CHEM 1110 – Fall 2021
Midterm Exam 2 Study Guide (Ch. 3-5)

This study guide is meant to provide only the barest direction as you study. Try to find practice problems from the textbook (both in the chapter text and in the end-of-chapter questions) rather than just relying on this guide. Note that most tables and equations will not be provided here or on the exam. You can find them in your textbook.

Chapter 3 – Ionic Compounds

- Recognizing ionic compounds
  - Metals bonded to non-metals
  - Polyatomic ions often don't include a metal
- Cations and anions
- Predicting ion charge based on position in the periodic table
  - Metals produce cations, non-metals produce anions
  - Elements will lose or gain enough electrons to satisfy the *octet* rule
  - Transition metals might take different charges in different compounds
- Ionic bonds create a crystal lattice (no discrete molecular unit)
- Ionic compounds are rigid, but brittle, and have high melting points
- Naming ions
  - Cations are just the metal name
  - For transition metals, include the charge in roman numerals, like this: Iron(III) for Fe^{3+}
  - Anions end in "-ide"
  - Polyatomic ions have their own names, but often end in "-ate" or "-ite" (table 3.3)
- Ions combine in definite ratios to form ionic compounds
  - The total positive and negative charges need to balance
  - Find the least common multiple to determine the right numbers of each ion
- Naming ionic compounds
  - Just combine the two ions names – Cation first, then anion
  - To determine the charge on a transition metal, look at the total amount of negative charge that it must balance

Chapter 4 – Molecular Compounds

- Covalent bonds involve *sharing* electrons
- Compounds with covalent bonds are called *molecules*
- Forces involved in a covalent bond:
  - The positively charged nuclei repulse each other when they are too close
  - There are no interactions at all when they are too far
  - At the right bond distance, the electrons are attracted to *both* nuclei, forming the bond
- Electrons are shared when orbitals from both bonding partners overlap
- Predicting the number of bonds based on position in the periodic table
  - Elements will form enough bonds to satisfy the octet rule
  - Each bond brings one additional electron
  - This trend is not a hard rule, and some compounds will form many more bonds than expected
- Multiple bonds
  - Sometimes molecules need to bond multiple times to the same partner in order to satisfy the octet rule
  - Consider O_2 (double bond) and N_2 (triple bond)
  - Each double bond shares 4 electrons, and each triple bond shares 6
- For coordinate covalent bonds, both electrons come from the same atom
- Molecular compounds have widely ranging properties, such as melting points and reactivities
- Ways of representing molecular compounds:
  - Chemical formula
  - Condensed structural formula
  - Lewis structure
- Drawing Lewis Structures
  - 5 steps to drawing a Lewis structure:
    a) Count up the total number of valence electrons (taking into account any net charge)
    a) Draw single bonds from the central atom to all peripheral atoms
    a) Fill the octets of the outer atoms with lone-pairs (H doesn't need any)
    a) Place any remaining electrons onto the central atom

        a) Satisfy the octet rule for the central atom, if needed, by converting outer lone-pairs into multiple-bonds with the central atom

  - Finding the geometry from the Lewis Structure
    a) Draw a valid Lewis structure
    a) Count the number of electron domains
    a) Count the number of bonds vs. lone pairs
    a) Consult Table 4.2 (or your memory)

- Electronegativity, and how it varies across the periodic table
- Polar vs non-polar bonds
- polar vs non-polar molecules
  - lone pairs disrupt symmetry and lead to polar molecules (NH_3)
  - Bonds to different atom types disrupt symmetry and lead to polar molecules (CH_2O)
- Naming binary molecular compounds
  - The first element named is the least electronegative
  - The second element named ends with "-ide"
  - Indicate how many of each element with the prefixes in table 4.3
  - Leave off "mono-" for the first element only

Chapter 5 – Classification and Balancing of Chemical Reactions

- How to balance chemical reactions
  - Start with elements that appear in *only* one place on each side of the equation
  - Often it is best to balance oxygen last
  - double everything if necessary to get rid of fractional coefficients
- Classes of chemical reactions
  - Precipitation reactions
  - Acid-base reactions
  - Oxidation-reduction reactions
- Precipitation reactions
  - Solubility rules in Table 5.1
  - Reactants will switch binding partners (cations and anions)
  - Check the solubility of the new products
  - Net ionic equations will eliminate the spectator ions

- Acid/base neutralization reactions
  - Acids produce H_3O^+ in water
  - Bases produce OH^- in water
  - Acids and bases react together to produce neutral solutions with salt, and sometimes water
- Reduction-oxidation reactions
  - Redox reactions involve the transfer of electrons
  - Electrons are tracked with oxidation numbers
    * Elemental forms have oxidation states of 0
    * H usually takes +1 and O usually takes -2
    * The sum of all oxidation numbers should equal the overall charge
  - OIL-RIG – Oxidation is losing electrons, reduction is gaining electrons
  - See how oxidation states change to identify which elements were oxidized and which were reduced
  - Oxidizing agents and reducing agents are identified by their effect on their reaction partner