to the K_p if He gas was introduced into the container?

3. Explain why changing the volume of a gaseous system at equilibrium shifts the equilibrium but does not change K_p, whereas a change in temperature changes both. Explain.

4. If you want to extract a compound from a water solution into an organic solvent, do you want a large or small partition coefficient. Explain you choice.

5. Describe the function of the stationary and mobile phases in chromatography.

6. What are synthesis gas and water gas, how are they made and what are they used for?

7. How does the partial pressure of O_2 relate to the saturation of hemoglobin and myoglobin?

Lesson 22: Acids and Bases I

text: 324-339

what to know:

-Bronsted theory of acids and bases, hydronium ion,(hydrated proton, H_3O^+), §8-1

-concept of the ion product of water, K_w, and its use, §8-2

-concept of pH, how to measure and use pH, §8-2

-weak acids and bases and their dissociation equilibria (K_a, K_b, pK_a & pK_b), §8-3

-what the magnitude of K_a tells us and how it relates to the % ionization, §8-3

-what indicators are and how they work, §8-3

-concept of amphoterism, §8-1

questions:

1.What is the conjugate acid of H§? HCO_3^- ? H_2O? CN^- ?

2.What is the conjugate base of HPO_4^{2-}? H§? OH^-? NH_4^+ ?

3.What is the a) $[OH^-]$ b) pH of a solution in which the $[H^+]$ is 2.0×10^{-4}?

4.What is the $[H^+]$ of a 0.0176 M HCl solution? a NaOH solution with a pH of 12.79?

5.What is the a) $[H^+]$ b) pOH of a coke if its pH is 2.77? c) the strongest base that can exist in water?

6.Which hydroxide is the strongest acid? $SO_2(OH)_2$, $Ba(OH)_2$, $SO(OH)_2$

7.Which of the following would make aqueous solutions basic? CN^- Cl^- NO_3^- CO_3^{2-}, SO_4^{2-}, NH_4^+

8.Acids HA, HB and HC have K_as of 10^{-3}, 10^{-4}, and 10^{-5} respectively.

a. Which is the weakest acid?

b. Which is the strongest base, A^-, B^- or C^-?

d. Which anion in b above would form the aqueous solution with the highest pH (be the most hydrolyzed)?

9.Which would form the most a. basic aqueous solution? b. acidic aqueous solution? c. neutral aqueous solution? NaCl, Na_2S, NH_4Cl, $(NH_4)_2S$

10.Relate the strength of an acid with the magnitude of its K_a. What is the K_a for aqueous HCl?

11.If acid, HA, has a K_a of 2.5×10^{-5}, what is the pK_b for A^-?

12.Given solution A with a pH of 6.0 and solution B with a pH of 8.0. What is the ratio of the hydrogen ion concentration in solution B as compared to solution A?

13.Given a water solution of bromothymol blue. Bromothymol blue has a K_a of about 1×10^{-7} and is yellow in acid and blue in base. What so you know about the ratio of [HIN] to $[In^-]$ and the pH when the solution is green? is blue? is yellow?

14.Given the following acids: HCl, HF($K_a = 7 \times 10^{-4}$), HCO_3^-($K_a = 5 \times 10^{-11}$)

a. Which is the strongest acid?

b.. Which anion is the strongest base, Cl^-, F^-, or CO_3^{2-}?

c Which acid has the highest pK_a?

d. Which anion has the smallest K_b, Cl^-, F^-, or CO_3^{2-}?

e. Write the reaction of F^- with water.

f. What is the value for the K_b for F^- ion?

Lesson 22 continued: Acids and Bases II

text: 339-351

what to know:

-weak acids and bases and their dissociation equilibria, §8-4

-acid-base properties of salts (ions)(hydrolysis), §8-4

-concept of buffers and how they work, §8-5

questions:

1.Propanoic acid has a K_a of 1.3×10^{-5}.

a. What is the pH of a 0.20 M solution?

b. What is the % ionization?

2.What is the a) K_a b) pK_a for 0.20 M HA(aq) with a pH of 3.00?

3.What is the a) pH b) K_a for a 0.30 M acid solution which is 0.20% ionized?

4.The K_a for acetic acid is 1.8×10^{-5}.

a.What is the $[H^+]$ of a 0.10 M sodium acetate solution?

b.How many moles of acetic acid would have to be added to 1.0 L of 0.100 molar sodium acetate solution to prepare a buffer of pH 5.00?

5.Is aqueous NH_4CN acidic or basic? Justify your answer.

6.Consider weak acid HB with a pK_a of 5.0. What is the ratio of $[B^-]$ to [HB] in each of the following cases?

a. 0.10 M HB

b. 50.0 mL of 0.30 M HB plus 50.0 mL of 0.20 M NaB

c. 50.0 mL of 0.30 M HB plus 50.0 mL of 0.20 M NaOH

d. 50.0 mL of 0.30 M NaB plus 50.0 mL of 0.20 M HCl

7.What is the $[H^+]$ of a solution prepared by mixing 50.0 mL of 0.100 M NaOH with each of the following assuming that volumes are additive?

a. 50.0 mL of 0.075 M HCl?

b. 50.0 mL of 0.100 M HCl?

c. 50.0 mL of 0.200 M acetic acid?

d. 50.0 mL of 0.100 M acetic acid?

e. 50.0 mL of 0.050 M acetic acid?

8.Which statement(s) is/are true regarding a 0.10 M solution of:

a. strong acid 1) pH = 1.00 2) $[H^+] = [A^-]$ 3) [HA] approximates 0. 4) K_a is large

b. very weak acid HA 1) pH is above 1.00 2) $[H^+] = [A^-]$ 3) [HA] approximates 0.10 4) $K_a < 1$

c. weak acid HA which has been exactly one half neutralized with NaOH.

1) $[H^+] = [A^-]$ 2) $[A^-] > [H^+]$ 3) $[A^-]/[HA] = 1.0$ 4) $[A^-]/[HA] = 2$ 5) $[H^+] = K_a$

9.The protonated form of pyridine, $C_5H_5NH^+$ (HP^+), has a K_a of 6×10^{-6}. Within what pH range could buffers be made from pyridine?

10.What would you add to an ammonium chloride solution to prepare a buffer?

11.Which combination of equal volumes would result in the formation of a buffer?

a. 0.1 M CH_3COOH and 0.05 M $NaCH_3COO$

b. 0.1 M CH_3COOH and 0.05 M NaOH

c.0.1 M $NaCH_3COO$ and 0.05 M HCl

Lesson 23: Titration curves

text: 351-364

what to know:

-acid-base titrations, related terms and diagrams, §8-6
-equilibrium (qualitative only) of polyprotic acids, §8-7
-concept of Lewis acid and bases, §8-8

questions:

1.Which would neutralize the most 0.100 M NaOH solution, 25.00 mL of 0.200 M HCl or 25.00 mL of 0.200 M $HC_2H_3O_2$?

2. Indicate whether the following aqueous solutions would have pHs <7, 7 or >7.

K_b for ammonia is 1.8×10^{-5} K_a for H§ is 1×10^{-19} K_a for formic acid, HFo, is 1.8×10^{-4}

a. 0.50 M NH_4NO_3	<7	7	>7
b. 0.50 M KCN	<7	7	>7
c. 0.50 M $(NH_4)_2S$	<7	7	>7
d. 0.50 M $Al(NO_3)_3$	<7	7	>7
e. mixture of 25 mL of 0.10 M $Ba(OH)_2$ and 25 mL of 0.05 M H_2SO_4	<7	7	>7
f. mixture of 25 mL of 0.10 M $Ba(OH)_2$ and 25 mL of 0.10 M H_2SO_4	<7	7	>7
g. mixture of 25 mL of 0.050 M $Ba(OH)_2$ and 25 mL of 0.10 M H_2SO_4	<7	7	>7
h. mixture of 25 mL of 0.10 M NaOH and 25 mL of 0.10 M H_2SO_4	<7	7	>7
i. mixture of 25 mL of 0.30 M formic acid (HFo) and 25 mL of 0.10 M NaOH	<7	7	>7
j. mixture of 25 mL of 0.30 M formic acid (HFo) and 25 mL of 0.30 M NaOH	<7	7	>7
k. mixture of 25 mL of 0.30 M formic acid (HFo) and 25 mL of 0.50 M NaOH	<7	7	>7
l. mixture of 25 mL of 0.30 M formic acid (HFo) and 25 mL of 0.50 M NaFo	<7	7	>7
m. mixture of 25 mL of 0.30 M HCl and 25 mL of 0.50 M NaFo	<7	7	>7
n. mixture of 25 mL of 0.50 M HCl and 25 mL of 0.50 M NaFo	<7	7	>7
o. 0.50 M NaFo	<7	7	>7

3.Which of the solutions in 2 above are buffers?

4.Which species which are present in 2.j above. HFo Fo^- H^+ OH^- H_2O

5.Given an indicator with a K_a of 1.0×10^{-5}, where the HIn species is red and the In^- species is yellow.

a. What color is the indicator in a pH 5.0 solution? red orange yellow
b. What color is the indicator in a pH 10 solution? red orange yellow
c. At what minimum pH will the indicator be about as yellow as it can be? pH 4 pH 5 pH 6
d. Would this indicator be appropriate for a HCl-NaOH titration? yes no

6.If phosphoric acid, H_3PO_4, has pK_a values of about 2, 7 and 12, what species would be present in the highest concentration in human blood at a pH of 7.4?

Lesson 24 & 25: Solubility Equilibrium I

text: 374-395

what to know:

-solubility terminology and effect of temperature on solubility, §9-1
-solubility, molar solubility and solubility product constant and their interconvertability, §9-2
-effect of common ion on solubility(qualitative), §9-3
-effect of pH on the solubility, §9-4
-concept of complex ions and their solubility(qualitative), §9-5
-lead all around you (page 403)

questions:

1.What is the effect of temperature on the solubility of salts?

2.Write the equilibrium equation for the dissolving of the following and the K_{sp} expression.
AgCl $PbBr_2$ Bi_2S_3 $Fe(OH)_3$

3.Given the solubility of CaF_2 (78 g/mole) as 0.0016 g/100 mL in water.
a. What is the $[Ca^{2+}]$ of a saturated solution of CaF_2?
b. What is the $[F^-]$ of a saturated solution of CaF_2?
c. What is the K_{sp} of CaF_2?

4.Calculate the molar solubility of $PbCl_2$ (278 g/mole) in water if its K_{sp} is 1.7×10^{-5}.

5.Circle those of the following which are soluble in aqueous HCl.
AgBr $CaCO_3$ $Fe(OH)_3$ $BaSO_4$ PbI_2 $AlPO_4$ ZnS

6.The molar solubility of $Mg(OH)_2$ in water is 1.4×10^{-4} M at 25 °C.
a. What is the K_{sp} for magnesium hydroxide at this T?
b. What is the pH of a saturated magnesium hydroxide solution?
c. Show with equations why the solubility of magnesium hydroxide is increased in acidic solutions.

7.Given the K_{sp} of Ag_2CrO_4 as 1.0×10^{-12}. What is its molar solubility?

8.The K_{sp} for $Ca(OH)_2$ is 8.0×10^{-6}. Show that a precipitate will not form when 2.00 mL of 0.200 M NaOH is added to 1.00 L of 0.100 M $CaCl_2$?

9.The solubility of Ag_3PO_4 is 6.7×10^{-3} g/L. The molar mass of Ag_3PO_4 is 419 g/mole. What is the K_{sp} of Ag_3PO_4?

10.Limestone caverns are formed when the limestone is dissolved by acidic water.
a. What is the chemical composition of limestone?
b. What <u>normal</u> substance in the atmosphere makes rain water acidic?
c. How is this process related to "hard" ground water?

11.Which of the following pairs of equations would demonstrate that zinc hydroxide is amphoteric?
a) $Zn(OH)_2 + 2H^+ \text{===>} Zn^{2+} + 2H_2O$, $Zn(OH)_2 + 2OH^- \text{===>} Zn(OH)_4^{2-}$
b) $Zn(OH)_2 + H^+ \text{===>} Zn(OH)^+$, $Zn(OH)^+ + H^+ \text{===>} Zn^{2+} + 2H_2O$
c) $Zn(OH)_2 + OH^- \text{===>} Zn(OH)_3^-$, $Zn(OH)_3^- + OH^- \text{===>} Zn(OH)_4^{2-}$

12.K_{sp}s for $Cu(OH)_2$ and $Mg(OH)_2$ are 2.2×10^{-20} and 6×10^{-12} respectively. When adding aqueous NaOH drop by drop to a solution containing equal molar concentrations of Cu^{2+} and Mg^{2+}, which hydroxide would begin to precipitate with the least number of drops?

Lesson 26: Thermochemistry I

text: 411-430

ask in class if we don't do it: Heating curve for water (16)

what to know:

-principles of heat flow and related terms, §10-1

-calorimetry, specific heat and heat capacity, §1-2

-concept of enthalpy and changes in enthalpy, §10-3

-thermochemical equations, Hess' law,-enthalpies of phase changes, enthalpies of formation, §10-3

-standard molar enthalpies, §10-4

questions:

1. Given the specific heats of $H_2O(l)$ and Hg(l) as 4.184 and 0.139 J/g°C, respectively, which liquid would require the greatest time to heat from 20 °C to 85 °C given identical heating conditions?
2. If 43.0 g of a substance at 98.0 °C is placed into a calorimeter containing 100.0 g of water at 25.0 °C, the temperature rises to 29.5 °C. If the heat capacity of the calorimeter is 35.0 J/°C, what is the specific heat of the substance?
3. Upon which of the following factors does the heat capacity of a substance depend? temperature, mass, composition
4. When 1.34 g of KBr (119 g/mole) dissolves in 74.0 g of water in a coffee-cup calorimeter, the temperature drops from 18.000 °C to 17.279 °C. Assume that all the heat absorbed in the solution comes from the water. The specific heat for water is 4.184 J/g °C. What is q for the dissolving of one mole of KBr in water?
5. A process occurring in a bomb calorimeter causes the change in the temperature of a calorimeter from 19.75 °C to 28.67 °C. The heat capacity of the calorimeter is 13.24 kJ/°C. What was q for the process involved?
6. Ten(10) g of water at 32 °C is heated to 100 °C, then cooled and frozen at 0 °C. The 10 g of ice is then melted and heated to 32 °C. Calculate ÄH for the process. The heat of fusion for water is 6.00 kJ/mole and the specific heat of water is 4.184 J/g °C. The processes are:
   a. $H_2O(l)(32\ ^oC) \text{ ===> } H_2O(l)(100\ ^oC)$ b. $H_2O(l)(100\ ^oC) \text{ ===> } H_2O(l)(0\ ^oC)$
   c. $H_2O(l)(0\ ^oC) \text{ ===> } H_2O(s)(0\ ^oC)$ d. $H_2O(s)(0\ ^oC) \text{ ===> } H_2O(l)(0\ ^oC)$
   e. $H_2O(l)(0\ ^oC) \text{ ===> } H_2O(l)(32\ ^oC)$
7. How much heat is given off when 1.26×10^4 g of ammonia is produced according to the equation, and under standard conditions? Ammonia is 17 g/mole. $N_2(g) + 3H_2(g) \text{ ====> } 2NH_3(g)$ $\Delta H^o = -92.6$ kJ
8. Which of the following are exothermic, and what is the sign on ÄH?
   a. $H_2O(g) \text{ =====> } H_2O(l)$ b. $2Na(s) + 2H_2O(l) \text{ ====> } 2Na^+(aq) + 2OH^-(aq) + H_2(g)$
   c. $H^+(aq) + OH^-(aq) \text{ ====> } H_2O(l)$ d. $CO_2(g) + 2H_2O(l) \text{ ====> } CH_4(g) + 2O_2(g)$
9. Acetylene burns in air according to the equation: $2C_2H_2(g) + 5O_2(g) \text{ ====> } 4CO_2(g) + 2H_2O(l)$ $\Delta H^o = -2598.9$ kJ. Would you expect the change in enthalpy to more or less negative if the water formed were a gas rather than a liquid?
10. The ΔH_f^o for $NO_2(g)$ is given as 33.9 kJ/mole. Write the thermochemical equation for which this is the standard change in enthalpy.
11. Which has the most positive ΔH_f^o? $Cl_2(g)$, $Cl_2(l)$, $Br_2(g)$, $Br_2(l)$?
12. If the heat of combustion for hydrogen gas is given by the equation, $2H_2(g) + O_2(g) \text{ ==> } 2H_2O(l)$ $\Delta H^o = -572$ kJ, what is the ΔH^o_f for liquid water?

lesson 26 continued on the next page

13.What is the standard heat of reaction (ÄH°) for the following reaction, $2C_2H_2(g) + 5O_2(g) \text{ ====> } 4CO_2(g) + 2H_2O(l)$, given standard heats of formation as follows: $C_2H_2(g) = +227$ kJ/mole, $CO_2(g) = -394$ kJ/mole, $H_2O(l) = -286$ kJ/mole

14.Calculate the standard enthalpy of formation of carbon disulfide (CS_2) from its elements, given that:
C(graphite) + $O_2(g)$ ===> $CO_2(g)$ ÄH° = -394 kJ
$S(s) + O_2(g)$ ====> $SO_2(g)$ ÄH° = -296 kJ
$CS_2(l) + 3O_2(g)$ ===> $CO_2(g) + 2SO_2(g)$ ÄH° = -1072 kJ

15.Given the thermochemical equation: $2Al_2O_3(s)$ ====> $4Al(s) + 3O_2(g)$ ÄH° = +3352 kJ
What is the standard heat of formation of aluminum oxide per mole?

Lesson 27: Thermochemistry II

text: 430-441

what to know:
-concept of bond energies and relationship to enthalpy changes, §10-5
-first law of thermodynamics, its use and related terms, §10-6
-relationship of ÄH and ÄE, §10-6
-food calorimetry, p-440-441

questions:
1.Why are bond-breaking processes always endothermic and bond-making processes always exothermic?

2.Using bond energies from the table in the text, estimate the ÄH for the:
a. formation of hydrazine, N_2H_4 from hydrogen gas and nitrogen gas.
b. combustion of propane, C_3H_8 to form $CO_2(g)$ and $H_2O(g)$

3.Consider the following changes and state whether work is done on the system, by the system or no work is done and give the sign for w. Consider these open to the atmosphere.

a. $HgO(s)$ ====> $Hg(l) + O_2(g)$

b. $2HI(g)$ =====> $H_2(g) + I_2(g)$

c. $3O_2(g)$ =====> $2O_3(g)$

4.Why is wrong to talk about a system of substance containing Awork@ or Aheat@? Or to talk of gasoline or petroleum as being Aenergy@?

5.A gas expands and does P-V work on the surroundings equal to 325 J. At the same time it absorbs 127 J of heat from the surroundings. Calculate the change in energy of the gas.

6.For the vaporization of one mole of water at 100 °C, determine q_v, w_p, and q_p. [If $ÄE = q_v$ and $ÄH = q_p$, $ÄE = q + w$ and $w = -P(ÄV)$, then $q_p = q_v + P(ÄV)$ and $ÄH = ÄE + P(ÄV)$.] (101.3 J/L atm)

7.Why are decomposition reactions usually endothermic?

Lesson 28: Thermodynamics I

text: 451-466

what to know:

- -what is meant by a spontaneous process and the two factors which determine spontaneity, §11-1
- -concept of entropy and how entropy changes during various processes, §11-2
- -third law of thermodynamics, absolute entropies, and the effect of temperature changes, §11-3
- -second law of thermodynamics, §11-4
- -Gibbs function (free-energy)and how its predicts spontaneity, §11-5, Trouton's Rule

questions:

1. Explain what is meant by a spontaneous process. Give two examples each of spontaneous and nonspontaneous processes.

2. In each pair of the substances listed, choose the one with the larger standard entropy per mole and give your reasons.
   a. Li(s) or Li(l)
   b. C_2H_5OH (l) or CH_3OCH_3 (l) (Hint: which one forms hydrogen bonds?)
   c. Ar(g) or Xe(g)
   d. CO or CO_2

3. Does the entropy increase, decrease or remain constant during the following processes?

   a. a solid melts.
   b. a liquid freezes.
   c. a liquid vaporizes.
   d. a solid sublimes.
   e. NaCl(s) ===> NaCl(aq)
   f. Fe(s)(25 ^oC) ===> Fe(s)(200 ^oC)
   g. $2H_2(g) + O_2(g) ==> 2H_2O(g)$
   h. $2HgO(s) ===> 2Hg(l) + O_2(g)$
   I. $H_2(g) + Cl_2(g) ==> 2HCl(g)$

4. Explain why entropy is zero at O K.

5. Entropy is expressed in what units?

6. What do you know about the ΔS_{univ} for all spontaneous processes?

7. Explain how an endothermic process can be spontaneous?

8. How does ΔG relate to ΔS_{univ} and ΔH_{syst} to $\Delta S_{surroundings}$?

9. Explain why ΔG is zero for a process at equilibrium.

10. What can you say about ΔH_{syst} and $\Delta S_{surroundings}$ for an exothermic process?

Lesson 29: Thermodynamics II

text: 466-473

what to know:

-standard Gibbs function and chemical reactions, §11-6
-temperature and Gibbs function, §11-6
-relationship between Gibbs free-energy and equilibrium, §11-7
-life does not violate the second law of thermodynamics p-476-477
-Clausius-Clapeyron Equation , section 11-8

questions:

1.Relate ÄG and $ÄG^o$.

When are they equal?
Are they both zero at equilibrium?
Do they always have the same sign?
Are both affected by a change in the concentrations of reactants and products?

2.For what are $ÄG_f^o$s useful?

3.For each of the following systems tell whether the thermodynamic function is zero, greater than zero or less than zero. For e, tell whether the equilibrium constant is one, greater than one or less than 1.

a. $ÄS^o_{sys}$ for the reaction, $NH_4Cl(s) \Longrightarrow NH_3(g) + HCl(g)$

b. $ÄS^o_{surr}$ for the reaction, $CH_4(g) + 2O_2(g) \Longrightarrow CO_2(g) + 2H_2O(g)$

c. $ÄS^o_{universe}$ for the process, $H_2O(s) \Longrightarrow H_2O(l)$:

1) at 0.0 ^oC.
2) at -20 ^oC.
3) at +30 ^oC.

d. $ÄG^o$ for the process, $H_2O(g) \Longrightarrow H_2O(l)$:

1) at 120 ^oC.
2) at 60 ^oC.

e. K_{eq} for the reaction, $2HI(g) \Longrightarrow H_2(g) + I_2(g)$ ($ÄG^o$ is -2.60 kJ)

f. $ÄG^o$ for the reaction, $3O_2(g) \Longrightarrow 2O_3(g)$($ÄH^o$ is positive):

1) at high temperatures.
2) at low temperatures.

Lesson 30: Metallurgy and Redox

text: 488-504

handout: occurrences of metals (17), alloys (18)

what to know:

-examples of the types of minerals used as ores, handout

-various steps involved in metallurgy, §12-1

- relative reactivity of metals, §12-1

-metallurgy of copper and iron (and steel formation), §12-1

-alloys, handout

-concept of half-reactions and balancing redox reactions, §12-2

questions:

1.What chemical property characterizes metals which can be mined as native metals?

2.Name several commonly used metals produced from ores containing:
a) oxide minerals. b) sulfide minerals. c) native metals.

3.Iron can be prepared by using carbon as a reducing agent but electrolysis is required for the preparation of sodium. Explain.

4.Describe the metallurgy of iron.

a. What is the actual reducing agent for Fe_2O_3 in a blast furnace?

b. The presence of what element makes pig iron so brittle?

c. Steel has a (lower, higher) carbon content than pig iron?

d. What element is largely responsible for the corrosion resistance of stainless steel?

e. Why is limestone added to the blast furnace?

5.What products are formed when Cu_2S is heated strongly in air?

6.What does it mean for a metal to be ductile? malleable?

7.What element is common to the alloys, brass and bronze as well as pennies. What is an amalgam?

8.Balance the following half reactions, classify each reaction as oxidation or reduction and classify each first named substance as either a reducing or oxidizing agent.

a. NO to NO_3^- in acid

b. SO_3^{2-} to SO_4^{2-} in base

c. MnO_4^- to Mn^{2+} in acid

d. MnO_4^- to MnO_2 in base

e. aluminum metal to aluminum ion

f. bromide ion to Br_2

g. SO_3^{2-} to SO_4^{2-} in base

h. $Cr_2O_7^{2-}$ to Cr^{3+} in acid

9.Write balanced net ionic equations for the following reactions. Assume aqueous solutions and use half reactions.

a. Sn^{2+} reacts with IO_4^- to form Sn^{4+} and I^- in acid.

b. $Na_2Cr_2O_7$ plus $Fe(NO_3)_2$ forms $Cr(NO_3)_3$ and $Fe(NO_3)_3$ in acid.

c. Cu + HNO_3 forms cupric nitrate and NO.

d. $Bi(OH)_3$ + SnO_2^{2-} forms SnO_3^{2-} and Bi in base.

e. Br_2 forms BrO_3^- and Br^- in base

f. Zn(s) + $AgNO_3$ forms $Zn(NO_3)_2$ and Ag(s)

Lesson 31: Electrochemistry I

text: 504-518

what to know:

-electrochemical cell terminology (electrodes, salt bridge, half-cells, current, voltage, coulomb, galvanic cell, electrolytic cell, anode, cathode), §12-3

-how to write abbreviated cells, anode to the left and anode to the right, e flowing L to R, p-508

-Faraday=s laws and their application, §12-4

-electrometallurgy, metallurgy of aluminum and magnesium, §12-5

questions:

1.Describe the basic features of a galvanic (voltaic) cell including what happens at the anode and cathode. Why are the two components separated from each other? What is the function of the salt bridge? Which direction do anions flow in the salt bridge on a operating voltaic cell?

2. An automobile battery delivers a current of 50 A for 1.5 sec to start an engine. Calculate the quantity of charge (in coulombs) passing through the battery. How many moles of electrons are passing per minute?

3.How many minutes would it take to deposit 1.00 oz (28.4 g) of gold from a $AuCl_3$ solution with a current of 2.00 amperes?

4.A current of 0.600 A deposits 2.42 g of a certain metal, M, in one hour from a M^+ solution. What is the atomic mass of the metal?

5.A constant electric current flows for 3.75 h through two electrolytic cells connected in series. One contains a solution of $AgNO_3$ and the second a solution of $CuCl_2$. During this time 2.00 g of Ag are deposited in the first cell.

a. How many grams of copper are deposited in the second cell?

b. What is the current expressed in amperes?

6.Galvanized steel consists of steel with a thin coating of zinc to slow corrosion. The zinc can be deposited electrolytically by making the steel object the cathode and a block of zinc the anode in an electrochemical cell containing a dissolved zinc salt. Suppose a steel garbage can is to be galvanized and requires a total mass of 7.32 g of zinc to be coated to the required thickness. How long should a current of 8.50 A be passed through the cell to achieve this?

7.Describe the metallurgy of aluminum. Why is it economically advantageous to recycle aluminum? (Each recycled aluminum can saves the equivalent of one cup of gasoline, every recycled aluminum can saves enough electricity to burn a 100 watt bulb for 3.5 years, two aluminum cans thrown away is a waste of more energy than is used daily by each of a billion people in poorer lands, every year Americans throw away enough aluminum to rebuild the entire American Airline fleet 71 times.)

8.Aluminum is not as good as copper in electrical conductance and yet it is often used instead of copper for that purpose. Why?

9.How is copper refined?

Lesson 32: Electrochemistry II

text: 526-544

what to know:

-Gibb's function and cell voltages, standard states, §13-1
-standard half-cell reduction potentials. §13-2, (Omit disproportionation, p-534)
-how to write balanced equations for electrode reactions, predict spontaneity of redox reactions and calculate standard cell emfs using the standard reduction potential table, p-532
-oxidizing and reducing agents, §13-3
-effect of concentrations (Nernst Equation) on cell voltages (qualitative), pH meters, §13-4
-relationship of Gibb's free-energy, cell voltages and equilibrium constants (qualitative), §13-5

questions:

1.Consider a voltaic cell in which the half-cell reactions Mg/Mg^{2+} and Ag/Ag^{+} are used.
a. What is the reducing agent?
b. What substance is reacting at the anode?
c. Write the overall balanced equation for the reaction occurring in the cell.
d. Calculate the standard cell voltage for the cell.

2.Consider the cell: Al Al^{+3}(aq,1M) KCl(sat'd) NO(g) NO_3^- (H^+) Pt(s)
a. What is the oxidizing agent?
b. What substance is reacting at the cathode?
c. Write the balanced net ionic equation for the reaction.
d. What is the E^o_{cell} for the reaction?
e. Would the E_{cell} for the reaction be (lower than, higher than, or unchanged from) E^o_{cell} if the aluminum ion concentration was 2 M?

3.Given the reaction, $2Au(s) + 3Ca^{2+}(aq) \Longrightarrow 2Au^{3+}(aq) + 3Ca(s)$
a. What is the E^o_{cell} for the reaction as written?
b. Is this reaction spontaneous?

4.Name:
a. a substance that will oxidize Fe^{2+} but not Br^-.
b. a halogen that will oxidize Cu but not Ag.
c. a substance that would oxidize Fe to Fe^{2+} but not to Fe^{3+}.

5.Which of the following:
a. would dissolve in nitric acid but not in hydrochloric acid under standard conditions?
Sn Hg Au Fe Mn Cu Cd Au
b. would oxidize Br^- to Br_2? I_2, I^-, Cl_2, Cl^-, Au^{3+}, Cu
c. is the strongest reducing agent? Al^{3+}, H_2, Cl_2, F^-

6.Relate ΔG^o, E^o_{cell}, and K_{eq} for redox reactions. If ΔG^o is very negative, what do you know about the magnitude of E^o_{cell} and K_{eq}?

7.Consider the reaction, $Zn^{++} + Cu \Longrightarrow Cu^{++} + Zn$. Is this reaction spontaneous under standard conditions? Under what conditions would it be spontaneous? Explain.

8.How does a pH meter work?

Lesson 33: Electrochemistry III

text: 907-909, 544-558

handout: Electrolysis (19)

what to know:

- -electrolysis of various substances in the molten or aqueous states, sup-19, §23-2
- -batteries and fuel cells, chemistry of the lead storage battery, §13-6
- -the chemistry of corrosion and its prevention, §13-7

questions:

1..Write the balanced half-equations for the anode and cathode reactions occurring when:

a. molten KBr is electrolyzed.
b. aqueous KI is electrolyzed.
c. aqueous HI is electrolyzed.
d. aqueous $CuSO_4$ is electrolyzed.
e. water is electrolyzed.

2.How does the ordinary dry cell work?

3.The fact that the lead storage battery is rechargeable is one reason why it is so useful. What reactions occur at the anode and the cathode during a recharge? Why can the status of the battery's "charge" be determined by measuring the density of the battery acid?

4.In a hydrogen-oxygen fuel cell, the reaction is very similar to the combustion of hydrogen, except that the anode and cathode reactions are carried out separately. The electrode reactions are $2H_2(g) + 4OH^-(aq) \Longrightarrow 4H_2O(l) + 4e^-$ and $O_2(g) + 2H_2O(l) + 4e^- \Longrightarrow 4OH^-(aq)$

a. What is the overall balanced equation?
b. What is the standard voltage of the cell?

5.Explain why aluminum cookware does not disintegrate in air but is not very useful for cooking acid foods.

6.The corrosion of iron can be prevented or retarded by a variety of means. Explain how each of the following work.

a. Keep the iron dry.
b. Galvanize it.
c. Plate it with tin (as in tin cans).
d. Paint it.
e. Connect it to a sacrificial active metal like magnesium.
f. Control the pH.

Lesson 34: Solids I

text: 792-809

what to know:
- -difference between crystalline and amorphous solids, p-792
- -how the structure of condensed phases are studied, §20-1
- -symmetry in crystals, §20-2
- -crystal systems, crystal lattice, unit cells, primitive cells, nonprimitive cells, face-centered, body-centered and side-centered cells, §20-2

questions:
Questions 9, 11, 19 on page 825 in the text

Lesson 35: Solids II

text: 809-823

handout: Types of solids (20), Properties and types of solids (21)

what to know:
- -the types of solids and their characteristic physical properties, §20-3, sup-20 & 21
- -effect of defects in solids, §20-4
- -liquid crystals & displays, §20-5, p-823

questions:

1.Which of the following statements are true?
- a. Molecular substances have relatively high melting points.
- b. Network covalent solids tend to be soluble in water as well as in nonpolar solvents.
- c. Ionic solids are good conductors of electricity.
- d. NaCl(s) is better conductor of electricity than silver metal because it is more polar.
- e. The "electron-sea" model is used to explain the electrical conductivity of network covalent solids.

2.Classify the following substances as nonpolar molecular, polar molecular, network covalent, ionic, metallic or amorphous solids. $CaCO_3$, SiO_2, KNO_3, H_2O(ice), Ca, CO_2, I_2, glass, diamond

3.Classify the following substances:
- a. A solid is hard, brittle and electrically nonconducting. Its melt (the liquid form of the substance) and an aqueous solution containing the substance do conduct electricity.
- b. A solid is soft and has a low melting point. The solid, its melt and an aqueous solution containing the substance are all nonconductors of electricity.
- c. A solid is very hard and has a high melting point. Neither the solid or its melt conducts electricity.

4. Questions 29, 31, 33 on page 826 of the text.

Lesson 36: Kinetics(rate laws) text: 567-581

what to know:

-what is meant by reaction rates, how they are measured and expressed, §14-1

-reactions rates and concentrations of reactants (rate laws, rate constants, reaction order, etc.) and how they are determined from experimental data, §14-2

-relationships between reactant concentration and time(half-life), §14-3, (omit second order)

questions:

1.Consider the experiment for the reaction, A + 2B ===> C + 2D.

Experiment	Initial [A]	Initial [B]	Rate(Ä[C]/s)
A.	0.0020	0.015	3.0×10^{-4}
B.	0.0050	0.015	3.0×10^{-4}
C.	0.0040	0.045	9.0×10^{-4}

What is the order of reaction with respect to A? with respect to B? overall

2.Consider the reaction, 2NOBr(g) ====> Br_2(g) + 2NO(g), which is 2nd order with respect to NOBr.

If -Ä[NOBr]/s is 3.2×10^{-2} M/s at 10^oC when [NOBr] is 0.20 M:

a. what is the rate of [Br_2] formation in molarity per minute?

b. what is the rate constant at 10^oC in terms of [NOBr]/s?

3.Given the following data for reaction: A + 2B ===> 2C

Experiment	Initial [A]	Initial [B]	Initial rate(Ä[C]/s)
A.	2.00×10^{-4}	3.00×10^{-3}	6.00×10^{-12}
B.	4.00×10^{-4}	9.00×10^{-3}	7.20×10^{-11}
C.	2.00×10^{-4}	9.00×10^{-3}	1.80×10^{-11}
D.	8.00×10^{-4}	9.00×10^{-3}	2.88×10^{-10}

The rate of reaction is given by which of the following?

$k[A]^o[B]^o$	$k[A]^1[B]^o$	$k[A]^2[B]^o$	$k[A]^3[B]^o$
$k[A]^o[B]^1$	$k[A]^1[B]^1$	$k[A]^2[B]^1$	$k[A]^3[B]^1$
$k[A]^o[B]^2$	$k[A]^1[B]^2$	$k[A]^2[B]^2$	$k[A]^3[B]^2$
$k[A]^o[B]^3$	$k[A]^1[B]^3$	$k[A]^2[B]^3$	$k[A]^3[B]^3$

4. Given the following data for the reaction, 2A + B === 2C.

Experiment	Initial [A]	Initial [B]	Ä[C]/s
A.	0.120	0.050	1.20×10^{-5}
B.	0.040	0.050	4.00×10^{-6}
C.	0.040	0.100	8.00×10^{-6}

a. What is the order of the reaction with respect to A? B?

b. What is the value for Ä[B]/s in the reaction for experiment A?

c. What is the value for the specific rate constant for the reaction in terms of C formed?

5.Given the relationship, $\ln([A]_0/[A]) = kt$ for a first order reaction.

a. What is the specific rate constant if the half-life is 80 minutes? Give your answer in s^{-1} units.

b. What is the time required for 90% of A to react if the specific rate constant is 2.0×10^{-5} §[1]

Given the relationship, $\ln([PH_3]_0/[PH_3]) = kt$, for the first order decomposition,

$4PH_3(g)$ ====> $P_4(g)$ + $6H_2O(g)$. k = 0.0198 §[1]

a. Calculate the half-life for the reaction in seconds.

b. What is the concentration of PH_3 after the reaction has gone for 100 seconds with an initial concentration of 0.20M?

Lesson 37: Kinetics (theory and catalysis)

text: 591-600

what to know:

-temperature and reaction rate(qualitative)

-steric factor, collision theory, activation energy, activated complex, diagrams and related terms, §14-6

-what a catalyst does, how an enzyme works, and how a catalytic converter functions, §14-7

questions:

1.Draw an activation energy diagram for an reaction which has an activation energy of 300 kJ and an enthalpy change of -150 kJ. Label the "activated complex" on the graph. How would the diagram differ if the reaction was endothermic?

2.Which of the following are true statements.

According to the collision theory, in a reaction between A and B, a reaction will occur every time an A molecule and a B molecule collide.

b. An increase in the temperature increases the number of collisions of gas molecules per second.

An increase in the temperature increases the number of collisions which result in the formation of the activated complex.

a. An increase in the temperature lowers the activation energy for a reaction thus increasing the rate of reaction.

e. Raising the T generally increases the fraction of collisions which result in product.

f. Raising the T increases the kinetic energy of colliding reactants so that a higher percentage achieve the "activated complex" status.

g. Raising the T decreases k, the specific rate constant for a reaction.

h. In general, the higher the activation energy, the faster the rate of a reaction at constant T..

i. The rate constant equals the reaction rate at $0^{o}C$ for most reactions.

j. The activation energy is lower for exothermic reactions than it is for endothermic reactions.

3.If kinetic energy is directly proportional to temperature, give two distinct reasons why the rate of a reaction increases as the temperature increases.

4.What measurements need to be taken in order to determine the energy of activation in the laboratory? What needs to be kept constant?

5.Methane (natural gas) burns readily in oxygen in a highly exothermic reaction. Yet a mixture of methane and oxygen can be kept indefinitely without any apparent change. Explain.

6.How does a catalyst increase the rate of a reaction?

a. Does it affect both the forward and reverse reactions?

b. Does a catalyst "get involved" in the reaction?

c. Does a catalyst increase the number of collisions per second?

d. Does a catalyst supply the activation energy?

e. Does a catalyst become part of the activated complex?

f. Does a catalyst become consumed during a reaction?

g. Does a catalyst alter the specific rate constant for a reaction?

7.Distinguish between homogeneous and heterogeneous catalysts

. 8.Describe the role of catalytic converters on automobiles.

9.Describe enzymes as catalysts including their specificity, origin, molecular structure, and thermal

Lesson 38: Kinetics (Mechanisms) and Atmospheric Chemistry

text: 582-591, 745-754

what to know:

-concept of reaction mechanisms and related terms, §14-4, §14-5, (omit relationship to equilibrium)
-atmospheric chemistry and air pollution, §18-5
-role of catalysis in stratospheric ozone depletion and photochemical smog, §18-5
-acid rain and greenhouse effect, p-752

questions:

1. If the rate law expression for reaction, $A + 2B_2 \text{ ===> } C$ is, rate = $k[A]^o[B_2]^2$, we can conclude

2.Given the reaction, $CO + NO_2 \text{ ===> } CO_2 + NO$ (all gases), with a rate law, rate = $k[NO_2]^2$.
Consider the following possible mechanisms:

A. $CO + NO_2 \text{ ===> } CO_2 + NO$

B. $2NO_2 \text{ ===> } N_2O_4$ (fast)
$N_2O_4 + 2CO \text{ ===> } 2CO_2 + 2NO$ (slow)

C. $2NO_2 \text{ ====> } NO_3 + NO$ (slow)
$NO_3 + CO \text{ ====> } NO_2 + CO_2$ (fast)

D. $2NO_2 \text{ ====> } 2NO + O_2$ (slow)
$2CO + O_2 \text{ ====> } 2CO_2$ (fast)

a. Which of the mechanisms is (are) consistent with this rate law for the overall reaction?
b. Write the rate law expression for both elementary steps in mechanism B.

3.Write the rate law expression for the elementary step, $2\ NO_2(g) \text{ ====> } 2NO(g) + O_2(g)$

4.Given the overall reaction, $2NO + O_2 \text{ ====> } 2NO_2$, and the mechanism:
$NO + O_2 \text{ <===> } NO_3$ (fast) $NO_3 + NO \text{ ====> } 2NO_2$ (slow)

a. Write the rate law expression for the rate determining step.
b. What substance is the common intermediate and does not appear in the rate law expression for the overall reaction?

5.Discuss the chemistry of ozone in the atmosphere.

Lesson 39: Silicon and Silicates

text: 831-842

what to know:
- -what semiconductors are and how they work, §21-1
- -doping, pigments and phosphors, §21-1
- -silicates, their properties and uses, §21-2
- -silicones, p-842

questions:
1. What are silicones?
2. Numbers 5, 13, 15 on page 859

Lesson 40: Ceramics

text: 843-857

what to know:
- -ceramics, properties and uses, §21-3, §21-4, §21-5
- -superconducting ceramics, p-854 & 855

questions:
1. Discuss the chemistry of glass and cement.

2. Numbers 19, 21, 27, 39, 49 on pages 860 and 861

Lesson 41: Organic I

text: 931-940

what to know:

-structure, nomenclature and properties of various alkanes and their alkyl groups, §24-1
-concept of structural isomers, §24-1
-chemistry of petroleum, §24-1
-alkenes, alkynes, cycloalkanes and aromatic hydrocarbons, §24-1

questions:

1. Why are hydrocarbons nonpolar molecular substances? Why do alkanes normally have relatively low boiling points? Why are alkanes so chemically inert?
2. Organic compounds always contain carbon but often contain hydrogen, oxygen, nitrogen or a halogen. In each case, state the number of covalent bonds each can exhibit and show the various possible arrangements(single, double and triple bonds).
3. Write structural and abbreviated formulas for the following.
   a. 2-methylhexane b. 2,2-dimethylbutane c. 3-ethylpentane d. 2,3,4-trimethylheptane
4. Draw and name all: a. five structural isomers with the formula, C_6H_{14}.
   b. nine structural isomers with the formula, C_7H_{16}.
5. Which has the lowest octane rating? hexane, 2,3-dimethylhexane, 2,2,4-trimethylpentane, toluene
6. What are the products of the fractional distillation of petroleum crude oil?
7. Write the balanced equation showing the complete combustion of heptane.
8. Illustrate what is meant by the term unsaturated?
9. Write structural formulas for the following:
   a. tran§2-butene
   b. 5,5-diethyl-3-propyl-1-octene
   c. 3-heptyne
   d. m-dibromobenzene
   e. 2,4-dimethyl-1-ethylbenzene
   f. 3-phenylbutyne
   g. 2,4,6-trinitrotoluene(1-methyl-2,4,6-trinitro-benzene)
10. Show the structures of compounds formed when the following are added to propene.
    a. H_2 b. H_2O c. HCl d. NH_3 e. Br_2

11. What happens when soybean oil is hydrogenated?

Lesson 42: Organic II

text: 940-949

supplement: organic compounds (22)

what to know:

-structures, nomenclature and properties of halides, alcohols, ethers, aldehydes, ketones, acids, esters, amines and amides, §24-2, sup-22

questions:

1.Describe two methods for making ethanol.

2.Write structural formulas for:

a. 3-pentanol
b. a 2 carbon aldehyde
c. a 3 carbon amine
d. acetic acid
e. acetone
f. an ester made from propanoic acid and ethanol
g. a 3 carbon ether
h. a soap with 16 carbon atoms
i. an amide
j. an amino acid

3.What product are made when:

a. ethanol is oxidized (dehydrogenated)?
b. a ketone is reduced?
c. an ester is hydrolyzed?
d. an aldehyde is oxidized?
e. acetic acid is treated with sodium hydroxide?
f. ethanol is dehydrated?

4.How does a soap or detergent work?

5.What is a fat?

6.How is 1) vinyl chloride and 2) ethylene glycol made from petroleum and what is it used for?

7.Distinguish between primary, secondary and tertiary alcohols.

8.What is the commercial synthesis of acetic acid?

Lesson 43: Organic III

text: 949-954

what to know:

-insecticides, herbicides, analgesics, antibacterial agents, steroids, §24-3

questions:

1.Distinguish between A first, second and third generation@ insecticides.

2.Which common insecticide is an organophosphate?

3.Describe the chemical structure (functional groups) of the 1) herbicide, atrazine and 2) penicillin G.

4.Note the chemical similarity of morphine, codeine and heroin.

5.Name some important cholesterol derivatives

Lesson 46 & 47 (continued): Synthetic Polymers I

text: 960-968, 982-991

supplements:

-Ethylene and propene (23), Types of plastics (24), Organic compounds (25)

what to know:

-properties and classifications of polymers, §25-1
-how to make addition and condensation polymers, §25-1
-concepts of copolymers and crosslinking, §25-1
-examples of synthetic polymers, their uses and monomers,§25-3
-fibers, plastics, rubber, §25-3
-paint, p-986

questions:

1.Distinguish between synthetic and natural polymers and between addition and condensation polymers. Give examples of each kind.

2.Think of organic polymers in your life and list the properties of organic polymers(plastics) which make them so useful.

3.The United States chemical industry produces more ethylene (ethene) than any other organic compound.

a. What is the source of the ethylene?
b. List the products made from ethylene.

4.Draw the structure for the monomer from which polypropylene is made.

5.If the structure of ethylene is $CH_2=CH_2$, polyethylene is best represented as:
$-(-CH=CH-)-_n$ or $-(-CH_2-CH_2-)-_n$

6.Polyethylene is stable in acid while Dacron(a polyester) and nylon (a polyamide) are not. Explain.

7.Look at the structure of Teflon and tell why it is so "unsticky".

8.Sketch a portion of the polystyrene molecule.

9.Which of the following could be used to form an addition polymer?
a condensation polymer?
a. C_2H_6 b. C_2H_4 c. $HOCH_2CH_2OH$ d. $HOCH_2CH_3$

10.Write a balanced equation to illustrate the polymerization of 1,1-dichloroethylene. The product of

11.Describe the various types of polyethylene.

12.Discuss the chemistry of rubber.

Lesson 49: Polymers II - Natural Polymers

text: 968-981

what to know:

- -importance of natural polymers, §25-2
- -what amino acids are, §25-2
- -concept of chiral centers, §25-2
- -structure, classification and function of proteins, §25-2
- -structure and function of carbohydrates and nucleic acids, §25-2

questions:

1. What is the difference between D- and L-alanine?
2. Determine the number of chiral centers in the structure of cholesterol (page 954).
3. How many different tripeptides can be formed using only three different amino acids?
4. Why would you expect many proteins to be water soluble?
5. What role does the disulfide bridge play in protein structure?
6. How do enzymes work?
7. How many chiral centers in á-D-glucose?
8. How do cellulose and starch differ?
9. Exactly what information is encoded on DNA?
10. What is the genetic code?
11. What is the difference between globular and fibrous proteins?