Things to Know for Exam 2

Chem 30A, Fall 2019

Chapter 4

1. Predict the charges of main-group monatomic ions (based on what group they are in).
2. Review the names of the elements (see the handout of element names and symbols).

Additional items:

- Estimate values from a graph accurately. Include the correct number of decimal places. Refer to the estimation handouts, the graphing assignment and answers, and the graph estimation problems on Exam 1.
- Determine the slope of a graph accurately. (Choose 2 points on the line far apart from each other, estimate the x and y values, and then find the slope, including units.)

Chapter 5

1. The properties of compounds are completely different than the properties of the elements in the compounds. (Compounds are **not** just mixtures of elements – the elements are chemically combined to make something new with different properties.)
2. Given a formula, state how many atoms of each element are present. [Ch. 5 # 31-40]
3. Given a molecular formula, write the empirical formula. (The empirical formula shows the lowest whole number ratio of atoms.) [Ch. 5 # 41-42]
4. Memorize the 7 diatomic elements.
5. Classify different elements as either atomic or molecular. (Some elements consist of individual atoms, and some consist of molecules.) [Ch. 5 # 43-44, 47-50]
6. Classify compounds as either ionic or molecular. (Ionic compounds consist of ions, and their smallest unit is the formula unit. Molecular compounds consist of molecules.) Ionic compounds contain a metal and a nonmetal (or recognizable ions), like NaCl or NH_4NO_3. Molecular compounds contain just nonmetals, like $C_{12}H_{22}O_{11}$. [Ch. 5 # 45-52]

7. Memorize the polyatomic ions: name, formula and charge. Given the name, write the formula and charge. Given the formula and charge, write the name.
8. Given the formulas and charges of ions, write the formula of the ionic compound containing those ions. (Look at the charges of each to figure out how many of each ion are in the formula. Compounds must be uncharged overall, so the total charges of all of the cations and anions in the compound have to add up to zero.) Example: if a compound contains Na^+ and Cl^-, the formula is NaCl. If it contains Mg^{2+} and Cl^-, the formula is $MgCl_2$.
9. Given the atoms in an ionic compound, predict their charges and then predict the formula of the compound. (This is a combination of items #1 and #9 above.) [Ch. 5 # 53-58]
10. Be able to recognize and name ionic compounds, binary molecular compounds, binary acids, and oxyacids. (First, decide what type of compound it is, then name it according to the appropriate rules.)
11. **Naming ionic compounds**: name the cation and then the anion. Example: $CaCl_2$ is calcium chloride, SnF_4 is tin (IV) fluoride. [Ch. 5 # 59-66]
12. **Naming binary molecular compounds**: Use prefixes. CO_2 is carbon dioxide, CO is carbon monoxide, S_2Cl_2 is disulfur dichloride. [Ch. 5 #71-72, 75-76]
13. **Naming binary acids**: "hydro___ic acid." HCl is hydrochloric acid. [Ch. 5 #77-78]
14. **Naming oxyacids**: what is the name of the anion in the acid? Change the ending. – "-ate" changes to "ic acid." "-ite" changes to "ous acid." HNO_3 is nitric acid. HNO_2 is nitrous acid. [Ch. 5 # 77-82]
15. **Given the name of a compound, be able to write the correct formula**. Be able to do this for all types of compounds: ionic, binary molecular, binary acids, and oxyacids. [Ch. 5 # 6970, 73-74, 81-82]
16. Calculate the formula mass for a compound, given the formula (add up all atomic masses.) [Ch. 5 # 83-84]

Chapter 6

1. Use Avogadro's number (6.022×10^{23}) to convert between number of atoms/molecules/formula units and number of moles. [Ch. 6 # 17-24]
2. Calculate molar mass for any substance. (Add up atomic masses.) Units are grams per mole.
3. Convert number of grams to number of moles of a substance (or the reverse) using molar mass as a conversion factor. [Ch. 6 #25-34]
4. Convert from grams of X $\leftrightarrow$ moles of X $\leftrightarrow$ molecules of X (or vice versa) [Ch. 6 # 35-56]

5. Use the chemical formula as a conversion factor to convert between number of moles of a compound and number of moles of an element in the compound. [Ch. 6 #59-66]
6. Given the mass of a compound, calculate the number of grams of an element in that compound, or vice versa. [Ch. 6 #67-70]
7. Calculate the mass percent of a given element in a compound, given the formula of that compound. [Ch. 6 #79-86]
8. Given the mass percent of an element in a compound, be able to write it as a conversion factor, and be able to use that conversion factor in a problem. If a compound is 17.5 % N, you can write the conversion factor 17.5 g N/100 g compound. For example, "What mass of nitrogen is present in 200 mg of the compound?"

Chapter 7

1. Given a chemical reaction, balance the equation. (Put coefficients in front of the formulas. Do not change any formulas. The number of atoms of each element on both sides of the equation must match.) [Ch.7 #47-56, 85-86]
2. Given a description of a chemical reaction in words, write the formulas and the equation, and then balance the equation. (In order to do this, you will need to remember which elements are diatomic, and you will need to be able to write the formulas of different types of compounds from their names. This is covered in Chapter 5.) [Ch. 7 #35-46]
3. Given the solubility rules for ionic compounds, be able to determine whether various ionic compounds are soluble or insoluble in water. [Ch. 7 #57-62]
4. Given the formulas or names of two ionic compounds, be able to determine the formulas of the products and the equation for the precipitation reaction. (First, what ions are present? Write them down, including charge. Second, switch the ions and write the new formulas for the products. Third, check the solubility rules to determine which ones are soluble or insoluble. Fourth, write and balance the equation.) [Ch. 7 #63-68]
5. Given a "molecular" equation for a precipitation reaction, write the complete/total ionic equation and the net ionic equation. (Soluble ionic substances are written as separate ions. For the net ionic equation, cancel out the spectator ions.) [Ch. 7 #71-72, 75-76]
6. Combine skills #4 and 5. [Ch. 7 #93-96]
7. Given the formulas of an acid and a base, predict the products and write the balanced equation for the reaction. (The ions switch, and one of the products is H_2O.) [Ch. 7 #77-80]
8. Be able to write the balanced equation for any of the gas-forming reactions (ions switch). You will probably need to memorize the patterns.

A carbonate compound + an acid $\rightarrow$ $H_2O_{(l)}$ + $CO_{2\,(g)}$ + an ionic compound
A bicarbonate compound + an acid $\rightarrow$ $H_2O_{(l)}$ + $CO_{2\,(g)}$ + an ionic compound
A sulfide compound + an acid $\rightarrow$ $H_2S_{(g)}$ + an ionic compound
A compound containing SO_3^{2-} or HSO_3^- + an acid $\rightarrow$ $H_2O_{(l)}$ + $SO_{2\,(g)}$ + ionic compound
A NH_4^+ compound + a OH^- compound $\rightarrow$ $NH_{3\,(g)}$ + $H_2O_{(l)}$
[Ch. 7 #81-82]

9. Recognize some types of oxidation-reduction reactions. If an (uncharged) element appears by itself somewhere in the reaction, it's a redox reaction. [Ch. 7 #83-84]
10. Identify/classify reactions as combination reactions, decomposition reactions, single-replacement reactions, double-displacement reactions, or combustion reactions. Some of these are sometimes also redox reactions. (There are other reaction types, but we haven't learned them yet.) For double displacement reactions, classify them further as precipitation, acid-base, or gas-forming reactions. [Ch. 7 #89-92, 99-100]
11. Be able to write and balance the equation for a combustion reaction for any organic compound (the compound contains C, H, and maybe O)
    Organic compound + $O_{2\,(g)}$ $\rightarrow$ $CO_{2\,(g)}$ + $H_2O_{(l)\text{ or }(g)}$
12. You still need to know the names and charges of the ions, including polyatomic ions. (If you don't know this, you probably won't be able to determine the new formulas in double-displacement reactions – items #4-8 on this list of Chapter 7 skills.)

Chapter 8

1. Using a balanced equation, you can relate moles of reactants and products to each other. * Balance any equation to be used in a stoichiometry problem! *
2. Convert moles of one substance to moles of another substance, using the mole ratio from the balanced equation. [Ch. 8 #15-30]
3. Stoichiometry problems: Given an unbalanced equation, balance it. Convert grams of one substance to grams of another substance using the balanced equation. First convert grams of X to moles of X using the molar mass of X as a conversion factor. Then convert moles of X to moles of Y using the coefficients of the balanced equation. Then convert moles of Y to grams of Y using the molar mass of Y as a conversion factor. [Ch. 8 #31-42]
   grams of X $\leftrightarrow$ moles of X $\leftrightarrow$ moles of Y $\leftrightarrow$ g of Y
4. Do variations of stoichiometry problems. You might use density to convert between volume and mass. You might use Avogadro's number to convert between number of moles and number of molecules or formula units. [Ch. 8 # 84, 91, 92]

5. Given the masses of both reactants, determine the amount of product made: this is a limiting reactant problem. Convert grams of the first reactant to grams of product. Then convert grams of the second reactant to grams of product. (This will be like two stoichiometry problems.) Compare the two answers. The smaller of the two answers is the amount of product that will be made in theory. The reactant that produces the smaller amount of product is the limiting reactant, and it will get used up first. [Ch. 8 #43-68]
6. Calculate percent yield, given the actual yield. The theoretical yield is the amount of product obtained via a calculation. [Ch. 8 #59-66]
7. If you are told the percent yield, be able to write it as a conversion factor and use it in a problem. For example, 91.5% yield means 91.5 g actual/100 g theoretical.
8. If a reaction is given with a value of ΔH next to it, you can write conversion factors that relate the amount of heat and the number of moles of reactant. Given the amount of reactant, calculate the amount of heat absorbed or given off. Be able to specify whether heat is absorbed or given off. OR Given the amount of heat, determine the mass of a reactant needed. [Ch. 8 # 69-76]

Chapter 16

1. What is oxidation? What is reduction? List several ways to tell. [Ch. 16 # 3]
2. Be able to find the oxidation number for different atoms in a given formula. [Ch. 16 # 45-56]
3. Given a reaction, determine what is being oxidized and what is being reduced. (Best way: find all oxidation numbers and see what is increasing and what is decreasing.) [Ch. 16 # 33-44, 57-60]
4. Identify the oxidizing agent and the reducing agent in a reaction (and define what that means). [Ch. 16 # 4, 37, 38 43, 44]
5. What is the "activity series?"
6. For a single replacement reaction $A + BC \rightarrow B + AC$, this reaction will happen if A is more active than B. If B is more active, the reaction will not occur.
7. Given info about relative activity of two elements, determine whether a single replacement reaction will occur or not. [Ch. 16 # 77, 78, 83, 84]
8. Given into about whether or not a single replacement reaction occurs, determine which element is more active.

Chapter 9

1. Electrons do not orbit the nucleus in a definite path. We can't know anything about the exact position of the electron, but we can determine probability maps for the location of electrons in atoms. These are called orbitals.
2. Sketch the general shape of an s orbital, a p orbital, and a d orbital.
3. Know how many orbitals are in a set. There is one orbital in any s subshell, 3 orbitals in a p subshell, and 5 orbitals in a d subshell.
4. What does the value of n represent? Higher n, larger orbital, higher energy, electrons average further from nucleus. [Ch. 9 #43-48]
5. Write the electron configuration or noble gas notation or arrow/orbital diagram for any atom. It's helpful to refer to the periodic table to remember the order of filling of orbitals. [Ch. 9 #49-76]
6. Determine the number of valence electrons for any main-group element.
7. Filled shells of electrons are especially stable.
8. Atoms tend to lose e^-, gain e^-, or share e^- so as to get a filled shell of electrons.
9. Rank atoms in order of size (look at where they are positioned on the periodic table. The closer they are to the upper right corner, the smaller they are. Bottom left corner: larger atoms. [Ch. 9 #81-84
10. Rank atoms in order of ionization energy. (Higher IE – harder to remove outermost electron.) [Ch. 9 #77-80]
11. Rank atoms in order of metallic character. [Ch. 9 #85-88]