Midterm Review

Unit 1 Atomic Structure & Nuclear Chemistry.

1. a. List the three subatomic particles and their charges, locations, and relative masses.

Subatomic particle	charge	location	relative mass
proton			
electron			
neutron			

b.Which subatomic particle is the lightest in mass?______________________________

2. Isotopes differ in the number of __________________ not __________________ (which always must stay the same as they indicate the atomic number!)

Ex. $^{A}_{Z}X$

Z represents the ____________
A represents the ____________

Ex. F-20

the Atomic Number is ____________
the Mass Number is ____________

3. Fill out the chart below for each of the three isotopes

Isotope	Atomic number	Protons	Electrons	Mass Number	Neutrons
^{238}U					
$^{16}O^{2-}$					
$^{23}Na^{1+}$					

4. What is the difference between Average Atomic Mass, Isotopic Mass, and Mass Number?
Average Atomic Mass:

Isotopic Mass:

Mass Number:

5. Balance the Nuclear Reactions (Remember Law of Conservation of Mass- both the mass and the charge have to be balanced on both sides

a. __________ $\rightarrow$ $^{37}_{17}Cl + ^{1}_{1}H$

b. $^{65}_{30}Zn$ + __________ $\rightarrow$ $^{64}_{29}Cu + ^{1}_{0}n$

6. a.Give the symbol and properties of the following decay symbols:

Nuclear Decay	Symbol	Mass	Charge	Penetration Power
Alpha				
Beta				
Gamma				

b. Which type of decay listed above is ENERGY, not a particle?

7. Balance the following nuclear reactions and label Alpha Decay or Beta Decay

a. $^{220}Rn \rightarrow {}^{4}He$ + ________ b. $^{216}Po \rightarrow {}^{0}_{-1}e$ + ________

8. Half Life Calculation: Tritium, ^{3}H, has a half-life of 12.3 years. How long would it take for a 40.0 g sample to decay down to 1.25 g?

9. Half Life Calculation: Fe-61 has a half-life of 6.00 min. Of a 100.0 g sample, how much will remain after 18.0 min?

10. Half Life Calculation: After 20.0 days, a 120 kg sample of Bi-210 decays down to just 7.5 kg. What is its half-life?

11. List the five indicators of a chemical change

G	
O	
P	
E	
C	

12. Label each of the following as a Physical or Chemical Property

Flammability	
Neutralize a Base	
Density	
Boiling Point	

13. Label each of the following as a Physical or Chemical Change

A Tire is Inflated with Air.	
Iron rusts, turning reddish brown in color.	
Food is Digested in the Stomach.	
Water is Heated and Changes to Steam.	

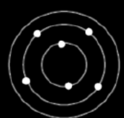
Unit 2: Electrons & Light

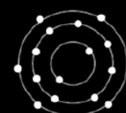
1. According to Bohr's Model of the Atom, where is the only place electrons can be located?

2. Describe the electron cloud.

3. Draw the Bohr Model of the ^{7}Li isotope.

4. Label each diagram with the element (***name and symbol***) they represent. list how many valence electrons each atom has in its outer shell.

   a. 

   b. 

5. a. Describe the difference between the ground state and an excited state (think about Bohr Hydrogen Model).

   ground state:

   excited state:

   b. When an electron moves from a higher E level to a lower E level, does it absorb or release energy?

   c. When an electron moves from a lower E level to a higher E level, does it absorb or release energy?

6. _______________: a packet of light that is released when an electron travels from a higher energy level to a lower energy level.

7. What is the relationship between
   a. Wavelength and Frequency INVERSE OR DIRECT

   b. Energy and Frequency INVERSE OR DIRECT

8. Use the Bohr model diagram on your reference sheet. What is the wavelength of a photon emitted when the electron falls from the third energy level to the second energy level? What type of electromagnetic radiation is it?

9. Use the Bohr model diagram on your reference sheet. What is the wavelength of a photon emitted when the electron falls from the sixth energy level to the third energy level? What type of electromagnetic radiation is it?

10. Niels Bohr produced a model of the hydrogen atom based on experimental observations. This model indicated that:

    a. An electron **surrounds** the nucleus only in fixed energy ranges called ______________________.

    b. The lowest energy orbit is ______________________ to the nucleus.

11. a. What is a "group" on the periodic table?

    b. How many are there on the periodic table?

    c. What does the roman numeral next to the letter A in the main group section on the table represent?

12. Name two things main group elements in the same family have in common.

13. a. What is a "period" on the periodic table?

    b. How many periods are on the periodic table?

14. Identify the elements based on their location: metals, nonmetals and metalloids

Right side of staircase	
Left side of staircase	
Along the border of the stair case with exception of Al	

15. Label the following elements as a Metal, Nonmetal, or Metalloid

a. Oxygen b. Lead c. Silicon d. Magnesium e. Boron f. Neon

16. Give the location (what group they are in) of the following families on the Periodic Table:

a. Representative/Main Group b. Halogens c. Alkaline Earth Metals

a. Alkali Metals e. Noble Gases f. Transition Elements

17. Define atomic radius.

18. What is the group and period trend for atomic radius and give an explanation for why?

	Does it increase or decrease?	Explanation
period trend		
group trend		

19. Put the following in order by decreasing atomic radius: Al, Na, S, K.

20. Define ionization energy.

21. What is the group and period trend for ionization energy?

	Does it increase or decrease?	Explanation
period trend		
group trend		

22. Put the following elements in order of decreasing ionization energy: Rb, Al, S, Mg

23. Define electronegativity.

24. What is the group and period trend for electronegativity?

	Does it increase or decrease?	Explanation
period trend		
group trend		

25. Put the following elements in order of increasing electronegativity energy: F, B, N, Li

26. The smaller an element is in size, the ________________ the ionization energy, and the __________________ the electronegativity.

27. Name the four sublevels, the shape of the orbital and the area it represents on the periodic table.

	Shape	Groups on the periodic table
s sublevel		
p sublevel		
d sublevel		
f sublevel		

28. Write the complete electron configuration for the following elements. Then, list the valence # of electrons for each.

a. Boron

b. Selenium

29. Draw the orbital notation for the following. then, list how many unpaired electrons are in each.

a. Sulfur

b. Nickel

30. Write the shorthand configuration for each.

a. Iron

b. Bromine

31. Identify the element represented by the configurations below:

a. $1s^2 2s^2 2p^4$

b. $[Ar] 4s^2 3d^{10} 4p^3$

32. Determine the number of valence electrons in the following configurations:

a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

b. $[Kr] 5s^1$

33. List the number of valence electrons for each of the following using the periodic table

a. Sodium

b. Nitrogen

c. Bromine

34. Write the ion configuration for:

a. Chloride ion

b. Sodium ion

35. Using the configurations below, first determine the element it represents, then determine the number of electrons lost or gained and the ionic charge(oxidation number) it will form when becoming an ion.

a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$

b. $1s^2 2s^2 2p^6 3s^2 3p^4$

Unit 3: Bonding & Naming and Formula Writing Compounds

1. Metallic bonds: are between all metals and when a Metallic bond forms, the electrons are all "piled together" in a sea of electrons. (Delocalized Electrons)

2. How are ions formed? Which arrangements are stable (What rule are we trying to get to…)?

3. Explain the difference between cation and anion. (what element do they stem form and do they lose or gain valence electrons)

   cation

   anion

4. Give the Ionic charge(oxidation number) for the following groups:

   a. Group 1 b. Group 2 c. Group 13 d. Group 15 e. Group 16 f. Group 17

5. a. What happens with the electrons when an ionic bond forms?

   b. What types of elements form Ionic compounds?

   c. Explain their electronegativity differences.

6. State 5 properties of Ionic Compounds
   a.
   b.
   c.
   d.
   e.

7. Write the chemical formula for each of the following:

   a. Magnesium Fluoride d. Calcium Nitride

   b. Sodium Carbonate e. Ammonium Phosphate

   c. Copper (III) Bromide f. Tin (IV) Oxide

8. Write the names for the following chemical formulas:

   a. FeS d. VBr_3

   b. NH_4NO_2 e. $Ba(OH)_2$

   c. Al_2S_3 f. $(NH_4)_3P$

9. a. What happens with the electrons when a covalent bond forms?

   b. What types of elements form Covalent (Molecular) compounds?

10. What are the 2 types of covalent bonds, their EN difference and the types of atoms that form them?

	Types of Atoms	EN difference
Polar Covalent Bond		
Nonpolar Covalent Bond		

11. Draw the Lewis Diagram for O_2, I_2, and N_2. ***Indicate if they contain single, double, or triple bonds***.

12. What is the relationship between bond energy and bond length of the single, double and triple bonds?

13. Write the chemical formulas for each of the following:

a. Carbon Tetrachloride
c. Dinitrogen Pentoxide
e. Carbon Monoxide
b. Nitrogen Monoxide
d. Phosphorus Pentafluoride
f. Nitrogen Trifluoride

14.. Write the names for the following chemical formulas:

a. PF_5
c. H_2SO_4
b. N_2O_3
d. $HC_2H_3O_2$

15. Complete the Chart

Molecule	**Draw the Lewis Structure**	**Geometry**	**Polarity**	**Predominant IMF's**
NH_3				
SO_2				
CF_4				
HBr				
CO_3^{2-}				

16. Explain why Intermolecular Forces are weaker than Ionic, Covalent or Metallic bonds (Intramolecular Forces).

17. Explain why Hydrogen Bonds are stronger than Dipole-Dipole Forces which are stronger than Dispersion Forces using the terms permanent and temporary dipole

Unit 4: Chemical Reactions

1. Using the “Guidelines of Chemical Reactions” on the reference sheet, predict the Type of Reaction (#/letter designation) in the margin and then write a complete balanced reaction for each of the following. Don’t forget that you never take subscripts across the yield sign unless they are part of the polyatomic ion. ALSO, you must check charges/criss cross if necessary for all new ionic compounds formed and remember the diatomic elements get a subscript of 2 in their formula (BrINClHOF)

    a. $Mg + O_2 \rightarrow$

    b. $Ag + Cl_2 \rightarrow$

    c. $ZnCl_2 \rightarrow$

    d. $H_2S \rightarrow$

    e. $Li + CuCl_2 \rightarrow$

    f. $Na + HBr \rightarrow$

    g. $Cl_2 + KF \rightarrow$

    h. $Pt + CaF_2 \rightarrow$

    i. $K_2S + ZnF_2 \rightarrow$

    j. $Ba(OH)_2 + HNO_3 \rightarrow$

    k. $NaCl + Pb(NO_3)_2 \rightarrow$

2. Use Activity Series to predict whether the following reactions take place

    a. $Cu + 2AgNO_3 \rightarrow$ Yes Reaction Occurs or No Reaction

    b. $2KBr + I_2 \rightarrow$ Yes Reaction Occurs or No Reaction

3. Use the Solubility Rules to determine the precipitate in each of the following reactions

a. $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$ Precipitate: ______________

b. $2NaOH + CuSO_4 \rightarrow Cu(OH)_2 + Na_2SO_4$ Precipitate: ______________

ANSWER KEY

[redacted]