Things to Know for Exam 3

Chem 30A, Fall 2019

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]

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]

Chapter 10

1. What is a covalent bond?
2. What are single, double, and triple bonds?
3. Memorize the preferred number of bonds for groups 4-7. (There are lots of exceptions.)
4. Draw Lewis structures for molecules or polyatomic ions. Check to make sure the number of valence electrons is correct, and give all atoms octets if possible. [Ch. 10 #47-60, additional problem 1]
5. Boron is often an exception to the octet rule. It often forms just 3 bonds and has no lone pairs. [Ch. 10 # 59-60]
6. Draw resonance structures when needed. (When there is a multiple bond and two or more equivalent locations for it, draw resonance structures.) [Ch. 10 #55-58]
7. Given a large organic molecule's skeleton structure, complete the Lewis structure: give all atoms their normal number of bonds and lone pairs. (No need to count the total number of valence electrons for these.) [Ch. 10, additional problem 2]
8. Determine the shape and bond angles for a molecule or polyatomic ion. (First, you need the Lewis structure.) Draw a 3-dimensional sketch of the molecule. [Ch. 10 #61-76, additional problem 3]
9. Given a large organic molecule's skeleton structure, complete the Lewis structure and state the shape and bond angle around each central atom. [Ch. 10 additional problem 4]
10. What is the difference between a polar bond and a nonpolar bond?
11. Given electronegativity values, state whether a given bond is nonpolar, polar, or ionic. [Ch. 10 #81-84]
12. Determine whether a molecule is polar or nonpolar. (First, you need the Lewis structure and need to figure out the shape.) If it's symmetric, it's nonpolar. Asymmetric molecules are often polar. The key question: do the dipole moments cancel out or not? [Ch. 10 #85-92, additional problem 3]

Chapter 11

1. Postulates of Kinetic Molecular Theory. What does the kinetic energy depend on? [Ch. 11 #3, 4]
2. Definition of pressure.
3. What is atmospheric pressure? Is it constant?
4. Converting between different units of pressure. (Equivalence statements will be given, for example 1 atm = 760 mmHg) [Ch. 11 #23-32]
5. Why do gases exert pressure?
6. Boyle's Law. $P_1V_1 = P_2V_2$ What happens to volume if pressure goes up? What is the underlying reason?

7. Be able to do problems involving Boyle's law. Solve for the missing variable. (P and V are changing, n and T are not changing.) Be able to explain why your answer makes sense, referring to what the molecules are doing. [Ch. 11 #33-38]
8. What is Charles's Law? What is the equation? Must use T in kelvins.
9. What happens to V if temp goes up? Explain the underlying reason.
10. For problems involving V and T changing (but n and P constant), be able to solve for the missing variable. Be able to explain why your answer makes sense, referring to what the molecules are doing. [Ch. 11 #39-44]
11. What is the combined gas law? Use the combined gas law to solve for the missing variable. (P, V, and T are all changing. Number of moles is constant.) [Ch.11 #51-58]
12. What is Avogadro's law? How does volume relate to moles of gas? Why? (P and T are constant)
13. Be able to use Avogadro's law to solve for V or n if P and T are constant. [Ch. 11 #45-50]
14. What is STP? (What T and P?)
15. If a gas is at STP, any gas has a "molar volume" of 22.4 L/1 mole. This can be used as a conversion factor if the gas is at STP. [Ch. 11 #83-88]
16. Ideal gas law: $PV = nRT$ (Given R.) Be able to solve for any variable. Sometimes you also need to convert units of things before plugging in to the equation. [Ch. 11 #59-68]
17. Finding molar mass of a gas: given grams. Given P, V, and T, use $PV = nRT$ to solve for the number of moles. Then divide the number of grams by the number of moles. The result is the # g/mol, which is the molar mass. [Ch. 11 #69-72]
18. Mixtures of gases: the pressure of each gas added up equals the total pressure. [Ch. 11 #73-78]
19. Pressure of a gas in a mixture = the fraction of molecules of that gas × total pressure. P_A = (fraction A)(P_{total}) [Ch. 11 #79-82]
20. Be able to do stoichiometry problems involving gases. Make sure to balance the equation. You can use $PV = nRT$ to find the number of moles of a gas or to solve for V, P, or any other gas property. Step A: balance equation. Step b: Find moles of whatever you can. Step c: use the mole ratios from the balanced equation to convert to moles of what you want. Step d: convert to the final units. (If the gas is at STP, you can use 22.4 L/1 mole as a conversion factor.) [Ch. 11 #89-98]

Chapter 12

1. Discuss the arrangement and movement of particles in liquids and solids. How are they different from gases? [Ch. 12 #4, 5]
2. Definition of intermolecular forces. [Ch. 12 #1]
3. What is surface tension? How does it relate to intermolecular forces? [Ch. 12 #7]
4. What is viscosity? How does it relate to intermolecular forces? [Ch. 12 #8]

5. What is evaporation? What does the rate of evaporation depend on? [Ch. 12 #10, 35]
6. Explain how and why evaporation depends on intermolecular forces. [Ch. 12 #13]
7. Explain how and why evaporation depends on temperature. [Ch. 12 #15]
8. Evaporation tends to cool the remaining liquid – why? What is a practical example of this? [Ch. 12 #16]
9. What is the meaning of the term "volatile"?
10. Define "vapor pressure." What does it depend on?
11. What happens when the vapor pressure equals the external pressure?
12. What is the "normal boiling point"?
13. Calculate the amount of energy absorbed or released during vaporization or condensation, given the value of ΔH_{vap}. OR calculate the amount of substance that can be vaporized/condensed, given some amount of energy. [Ch. 12 #49-56]
14. Phase changes happen at a constant temperature. Be able to sketch the graph for temperature vs. heat added for a substance (usually, we draw it for water). [Ch. 12 #39-40]
15. Sketch the shape of the curve in a graph of vapor pressure vs. temperature for a liquid. Explain the significance. [Ch. 12 #47-48]
16. Calculate the amount of energy absorbed or released during melting or freezing, given the value of ΔH_{fus}. OR calculate the amount of substance that can be melted/frozen, given some amount of energy. [Ch. 12 #57-62]
17. What is "sublimation"? What is the opposite of sublimation?
18. Know the names of each of the different phase changes, and know whether each process is exothermic or endothermic.
19. List and explain all of the different types of intermolecular forces. (See the yellow handout on IMFs!)
20. Which types of molecules have London dispersion forces? Which types of molecules have dipole-dipole forces? Which types are able to hydrogen bond? [Ch. 12 #63-68]
21. Given structures of molecules, state what types of intermolecular forces they have. [Ch. 12 #63-68]
22. Given a list of substances, rank them in order of strength of intermolecular forces or in order of boiling point or in order of heat of vaporization or in order of vapor pressure. Explain your reasoning thoroughly. [Ch. 12 #69-74, 87-88, practice problems on IMFs #1, 2, 3]
23. Recognize whether a solid is molecular, ionic, or metallic. [Ch. 12 #79-82]
24. For each of the types of solids (molecular, ionic, metallic, or covalent network), what are the particles and what holds things together?
25. For each of the types of solids, would you expect it to have a low, medium, or high melting point? Why? [Ch. 12 #83-86]
26. List all of the unusual properties of water.

Chapter 13

1. A solution is a homogeneous mixture.
2. What is the difference between the solute and the solvent? [Ch. 13 #3, 25, 26]
3. Like dissolves like [Ch. 12 #75-78, Ch. 13 #23-24, 27-28, Practice Problems on IMFs #4]
4. Given two substances, predict whether or not they will mix/form a solution. [Ch. 12 #75-78, Ch. 13 #23-24, 27-28, Practice Problems on IMFs #4]
5. Given a list of substances, rank them in order of solubility in water or in a nonpolar solvent. (More hydrogen bonding groups per carbon = more soluble in water. The order or ranking of solubility in a nonpolar solvent will be the opposite order.) [Practice Problems on IMFs #5]
6. If the solubility is given as a number with units, then this is the maximum that can dissolve.
7. Difference between saturated, unsaturated, and supersaturated solutions. How are each of them made? [Ch. 13 #6, 31-36]
8. What is the difference between an electrolyte solution and a nonelectrolyte solution? What happens when ionic substances dissolve? What about molecular substances? [Ch. 13 # 29, 30]
9. How does the solubility of a solid usually depend on temperature?
10. How does the solubility of a gas depend on temperature? How does the solubility of a gas depend on pressure? [Ch. 13 # 13, 39, 40]
11. Be able to calculate the mass percent of a solute in a solution. [Ch. 13 # 41-46]
12. Given the mass percent, be able to write it as a conversion factor. (10.0 % $NaCl_{(aq)}$ means 10.0 g NaCl / 100 g solution.)
13. Given the mass percent, be able to calculate either the amount of solute in a specific solution, or the amount of solution containing a given amount of solute. [Ch. 13 # 45-58]
14. Be able to calculate the molarity of a solution. [Ch. 13 #59-64]
15. Given the molarity of a solution, be able to write it as a conversion factor.
16. Given the molarity of a solution, determine either the volume of solution or the amount of solute. (Use M as a conversion factor). [Ch. 13 #65-80]
17. Dilutions: add water to make a less concentrated solution.
18. Be able to do problems involving dilutions. Use $M_1V_1 = M_2V_2$. Remember that V_2 is the total final volume. (find the concentration after a dilution, find the volume needed, find the volume of water needed, etc.) [Ch. 13 #81-88]
19. Volume (in liters) × Molarity (moles/L) = moles of solute.
20. Stoichiometry problems involving solutions: there are many variations of this type of problem. Step 1: balance the equation. Step 2: find moles of whatever you can. Step 3: Use the balanced equation to convert to moles of the desired substance. Step 4: Convert moles to the desired unit. In this chapter, you will use molarity as a conversion factor between volume of solution and moles of solute. [Ch. 13 #89-96, 115]

21. Osmosis: predict which way water will flow across a semipermeable membrane. (Water flows to try to equalize the concentration of solute particles on each side of the membrane.) [Ch. 13 #131]
22. Blood and body fluids 0.307 osM. Will cells shrink, swell, or stay the same in different solution concentrations?

Chapter 14

1. Know the general properties of acids and bases. [Ch. 14 #2, 5]
2. State some examples of acids and bases.
3. What are the Arrhenius definitions of an acid and a base? [Ch. 14 #7, 8]
4. What are the Brønsted-Lowry definitions of an acid and a base? [Ch. 14 #9]
5. Given a proton-transfer equilibrium, label which substances are acids and which are bases (in both the forward and the reverse directions). [Ch. 14 #31-32]
6. Given the conjugate acid, state the formula of the conjugate base and vice versa. [Ch. 14 #33-38]
7. Given the formulas of an acid and a base, write and balance the equation for the neutralization reaction that would occur between them. (This was covered in section 7.8. What are the ions present? Switch the ions and get new formulas.) [Ch. 14 #39, 40]
8. Write the equation for the gas forming reaction between a carbonate or bicarbonate compound and an acid.
9. Write the equation for an acid reacting with a metal. [Ch. 14 #41, 42]
10. Titration problems: Write the overall equation and balance it. Find moles of the first substance. Convert to moles of the other substance. Finish the problem, calculating either M or V of the new solution. (You will use M as a conversion factor. You might also need to calculate M of one of the substances.) Important: do not use the dilution formula. It does not apply to titrations or reactions in general. [Ch. 14 #47-52]
11. What is the difference between a strong acid and a weak acid? State some examples of each. [Ch. 14 #53, 54]
12. Determine the hydronium concentration in a solution of a strong acid. [Ch. 14 #55, 56]
13. What is the difference between a strong base and a weak base? State some examples of each. [Ch. 14 #57, 58]
14. Determine the hydroxide concentration in a solution of a strong base. [Ch. 14 #59, 60]
15. What is autoionization?
16. Know that $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$. The number will be given.
17. Given the hydronium ion concentration, calculate the hydroxide concentration and vice versa. [Ch. 14 #63-66]

18. Determine if a solution is acidic, basic, or neutral from the hydronium concentration or the pH or the pOH or the hydroxide concentration. [Ch. 14 #61-68]
19. Given the hydronium concentration, calculate the pH of the solution, and vice versa. [Ch. 14 #69-78]
20. Be able to calculate the pOH. [Ch. 14 #79-84]
21. What does a buffer do? What are the components of a buffer? [Ch. 14 #85, 86, 89, 90]
22. Given the buffer components, write the equation for the reaction that happens when a strong acid is added to it. (Strong acid reacts with the base component of the buffer.) [Ch. 14 #87]
23. Given the buffer components, write the equation for the reaction that happens when a strong base is added to it. (Strong base reacts with the acid component of the buffer.) [Ch. 14 #88]
24. Explain how a buffer is able to keep the pH of a solution from changing too much.