CHAPTER ONE HOMEWORK PROBLEMS

Assigned problems: 38, 40, 44, 52, 56, 58, 64, 68, 72, 76, 78, 84, 88, 98, 104

Helpful hints:

1. Table 1.1, page 16; Seven fundamental SI units of measure
2. Table 1.2, page 19; Prefixes for SI units
3. Page 25-26, significant figure rules (NOTE: Not covered in lecture)
4. Page 17-18; Temperature conversion formulas
5. Page 20; density formula

CHAPTER TWO HOMEWORK PROBLEMS

Assigned problems: 56, 58, 60, 64, 66, 70, 72, 74, 80, 82, 98

Helpful hints:

1. Section 2.6; Atomic number, isotopes and ions discussion
2. Section 2.7; Using the periodic table to classify elements
3. Section 2.8 and 2.9; Atomic mass, moles and molar mass

$${}^{A}_{Z}X$$

A- Mass # (# protons + neutrons)
Z- Atomic # (# protons)
X- Chemical symbol (e.g. C for carbon)

CHAPTER THREE HOMEWORK PROBLEMS-PART ONE-

Assigned problems: 30, 32a, 32c, 34, 40, 44, 46, 50, 54, 56

Some Helpful hints:

1. Ionic compounds are formed from one of the following
    a. Metal plus nonmetal
    b. Polyatomic ion plus something (metal or nonmetal)
2. Molecular (covalent) compounds are formed from two nonmetals

Nomenclature rules

1. Oxygen series of polyatomic ions
    a. Based upon the –ate form
    b. Add per in front if one more oxygen than –ate form
    c. Replace –ate with –ite if one fewer oxygen than –ate
    d. Add hypo- in front of the –ite form if one fewer oxygen than –ite form
    e. No oxygens= end in -ide
2. Acids and bases are considered to be ionic compounds
    a. Bases are named as a metal hydroxide (e.g. sodium hydroxide)
    b. Acids are named by applying following rules to the oxygen series name of the acid
        i. -ate is replaced by –ic in the acid name
        ii. -ite is replaced by –ous in the acid name
        iii. -ide is replaced by hydro- in front **and** –ic at the end

CHAPTER THREE HOMEWORK PROBLEMS-PART TWO-

Assigned problems: 58, 66, 76a, 76b, 80, 86a, 86b, 90, 92, 94, 98, 100b

Some helpful hints:

1. In formulas, treat each subscript of each element as a multiplier to the atomic mass of each element. Possible applications are below:
    A. Formula mass
    B. Percent composition
    C. Number of moles of an element in a compound
2. Empirical formula is the simplest formula for a compound
3. Molecular formula is the true formula of a compound

Balancing equations

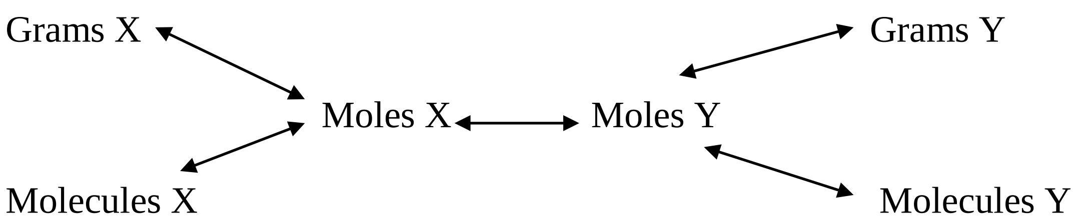
1. Balancing equations method 1
    a. Balance those atoms which occur in only one compound on each side
    b. Balance the remaining atoms
    c. Reduce coefficients to smallest whole integers
    d. Check your answer
2. Balancing equations method 2
    a. Identify most complex compound
    b. Balance this compound by placing 1 before it
    c. Balance remaining compounds using fractions
    d. Multiply fractions to obtain integers
3. Balancing equations method 3
    a. Create an element table to help identify what is not balanced

CHAPTER FOUR PROBLEMS – PART ONE -

Assigned problems: 26, 28, 32, 34, 40, 42, 44, 46, 50, 54, 60

Helpful Hints:

1. Section 4.2; Stoichiometry
2. Section 4.3; Limiting reactant, theoretical yield and percent yield
3. Section 4.4; Molarity and solution stoichiometry; dilution (M1V1=M2V2)
4. Mole highway



CHAPTER FOUR HOMEWORK PROBLEMS – PART TWO -

Assigned problems: 64, 66, 68, 72, 76, 80, 84, 86, 88, 122

Helpful hints:

1. Electrolytes
    a. Weak electrolytes dissolve sparingly in water
    b. Strong electrolytes produce many ions in water
        i. Strong acids
        ii. Strong bases (NOTE: Alkaline bases are not very soluble, but what does dissolve ionizes completely)
        iii. Soluble ionic compounds
    c. Nonelectrolytes do not produce ions in solution when they dissolve
2. Identifying types of reactions
    a. Precipitation reactions produce a solid as a product
    b. Acid-base reactions produce water and a salt (ionic compound) as a product
    c. Oxidation-reduction reactions involve changes in oxidation numbers **in both** directions
3. Types of equations
    a. Molecular equations- the ones that you are already familiar with
    b. Ionic equations- all dissolved strong electrolytes are separated into their ions
    c. Net ionic equations- ions that appear on both side of an equation are cancelled out
5. Section 4.9, Oxidation-reduction reactions and oxidation number rules
6. Solubility rules

Compounds Containing the Following Ions are Generally Soluble	**Exceptions** **(when combined with ions on the left the compound is insoluble)**
Li^+, Na^+, K^+, NH_4^+	none
NO_3^-, $C_2H_3O_2^-$	none
Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
SO_4^{2-}	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Compounds Containing the Following Ions are Generally Insoluble	**Exceptions** **(when combined with ions on the left the compound is soluble or slightly soluble)**
OH^-	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
S^{2-}	Li^+, Na^+, K^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}
CO_3^{2-}, PO_4^{3-}	Li^+, Na^+, K^+, NH_4^+

CHAPTER FIVE HOMEWORK PROBLEMS

Assigned problems: 30, 36, 38, 42, 44, 52, 54, 58, 66, 68, 72, 76

1. 1 atm= 760 mm Hg= 101, 325 Pa= 760 torr = 14.7 lb/in^2 (psi)
2. Combined gas law:

$$\frac{P1\ V1}{T1} = \frac{P2\ V2}{T2}$$

3. Ideal Gas Law: PV = nRT
4. density = $\frac{n \times M}{V} = \frac{P \times M}{R \times T}$
5. Dalton's law of partial pressures:

   $Xa + Xb = 1$; $Xa = n_A / n_A + n_B$

   $Pi = Xi\ P_{total}$; Pi= partial pressure, Xi= mole fraction

CHAPTER SIX HOMEWORK PROBLEMS

Assigned problems: 32, 40, 44, 46, 50, 56, 58, 64, 68, 72, 78, 82

Helpful hints:

1. 1 calorie = 4.184 J (exactly); 1 calorie is the amount of energy needed to raise the temperature of 1 gram of water 1 deg. Celsius
2. kinetic energy = $\frac{1}{2} mv^2$
3. work = force x distance= $- P\Delta V$
4. $\Delta E = q + w = q - P\Delta V$
5. $C = q / \Delta T$ (heat capacity)
6. $q = C \times m \times \Delta T$ $\quad\quad$ $\Delta T = T_{final} - T_{initial}$
7. Hess' law/enthalpy rules
    a. Reversing a reaction changes the sign of its enthalpy
    b. Multiplying a reaction by a number increases enthalpy by the same factor
8. Helpful tables
    a. Appendix II (Heats of formation); back of book
    b. Table 6.4, Specific heat and molar heat capacities
    c. Table 6.5, Standard heats of formation of substances

CHAPTER SEVEN HOMEWORK PROBLEMS

Assigned problems: 38, 40, 42, 48, 56, 58, 60, 62

Helpful Hints

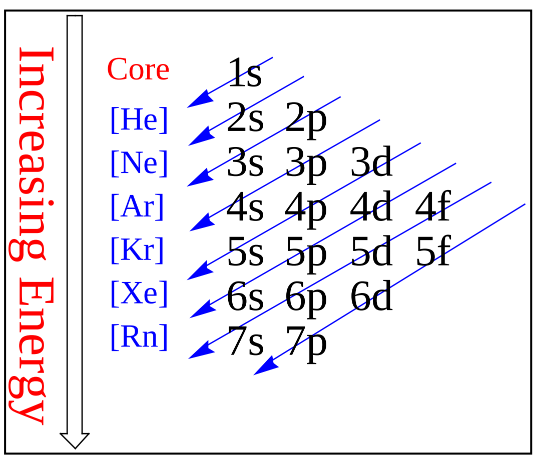
1. Constants used in this chapter
   a. $c = 3.0 \times 10^{8}$ m/s
   b. $h = 6.626 \times 10^{-34}$ J/s
   c. $J = kg \bullet m^2/s^2$
2. Equations used in this chapter
   a. $\lambda \times \nu = c$
   b. $E = h \times \nu$
3. Quantum number rules
   a. n; any positive integer
   b. l; 0 to n-1
   c. ml; - l to +1
   d. ms; ½ or -½

CHAPTER EIGHT HOMEWORK PROBLEMS

Assigned problems: 42, 44, 46, 48, 52, 56, 60a, 60b, 60d, 62, 64, 68, 72b, 72c, 72d 74

Helpful Hints:

1. Rules of Aufbau Principle:
   a. Lower n orbitals fill first.
   b. Each orbital holds two electrons; each with different ms.
   c. Half-fill degenerate orbitals before pairing electrons
2. Filling electron orbitals (see diagram below)
3. Ionic radii
   a. Cations are smaller than their parent elements
   b. Anions are larger than their parent elements
4. First ionization energy increases up and to the right
5. Beyond the first ionization energy, factors such as removing an electron from a filled valence shell will affect ionization energy
6. Electron affinity increases up and to the left, half-filled subshells and filled octets are examples of factors that do influence this observation



CHAPTER NINE HOMEWORK PROBLEMS

Assigned Problems: 40, 52, 54, 56, 60, 62, 64, 68, 72, 74

Helpful Hints

1. Figure 9.10, electronegativity values
2. % Ionic Character
    a. If ΔEN = 0; the bond is pure covalent
    b. If ΔEN 0.1-0.4 the bond is nonpolarcovalent
    c. If ΔEN 0.5-1.9 the bond is polar covalent
    d. If ΔEN ≥ 2.0 the bond is ionic
3. Drawing Lewis Dot Structures
    a. Rule 1: Count the total valence electrons
    b. Rule 2: Draw the structure using single bonds
    c. Rule 3: Distribute the remaining electron pairs around the peripheral atoms
    d. Rule 4: Put remaining pairs on central atom
    e. Rule 5: Share lone pairs between bonded atoms to create multiple bonds.
4. Formal charges= # valence electrons - ½ # bonding electrons - # nonbonding electrons

CHAPTER TEN HOMEWORK PROBLEMS

Assigned problems: 30, 32, 34, 36, 40 (don't sketch), 46, 50, 58, 66 (do it for carbon atoms only)

Helpful Hints

1. Table 10.1 (414-415); Molecular Shapes
2. Table 10.2 (421); Helpful for predicting molecular polarity
3. Table 10.3 (p. 436); Helpful for predicting hybridization schemes
4. Determining molecular shapes table:

Electron Groups	Lone Pairs	Bonds	Geometry	Examples
2	0	2	Linear	$BeCl_2$
3	0	3	Trigonal planar	BF_3
3	1	2	Bent	SO_2
4	0	4	Tetrahedral	CH_4
4	1	3	Trigonal pyramidal	NH_3
4	2	2	Bent	H_2O
5	0	5	Trigonal bipyramidal	PCl_5
5	1	4	See-saw	SF_4
5	2	3	T-Shaped	ClF_3
5	3	2	linear	I_3^-
6	0	6	Octahedral	SF_6
6	1	5	Square pyramidal	$SbCl_5^{2-}$
6	2	4	Square planar	XeF_4