Chemistry 2nd Semester Final Exam Review

Chemical Bonds

1. Give a physical description of how the atoms and molecules are arranged in solids, liquids, and gases.
   A: *In a liquid, the forces between the molecules are weak, so the molecules are free to move around. In a solid the forces between the molecules (intermolecular forces) are stronger so they are held in place.*

2. Use the information from the previous question to describe the shape and volume for each phase (state) of matter.
   A: *Solids have a repeating geometric pattern (lattice structure) in which the particles vibrate in place. They have a definite shape and a definite volume.*
   Particles within a liquid slide past each and are close, but are not held in place. They have a definite volume, but no definite shape.
   Gases particles move chaotically bouncing off one another. They have neither a definite shape nor a definite volume.

3. List the three phases of matter in order of increasing intermolecular attractions.
   A: *Gas, liquid, solid.*

4. Why do the atoms and molecules in liquids move in a random pattern relative to one another instead of being in a solid form?
   A: *Their kinetic energy (temperature) is high enough to overcome the attractive forces that might otherwise hold them in a repeating pattern. The attractive forces aren't strong enough to hold them together at that temperature.*

Conservation of Matter and Stoichiometry

5. The atomic mass unit and the mole are defined using what isotope? A: *Carbon-12*

6. Balance the following equations:
   a. $Al_2O_3 + Cl_2 + C \rightarrow AlCl_3 + CO$
      A: $Al_2O_3 + 3Cl_2 + 3C \rightarrow 2AlCl_3 + 3CO$
   b. $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$
      A: $2C_4H_{10} + 13O_2 \rightarrow 8\,CO_2 + 10H_2O$
   c. $P + O_2 \rightarrow P_2O_5$
      A: $4P + 5O_2 \rightarrow 2P_2O_5$

7. What isotope is used to define one mole? A: *Carbon-12*

8. How many particles (atoms or molecules) are in one mole? A: 6.02×10^{23}

9. Find the molar mass for the following chemical formulas:
   a. ZnF_2 A: *103.377 g/mol*
   b. $Al_2(SO_4)_3$ A: *342.151 g/mol*
   c. NH_4IO_3 A: *192.941 g/mol*

10. How many moles are in 4.37 kg of NaOH? A: *109.3 mol*

11. How many atoms are in 2075 g of He? A: 3.12×10^{26} *atoms*

12. What volume will 92.31 g of CO_2 occupy at STP? A: *46.985 L*

13. How many grams of sodium hydroxide are needed to react with 6.23 g of barium bromide?
    $NaOH + BaBr_2 \rightarrow NaBr + Ba(OH)_2$
    A: *1.68 gNaOH*

14. In the reaction, $2\,Mg + O_2 \rightarrow 2\,MgO$, if 100.0 g of magnesium reacts with 50.0 g of oxygen, what mass of product is produced?
    A: *126 g MgO*

Gases and Their Properties

15. What is the kinetic molecular theory (KMT)?
    A: *The Kinetic Molecular Theory is used to explain the behavior of gases and is based upon the following postulates:*
    - *Gases are composed of a many particles that behave like hard spherical objects in a state of constant, random motion.*
    - *These particles move in a straight line until they collide with another particle or the walls of the container.*
    - *These particles are much smaller than the distance between particles, therefore the volume of a gas is mostly empty space and the volume of the gas molecule themselves is negligible.*
    - *There is no force of attraction between gas particles or between the particles and the walls of the container.*
    - *Collisions between gas particles or collisions with the walls of the container are elastic. That is, none of the energy of the gas particle is lost in a collision.*

16. How do gases create pressure, use KMT to support your answer.
A: *Gas molecules have random motion. The collisions of the gas particles with a surface create pressure.*

17. Explain diffusion. Use KMT to support your answer.
A: *The random motion of the gas particles result in many collisions which push each other around until they are evenly distributed throughout the volume of gas. Smaller particles require less energy to be pushed around and therefore diffuse faster.*

18. Is Boyle's law direct or inverse? Charles's Law? Gay-Lussac's Law?
A: *Boyle's Law is inverse (as pressure goes up, volume goes down). CharlesLaw is direct (as temperature goes up, volume goes up). Gay-Lussac's Law is direct (as temperature goes up, pressure goes up).*

19. If 735 L of a gas is at 3.11 atm and 34°C, what is its temperature at 6.11 atm and 235 L?
A: $P_1V_1/T_1 = P_2V_2/T_2$ *(3.11 atm • 735 L) / 307 K = (6.11 atm • 235 L) / X* *X = 193 K (-80.2 °C)*

20. If 12.2 mL of a gas is at 178°C, what is its volume at 53.0°C?
A: $T_1V_2 = T_2V_1$ *((326 K • 12.2 mL) / (451 K • X))* *X = 8.82 mL*

21. What are the values of STP? What is the meaning behind STP?
A: *Standard Pressure = 1.0 atmosphere. Standard temperature = 273K = 0.0°C. Standard temperature and pressure (STP) allows scientists to have a common reference point when describing gases.*

22. What is the lowest temperature possible? What is it called?
A: *0.0K or -273°C It is called absolute zero because at that temperature the particles have no kinetic energy (aren't moving).*

Acids and Bases

23. What are the major observable properties of acids, bases, and salt solutions?
A: *Acids: taste sour, sting in cuts, react with metals, turn blue litmus red, is clear in phenolphthalein, has a pH below 7*
Bases: taste bitter, feel slippery, don't react with metals, turns red litmus blue, is pink in phenolphthalein, has a pH above 7.
Salt Solutions: very boring, don't do anything with indicators, may sting a little in a cut, may be neutral or slightly acidic or basic, depending upon the acid/base combination that created them.

24. What gas is produced when an acid reacts with a metal?
A: *Hydrogen gas (H_2)*

25. What is the Arrhenius definition for acids and bases?
A: *An Arrhenius acid ionizes in water to produce hydrogen ions (H^+). Arrhenius acid formulas begin with "H". An Arrhenius base dissociates to produce hydroxide ions (OH^-) in water. Arrhenius bases end with "OH".*

26. What is the Brønsted-Lowry definition for acids and bases?
A: *A Brønsted-Lowry acid donates protons (H^+ ions) in water. The definition is not far from the Arrhenius definition. A Brønsted-Lowry base accepts protons. By accepting protons from water, a Brønsted-Lowry creates an imbalance between H^+ and OH^-.*

27. Describe the dissociation (ionization) of strong acids and bases versus weak acids and bases.
A: *Strong acids and strong bases completely dissociate in water. Weak acids and bases only partially dissociate.*

28. List the 6 strong acids and state the rule for strong bases.
A: *HCl , HNO_3 , H_2SO_4 , HBr , HI , $HClO_4$ (sometime $HClO_3$ is included as a 7th).*
The rule of thumb is that a strong acid is 100% dissociated in solutions of 1.0 M or less.

29. What are the pH values for acids? Bases? A: *Acids have pH below 7. Bases have pH above 7.*

30. What is more acidic, a solution with a pH of 2 or 5?
A: *A solution with pH of 2 is 1000 times more acidic than a solutuion with pH of 2. (Each pH point represents 10x change.)*

What is more basic, a solution with a pH of 8 or 13?
A: *A solution with pH of 13 is 100000 times more basic than a solutuion with pH of 8.*

31. What is the pH value for a neutral substance? A: *Neutral substances have a pH of 7.*

32. List the color for each indicator in an acid solution and a base solution: phenolphthalein, red litmus, and blue litmus.
    A: *phenolphthalein – colorless in acid, pink in base*
    red litmus – red in acid, blue in base
    blue litmus – blue in base, red in acid

Solutions

33. What is a solute? Solvent?
    A: *Solute: the stuff that dissolves; may change state, as in salt changes from a solid to a liquid state when dissolved in water.*
    Solvent: the stuff the solute dissolves in; water is the "universal solvent"; stays the same state
    When the solute and solvent are in the same state, the one in greater volume is considered the solvent.

34. Describe the dissolving process at the molecular level by using the concept of random molecular motion.
    A: *Using salt water as an example: water molecules move around randomly. They crash into the salt crystals, causing some of the salt ions to break loose. The random motion of the water molecules and salt ions cause them to disperse around the entire liquid.*
    There must be attraction between the solute and solvent for them to attach (solvate).
    Polar solvents dissolve polar solutes. Non-polar solvents dissolve non-polar solutes.

35. What three factors affect the dissolving process?
    A: *Temperature, surface area, amount of solute already in solution, and stirring affect rate of solution formation.*
    Nature of the solute and solvent, temperature, and pressure affect the amount of solute that will ultimately dissolve.

36. When 5.20 g of salt is added to 5000 g of water, what is the concentration in parts per million (ppm)? A: *1040ppm*

37. How many grams of KOH would you need to make 750. mL of solution with a concentration of 5.5% by mass?
    A: *41.25 grams*

38. If you add 25 g of $CaCl_2$ to 1000. mL of water, what would the concentration of the solution be in grams/liter? A: *25g/L*

39. What is the molarity of a solution that contains 78.2 grams of NaCl in 4.25 liters of solution? A: *0.316 M*

Chemical Thermodynamics

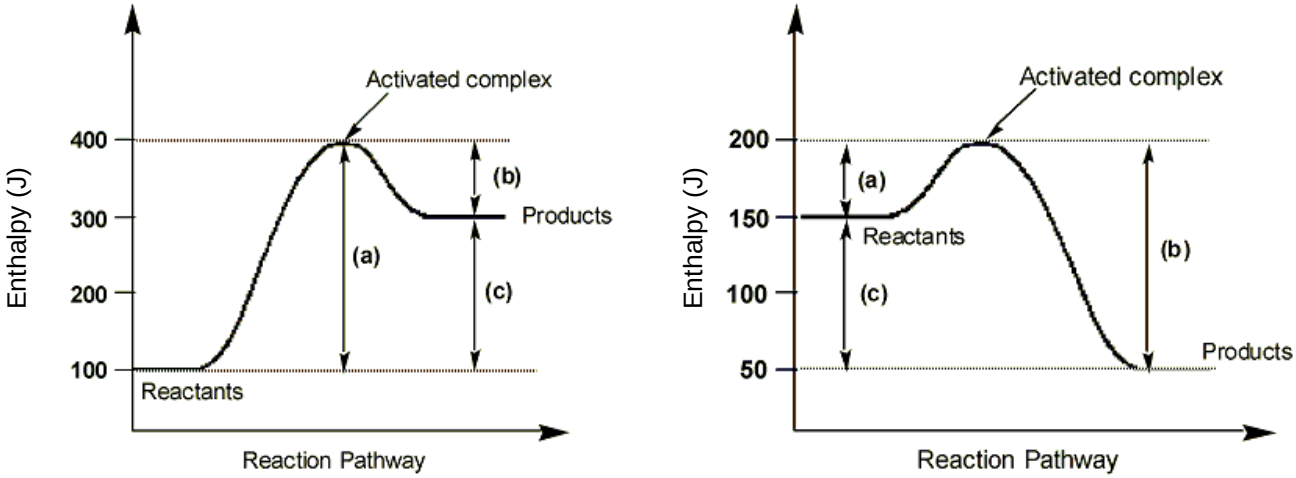
40. What happens to atoms or molecules as their temperature is increased?
    A: *They move faster.*

41. Describe heat flow.
    A: *Heat flows from something that's hot to something that's cold. The hot molecules are moving faster and when they collide with the slower moving particles of the cooler substance, some of the energy is transferred, making the slower particle to speed up (raise temperature) and the faster one to slow down (lower temperature.)*

42. What is the difference between an exothermic reaction and an endothermic reaction?
    A: *Exothermic gives off heat, endothermic absorbs heat. Exothermic reactions have more stored energy in the reactants than in the products (ΔH is negative). Endothermic reactions have more energy in the products than the reactants (ΔH is positive).*

43. Draw an energy diagram for an endothermic and exothermic reaction.

A:



44. Are the following changes of state exothermic or endothermic:
    a. An ice cube melting
       A: *Endothermic*
    b. Dry ice subliming to carbon dioxide
       A: *Endothermic*
    c. Water vapor condensing on a mirror
       A: *Exothermic*
    d. Water freezing into an ice cube
       A: *Exothermic*

45. How many calories are needed to raise 450. grams of water from 21.0°C to 85.5°C?
    A: *($q = m \times C_p \times \Delta T$) q = 450. grams x 1 cal/(gram • °C) x 64.5 °C = 29 000 cal*

46. How many grams of water can be heated 46.0°C by 34.8 kJ?
    A: *($m = q / (C_p \times \Delta T$) m = 34 800 J / (4.184 J/(g•°C) x 46 °) = 181 grams*

Reaction Rates

47. Define reaction rate.
    A: *The rate at which reactant is transformed into product measured as the change in the concentration of a reactant or product divided by change in time.*

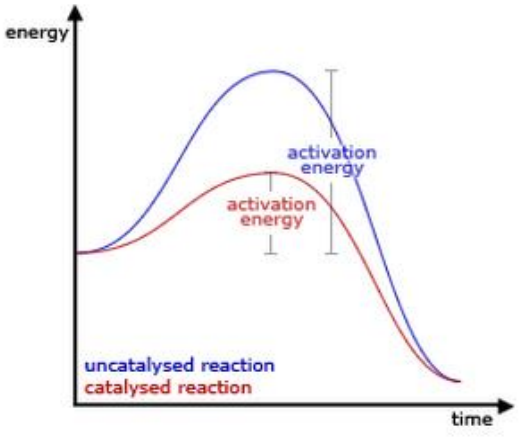
48. What are the units for reaction rate?
    A: *Molarity / second = M/s (or mol/(L • s))*

49. Describe what happens to the concentration of reactants during a chemical reaction. Products?
    A: *The concentration of reactants starts high and immediately begins to decrease as reactant is converted into product. The concentration of product starts at zero and immediately begins to increase as product is formed from reactant.*

50. What happens to the reaction rate when there is an increase in concentration? Temperature? Pressure?
    A: *These changes generally speed up reactions by creating more effective collisions between reactants. Increasing concentration crowds particles together making more collisions occur. Increasing temperature increases the kinetic energy of collisions and makes collisions occur more frequently. Increasing pressure increases the concentration of gases.*

51. What is a catalyst? How does the addition of a catalyst affect the rate of a reaction?
    A: *A catalyst is a substance that increases the rate of a chemical reaction by lowering the activation energy of the reaction. It is not consumed in the reaction and may be used again and again.*

52. Draw an energy diagram for a catalyzed reaction.
    A:



Chemical Equilibrium

53. Define equilibrium.
    A: *Equilibrium occurs in reversible reactions when the products of the reaction begin to revert back into reactants. Equilibrium exists when the forward changes in the reactions occur at the same rate as the reverse changes in the reaction. At this point, the concentrations of the reactants and products no longer change.*

54. What is Le Chatelier's Principle?
    A: *If a system at equilibrium is stressed/changed, the system will shift one way or the other in order to minimize the stress. The three stresses are: change in concentration, change in temperature, and change in pressure (for gases only).*

55. Which direction will the following reaction shift if: $H_2(g) + I_2(g) + 50.9\ kJ \leftrightarrows 2\ HI(g)$
    a. Add HI A: *Left*
    b. Decrease temperature A: *Left*
    c. Increase pressure A: *No change*
    d. Remove I_2 A: *Left*

56. Which direction will the following reaction shift if: $PCl_3(g) + Cl_2(g) \leftrightarrows PCl_5(g) + heat$
    e. Add PCl_3 A: *Right*
    f. Remove Cl_2 A: *Left*
    g. Add heat A: *Left*
    h. Decrease pressure A: *Left*

True/False

57. The quantity *one mole* is set by defining one mole of carbon 12 atoms to have a mass of exactly <u>12 grams</u>.
    A: *True*
58. Acids are hydrogen-ion-<u>accepting</u> and bases are hydrogen-ion-<u>donating</u> substances.
    A: *False ("donating", "accepting")*
59. Energy is <u>absorbed</u> when a material condenses or freezes and is absorbed when a material evaporates or melts.
    A: *False ("released")*
60. The rate of a reaction is the <u>decrease</u> in concentration or reactants or the <u>increase</u> of products with time.
    A: *True*
61. A catalyst <u>increases</u> the rate of a reaction by <u>raising</u> the activation energy of a reaction.
    A: *False ("increases", "lowering")*