

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	
${}_{15}^{31}\text{P}$	0				
	+3	27		30	
	-2	34			79
${}_{43}^{99}\text{Tc}^{2+}$					
		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/cm³ 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

- HW2 and recitation worksheet on Chapter 8
- 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

- Inner shell vs. valence shell electrons
- 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.
- Lewis electron dot symbols
- Ionic bonding – illustrate using Lewis electron dot symbols
- Covalent bonding (single, double, triple) – illustrate using Lewis electron dot symbols
- Bonding electron pairs vs. “lone” electron pairs
- 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)
- Formal charges

2. Practice problems

- Drawing Lewis structures and assigning formal charges examples in Recitation
- 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 Cl₂CO), 63 b & d, 65, 71 and 73 a & c.

Chapter 10: Chemical Bonding II

- Electronegativity (EN) – table of ENs will be provided; use difference in EN between bonded atoms to determine if bond is polar or not
- Polar and nonpolar bonds
 - Nonpolar if $\Delta EN = 0$ to 0.40
 - Polar if $\Delta EN > 0.4$
- Polar and nonpolar molecules
 - Polar** molecule only if **BOTH** conditions below are met:
 - Presence of **polar bond** (based on ΔEN) AND
 - **Nonsymmetrical** geometry – see molecular geometry below
 - Nonpolar** otherwise
- 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°)
- ii. Trigonal planar (120°) - Bent (1 lone pair, 2 bonds)
- iii. Tetrahedral (109.5°) – Trigonal pyramid (1 lp; 3 bonds); Bent (2 lp; 2 bonds)
- iv. Trigonal bipyramid (axial-equatorial 90°) and (equatorial-equatorial 180°) – See-saw (1 lp; 4 bonds); T-shaped (2 lp; 3 bonds); Linear (3 lp; 2 bonds)
- v. Octahedral (90°)

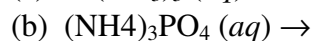
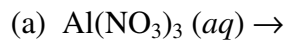
2. Chapter 11: Liquids, Solids and Intermolecular Forces (IMFs)

- 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: Stoichiometry

- a. Conversions and Avogadro’s number ($= 6.022 \times 10^{23}$ 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:

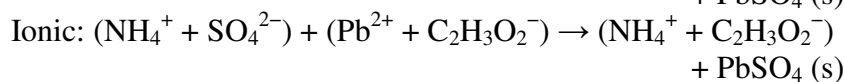
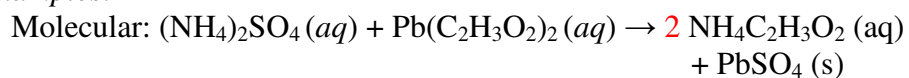
$$\text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{H}_2\text{O}(\text{g})$$
 Which is the limiting reactant? What is the maximum yield in grams of NO can be produced? If the actual yield of NO is 1.78 g, what is the % yield?
 (*Answer:* After balancing the reaction, 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.



g. **Precipitation** reactions

- i. Complete and balance equations; Use solubility table to determine the precipitate 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:



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