The midterm exam covers chapters 1 - 4 & 9 - 11. You should read through each chapter, look over old tests you still have, answer the following questions and do the calculations in order prepare yourself for the mid-term.

1. Define the following terms and describe where each is located.
   - **Proton**: Positive subatomic particle (inside nucleus)
   - **Neutron**: Neutral subatomic particle (inside nucleus)
   - **Electron**: Negative subatomic particle (outside nucleus)

2. Complete the following table:

<table>
<thead>
<tr>
<th>Element name</th>
<th>Atomic #</th>
<th>Mass #</th>
<th># of protons</th>
<th># of Neutrons</th>
<th># of Electrons</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>1</td>
<td>1</td>
<td>0</td>
<td>1</td>
<td>H</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>6</td>
<td>6</td>
<td>6</td>
<td>6</td>
<td>C</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>11</td>
<td>11</td>
<td>12</td>
<td>11</td>
<td>Na</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
<td>20</td>
<td>20</td>
<td>20</td>
<td>20</td>
<td>Ca</td>
</tr>
<tr>
<td>Fe</td>
<td>Fe</td>
<td>26</td>
<td>26</td>
<td>30</td>
<td>26</td>
<td>Fe</td>
</tr>
</tbody>
</table>

3. Define the following and give an example of each using chemical symbols.
   - **Atomic number**: Whole # on periodic table (# of protons) / H
   - **Atomic mass**: Decimal # on periodic table (# of p + n) / H
   - **Isotope**: Same # of p different # of neutrons / H and H<sup>+</sup>

4. Which of the following are isotopes of the same element?
   - \(^{19}\text{Y}, ^{20}\text{Y}, ^{19}\text{Y}\)

5. Describe Rutherford's experiment:
   - **Gold Foil Experiment**: He shot alpha particles at gold foil. He expected the particles to go straight through, but some were deflected. He discovered the positively charged nucleus.

6. Explain all the major parts of Dalton's Atomic Theory.
   - **a. Law of constant composition**: Compounds are formed in whole # ratios
   - **b. All atoms of the same element are identical** (false b/c of isotopes)
   - **c. Atoms are indivisible** (false b/c of subatomic particles)
   - **d. All elements are composed of atoms**

7. Balance the following chemical equations:
   - \(3 \text{CO} + \text{Fe}_2\text{O}_3 \rightarrow 2 \text{Fe} + 3\text{CO}_2\)
   - \(3\text{Zn(OH)}_2 + 2\text{H}_3\text{PO}_4 \rightarrow \text{Zn}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}\)

8. Define Ionic and Molecular compounds, and tell how each is formed.
   - **Ionic**: Metal bonded to 1 or more non-metals
   - **Molecular**: Two or more nonmetals bonded together

9. Name the following compounds and state if it is ionic or molecular in nature:
   - **Ionic**: Al(OH)<sub>3</sub>, Aluminium hydroxide, MgCl<sub>2</sub>, Magnesium chloride, Cl<sub>2</sub>O<sub>7</sub>, Dichlorine heptoxide
   - **Molecular**: N<sub>2</sub>O<sub>3</sub>, dinitrogen pentoxide, MgH<sub>2</sub>, Magnesium hydride
10. Write a chemical formula for each name given and tell whether it is an ionic (i) or molecular (m) compound:

- **a. Ammonium Phosphate** \((\text{NH}_4)_3\text{PO}_4\) – Ionic
- **b. Magnesium Nitride** \(\text{Mg}_3\text{N}_2\) – Molecular
- **c. Oxygen Difluoride** \(\text{OF}_2\) – Molecular
- **d. Carbon Dioxide** \(\text{CO}_2\) – Molecular
- **e. Sulfur Dioxide** \(\text{SO}_2\) – Molecular

11. Name and describe the 6 types of chemical reactions. Give an example of each:

- **a. Combustion** \(\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}\)
- **b. Synthesis (Combination)** \(\text{Na} + \text{Cl}_2 \rightarrow \text{NaCl}\)
- **c. Decomposition** \(\text{H}_2\text{O} \rightarrow \text{H}_2 + \text{O}_2\)
- **d. Single Replacement** \(\text{Na} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\)
- **e. Double Replacement** \(\text{NaCl} + \text{HgO} \rightarrow \text{Na}_2\text{O} + \text{HgCl}_2\)
- **f. Oxidation-Reduction** \(\text{Na} + \text{Cl}_2 \rightarrow \text{NaCl}\)

12. Define and give an example of each:

- **Element** – substances that contain only one type of atom
- **Mixture** – a physical blend of two or more components
- **Compound** – two or more elements chemically combined
- **Homogeneous substance** – a mixture with uniform composition throughout
- **Heterogeneous substance** – a mixture that does not have uniform composition
- **Physical property** – a property that can be observed
- **Chemical property** – ability to undergo a change in chemical composition
- **Physical change** – change that produces matter with a different composition
- **Chemical change** – change that produces matter with a different composition
- **Qualitative measurement** – measurements made by observations
- **Quantitative measurement** – measuring temperature

13. Which of the following is a homogeneous mixture?
   - **a. oil in water**
   - **b. soot in water**
   - **c. alcohol in water**

14. Which of the following could be considered a physical change?
   - **a. cooking a pancake**
   - **b. burning a tree**
   - **c. melting an ice cube**

15. Which of the following is considered a heterogeneous mixture?
   - **a. salt and sugar**
   - **b. flour and baking powder**
   - **c. salt and pepper**

16. Classify each as a physical or chemical change.
   - **P** a. instant coffee is combined with hot water to produce a brown liquid
   - **C** b. from exposure to air and moisture, iron turns reddish and cannot conduct electricity
   - **P** c. iron is heated, turns red and then melts
   - **C** d. sugar is heated to produce steam and a black solid
1. Classify each as an element, mixture, ionic compound or molecular compound.
   a. sodium  
   b. water  
   c. table salt  
   d. sugar  
   e. oxygen  
   f. air  
   g. soil  
   h. lemon soda

2. Classify each as a qualitative or quantitative observation:
   a. the liquid solution was blue  
   b. the reaction gave off smoke  
   c. 5 grams of the chemical was used  
   d. the temperature was 87 degrees  
   e. the metal was smooth

3. List the diatomic molecules:
   a. H  
   b. O  
   c. Br

4. List the names & formulas of the six common acids:
   a. H2SO4 - sulfuric acid  
   b. HCl - hydrochloric  
   c. HNO3 - nitric  
   d. H2CO3  
   e. HC2H307  
   f. H3PO4 - phosphoric

5. Define:
   Metal (left of staircase)  
   Non-metal (right of staircase)  
   Metalloid

   Metal: 1 - 3 valence e's, become cations in ionic compounds, lose e's, luster, malleable, conductors
   Non-metal: opposites of above...
   Metalloid: has properties of both metals & nonmetals

6. Classify each element as a metal, non-metal, or metalloid.
   a. aluminum  
   b. gold  
   c. silicon  
   d. hydrogen  
   e. argon

7. Define groups and periods
   Describe how elements are arranged on the periodic table:
   by atomic mass or atomic charge

8. What are the main groups of elements on the periodic table and where are they located?
   Alkali metals / 1st column  
   Halogens / 17th column  
   Noble gases / 18th column

9. What is special about the elements in a particular group on the periodic table?
   Some chemical properties b/c form same ions

10. What is the oxidation (nuclear) charge of each substance (ion) given? Answers Given
    a. Al 3+  
    b. S 2-  
    c. Cl 1-  
    d. phosphorus 3-  
    e. nitrate 1-  
    f. carbonate 2-  
    g. lithium 1-  
    h. Ag 1-
11. What is the total positive charge on the Aluminum ion in the following compounds?
a. Al(ClO₄)₃ b. Al₂(SO₄)₃ c. AlPO₄

12. For each compound in question # 25 give the following information:

<table>
<thead>
<tr>
<th></th>
<th>Al(ClO₄)₃</th>
<th>Al₂(SO₄)₃</th>
<th>AlPO₄</th>
</tr>
</thead>
<tbody>
<tr>
<td># of moles of atoms for each element</td>
<td>Al 1 mole</td>
<td>Al 2 moles</td>
<td>Al 1 mole</td>
</tr>
<tr>
<td></td>
<td>Cl 5 mole</td>
<td>S 3 moles</td>
<td>P 4</td>
</tr>
<tr>
<td></td>
<td>O 12 mole</td>
<td>O 6</td>
<td>O 4</td>
</tr>
<tr>
<td>the total number of atoms in the entire compound</td>
<td>16 atoms</td>
<td>11 atoms</td>
<td>16 atoms</td>
</tr>
<tr>
<td>the gram molecular mass of compound</td>
<td>325 g</td>
<td>246.1 g</td>
<td>121.8 g</td>
</tr>
</tbody>
</table>

13. What is Avogadro's number? 6.02 x 10²³

14. Define the following:

- Molecule: group of bonded atoms
- Atom: smallest particle of matter that retains individual properties
- Ion: atom w/ a charge
- Cation: ion w/ positive charge
- Anion: ion w/ negative charge

15. From what type of elements are cations and anions formed and explain how each is formed.

Cations: metal - loss of electron(s) (-)
Anions: nonmetal – gain electron(s)

16. Calculate the % composition by mass of the compounds formed from these reactions.

a. 8.2 g of Mg combine with 5.4 g of oxygen
   Mg = 60.3% - O = 39.7%

b. 29 g of Ag combine with 4.3 g of sulfur
   Ag = 87.1% - S = 12.9%

17. Calculate the % composition by mass of:

- Propane C₃H₈
  C = 81.2% H = 18.9%

- Water H₂O
  H = 11.1% O = 88.9%

18. Element X has two isotopes. The first isotope has a mass of 10.012 amu with a relative abundance of 19.91%. The second has a mass of 11.009 and has a relative abundance of 80.09%. Calculate the atomic mass of this element, and name it.

19. The four isotopes of lead are given below, each with its percent by mass abundance and the composition of its nucleus. Using this data, calculate the atomic mass of lead.

<table>
<thead>
<tr>
<th></th>
<th>Pb</th>
<th>Pb</th>
<th>Pb</th>
<th>Pb</th>
</tr>
</thead>
<tbody>
<tr>
<td>p+ = 82</td>
<td>n= 122</td>
<td>137%</td>
<td>26.26%</td>
<td></td>
</tr>
<tr>
<td>p+ = 82</td>
<td>n= 124</td>
<td>125</td>
<td>20.82%</td>
<td></td>
</tr>
<tr>
<td>p+ = 82</td>
<td>n= 126</td>
<td>126</td>
<td>51.55%</td>
<td></td>
</tr>
<tr>
<td>Mass&gt; 208 = 92 x 122 = 104</td>
<td>206 + 208</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(204 x 137% + 206 x 26.26% + 207 x 20.82%) + (208 x 51.55) = 207.2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Hint: for #17, and 18 use the formula: % mass (of each element) = grams of element grams of compound x 100%
Chemistry - Mid-Term Exam
Study Guide: 3

1. A copper penny has a mass of 3.1 g and a volume of 0.35 cm$^3$. What is the density of copper?
2. A liquid has a density of 4.8 g/ml. What is the mass of a 2 liter sample?
3. What is the volume of a substance that has a mass of 80 g and a density of 10 g/cm$^3$?
4. Indicate the meaning (as a power of 10) for each of the following metric prefixes:
   a. kilo
   b. centi
   c. milli
   d. deci
   e. nano
   f. Micro
5. Calculate the following quantities:
   a. 1,100 cm = __________ m
   b. 1 m = __________ mm
   c. 10 m = __________ cm
   d. 2.5 km = __________ m
   e. 4.05 kg = __________ g
   f. 0.5 g = __________ mg
6. Indicate the number of significant figures in each of the following:
   a. 12600 __________
   b. 0.09 __________
   c. 2001 __________
   d. 0.00500100 __________
   e. 1000 __________
7. Define:
   a. accuracy
   b. precision
8. The accepted value or true value for the density of lead (Pb) is 11.35 g/ml. Your experimental value or observed value found during a class lab is 9.65 g/mL. What is the error of your measurement?
   What is the percent error of your measurement?
9. Define:
   a. Meter
   b. Liter
   c. Volume
   d. Mass
   e. Gram
   f. Temperature
10. Name the two temperature scales used in science? Give the freezing pt. and boiling pt. of water for each of them.
11. Which type of particle (atom, ion, or molecule) goes with each of the following substances?
   a. Na __________ atom
   b. Ca$^{2+}$ __________ ion
   c. N$_2$ __________ molecule
   d. Cl$_2$ __________ molecule
   e. H$_2$O __________ molecule
   f. CO __________ molecule
12. Define:
   a. empirical formula
   b. molecular formula
13. Which of the following are empirical formulas and which are molecular formulas?
   a. CH$_4$N __________
   b. NaO __________
   c. C$_6$H$_3$O$_3$ __________
14. Find the empirical formula of each compound from its % composition.
   a. 72.4% Fe and 27.6% O
   b. 94.1% O and 5.9% H

15. If given the empirical formula and gram formula mass for a compound, calculate the compound's molecular formula?
   a. CH₄O, mass = 90 g/mol
   b. C₃H₆O₃, mass = 146 g/mol

16. Find the missing density, mass of volume of the following:
   a. The mass of a substance is 45.6 g and the volume is 15 cm³:
   Density = 
   b. The volume of a substance is 2.9 ml its density is 6 g/ml:
   Mass =
   c. The density of a substance is 7.8 g/cm³ and the mass is 125 g:
   Volume =
   (Hint: D = M/V (Given any two of the numbers; D, M or V, you can cross multiply and divide to find what's missing)

17. If you have 6.7 L of O₂ at STP, how many moles do you have

18. What is the molar mass of Sn₃(PO₄)₂?

19. How many moles are in 137.5 g of Mn?

20. What is the mass of 3 moles of Sc?

21. What is the mass of 2 moles of C₂H₄?

22. What are the correct formulas for the following compounds?
   a. potassium sulfate K₂SO₄
   b. calcium phosphate Ca₅(PO₄)₂

23. How many moles of CaCl₂ are in 12 g of CaCl₂?
   \[
   \frac{12 \text{g}}{\text{mol}} = \frac{X \text{mol}}{1 \text{mol}}
   \]

Finding % composition from Mass of elements in a compound:
   What is the percent mass of each element in K₂O if the mass of the compound is 188 g and the mass of oxygen is 32 g? (Hint: Mass of K must be 188 − 32 = 156 g)
   K = 156/188 = 83%
   O = 32/188 = 17%

Finding % composition from the chemical formula of elements in a compound:
   What is the percent mass of the elements in C₃H₈? (Hint: Find molar mass of each element and divide by molar mass of compound).
   Molar Mass of C₃H₈ = 44 g
   Mass of 3 moles C = 36 g
   Mass of 8 moles H = 8 g
   = 18%

Finding empirical formulas by % mass of a compound:
   A compound consists of 80% carbon and 20% Hydrogen. What is its empirical formula? (Hint divide each % by the molar mass of the element)
   C = 80/12 = 6.7
   H = 20/1 = 20
   The ratio of 20 to 6.7 is 3 to 1 (20/6.7 = 2.99) so there are 3 times as many H as C atoms.
   The empirical formula is CH₃