Chemistry Semester 2 Review

Unit 6 and 7: KMT and Gases, Thermochemistry

Vocab:

STP
Molar volume
Ideal gas law
Ideal gas constant
Kinetic-molecular theory
Pascal
Pressure
calorimeter
chemical potential energy
energy

enthalpy
enthalpy (heat) of combustion
enthalply (heat) of reaction
entropy
heat
Hess's law
Joule
Law of conservation of energy
Law of disorder
Molar enthalpy (heat) of fusion

Molar enthalpy (heat) of vaporization
Specific heat
Spontaneous process
Standard enthalpy (heat) of formation
Surroundings
System
Thermochemistry
Universe

1. Temperature is a measure of the _____ _____ energy of the molecules in a sample.

2. A gas exerts pressure on its container because the molecules _____ with the walls.

3. What are the four variables that describe a gaseous system?

     1.   2.   3.   4.

4. Temperature must *always* be in _____ when calculating gas law problems.

5. Standard pressure = _____ atm

6. Standard temperature = _____ K = _____ degrees Celsius.

7. Answer the following questions with INVERSELY or DIRECTLY

       a) How are pressure and temperature related? _____

       b) Pressure and Volume? _____

       c) Volume and Temperature? _____

8. What will happen to a balloon filled with helium gas when you take it from outside on a hot day to inside an air conditioned house?

9. How many moles of a gas will occupy 2.50L at STP?

10. Calculate the volume that 3.60 grams of H_2 gas will occupy at STP.

11. Use the reaction shown to calculate the mass of iron that must be used to obtain .500L of hydrogen at 24.3 degrees Celsius and 100.0 kPa of pressure. $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

12. What does each of the symbols below represent?
    a) ΔH
    b) ΔS

13. Define entropy and enthalpy.

14. Describe an endothermic and exothermic reaction.

15. Determine if the letter below is supporting an exothermic or endothermic reaction:
    a) Products have more energy than the reactants _____
    b) Reactants have more energy than the products _____
    c) H_2O (l) → H_2O (g) _____
    d) $+\Delta H$ _____
    e) $-\Delta H$ _____
    f) Water freezing _____

18. In nature, do things tend to become more organized or more disordered? What law of thermodynamics is this?

19. Determine whether the change below supports an increase or decrease in disorder ($-\Delta S$ or $+\Delta S$)
    a) CH_3OH (l) → CH_3OH (g) _____
    b) $2KClO_3$(s) → $2KCl$ (s) + $3Cl_2$ (g) _____
    c) $2SO_2$ (g) + O_2 (g) → $2SO_3$(g) _____

20. The enthalpy of the products is 255 kJ and the enthalpy of the reactants is 335 kJ. Calculate the change in enthalpy and determine if the reaction is exothermic or endothermic.

21. Predict the sign of ΔS_{system} for the following changes and explain your answer:
    a) ClF (g) + F_2 (g) → ClF_3 (g)

    b) $C_{10}H_8$ (l) → $C_{10}H_8$ (s)

Unit 9: Rates and Equilibrium

Vocab:

Activated complex	Catalyst	heterogeneous equilibrium
Activation energy	Dissociation equations	Le Chatelier's Principle
Collision theory	reversible reaction	Ksp
Reaction rate	completion reaction	Keq
Transition state	chemical equilibrium	
	homogeneous equilibrium	

22. List the factors that affect the RATE of a chemical reaction _and_ tell HOW they affect the rate.

5 Factors that affect the reaction rate:	How the factors alter the rate:

23. What is a catalyst? How is an enzyme like a catalyst? How do catalysts work?

24. In order for a reaction to occur, the reactants must _____ with enough _____ and the correct _____.

25. The amount of energy needed for an effective collision is called the _____ _____.

26. What happens to the rate of a chemical reaction over time? _____

27. In a chemical reaction that produces hydrogen 14.3 ml of gas was collected over a 20.0 second period. Calculate the rate of the reaction in ml/sec.

28. A double arrow signifies a _____ reaction, while a single arrow signifies a _____ reaction.

29. What causes a reaction to go to completion? The evolution of a _____ or the formation of a _____. Describe chemical equilibrium. Give an example.

30. Write the equilibrium constant expression for $4HCl_{(g)} + O_{2(g)} \leftrightarrow 2Cl_{2(g)} + 2H_2O_{(g)}$

31. If you calculate a number less than 1 for the constant expression above, what does that tell you?

32. At 773K, the reaction $2NO(g) + O_2(g) \leftrightarrow 2NO_2(g)$ produces the following concentrations: $[NO]=3.49 \times 10^{-4}$ M; $[O_2]=0.80M$; $[NO_2]=0.250M$. Write the equilibrium constant expression for the reaction, & calculate the value of the equilibrium constant.

33. _____ explains how an equilibrium system will respond to stress.

34. For the reaction given, complete the following table: **C (s) + H₂O (l) + heat $\leftrightarrow$ CO (g) + H₂ (g)**

Stress applied	Shift left, shift right, or no change?	What happens to the concentration of CO?
Cooling		
Adding water		
Adding a catalyst		
Removing H₂		
Decreasing volume		

35. For the reaction; **Heat + H₂(g) + I₂(g) $\leftrightarrow$ 2HI(g)**

    A. How will an increase in temperature change the concentration of Hydrogen gas? _____

    B. How will an increase in pressure affect the system? _____

    C. Which direction will the addition of Iodine gas shift the system? _____What does this do to the concentration of Hydrogen gas? _____

36. For the reaction $N_2O_4(g)$ + heat $\longleftrightarrow$ 2 $NO_2(g)$
   a. List 2 stresses that you could apply to the equilibrium system to increase the 2 $NO_2(g)$:

   b. List 2 stresses that you could apply to the equilibrium system to increase the $N_2O_4(g)$:

Unit 8: Solutions

Vocab:

Dissociation equations	Keq	Solution
chemical equilibrium	Concentration	Solvation
homogeneous equilibrium	Insoluble	Solvent
heterogeneous equilibrium	Molarity	Solute
Le Chatelier's Principle	Saturated solution	Supersaturated solution
Ksp	Solubility	Unsaturated solution
	Soluble	Dilution

37. Describe solute and solvent.

38. List the factors that affect solubility.

39. In general, the solubility of most solid substances _____ as temperature increases. The solubility of gases, however, _____ as temperature increases. Pressure only changes the solubility of _____.

   a) Describe the rule "Likes dissolves Like".

   b) What type(s) of compounds are soluble in water.
   _____

   c) Circle the chemical(s) that are soluble in water. Cross out the one(s) that are not.

   Fe (iron) $MgCl_2$ C_5H_{10} SiO_2

40. Describe the three types of solutions. Include how you could determine which solution is which.
   Saturated:

   Unsaturated:

   Supersaturated:

41. When you add more solvent to a solution, the solution becomes more _____.

42. What unit do chemists use most often to describe concentration? _____

43. Calculate the molarity for each of the following solutions:
   a. 3.4 moles of NaCl dissolved in 0.956 L of water

   b. 1.28 g of $CuSO_4$ dissolved in 150 mL of water

44. **How would you prepare** 500 mL of 1.5 M NaCl from solid NaCl? Show any calculations needed.

45. A .600 L sample of a 2.50 M solution of KI contains what mass of KI?

46. What is the volume of 0.1250 M solution of $AgNO_3$ that contains 1.75 moles of solute.

47. How many mL of 2.0 M KOH stock solution do you need to prepare 100 mL of 0.40 M KOH.

48. What would be the new molarity if you diluted 250 mL of 6.0 M HCl and up to 800 mL?

Unit 10: Acids and Bases
Vocab:

Arrhenius model	Conjugate acid-base pair	hydronium ion
Acid-base indicator	Conjugate base	neutralization reaction
Amphoteric (amphiprotic)	End point	pH
Bronsted-Lowry model	Equivalence point	pOH
Conjugate acid	K_w	titration

49. List 5 properties of acids and 5 properties of bases.

1)   1)

2)   2)

3)   3)

4)   4)

5)   5)

50. Describe the differences between an Arrhenius and a Bronsted-Lowry acid and base.

51. Identify the Bronsted-Lowry acid-base pairs in each of the following reactions. Label each substance (BB, BA, cb, ca).
   a. $NH_3 + H_2O \longleftrightarrow NH_4^+ + OH^-$ b. $HC_2H_3O_2 + H_2O \longleftrightarrow C_2H_3O_2^- + H_3O^+$

52. Answer the following questions about electrolytes:
   a) What is a strong electrolyte? _____
   b) What is a non-electrolyte? _____
   c) Give an example of each: strong electrolyte - _____ Non-Electrolyte - _____

53. _____ acids & bases dissociate (ionize) completely. _____ acids & bases only slightly dissociate (ionize).

54. Circle the strong base and put a box around the strong acid.
            HCl NH_3 CH_3COOH NaOH

55. What are the formulas for hydroxide _____ and hydronium _____?

56. If the hydronium concentration of a solution is 2.34×10^{-3} M, what is the pH?

57. **If** the concentration of $HClO_4$ is 0.00025 M, calculate the pH and pOH.

58. What is the $[H^+]$ concentration of a solution with a pH of 2.687?

59. Calculate the pH and the pOH for a 6.57 x 10^{-5} M solution of LiOH.

60. An acid + a base yields a _____ + _____. This type of reaction is called _____.
61. The process used to find the concentration of an acid or a base is a _____.
62. What do we call it when the moles of acid = moles of base? _____
63. What do we add to signal the end of the titration? _____
64. For the following types of titrations, give the pH range for the equivalence point:
   a) Weak acid-strong base   b) Strong acid-strong base   c) Strong acid- weak base

      _____   _____   _____
65. Which indicator is best for a strong acid-strong base titration? Why?

66. If 25 mL of 0.20 KOH were used to titrate 15 mL of H$_2$SO$_4$, what is the molarity of the acid? You must first
   complete and balance the equation. Show your work for the calculation.
         _____KOH + _____H$_2$SO$_4$ →

67. 75.0 ml of .250M nitric acid, HNO$_3$, reacts with 25.5 ml of potassium hydroxide, KOH. What is the molarity of
   the base? Be sure to write a balanced chemical equation first.

68. What is the molarity of Ca(OH)$_2$ solution if 30.5 ml of the solution is neutralized by 36.6 ml of .250 M HBr?
Ca(OH)$_2$ + 2HBr → 2H$_2$O + CaBr$_2$