Name:______________________________Block:_____Date:________________________

Review for Chemistry Final Exam
[Chapters 1-9 & 12]

<u>Chapter 1 – Matter and Change</u>

1-1. Define the terms “matter” and “atom”.

1-2. Define the terms “element” and “compound” and list some examples of each.

1-3. Define the term “mixture” and explain the two different types of mixtures.

1-4. Compare and contrast a “mixture” to a “pure substance”.

1-5. Compare and contrast a “chemical change” to a “physical change”. List three examples of each type of change.

1-6. Define the "Law of Conservation of Energy".

Unit 2 – Measurements and Calculations

2-1. Define the terms "qualitative measurement" and "quantative measurement".

2-2. List the seven "SI quantities" and their "units of measurement".

2-3. Know the metric prefixes on your metric line handout (Tera through pico).

2-4. How many grams are in 882 μg?

2-5. How many kilometers are in 92.25 meters?

2-6. Define "density" and state the equation used to calculate density.

2-7. Find the density of a material, given that a 5.03 g samples occupies 3.24 mL.

2-8. What is the mass of a sample of material that has a volume of 55.1 cm^3 and a density of 6.72 gm/cm^3?

2-9. Know the five rules for identifying significant figures.

2-10. Know the rules for using significant figures when performing mathematical operations.

2-11. Perform the calculations involving significant figures.

a. Identify the number of significant figures in each of the following measurements. Write the number in the space provided.

1.) 520 mL ________ 3.) 10.002 ns ________

2.) 0.0102 ms ________ 4.) 0.451 Pa ________

b. Perform the following calculations, round off the answer to the correct number of significant figures, and include the appropriate units.

1.) 72.60 m + 0.0950 m = ____________________

2.) 0.0890 cm – 0.066 cm = ____________________

3.) 120 km^2 ÷ 8.56 km = ____________________

2-12. Define "direct relationship" and give the equation used.

2-13. Define "inverse relationship" and give the equation used.

Unit 3 – Atoms – The Building Blocks of Matter

3-1. Define the following laws:

a. Conservation of Mass

b. Law of Definite Proportions

c. Law of Multiple Proportions

3-2. List the five statements of “Dalton’s Atomic Theory”.

3-3. What scientist discovered the electron? Describe his experiment.

3-4. What scientist discovered the nucleus of an atom? Describe his experiment.

3-5. Define the terms “atomic number” and “mass number”.

3-6. Define the terms "ion" and "isotope".

3-7. Know how to determine the number of protons, neutrons, and electrons using mass and atomic numbers.

3-8. Determine the number of protons, electrons, and neutrons in each of the following isotopes:

a.) sodium – 23 p:_____ n:_____ e:_____

b.) calcium – 40 p:_____ n:_____ e:_____

c.) ${}^{64}_{29}Cu$ p:_____ n:_____ e:_____

d.) ${}^{108}_{47}Ag$ p:_____ n:_____ e:_____

3-9. Write the nuclear symbol and hyphen notation for each of the following isotopes.

a.) mass number of 28 and atomic number of 14

b.) 26 protons and 30 neutrons

3-10. Define "Avogadro's number" and "mole".

3-11. Know how to use molar mass and Avogadro's number to convert units.

3-12. Determine the mass of 2.00 mols of Carbon.

3-13. Determine the mass of 3.01 x 10^{23} atoms of Sulfur.

3-14. Determine the number of moles of 1.50 x^{23} atoms of Phosphorus.

Chapter 4 – Arrangement of Electrons in Atoms

4-1. State the "wave equation" and explain how the variables wavelength and frequency are related to one another.

4-2. State the "energy equation" and explain how the variables energy and frequency are related to one another.

4-3. Use the equations listed above to solve the following problems:

a. Calculate the wavelength of the electromagnetic radiation whose frequency is 7.50 x 10^{12} Hz.

b. What is the frequency of a radio wave whose energy is 1.55 x 10^{-24} J per photon?

c. Calculate the wavelength of the electromagnetic radiation with an energy of 3.37 x 10^{-19} J.

4-4. Describe the "Bohr Model" of the atom and include a diagram.

4-5. What was the conclusion of Louis de Broglie's "Double Slit Experiment"?

4-6. What name is given to the current model of the atom? What were Heisenberg's and Schrodinger's contributions to this model?

4-7. Compare and contrast the terms "orbit" and "orbital" in regards to electrons.

4-8. Define the four "quantum numbers" and list the symbols used for each.

4-9. Know how to determine the number of orbitals and number of electrons for any given energy level (n = 1-7).

a. What sublevels are present in n= 4?

b. What sublevels are present in n = 3?

c. What is the total number of orbitals in n = 3?

d. What is the total number of orbtitals in $n = 4$?

e. What is the maximum total number of electrons that can be held by the orbitals in $n = 1$?

f. What is the maximum total number of electrons that can be held by the orbitals in $n = 2$?

4-10. Describe the 3 rules that govern the filling of atomic orbitals by electrons.

4-11. You should be able to write electron configurations, orbital diagrams, Lewis dot diagrams, and nobel gas configurations for atoms with atomic numbers from 1 to 20.

For each of the following elements, (a) write the electron configuration, (b) draw the orbital diagram, (c) draw the Lewis dot diagram, and (d) write the nobel gas configuration.

A. magnesium

a.____

b.____

c._____________

d.____

B. sulfur

a.____

b.____

c._______________

d.____

Chapter 5 – The Periodic Law

5-1. What scientist came up with the first periodic table? How were the elements arranged?

5-2. What scientist came up with the current periodic table? How are the elements arranged?

5-3. State the location, using group numbers, for the following:

a. Alkali metals: ______

b. Alkaline-earth metals: ______

c. Transition metals: ______

d. Inner Transition metals: ______

e. Halogens: ______

f. Noble gases: ______

g. Lanthanides: ______

h. Actinides: ______

i. *s* block ______

j. *p* block ______

k. *d* block ______

l. *f* block ______

5-4. Those electrons that are largely responsible for an atom’s chemical behavior are called what?

5-5. Define the term “atomic radius” and describe the trend seen on the periodic table.

5-6. Define the term “ionization energy” and describe the trend seen on the periodic table.

5-7. Define the term “electron affinity” and describe the trend seen on the periodic table.

5-8. Define the term “electronegativity” and describe the trend seen on the periodic table.

5-9. Which family of metals on the periodic table are the most reactive metals?

5-10. Which family of elements on the periodic table are the least reactive?

5-11. Which block of the periodic table contains the semimetals?

5-12. Which block of the periodic table contains the alkali metals?

Chapters 6 & 7: Chemical Bonding & Chemical Formulas

6-1. Define the following types of bonds:

a. Ionic bond:

b. Covalent bond:

c. Metallic bond

6-2. Compare the terms "empirical formula", "molecular formula", and "structural formula".

6-3. List all of the elements that form diatomic molecules.

6-4. All compounds that are hydrated contain what chemically attached to them?

6-5. Define the term "acid" and describe the two classifications of acids.

6-6. Know all of the rules for naming and formula writing for ionic compounds, molecular compounds, hydrates, and acids.

a. Write the formulas for the following compounds

1. carbon monoxide
2. sodium chlorate
3. carbon tetrachloride
4. magnesium bromide
5. phosphoric acid
6. iron (III) oxide
7. magnesium sulfate
8. sodium phosphate trihydrate
9. dinitrogen hexoxide
10. phosphorus trichloride

b. Write the names for the following compounds

1. $NaClO_2$
2. MnS
3. K_2O
4. $SnBr_2$
5. Na_2CO_3 º $10H_2O$
6. $Ca(OH)_2$
7. NH_4Cl
8. NH_3
9. $HC_2H_3O_2$
10. H_2S

Chapter 8: Chemical Equations and Reactions

8-1. Define the terms “reactant” and “product”.

8-2. Know how to write balanced formula equations for various reactions.

a. Iron reacts with sulfuric acid to form iron II sulfate and hydrogen gas.

b. Mercury II hydroxide reacts with phosphoric acid to form mercury II phosphate and water.

8-3. Define synthesis, decomposition, single replacement and double replacement reactions using the general equations.

Synthesis:

Decomposition:

Single Replacement:

Double Replacement:

8-4. Know the rules of predictions for reactions listed above, and be able to write a balanced equation given the reactants of a reaction.

a. solid Silver oxide exposed to heat

b. zinc reacts with lead II nitrate

c. sodium and oxygen gas

d. potassium iodide and lead II nitrate

e. magnesium and fluorine

f. aluminum reacts with mercury II acetate

8-5. The activity series ranks elements in order of what? How is this used to predict the products in single replacement reactions?

Chapter 9 – Stoichiometry

9-1. Define the term "mole ratio" and explain how to write one.

9-2. In a balanced equation, the coefficients represent the relative numbers of what?

9-3. What temperature is defined as "standard temperature" under "ideal conditions"?

9-4. What pressure is defined as "standard pressure" under "ideal conditions" ?

9-5. Define the terms "limiting reactant" and "excess reactant".

9-6. Define the term "percentage yield" and explain how it is calculated.

9-7. Be able to perform any of the stoichiometry calculations. Use the following balanced chemical equation to answer the following problems.

$$C_3H_8\ (g) + 5O_2\ (g) \longrightarrow 3CO_2\ (g) + 4H_2O\ (l)$$

a. How many moles of propane are given in the equation?

b. How many moles of carbon dioxide gas would be produced from 1.89 mol of oxygen?

c. How many moles of water would be produced from 7.35 g of propane?

d. What mass of water would be produced from 16.3 g of propane?

9-8. Be able to determine the limiting reactant.

a. Zinc and sulfur react to form zinc sulfide according to the following equation:

$$8Zn\ (s) + S_8\ (s) \longrightarrow 8ZnS\ (s)$$

If 2.00 moles of zinc are heated with 1.00 mol of sulfur, identify the limiting reactant.

9-9. Be able to determine the percentage yield.

a. If 75.0 g of CO reacts to produce 68.4 g CH_3OH, what is the percentage yield of CH_3OH?

$$CO\ (g) + 2H_2\ (g) \longrightarrow CH_3OH\ (l)$$

Chapter 12 – Solutions

12-1. List and give examples of the 6 types of solute-solvent combinations for solutions.

12-2. Compare and contrast properties of solutions to those of suspensions and colloids.

12-3. Define the following terms:
a. solute.

b. solvent.

12-4. Define the following terms:
a. soluble.

d. insoluble.

12-5. Compare electrolytes to nonelectrolytes and provide examples.

12-6. Compare unsaturated, saturated and supersaturated solutions.

12-7. Calculate the molarity of a solution containing 0.2 mol of sodium hydroxide dissolved in enough water to make a 0.5 L solution.

12-8. What is the molarity of a methanol solution that contains 25 g of methanol in 3.5 L of a solution? The molar mass of methanol is 32 g/mol.

12-9. Calculate the molality of a solution containing 44 g of formic acid dissolved in 470 g of nitrobenzene. The molar mass of formic acid is 36 g/mol.