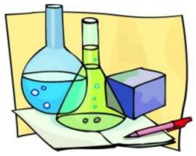

Summer Assignment & Class Information 2018-19

Dear AP Chemistry Student and Parent:

I am very excited to have you in class next year! As you probably already know, AP Chemistry is a very challenging course with an equally challenging AP test. The material is detailed and we have several topics to cover. To aid in getting the most out of your AP Chemistry course, you will need to refresh your memory on information that was taught in PAP Chemistry. This will allow us to review the basics of chemistry within the first week of class, and then move on to expand your knowledge base. **You are required to complete the summer assignment below and bring it and any questions you may have to class the first day.** We will reinforce this review with some lab activities during the beginning of the semester. **You should anticipate a test within the first week of class.** You need to master the formulas, charges, and names of common ions. You will be asked to complete a few tasks involving naming ions and compounds, determining formulas for compounds, use the rules for solubility, oxidation reduction, and transition metal charges.

I have included several resources in this packet. First, there is a list of the ions that you must know. This list also has some suggestions for making the process of memorization easier. For instance, many of you will remember that most of the monatomic ions have charges that are directly related to their placement on the periodic table. There are naming patterns that greatly simplify the learning of the polyatomic ions as well. Also included is a copy of the periodic table used in AP Chemistry. Notice that this *is **not*** the table used in first year chemistry. The AP table is the same that the College Board allows you to use on the AP Chemistry test. Notice that it has the symbols of the elements but ***not*** the written names. You need to take that fact into consideration when studying! Doubtless, there will be some students who will procrastinate and try to do all of this studying just before the start of school. Those students may even cram well enough to do well on the initial test. However, they will quickly forget the ions, and struggle every time that these formulas are used in lecture, homework, quizzes, tests and labs. All research on human memory shows us that frequent, short periods of study, spread over long periods of time will produce much greater retention than long periods of study of a short period of time. I could wait and throw these at you on the first day of school, but I don't think that would be fair to you. Use every modality possible as you try to learn these – speak them, write them, visualize them.

I look forward to seeing you all at the beginning of the next school year. If you need to contact me during the summer, you can email me and I will get back to you. You may check my google classroom for any updates and also to access review material that will give you an idea whether you have adequately mastered the summer assignment.

******Before summer begins, please come by my room (S-211) with a flash drive so that I can give you an electronic copy of your textbook. Also, sign up for my google classroom. Your user name is your DISD email (DISDstudentid@apps.dickinsonisd.org) and your password is DisdMMDDYYYY(your birthday). The classroom code is ____________________.***
Also join REMIND for AP Chemistry. Text 81010 with the message @201819apc.

AP Chemistry Test Thursday, May 9, 2019

Sincerely,
Mrs. Wagner and Mrs. Everist

Recommended supplementary study materials:

AP Chemistry Princeton/Barron's Review – This book is excellent when preparing for the actual AP Chem exam. It gives provides many sample questions and practice questions that will help you prepare for the AP exam.

5 Steps to a 5 – is a more basic review that you will find useful throughout the year.

Study tools - please review these documents and begin to memorize! AP Chemistry is a challenging course. While it is not all about memorization, having these items memorized is essential for success in learning the concepts covered in the course.

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

*Lanthanide Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Know element symbols and names for elements 1-38 and Ag, Cd, I, Xe, Cs, Ba, W, Hg, Pb, Sn, Rn, Fr, U, Th, Pu, and Am written correctly. Students should be able to locate these elements quickly on the periodic table provided on the exam. This table will not have the element names written on it.

Ions to Memorize			
Cations	**Name**		
Ag^{+}	Silver		
Zn^{2+}	Zinc		
Hg_2^{2+}	Mercury(I)		
NH_4^{+}	Ammonium		
Cadmium	Cd	+2	Cadmium
Anions	**Name**		
NO_2^{-}	Nitrite		
NO_3^{-}	Nitrate		
SO_3^{2-}	Sulfite		
SO_4^{2-}	Sulfate		
HSO_4^{-}	Hydrogen sulfate (bisulfate)		
OH^{-}	Hydroxide		
CN^{-}	Cyanide		
PO_4^{3-}	Phosphate		
HPO_4^{2-}	Hydrogen phosphate		
$H_2PO_4^{-}$	Dihydrogen phosphate		
NCS^{-}	Thiocyanate		
CO_3^{2-}	Carbonate		
HCO_3^{-}	Hydrogen carbonate (bicarbonate)		
ClO^{-}	Hypochlorite		
ClO_2^{-}	Chlorite		
ClO_3^{-}	Chlorate		
ClO_4^{-}	Perchlorate		
BrO^{-}	Hypobromite		
BrO_2^{-}	Bromite		
BrO_3^{-}	Bromate		
BrO_4^{-}	Perbromate		
IO^{-}	Hypoiodite		
IO_2^{-}	iodite		
IO_3^{-}	iodate		
IO_4^{-}	Periodate		
$C_2H_3O_2^{-}$	Acetate		
MnO_4^{-}	Permanganate		
$Cr_2O_7^{2-}$	Dichromate		
CrO_4^{2-}	Chromate		
O_2^{2-}	Peroxide		
$C_2O_4^{2-}$	Oxalate		
NH_2^{-}	Amide		
BO_3^{3-}	Borate		
$S_2O_3^{2-}$	Thiosulfate		

Diatomic Elements (7,7,7 rule) H_2, N_2, F_2, Br_2, Cl_2, O_2, I_2

- BrINClHOF (say "Brinkelhof)

Solubility Rules:
Learn the following solubility rules:
Salts containing the following ions are normally **soluble**:
○ All salts of Group IA (Li^+, Na^+, etc.) and the ammonium ion (NH_4^+) are *soluble.*
○ All salts containing nitrate (NO_3^{1-}), acetate (CH_3COO^{1-}) and percholates (ClO_4^{1-}) are *soluble*
○ All chlorides (Cl^{1-}), bromides (Br^{1-}), and iodides (I^{1-}) are *soluble* except those of Cu^+, Ag^+, Pb^{2+}, and Hg_2^{2+}.
○ All salts containing (SO_4^{2-}, are *soluble* except those of Pb^{2+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}).
Salts containing the following ions are normally insoluble:
○ Most carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are *insoluble* except those of Group IA and the ammonium ion.
○ Most sulfides (S^{2-}) are *insoluble except* those of Group IA and IIA and the ammonium ion.
○ Most hydroxides (OH^{1-}) are *insoluble* except those of Group IA, calcium, and barium.
○ Most oxides (O^{2-}) are *insoluble* except for those of Group IA and IIA which react with water to form the corresponding hydroxides

TIPS FOR LEARNING THE IONS

1. Their place on the table suggests the charge on the ion, since the neutral atom gains or loses a predictable number of electrons in order to obtain a noble gas configuration. This was a focus in first year chemistry, so if you are unsure what this means, get help BEFORE the start of the year.

   a. All Group 1 Elements (alkali metals) lose one electron to form an ion with a 1+ charge
   b. All Group 2 Elements (alkaline earth metals) lose two electrons to form an ion with a 2+ charge
   c. Group 13 metals like aluminum lose three electrons to form an ion with a 3+ charge
   d. All Group 17 Elements (halogens) gain one electron to form an ion with a 1- charge
   e. All Group 16 nonmetals gain two electrons to form an ion with a 2- charge
   f. All Group 15 nonmetals gain three electrons to form an ion with a 3- charge
   Notice that cations keep their name (sodium ion, calcium ion) while anions get an "-ide" ending (chloride ion, oxide ion).

2. Metals that can form more than one ion will have their positive charge denoted by a roman numeral in parenthesis immediately next to the name of the

Polyatomic Anions: Most of the work on memorization occurs with these ions, but there are a number of patterns that can greatly reduce the amount of memorizing that one must do.

1. "ate" anions have one more oxygen then the "ite" ion, but the same charge. If you memorize the "ate" ions, then you should be able to derive the formula for the "ite" ion and vice-versa.
   a. sulfate is SO_4^{2}-, so sulfite has the same charge but one less oxygen (SO_3^{2-})
   b. nitrate is NO_3^-, so nitrite has the same charge but one less oxygen (NO_2-)

2. If you know that a sulfate ion is SO_4^{2-} then to get the formula for hydrogen sulfate ion, you add a hydrogen ion to the front of the formula. Since a hydrogen ion has a 1+ charge, the net charge on the new ion is less negative by one.
   a. Example: $PO_4^{\ 3-}$ --> HPO_4^{2-} --> $H_2PO_4^{-1}$
   phosphate hydrogen phosphate dihydrogen phosphate

3. Learn the hypochlorite --> chlorite --> chlorate --> perchlorate series, and you also know the series containing iodite/iodate as well as bromite/bromate.
   a. The relationship between the "ite" and "ate" ion is predictable, as always. Learn one and you know the other.
   b. The prefix "hypo" means "under" or "too little" (think "hypodermic", "hypothermic" or "hypoglycemia")
      i. Hypochlorite is "under" chlorite, meaning it has one less oxygen
   c. The prefix "hyper" means "above" or "too much" (think "hyperkinetic")
      i. the prefix "per" is derived from "hyper" so perchlorate (hyperchlorate) has one more oxygen than chlorate.
   d. Notice how this sequence increases in oxygen while retaining the same charge:

ClO^-	-->	ClO_2^-	-->	ClO_3^-	-->	ClO_4^-
hypochlorite		chlorite		chlorate		perchlorate

Nomenclature Review

Forming binary ionic compounds

A. In a binary ionic compound the total positive charges must equal the total negative charges. The best way to write correct formula units for ionic compounds is to use the "Criss-cross method".
B. Sample Problem: What ionic compound would form when calcium ions combine with bromide ion?

   Steps to the Criss-cross Method:
   1. Write the ions with their charges, cations are always first. Ca^{2+} Br^{1-}
   2. Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion. Ca_1Br_2
   3. Check to make sure the subscripts are in the lowest whole number ratio possible. Then write the formula. $CaBr_2$

Naming binary ionic compounds

A. Combine the names of the cation and the anion, and change the second half of the anion to -ide.

B. Example: $BaBr_2$ is named barium bromide

Naming binary ionic compounds that contain polyatomic ions.

A. The polyatomic ions on your common ion list should be memorized.

B. The most common oxyanions (polyatomic ions that contain oxygen) end in –ate.
Oxyanions with one less oxygen end in –ite. For example:

NO_3^{1-} is nitrate	SO_4^{2-} is sulfate
NO_2^{1-} is nitrite	SO_3^{2-} is sulfite

C. Anions with one less oxygen than the –ite ion are given the prefix hypo-.

D. Anions with one more oxygen than the –ate ion are given the prefix per-.

ClO^{1-} is hypochlorite	ClO_3^{1-} is chlorate
ClO_2^{1-} is chlorite	ClO_4^{1-} is perchlorate

E. Naming compounds with polyatomics is the same as naming other compounds, just name the cation and then the anion. If there is a transition metal involved, be sure to check the charges to identify which ion (+1, +2, +3, +4…) it may be so that you can put the correct Roman numeral in the name.

Some preliminary notes from Chapters 1-3

Chapter 1 – Introduction: Matter and Measurement

A. Classification of Matter

1. States of Matter
   a. Gas(vapor)
      i. Has no fixed volume or shape
      ii. Takes the shape of its container
      iii. Can be compressed or expanded
      iv. Molecules are far apart and moving at high speeds
   b. Liquid
      i. Definite volume, cannot be compressed
      ii. Takes the shape of its container
      iii. Molecules are much closer than in a gas but still move rapidly (they can slide past each other)
   c. Solid
      i. Definite shape and volume, cannot be compressed
      ii. Molecules are held tightly together, typically in definite arrangements

2. Pure substances and mixtures
   a. Pure substances – matter that has a fixed composition and distinct properties
      i. Two types
         1. Elements – substances that cannot be decomposed into simpler substances
         2. Compounds – composed of two or more elements chemically bonded together
            a. Law of Constant Composition – (Joseph Proust) the makeup of compounds is always the same

b. Mixtures – combination of two or more substances in which each substance retains its own chemical identity and properties
    i. Properties can vary
        1. Example – adding sugar to coffee is a mixture, you can make it very sweet, add a little, or none at all.
    ii. Two types
        1. Heterogeneous – different composition throughout
            a. Rocks, sand, wood, chocolate chip cookies
        2. Homogeneous (aka solutions) – uniform composition throughout
            a. Air(gaseous solution), gasoline(liquid solution), brass(solid solution)

c. Separation of Mixtures
    i. Filtration – separating a solid component from a liquid component using a funnel, filter paper, and gravity
    ii. Distillation – separating liquid components utilizing different boiling points
    iii. Chromatography – separating substances by how they adhere to surfaces (used frequently for ink)

3. Properties of Matter
    a. Physical properties – description of what something looks like
        i. Color, odor, density, melting point, boiling point, hardness
    b. Chemical properties – how a chemical reacts with other chemicals
        i. Flammability, reactivity with other chemicals
4. Changes in Matter
    a. Physical changes – physical appearance is changed
        i. Ripping paper, melting wax, ALL CHANGES OF STATE (BOILING, EVAPORATING)
    b. Chemical changes (reactions) – chemically transformed into a new substance
        i. Sodium metal reacts with chlorine gas to form salt

B. Units of Measurement

I. Metric System

SI Prefixes

The Great Monarch King Henry Died By Drinking Chocolate Milk Made Near Poland

- The – Tera
- Great – Giga
- Monarch – Mega
- King – Kilo
- Henry – Hecto
- Died – Deca
- By – base unit – m, l, g, cal, J
- Drinking – deci
- Chocolate – centi
- Milk – milli
- Made – micro - μ
- Near – nano
- Poland – pico

II. Significant Figures

Leading Zeros –

Captured Zeros-

Trailing Zeros-

Adding and Subtracting Rules:

Multiplying and Dividing Rules:

Chapter 2 – Atoms, Molecules, Ions

A. The Atomic Theory of Matter

1. History of the Atom

a. Democritus – first person to speculate that matter was mass of atoms. Greek philosopher

i. Plato and Aristotle refuted this idea, atomic theory faded for many centuries

b. John Dalton – came up with first atomic theory, English school teacher

i. Each element is composed of extremely small particles called atoms

ii. All atoms of a given element are identical; the atoms of different elements are different and have different properties (including different masses)

iii. Atoms of an element are not changes into different types of atoms by chemical reactions; atoms are neither created nor destroyed in chemical reactions

iv. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atom.

Dalton thought that atoms could not be broken down any further, this was expressed in the atomic model – Billiard Ball Model

Laws from the time period

a. Law of Constant Composition

b. Law of Conservation of Mass (LeChatelier) – matter and energy cannot be created or destroyed

c. Law of Multiple Proportions – if elements combine to form more than one compound they must be different by whole numbers.

i. Carbon monoxide, CO, carbon dioxide, CO_2

c. Cathode Rays – a high voltage electricity passed through partially evacuated tubes produced radiation and mass glass fluoresce, called cathode rays because they originated from the cathode

i. Rays were deflected by electric and magnetic fields, suggesting that the rays were charged

ii. J.J. Thomson – observed that the rays were the same no matter what type of material was used, concluded that the rays were actually particles with mass, these particles were called electro

1. Able to calculate the charge to mass ration of an electron, 1.76×10^8 Coulombs/gram
2. *Came up with second atomic model – Plum Pudding Model*

d. Robert Millikan – performed the oil drop experiment and determined the charge of an electron (1.60×10^{-19}) and then determined the mass of an electron (9.11×10^{-28}g)

e. Henri Becquerel – studied an ore of Uranium called pitchblende and discovered the spontaneous emission of radiation called radioactivity

i. Marie Curie and her husband, Pierre also studied this

f. Ernest Rutherford – studied radiation and discovered three types of radiation: alpha, beta, and gamma

i. Utilizing alpha particles, Rutherford performed the Gold Foil Experiment and determined that the atom had a nucleus

ii. Also discovered protons

g. James Chadwick – discovered neutrons

2. Modern View of Atomic Structure

a. Atoms are made of protons, neutrons, and electrons

b. Electronic charge is measured in Coulombs (C)

i. Electrons have a charge of -1.60×10^{-19} C

ii. Protons have a charge of $+1.60 \times 10^{-19}$ C

iii. For simplicity we change this to +1 and -1, but you should still know what the value is

iv. Neutrons have no charge

c. Atoms are typically neutral, which means they have the same number of protons and electrons

d. Protons and neutrons are in the nucleus, electrons circle around

e. Vast majority of an atom's volume is the space where the electrons are found

f. Isotopes – atoms of a given element that differ in the number of neutrons
g. Protons – all atoms of an element have the same number of protons in the nucleus, aka atomic number
h. Mass number – number of protons + number of neutrons

3. Periodic Table
    a. You should know the general layout of the periodic table (groups, rows, where the metals, nonmetals, and metalloids are)

4. Writing chemical formulas (reviewed earlier in packet)

Chapter 3 – Stoichiometry: Calculations with Chemical Formulas and Equations

1. All chemical equations need to be written correctly and balanced appropriately (kind of redundant I know)
2. We will go over all of the types of chemical reactivity but below are some for review
    a. Most common involve oxygen as a reactant
    b. Often involve hydrocarbons (compounds that contain hydrogen and carbon)
3. Atomic and Molecular Weight
    a. Atomic Mass Scale – is based off of Carbon – 12, mass of carbon – 12 = 12 amu
    b. Amu = atomic mass unit, 1g = 6.022×10^{23} amu

4. Average Atomic Masses
    a. the masses listed on the periodic table are weighted averages based on the abundance in nature
    b. see example problems in book
5. Percent Composition from Formulas
    a. part/whole x 100%
    b. used to determine how much of a compound is a particular kind of element
6. The Mole
    a. used to convert between the microscopic and the macroscopic
    b. Avogadro's number = 6.02×10^{23}
7. Problems – work through chapter problems if you need extra help

Other Helpful Equations for the Summer Assignment

- Density = mass/volume
- n = mass/ MM
- PV= nRT
- Percent Composition = (total mass of the element/total mass of the compound) x100
- Empirical Formula = Reduced Formula

    Find it by using the 4 steps
    1. Find the grams for each element
    2. Determine the number of moles for each element
    3. Divide by the smallest mole number to get a whole number.
    4. If all are whole numbers, write the empirical formula
- Molecular Formula = the Real Formula

    The molecular formula is a multiple of the empirical formula. To determine the multiple use the following equation.
    - Multiple = Molecular Molar Mass / Empirical Formula Mass
- Molarity = moles/Liters
- Moles = Molarity x Liters

AP CHEMISTRY SUMMER ASSIGNMENT

Show ALL work!!!! Use dimensional analysis. Report answers using the correct number of significant figures. Due 7 days after school starts! The answers are given in parenthesis at the end of each problem. Good Luck and Have a Great Summer!

1. When Mount St. Helens erupted in 1980, analysts collected ash samples to determine the amount of toxic heavy metals that were introduced to the environment. The ash was determined to contain 0.0018% mercury. How many moles of mercury are contained in 500.0g of ash from the volcano? (*$4.5*10^{-5}$ mol Hg*)
2. One thyroid hormone is thyroxine, $C_{15}H_{11}I_4NO_4$. How many mg of thyroxine could be produced from 210 mg of iodine atoms, the amount that a typical adult consumes per day? How many molecules is this? *(320mg; $2.5*10^{20}$ molecules)*
3. A sample thought to be caffeine, the stimulant found in coffee, tea, and cola, gave the following elemental analysis: 49.5%C, 5.2%H, 28.8%N and 16.5%O. Is it possible that this substance is caffeine which is $C_8H_{10}N_4O_2$? *(Yes!)*
4. Analysis of ibuprofen shows that is contains 75.7% carbon, 8.8% hydrogen, and 15.5% oxygen. The mass spectrum of ibuprofen shows that its molar mass is approximately 210g/mol. What is the chemical formula for ibuprofen? *($C_{13}H_{18}O_2$)*
5. Butyric acid (a hydrocarbon containing C, H, and O), a component of rancid butter, has a vile stench. Burning 0.440g of butyric acid in excess oxygen yields 0.882 g of CO_2 and 0.360 g of H_2O. The molar mass of butyric acid is 88 g/mol. What are its empirical and molecular formula? *(EF =C_2H_4O; MF = $C_4H_8O_2$)*
6. An ore of copper is 80% copper (II) oxide. How many kg of copper could be recovered from 846 kg of ore?*(5.4×10^2 kg Cu)*
7. A solution of sodium hydroxide is made by dissolving 56.0 g of 92.0% pure sodium hydroxide into 750.0 mL of water. What is the molarity of the sodium hydroxide solution?*(1.71M NaOH)*

Density:

8. A solid block of substance is 74.0 cm by 55.0 cm by 29.0 cm and it weighs 625 kg. Assuming that it did not chemically react with water nor dissolve in it, would it float in water? *(No, it would sink!)*
9. A certain liquid has a density of 0.855 g/ml. How many LITERS would weigh 1.00 kg? *(1.17 L)*

Dimensional Analysis and Stoichiometry

10. How many milligrams are there in a 5.00 grain aspirin tablet? 1 grain = 0.00229 ounces. There are 454 grams/pound and 16.0 ounces/pound. *(325mg)*
11. In 1980, the US produces 18.4 billion pounds (1.84×10^{10} lbs) of phosphoric acid to be used in the manufacture of fertilizer. The average cost of the acid is $318/ton. (1 ton = 2000 lbs) What was the total value of the phosphoric acid produced? *($2.93 billion)*
12. In the decomposition of sodium hydroxide, water and sodium oxide are produced. How many moles of sodium hydroxide are needed to produce 30.0 moles of water? *(60.0 moles NaOH)*
13. In the single replacement reaction of lithium and magnesium nitrate, what mass of lithium combines with 75.0 grams of magnesium nitrate? *(7.02 g Li)*
14. How many grams of lead (II) nitrate are needed to produce 60.0 grams of potassium nitrate in the double replacement reaction of potassium iodide and lead (II) nitrate. *(98.2 g lead (II) nitrate)*
15. In the synthesis reaction of zinc(II) and sulfur, what mass of zinc (II) sulfide is produced from 100.0 grams of sulfur? *(303.8 g ZnS)*
16. A synthesis reaction of calcium and oxygen was completed in a lab and 234.9 grams of calcium oxide were produced from 75.00 grams of oxygen. What is the percent yield? *(89.35%)*
17. If you begin with 1250 g of N_2 and 225g of H_2 in the reaction that forms ammonia gas (NH_3), how much ammonia will be formed? What is the limiting reagent? How much of the reagent is left when the maximum amount of ammonia is formed? *(1265g NH_3; LR is H_2; 208.2g N_2 left)*

% Composition, Empirical and Molecular Formulas

18. A 0.941 gram piece of magnesium metal is heated and reacts with oxygen. The resulting oxide weighed 1.560 grams. Determine the percent composition of each element in the compound. *(60.3% Mg ;39.7% O)*
19. 42.0 g of a mixture of sodium bicarbonate that is 77% sodium bicarbonate is decomposed into sodium carbonate, water, and carbon dioxide. How many liters of carbon dioxide gas will be provided from the decomposition at 25°C and 722.0 mmHg? *(4.96L)*
20. How many moles of carbon atom are produced from the combustion of 82.0 mL of ethanol (C_2H_6O)? The density of ethanol is 0.793 g/mL. *(2.82 mol)*
21. 8.24g of calcium carbonate is decomposed at STP and carbon dioxide and calcium oxide are formed. What is the volume of carbon dioxide that will be produced from the decomposition? *(1.85L)*
22. A compound has the empirical formula of CH_3Br and a vapor density of 6.00 g/L, at 375 K and 0.983 atm. Using these data, determine the following: (Hint: Derive a new equation for density using your n=mass/MM and PV=nRT)
    a) The molar mass of the compound. *(187.9 g/mol)*
    b) The molecular formula of the compound. *($C_2H_6Br_2$)*
23. One carrot may contain 0.75 mg of vitamin A ($C_{20}H_{30}O$). How many moles of vitamin A is this? How many molecules of vitamin A? How many atoms of hydrogen? What mass of hydrogen is this? *($2.6*10^{-6}$ mol; $1.6*10^{18}$ mole'c; $4.73*10^{19}$ atoms H; $7.86*10^{-5}$ g H)*
24. If 200.mL of 0.200 M potassium phosphate is mixed with 300.mL of 0.250 M calcium chloride.
    a. Write a balance chemical equation for the reaction.
    b. Determine the number of moles of each reactant.
    c. Determine which reactant is the limiting reactant.
    d. Determine the mass of the precipitate formed from the reaction. *(6.20g ppt)*
    e. Determine the number of moles of all ions remaining in solution.
    f. Determine the concentration of all ions remaining in solution. *(0.240M K^+; 0.0300M Ca^{2+} ; 0.300M Cl^-)*

<u>Naming Chemical Compounds</u>

25. NaBr ____________________________
26. $Ca(C_2H_3O_2)_2$ ____________________________
27. P_2O_5 ____________________________
28. $Ti(SO_4)_2$ ____________________________
29. $FePO_4$ ____________________________
30. K_3N ____________________________
31. SO_2 ____________________________
32. CuOH ____________________________
33. $Zn(NO_2)_2$ ____________________________
34. V_2S_3 ____________________________

<u>Write the formulas for the following chemical compounds:</u>

35. silicon dioxide _____
36. nickel (III) sulfide _____
37. manganese (II) phosphate _____
38. silver acetate_____
39. diboron tetrabromide _____
40. magnesium sulfate _____
41. potassium carbonate _____
42. ammonium oxide _____
43. tin (IV) selenide _____
44. carbon tetrachloride _____

<u>Predicting Reaction Products:</u> Balance the equations and predict the products and states of matter (assume reactions are taking place at STP) for the following reactions. If no reaction occurs write 'NR'.

45. ____ *Na* + ____ *$FeBr_3$* →
46. ____ *NaOH* + ____ *H_2SO_4* →
47. ____ *$C_2H_4O_2$* + ____ *O_2* →
48. ____ *$PbSO_4$* + ____ *$AgNO_3$* →
49. ____ *HBr* + ____ *Fe* →
50. ____ *O_2* + ____ *$C_5H_{12}O_2$* →