CHEM141 – Final Exam Study Guide
Fall 2012

Chapters 1-2

1. *Isotope symbols and subatomic particles:* Complete the table below by writing the missing information.

Isotope symbol (include charge if applicable)	Charge	Number of protons	Number of electrons	Number of neutrons	Mass number
	0	9		10	
$^{31}_{15}P$	0				
	+3	27		30	
	-2	34			79
$^{99}_{43}Tc^{2+}$					
		16	18	16	

2. Use the periodic to answer the following questions.
   (a) Identify (write the name and symbol) four *transition metals*
   (b) Identify (write the name and symbol) two *alkaline-earth metals*
   (c) How many *metals* are there in group 13 or IIIB?
   (d) How many *metalloids* are there in group 14 or IVB?
   (e) Identify (write the name and symbol) four halogens
   (f) Write the name and symbol of the lightest metal
   (g) Name two groups in the periodic table that do not have any metals
3. The density of an object that has a mass of 38.47 g and displaces 14.62 mL of water when placed in a graduated cylinder is ________.

4. The volume of an object that has a mass of 125.4 g and a density of 7.9 g/cm^3 is ______.

5. If the outside temperature is 17.3 °C, what is the temperature in °F?

6. How many L are in 72.6 mL?

Chapter 7: Quantum Mechanical Model of the Atom

1. Key Concepts
   (a) Electromagnetic radiation and electromagnetic spectrum
      - Types of EMR and relative energies/frequencies/wavelength
        i. Decreasing energy and frequency
           Gamma rays > X-rays > Ultraviolet > Visible > Infrared > Microwave > Radiowaves
        ii. The trend is opposite for wavelength (E.g. Gamma rays have the shortest wavelength while radiowaves have the longest)

   (b) Properties of light as a wave
      - Wavelength, frequency, speed
      - Relationship between wavelength and frequency

   (c) Properties of light as a particle
      - Blackbody radiation
      - Einstein's Photoelectric effect
      - Planck's Quantum Theory
      - Relationship between energy, frequency and wavelength

   (d) Atomic emission spectra (also called line spectra) and Bohr's atomic model – meaning of n; relationship between n, energy of electron and atomic emission spectra

   (e) Atomic orbitals, shapes (be able to sketch s, p and d) and relative sizes

2. Practice problems
   (a) HW 2 and recitation worksheets on Chapter 7
   (b) Chapter 7 Problems by Topic (pp. 311-) 39, 41a, 43a, 59, 61, 63 and 65

Chapter 8: Periodic Properties of the Elements

1. Key Concepts
   (a) Orbital diagram – use blocks and arrows
      Goal: Be able to write the orbital diagram of an atom or ion given a copy of the Periodic Table of Elements
   (b) Electron configuration (Ex. Kr = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$)
      Goal: Be able to write the electron configuration of an atom or ion given a copy of the Periodic Table of Elements
   (c) Abbreviated electron configuration (Ex. Ca = [Ar] $4s^2$
      Goal: Be able to write the electron configuration of an atom or ion given a copy of the Periodic Table of Elements
   (d) s-block, p-block and d-block elements – where are they in the Periodic Table?
   (e) Isoelectronic = elements/ions with the same electron configuration (Ex. Ar and S^{2-})
   (f) Periodic trends
      - Atomic size (or radius) and ionic size
      - (First) ionization energy, IE

- Electronegativity, EN

  Goal: Be able to explain periodic trends (size, IE, EN) within a group and within a period

2. Practice problems
   a) HW2 and recitation worksheet on Chapter 8
   b) Chapter 8 Problems by Topic (pp. 354-) 43, 45, 47a, b & c, 49a, b & c, 51a, b & d, 53a, b & d, 55, 61, 63, 65 a & b, 73 and 75

Chapter 9: Chemical Bonding I – Lewis Theory

1. Key Concepts
   (a) Inner shell vs. valence shell electrons
   (b) Valence electrons and the Periodic Table of Elements
   Goal: Given an element, determine the number of valence e- from its location in the P.T.
   (c) Lewis electron dot symbols
   (d) Ionic bonding – illustrate using Lewis electron dot symbols
   (e) Covalent bonding (single, double, triple) – illustrate using Lewis electron dot symbols
   (f) Bonding electron pairs vs. "lone" electron pairs
   (g) Lewis structures of molecules and polyatomic ions
   - Strictly octet atoms (F, O, N, C)
   - Resonance structures
   - Exceptions to Octet Rule
     - Molecules with an odd number of electrons
     - Molecules where the central atom has less than octet
     - Molecules/ions where the central atom has expanded octet (period 3 nonmetals and beyond)

   (h) Formal charges

2. Practice problems
   a) Drawing Lewis structures and assigning formal charges examples in Recitation
   b) Chapter 9 Problems by Topic (pp. 394-) 37, 39 a & b, 49, 51, 53 a & b, 59 a &c (C is the central atom in Cl_2CO), 63 b & d, 65, 71 and 73 a & c.

Chapter 10: Chemical Bonding II

a. Electronegativity (EN) – table of ENs will be provided; use difference in EN between bonded atoms to determine if bond is polar or not
b. Polar and nonpolar <u>bonds</u>
   i. Nonpolar if $\Delta EN = 0$ to 0.40
   ii. Polar if $\Delta EN > 0.4$
c. Polar and nonpolar <u>molecules</u>
   i. ***Polar*** molecule only <u>if **BOTH** conditions below are met</u>:
      - Presence of polar bond (based on ΔEN) AND
      - Nonsymmetrical geometry – see molecular geometry below
   ii. ***Nonpolar*** otherwise
d. Molecular geometry and ideal bond angles – perfect symmetry (in blue) and not-so-perfect (with lone pairs; different atoms attached to the central atom – in orange below) symmetry

i. Linear (180^0)
ii. Trigonal planar (120^0) - Bent (1 lone pair, 2 bonds)
iii. Tetrahedral (109.5^0) – Trigonal pyramid (1 lp; 3 bonds); Bent (2 lp; 2 bonds)
iv. Trigonal bipyramid (axial-equatorial 90^0) and (equatorial-equatorial 180^0) – See-saw (1 lp; 4 bonds); T-shaped (2 lp; 3 bonds); Linear (3 lp; 2 bonds)
v. Octahedral (90^0)

2. Chapter 11: Liquids, Solids and Intermolecular Forces (IMFs)

a. General properties of solids and liquids
b. Intermolecular Forces (***IMFs***) of Attraction – holds molecules together in a liquid or solid
   - Not the same as chemical bonds (hold atoms together in a compound)
      i. H-bonding in polar molecules (only if **H** is attached to either F, O or N)
      ii. Dipole-dipole attractions in polar molecules
      iii. Dispersion forces in nonpolar molecules
c. Effect of IMFs on boiling point and vapor pressure of liquids
d. IMFs and solubility – "Like dissolves like" (Ex. Polar molecules dissolve in polar solvents)
   - Polar-to-polar mix because of dipole-dipole or H-bonding-H-bonding IMFs
   - Nonpolar-to-nonpolar mix because of dispersion-dispersion IMFs
   - Ionic compounds (many) dissolve in water due to ion-dipole attractions

3. Chapters 3 (Part) and 4: Stoichiomtery

a. Conversions and Avogadro's number (= 6.022 x 1023 particles/mol)
   i. g to mol; g to # molecules or atoms; mol to # molecules and vice versa
   ii. Molar mass
b. Balancing equations – practice and practice
c. Mass-to-mass stoichiometry, including ***limiting reactants*** and ***% yield***
   - Read notes and sample problems in the notes
   - *Example:* A 2.00 g sample of ammonia is mixed with 4.00 g of oxygen. The reaction involved is:

$$\mathbf{NH_3\,(g) + O_2\,(g) \rightarrow NO\,(g) + H_2O\,(g)}$$

Which is the limiting reactant? What is the maximum yield in grams of NO can be produced? If the actual yield of NO is 1.78 g, what is the % yield?
(*Answer:* After balancing the reaction, and determining the actual moles of each reactant, O2 should be found limiting. The theoretical yield of NO is 3.00 g and the % yield is 59.3.)

d. Work on Problems 49 and 51 (mass-to-mass stoichiometry), Tro, p. 173. Answers are found on page A-19.
e. ***Molarity*** (M) as a unit of concentration; Molarity and stoichiometry
   i. Calculate M given the mass or moles of solute and solution volume
   ii. Determine the mass or moles of solute present given M and V of a solution
   iii. Work on Problems 53, 55, 57 and 63. Again, answers are found on page A-19.
f. ***Dissociation*** (ionic compounds)
   i. Write the dissociation reaction for a given ionic compound
   ii. Examples: Complete and balance the following dissociation and ionization reactions.

(a) $Al(NO_3)_3$ (*aq*) $\rightarrow$

(b) $(NH4)_3PO_4$ (*aq*) $\rightarrow$

g. ***Precipitation*** reactions

   i. Complete and balance equations; Use solubility table to determine the precipítate or solid product, (s)

   ii. Work on Problem 71, p. 174. Check answers on page A-19.

   iii. Given a reaction or at least the reactants, write the following equations:

   - ***Molecular*** – write all reactants and products undissociated
   - ***Ionic*** – do not dissociate solid products
   - ***Net ionic*** – cancel spectator ions

   Examples:

   Molecular: $(NH_4)_2SO_4$ (*aq*) + $Pb(C_2H_3O_2)_2$ (*aq*) $\rightarrow$ 2 $NH_4C_2H_3O_2$ (aq) + $PbSO_4$ (s)

   Ionic: $(NH_4^+ + SO_4^{2-}) + (Pb^{2+} + C_2H_3O_2^-) \rightarrow (NH_4^+ + C_2H_3O_2^-)$ + $PbSO_4$ (s)

   Net ionic: Pb^{2+} (aq) + SO_4^{2-} (aq) $\rightarrow$ $PbSO_4$ (s)

   iv. Work on Problems 73, 75 and 77, p. 174. Check answers on page A-19