Chemistry Final Exam Study Guide

This final exam study guide is NOT worth bonus points. If you complete the study guide, then I will use your final exam grade to replace low test grades. **You must put your answers on another sheet of paper.

***Unit 1 (Chapter 1, 2, and 3):**

1. What are the 5 steps of the scientific method (IN ORDER)?
2. Define pure substance. What are the two categories of pure substances?
3. What is the difference between a compound and an element? List an example of each.
4. List and define the two methods for separating mixtures.
5. What is the law of conservation of mass and how does it apply to chemical reactions?
6. Classify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
   a. carbon monoxide c. mushroom pizza
   b. zinc metal d. tap water
7. What are the 5 signs that a chemical reaction has taken place?
8. Express the following number in scientific notation:
   a. 4380000 s b. 0.000274 m
9. List the SI units for the following quantities: length, mass, temperature, time, amount of substance, luminous intensity, and electric current.
10. How many significant figures are in the following numbers:
    a. 702000 m b. 40 crayons c. 0.00630100 g d. 170.4380 s
11. Convert 14.8 g to micrograms.
12. Convert 867,329 s to hours.
13. Convert 329°C to Kelvin.
14. What is the mass of 180.3 cm^3 of lead if the density is 11.4 g/cm^3?
15. Define accuracy and precision.
16. Complete the following calculations with the correct number of significant figures:
    a. 1.23 kg + 4.082 kg
    b. 16.04 s – 5 s
    c. 0.070 cm x 1.08 cm
17. What is the lowest temperature on the Kelvin scale?

***Unit 2 (Chapter 4, 5, and 6):**

18. What did the following scientists contribute to Chemistry: Democritus, J. J. Thomson, Eugen Goldstein, and James Chadwick?
19. List the charge, relative mass, and location for the following: protons, electrons, and neutrons.
20. Which subatomic particle is the most responsible for the chemical behavior of an atom?
21. Define isotope. Would isotopes have different atomic numbers or mass numbers?
22. Calculate the number of neutrons in the following:
    a. $^{7}_{3}Li$ b. $^{45}_{21}Sc$

23. Define the Aufbau Principle, Pauli Exclusion Principle, and Hund's Rule.
24. List the number of orbitals and maximum number of electrons that are in the following sublevels: s, p, d, f, and g.
25. What is the standard electron configuration for zinc?
26. What element has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$?
27. Explain how atoms emit light.
28. What are the following groups called: Group 1, 2, 3 – 12, 17, and 18?
29. What does each row on the periodic table represent?
30. How did Mendeleev arrange his periodic table? How is the modern periodic table arranged?
31. What determines an element's chemical properties?
32. Define cation and anion.
33. The radius of a cation is _____ than its neutral atom. The radius of an anion is _____than its neutral atom.
34. What are 2 characteristics of noble gases?
35. What charges do the following elements have when they form ions?
    a. S b. Ne c. Ca d. P
36. Draw the periodic trends for atomic size, ionization energy, and electronegativity.

***Unit 3 (Chapter 7, 8, and 9):**

37. Write the formula for the compound formed between the following elements:
    a. lithium and sulfur
    b. oxygen and aluminum
    c. phosphorus and calcium
    d. rubidium and chlorine
    e. oxygen and strontium
38. Define electron sea model.
39. Define ionic bond, covalent bond, and metallic bond.
40. What is a diatomic molecule? List the 7 naturally occurring diatomic elements.
41. Metals have _____ electronegativities. Nonmetals have _____ electronegativities.
42. Do metals tend to gain or lose electrons? Do nonmetals tend to gain or lose electrons?
43. Tell if the following molecules contain nonpolar, polar, or ionic bonds.
    a. CH_4 b. HCl c. H_2O d. Li_3N e. F_2
44. How many lone pairs of electrons are in the Lewis dot structure for H_2O?
45. Draw the Lewis dot structures for the following: CO, CO_2, N_2, and O_2.
46. Define London dispersion forces, dipole-dipole attractions, and hydrogen bonding.
47. Identify the major intermolecular forces in the following molecules: NH_3, Cl_2, and CO_2.
48. What are the rules for naming a compound that starts with the following type of element:
    a. regular metal
    b. transition metal
    c. nonmetal
    d. hydrogen
49. Name the following compounds:

a. Li_2CO_3
b. H_2SO_3
c. N_2O_4
d. $Mg(NO_3)_2$
e. HF
f. Cu_2O
g. H_3PO_4
h. SO_3
i. $FeCl_3$
j. PbO_2

50. Write the formulas for the following compounds:
    a. nitric acid
    b. dichlorine heptoxide
    c. zinc (II) acetate
    d. calcium nitride
    e. aluminum fluoride
    f. hydrobromic acid

***Unit 4 (Chapter 10, 11, and 12):**

51. How many moles are in 348 g of iron (III) fluoride?
52. What is the mass of 17.36 mol of calcium nitrate?
53. Calculate the molar mass of the following:
    a. aluminum bromide
    b. calcium phosphate
    c. zinc (II) oxide
    d. nitrous acid
54. What is the mass percent of carbon in strontium acetate?
55. Calculate the percent composition of all elements in sodium carbonate.
56. Write the empirical formula for $C_6H_{12}O_6$.
57. Calculate the empirical formula for a compound that is 62.1 %C, 13.8 %H, and 24.1 %N.
58. What is the molecular formula of a compound whose molar mass is 60 g/mol and empirical formula is CH_4N?
59. Balance the following equations:
    a. $AgNO_3 + H_2S \rightarrow Ag_2S + HNO_3$
    b. $Zn(OH)_2 + H_3PO_4 \rightarrow Zn_3(PO_4)_2 + H_2O$
    c. $Fe_2O_3 + H_2 \rightarrow Fe + H_2O$
    d. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
    e. $PbO_2 \rightarrow PbO + O_2$
    f. $(NH_4)_2CO_3 \rightarrow NH_3 + H_2O + CO_2$
    g. $C_6H_6 + O_2 \rightarrow CO_2 + H_2O$
    h. $Fe(OH)_3 \rightarrow Fe_2O_3 + H_2O$
60. Identify the type of the following reaction:
    a. $K_2CO_3 + BaCl_2 \rightarrow KCl + BaCO_3$
    b. $C_3H_6 + O_2 \rightarrow CO_2 + H_2O$
    c. $Al(OH)_3 \rightarrow Al_2O_3 + H_2O$
    d. $Li + O_2 \rightarrow Li_2O$
    e. $Zn + AgNO_3 \rightarrow Ag + Zn(NO_3)_2$
61. Calculate the moles of hydrogen needed to react to form 17.8 moles of ammonia.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

62. How many grams of nitrogen dioxide must react to produce 5.00 x 10^{22} m/c of nitrogen monoxide?

$$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

63. How many grams of oxygen are required to burn 3.86 L of carbon monoxide?

$$2CO + O_2 \rightarrow 2CO_2$$

64. How many grams of copper (I) sulfide can be formed when 80.0 g of copper reacts with 25.0 g of sulfur?

$$2Cu + S \rightarrow Cu_2S$$

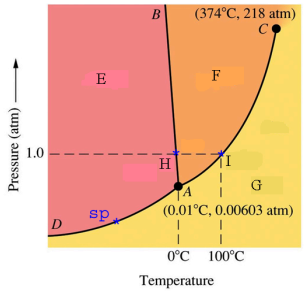
65. When 84.8 g of iron (III) oxide reacts with excess carbon monoxide in the lab, 47.8 g of iron is produced. What is the percent yield?

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

***Unit 5 (Chapter 13, 14, and 15):**

66. Write the 4 assumptions of the kinetic theory.
67. Define the following phase changes: melting, condensation, sublimation, vaporization, solidification, and deposition.
68. On a phase diagram, what does a line represent? What does point A represent?

69. Label the states of matter on the phase diagram. (areas E, F, and G)



70. Convert 782.3 kPa to torr.
71. What is STP and what are its values?
72. A balloon with 5.00L of air at -50.0°C is warmed to 100°C, what is the new volume?
73. Determine the volume occupied by 20.3 g of carbon dioxide gas at 15°C if the pressure is 81.8 kPa.
74. A gas with a temperature of 78.3°C is warmed to 169.8°C and the pressure increases to 1.5 atm, what was the original pressure?
75. A 5.00 L air sample has a pressure of 107 kPa at a temperature of -50.0°C. If the temperature is raised to 102°C and the volume expands to 7.00 L, what will the new pressure be?
76. If the temperature of a gas increases from 150K to 300K, what happens to the volume?
77. What is the gas laws poem that I wrote?
78. What causes water's low vapor pressure and high surface tension?
79. Draw the polarity of a water molecule.

80. Draw a diagram of an ionic solid dissolving in water.

***Unit 6 (Chapter 16, 17 and 19):**

81. Define dilute and concentrated.
82. How many moles of ammonium nitrate are in 335 mL of 0.425 M NH_4NO_3?
83. How many milliliters of a solution of 4.00 M KI are needed to prepare 250 mL of 0.760 M KI?
84. How many grams of K_2SO_4 would you need to prepare 1500 g of 5.0 % (m/m) K_2SO_4 solution?
85. What are 3 ways to increase the rate at which a solid dissolves in water?
86. What is a colligative property?
87. In Kool-aid, what is the solute and what is the solvent?
88. What is the difference between and exothermic and endothermic reaction? What does each feel like?
89. What is the specific heat of 560 g of a substance that releases 6700 J of heat when it cools from 100°C to 75°C?
90. Does the temperature of a substance change during a phase change?
91. Define Arrhenius acid and base.
92. Define Bronsted-Lowry acid and base.
93. Which ion indicates an acid? Which ion indicates a base?
94. Define monoprotic, diprotic, and triprotic acid.
95. What pH values are acidic? basic? neutral?
96. If $[H^+] = 3.7 \times 10^{-5}$M, what is the $[OH^-]$?
97. If $[H^+] = 2.1 \times 10^{-9}$M, what is the pH?
98. If pH = 2.7, what is the $[H^+]$?
99. If $[OH^-] = 1.72 \times 10^{-8}$M, what is the pOH?
100. If $[OH^-] = 6.33 \times 10^{-11}$M, what is the pH?