Greetings AP Chemistry Class of 2018,
The summer assignment is enclosed in the following 51 page attachment. Please print out the information and read carefully. Basically memorize **Part A** and be prepared to take a quiz on this section the second day of school. Hence, spend a little time each day on this section and you will breeze through it in the fall. **Part B** is the application process that you will hand in to me. It ranges from pages 7-33. Some material will come back to you and some you will need to use the internet to learn from (YouTube, etc.). I will also be happy to answer any questions you may have throughout the summer, as I check email often. I will also be happy to meet with you all if you desire a help session at the Tiverton Library – just let me know and I can set it up and send out the information to the class via email. So please check your email often throughout the summer and please respond to my emails by saying such things as "Got the email". This way I know everyone knows what is going on. On page #35 is a list of supplies, please has them for day 1. On pages 35-51 is a beginning glossary that we will use throughout the course and add terms as we go on. It is imperative that chemists have a working knowledge of the vocabulary. Do not memorize these terms for day 1, we need these terms to develop understanding.

Page #2 - AP Intro Letter

Pages #3 –6 **Summer Assignment - Part A** . The information in these 4 pages must be committed to memory for a **QUIZ ON THE SECOND DAY OF SCHOOL.**

Pages #7- 33 **Summer Assignment - Part B** - A series of exercises asking you to intensely review your sophomore year of chemistry - **due by 2 PM on Friday August 18, 2017 in the Main Office. THIS IS 10% OF YOUR 1st Quarter Grade – 20% off for each day late.**

Page #34 - **List of School Supplies** - please have these on the first day of class

Pages # 35-51 –A 16 page **List of Vocabulary** and information to printout and place in your 3 ring-binder and bring to class the first day. You are not responsible for this material at this time, but it is a beginning glossary that you will refer to throughout the course.

Welcome to AP Chemistry,

I am very excited for an awesome year ahead! In order to ensure the best start for you next fall, I have prepared a **Summer Assignment** that reviews many basic chemistry concepts. These problems will help you build a foundation in chemistry and contribute to your success. This is a required assignment, and your first few assessments will be drawn from these topics.

The Advance Placement Chemistry experience covers a full year of freshman college-level chemistry, so it places heavy demands on the student, especially in terms of the time commitment required. In fact, the College Board suggests that students devote a ***minimum of five hours per week*** for individual study outside of the classroom. The ultimate objective, of course, is to prepare you to take the AP Chemistry test in May 2018, and in order to accomplish this, topics are covered very quickly. However, if anytime during the year you fall behind or are absent, please seek help immediately.

It is also important that you realize up front how your performance in this course will be measured. Although homework does not count as much as tests or quizzes, doing homework on a regular basis is vital to success on the assessments. Homework is practice. So PLEASE practice, practice, practice. All assessments will be graded using an AP rubric. The good news is that there is a vast amount of chemistry available on the Internet. I encourage you to seek these resources often. With ready access to these websites as well as the classroom experiences and a good work ethic, I am confident that you will have everything you need to learn chemistry at the AP level.

Finally, I recommend that you spread out the Summer Assignment, rather than trying to complete it in the final week of the summer! It takes time for a student to process, practice and subsequently learn chemistry at the level necessary for success in AP Chemistry. Taking a college level course in high school is difficult, and it **requires commitment, hard work and time**, but completion of a class like this is a **great investment** in your education. Prepare yourself and arrive ready to learn!

Please let me know via e-mail, lcusumano@tivertonschools.org if you have any questions or concerns. Please do not hesitate to contact me.

Have a great vacation!

Ms. Cusumano

Part A–Vital Information

Although this is a problem-solving course, memorization of some topics/rules is necessary. This information in Part A has to be second nature to you to ensure your success in this course. Master the memorization material listed below. Do whatever it takes to commit this information to memory. **YOU WILL HAVE A QUIZ ON ALL OF *PART A* MATERIAL THE SECOND DAY OF SCHOOL.**

Subject #1 – Key Elements and Symbols in the Periodic Table

- Know the names and symbols of element 1 to 38
- Also know the names and symbols of the following: Ag, Cd, I, Xe, Cs, Ba, W, Hg, Pb, Sn, Rn, Fr, U, Th, Pu, and Am as well as quickly locate these elements on the periodic table since the periodic table provided on the exam does not include element names

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Subject #2 – The Six Strong Acids

- HCl – Hydrochloric Acid
- HBr – Hydrobromic Acid
- H_2SO_4 – Sulfuric Acid
- $HClO_4$ -Perchloric Acid
- HI - Hydroiodic Acid
- HNO_3 - Nitric Acid

Subject #3 –Strong Bases

- Group 1 metal hydroxides (NaOH, KOH, etc.)
- Group 2 metal hydroxides - $Ba(OH)_2$, $Sr(OH)_2$,$Ca(OH)_2$ -these are only slightly soluble – others are insoluble
- Ammonia is a weak base (NH_3)

Table # 1- Monatomic Ions

Ions usually with one oxidation state			
Li^{+}	Lithium Ion	N^{3-}	Nitride
Na^{+}	Sodium Ion	P^{3-}	Phosphide
K^{+}	Potassium Ion	O^{2-}	Oxide
Mg^{2+}	Magnesium Ion	S^{2-}	Sulfide
Ca^{2+}	Calcium Ion	F^{-}	Fluoride
Sr^{2+}	Strontium Ion	Cl^{-}	Chloride
Ba^{2+}	Barium Ion	Br^{-}	Bromide
Al^{3+}	Aluminum Ion	I^{-}	Iodide

Cations with more than one oxidation state	
1^{+}	2^{+}
Cu^{1+} Copper (I)	Cu^{2+} Copper (II)
2^{+}	3+
Fe^{2+} Iron(II) Cr^{2+} Chromium(II) Co^{2+} Cobalt (II)	Fe^{3+} Iron(III) Cr^{3+} Chromium(III) Co^{3+} Cobalt (III)
2^{+}	4^{+}
Sn^{2+} Tin(II) Pb^{2+} Lead (II)	Sn^{4+} Tin(IV) Pb^{4+} Lead (IV)

Note that silver and zinc is always assumed to be the following:

Silver ion = Ag^{+} **Zinc Ion = Zn^{2+}**

Table #2 – The Nine Most Important Polyatomic Ions

Polyatomic Ion	Name
NH_4^{+}	Ammonium (Only Positively Charged Polyatomic Ion)
NO_3^{-1}	Nitrate
MnO_4^{-1}	Permanganate
CH_3COO^{-1}	Acetate
OH^{-1}	Hydroxide
CN^{-1}	Cyanide
CO_3^{2-}	Carbonate
SO_4^{2-}	Sulfate
PO_4^{3-}	Phosphate

Table #3 – Diatomic

Elements	Formula
Hydrogen	H_2
Nitrogen	N_2
Oxygen	O_2
Fluoride	F_2
Chloride	Cl_2
Bromide	Br_2
Iodide	I_2

Table #4 – Solubility Rules

These rules you will see inside this packet where you will get a chance to apply them. You need to know these cold. Remember that soluble means an aqueous compound is made (aq) and insoluble means that a precipitate is made (ppt).

- **Soluble with NO exceptions** – all Group 1A (Li^+, Na^+, etc.), ammonium ion (NH_4^+), nitrate (NO_3^{1-}), acetate (CH_3COO^{1-})

- **Soluble with exceptions**
  - 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 sulfates (SO_4^{2-}) are soluble except those of Pb^{2+}, Ca^{2+}, Sr^{2+}, and Ba^{2+}

- **Insoluble with exceptions**
  - 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

Table #5 – Colors of Common Ions In Aqueous Solution

Most common ions are colorless in solution, however, some have distinctive colors. These colors have appeared in questions on the AP Exam.	
Fe^{2+} and Fe^{3+}	Various Colors
Cu^{2+}	Blue to Green
Cr^{2+}	Blue
Cr^{3+}	Green or Violet
Mn^{2+}	Faint Pink
Ni^{2+}	Green
Co^{2+}	Pink
MnO_4^{1-}	Dark Purple
CrO_4^{2-}	Yellow
$Cr_2O_7^{2-}$	Orange

Table #6 – Molecular Geometry

Compound	Bonding Pairs	(Lone Pairs) Non-Bonding Pairs	Molecular Geometry	Hybridization	Angle
BeF_2	2	0	Linear	sp	180°
BF_3	3	0	Trigonal Planar	sp^2	120°
SO_2	2	1	Bent	sp^2	120°
CH_4	4	0	Tetrahedral	sp^3	109.5°
NH_3	3	1	Trigonal Pyramidal	sp^3	107°
H_2O	2	2	Bent	sp^3	105°
PF_3	5	0	Trigonal Bipyramidal	sp^3d	XXXXXX
SF_4	4	1	See-Saw	sp^3d	XXXXXX
ClF_3	3	2	T-shaped	sp^3d	XXXXXX
XeF_2	2	3	Linear	sp^3d	XXXXXX
SF_6	6	0	Octahedral	sp^3d^2	XXXXXX
BrF_3	5	1	Square Pyramidal	sp^3d^2	XXXXXX
XeF_4	4	2	Square Planar	sp^3d^2	XXXXXX

Name ______________________ Date ______________ CUSUMANO

Part B- Nomenclature, Balancing Equations, Oxidation Numbers, Solubility Rules, and Problem Solving

1. Nomenclature – Naming Compounds

- **Forming binary ionic compounds** - 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".

Example: What ionic compound would form when calcium ions combine with bromide ions?

Step One: Ca^{2+} and Br^{1-}

Step Two: Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion. So the absolute value of +2, (which is 2) of Ca becomes the subscript of Br and the absolute value of -1(which is 1) becomes the subscript of Ca.

Step Three: Check to make sure the subscripts are in the lowest whole number ratio possible. Then write the formula, $CaBr_2$

- **Naming binary ionic compounds (Ionic means that a metal and a nonmetal combine)**

  - Name $BaBr_2$
  - Combine the names of the cation (positive ion) and the anion (negative ion).
  - Ba is barium – the metal and makes the cation (Ba^{2+})
  - The metal's name is used first and is not changed in anyway
  - Br is bromine – the nonmetal and makes the anion (Br^{1-})
  - Notice that bromine is changed to brom**ide**
  - Hence $BaBr_2$ becomes barium bromide

  - Name $FeCl_2$
  - Notice that Fe is a transition element so it can produce more than one possible ion (Fe^{2+} or Fe^{3+})
  - Uncriss-cross $FeCl_2$ and you should see that iron is Fe^{2+} and chlorine is Cl^{1-}
  - Since the metal is a transition metal and has more than one choice, it must be distinguished by a Roman Numeral
  - So this compound is Iron (II) chloride

- **<u>Naming binary ionic compounds that contain polyatomic ions</u>**
    - The polyatomic ions on your common ions list (Tables 2, 3, and 4) should be memorized.
    - 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. Name the following:

Example – Na_2SO_4

- Uncriss-cross the compound
- Na^{1+} and SO_4^{2-}
- Sodium sulfate (the metal's name stays the same and the polyatomic ion is what you know from mastering Tables 2, 3,and 4)

- **<u>Naming binary molecular compounds</u>(Molecular compounds are usually formed between two nonmetals<u>)</u>**

    - Molecular Compounds use a prefix system.

Mono -1	Hexa – 6	Undeca - 11
Di – 2	Hepta – 7	Dodeca - 12
Tri -3	Octa - 8	
Tetra -4	Nona - 9	
Penta – 5	Deca - 10	

- The less electronegative element is always written first. It only gets a prefix if it has more than one atom in the molecule.
- The second element also gets the prefix and the ending –ide.
- The letter "O" or "A" at the end of the prefix is dropped when the word following the prefix begins with another vowel, for example monoxide or pentoxide.

Example – N_2O_4 = Dinitrogen tetroxide

Exercise 1 - Nomenclature: Simple Inorganic Formulas

I. In the first column, classify each of the following as molecular (M) or ionic (I). In the second column, name each compound:

	M or I	Name		M or I	Name
1) CaF_2			**10) SrI_2**		
2) P_4O_{10}			**11) CO**		
3) K_2S			**12) Cs_2O**		
4) NaH			**13) ZnI_2**		
5) Al_2Se_3			**14) P_2S_3**		
6) N_2O			**15) AgCl**		
7) O_2F			**16) Na_3N**		
8) SBr_6			**17) Mg_3P_2**		
9) Li_2Te			**18) XeF_6**		

II. In the first column, write the chemical formula for the compound formed. In the second column, write the compound's name:

	Elements	Chemical Formula	Compound Name
1	magnesium and iodine		
2	potassium and sulfur		
3	chlorine and aluminum		
4	zinc and bromine		
5	strontium and oxygen		
6	calcium and nitrogen		
7	oxygen and calcium		
8	copper(I) and oxygen		
9	copper (II) and chlorine		
10	mercury (II) and oxygen		
11	nitrogen and aluminum		
12	sulfur and cesium		

Exercise 2 - Oxidation Numbers: Anions and Cations
Summary of Rules for Oxidation Numbers:

<u>Rule 1</u>: Atoms in a pure element have an oxidation number of zero.

<u>Rule 2</u>: The more electronegative element in a binary compound is assigned the number equal to the negative charge it would have as an anion. The less electronegative atom is assigned the number equal to the positive charge it would have as a cation.

<u>Rule 3</u>: Fluorine has an oxidation number of -1 in all of its compounds because it is the most electronegative element.

<u>Rule 4</u>:Oxygen has an oxidation number of -2 in almost all compounds.
Exceptions:

- Peroxides, such as H_2O_2, in which its oxidation # is -1
- When oxygen is in compounds with halogens, such as OF_2, its oxidation # is +2.

<u>Rule 5</u>: Hydrogen has an oxidation # of +1 in all compounds that are more electronegative than it; it has an oxidation # of -1 in compounds with metals.

<u>Rule 6</u>: The algebraic sum of the oxidation numbers of all atoms in a neutral compound is zero.

<u>Rule 7</u>: The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.

<u>Rule 8</u>: Rules 1-7 apply to covalently bonded atoms; however, oxidation numbers can also be assigned to atoms in ionic compounds.

Determine the Oxidation Number of each underlined element in the table below

1) $K_2\underline{S}$	**6)** $\underline{S}_8$	**11)** $\underline{C}_{60}$
2) $Na\underline{Cl}O_4$	**7)** $\underline{Mg}$	**12)** $\underline{Zr}O_2$
3) $\underline{Br}Cl$	**8)** $K_2\underline{W_4}O_{13}$	**13)** $K_2\underline{Cr_2}O_7$
4) $Li_2\underline{C}O_3$	**9)** $Mg(\underline{B}F_4)_2$	**14)** $Al_2(\underline{Cr}O_4)_3$
5) $\underline{O}F_2$	**10)** $\underline{Au}_2O_3$	**15)** $Cs_2\underline{Te}F_8$

Exercise 3 – More Nomenclature Including Acids and Salts. If you are not provided with enough information, use the internet to search for information.

I. Name the following substances:

Formula	Name	Formula	Name
$FeSO_4$		Fe_2O_3	
$Cu(NO_3)_2$		$(NH_4)_2SO_3$	
HgCl		$Ca(MnO_4)_2$	
AgBr		PF_5	
$KClO_3$		LiH	
$MgCO_3$		HIO_3	
BaO		$Ca_3(PO_4)_2$	
K_2O		HIO_4	
SnO_2		$NaBrO_2$	
$Ni_3(PO_4)_2$		$Fe(IO_2)_3$	
$Pb(OH)_2$		HAt	
$CuCH_3COO$		H_3PO_4	
N_2O_4		NH_4BrO_3	
Rb_3P		S_8	

II. Write the formula for the following substances:

Name	Formula	Name	Formula
Vanadium (V) oxide		Francium dichromate	
Dihydrogen monoxide		Calcium carbide	
Ammonium Oxalate		Mercury (I) nitrate	
Polonium (VI) thiocyanate		Carbonic acid	
Tetraphosphorus decoxide		Calcium hypochlorite	
Zinc hydroxide		Copper (II) nitrite	
Potassium cyanide		Nitrous acid	
Cesium thiosulfate		Cyanic acid	
Oxygen molecule		Tin(IV) chromate	
Mercury (II) acetate		Manganese (VII) oxide	
Silver chromate		Sodium bicarbonate	
Tin (II) carbonate		Copper (II) dihydrogen phosphate	

Exercise 4 – Balancing Equations

Balance the following equations by adding coefficients as needed.

1) _____ C_6H_6 + _____ $O_2 \rightarrow$ _____ H_2O + _____ CO_2

2) _____ NaI + _____ $Pb(SO_4)_2 \rightarrow$ _____ PbI_4 + _____ Na_2SO_4

3) _____ NH_3 + ____ $O_2 \rightarrow$ _____ NO + _____ H_2O

4) _____ HNO_3 + _____ $Mg(OH)_2 \rightarrow$ _____ H_2O + _____ $Mg(NO_3)_2$

5) _____ H_3PO_4 + _____ NaBr $\rightarrow$ _____ HBr + _____ Na_3PO_4

6) _____ CaO + _____ $MnI_4 \rightarrow$ _____ MnO_2 + _____ CaI_2

7) _____ C_2H_2 + _____ $H_2 \rightarrow$ _____ C_2H_6

8) _____ OsO_4 + _____ $PtCl_4 \rightarrow$ _____ PtO_2 + _____ $OsCl_8$

9) _____ Hg_2I_2 + _____ $O_2 \rightarrow$ _____ Hg_2O + _____ I_2

Exercise 5 – Solubility Rules –See Table #4 in Part A to help you answer the following questions.

I. **For the compounds in the table, write the formula for each compound in the first column and then use the solubility rules on the previous page to determine if each compound is soluble or insoluble in water. In the second column write an (S) for those that are soluble and an (I) for those that are insoluble in water. Remember soluble means dissolves and is aqueous (aq) and insoluble forms a ppt and is given the symbol of (s).**

Name	Formula	(S) or (I)
Silver nitrate		
Cobalt (II) sulfate		
Zinc hydroxide		
Iron (III) iodide		
Nickel (II) chloride		
Lead (II) iodide		
Sodium carbonate		
Barium sulfate		
Lead (II) sulfide		
Silver phosphate		
Lithium phosphate		
Nickel (II) carbonate		
Copper (II) hydroxide		
Tin(IV) sulfate		
Lead (II) nitrate		

Exercise 6 – Reaction Prediction Practice

All chemical reactions can be placed into 5 categories.

1) **Combustion**: A combustion reaction is when oxygen combines with another compound to form water and carbon dioxide. These reactions are exothermic, meaning they produce heat. An example of this kind of reaction is the burning of naphthalene:

$$C_{10}H_8 + 12\ O_2 \text{ ---> } 10\ CO_2 + 4\ H_2O$$

2) **Synthesis**: A synthesis reaction is when two or more simple compounds combine to form a more complicated one. These reactions come in the general form of:

$$A + B \text{ ---> } AB$$

One example of a synthesis reaction is the combination of iron and sulfur to form iron (II) sulfide:

$$8\ Fe + S_8 \text{ ---> } 8\ FeS$$

3) **Decomposition**: A decomposition reaction is the opposite of a synthesis reaction - a complex molecule breaks down to make simpler ones. These reactions come in the general form:

$$AB \text{ ---> } A + B$$

One example of a decomposition reaction is the electrolysis of water to make oxygen and hydrogen gas:

$$2\ H_2O \text{ ---> } 2\ H_2 + O_2$$

4) **Single displacement**: This is when one element trades places with another element in a compound. These reactions come in the general form of:

$$A + BC \text{ ---> } AC + B$$

One example of a single displacement reaction is when magnesium replaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

$$Mg + 2\ H_2O \text{ ---> } Mg(OH)_2 + H_2$$

5) **Double displacement**: This is when the anions and cations of two different molecules switch places, forming two entirely different compounds. These reactions are in the general form:

AB + CD ---> AD + CB

One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

$$\mathbf{Pb(NO_3)_2 + 2\ KI \text{ ---> } PbI_2 + 2\ KNO_3}$$

In the following reactions below:

- **Write the products in words.**
- **Then below that write the equation with the correct formulas**
- **Then balance the equation.**

<u>COMBUSTION</u>

1. C_4H_9OH + oxygen →

2. C_7H_{14} + oxygen →

SYNTHESIS

1. Sodium + oxygen →

2. Calcium + nitrogen →

3. Potassium + bromine →

DECOMPOSITION

1. Strontium carbonate →

2. Mercury(II) oxide →

3. Aluminum carbonate→

DOUBLE REPLACEMENT (Note that the reactants in these problems are aqueous – so use the solubility rules table provided in Table #4 in Part A to determine which products are soluble and which are insoluble)

1. Iron (III) sulfate + calcium hydroxide →

2. Sodium hydroxide + sulfuric acid →

3. Sodium sulfide + manganese (VI) acetate →

4. Chromium (III) bromide + sodium sulfite →

5. Barium hydroxide + chlorous acid →

SINGLE REPLACEMENT

1. Chlorine gas + aluminum iodide →

2. Potassium metal + water→

3. Zinc + hydrochloric acid →

Exercise 7 – Significant Figures

(1) After solving the problems, express the answer with the correct significant figures AND in scientific notation.

a. $(5.03 \times 10^{-8})\ (3.05 \times 10^{7}) =$

b. $(5.5 \times 10^{7})\ (6.7435 \times 10^{3}) =$

c. $\dfrac{(4.871 \times 10^{4})\ (7.53 \times 10^{6})}{(7.4 \times 10^{-9})\ (2.982 \times 10^{7})} =$

(2) Identify the # of sig figs in the following

(a) 15.12 ________________

(b) 0.0155________________

(c) 5,677,000.________________

(d) 0.0005976________________

(e) 345,690,000,000________________

(f) 0.088809________________

(g) 505________________

(h) 0.005________________

(i) 500.0________________

(j) 0.050________________

(k) 50________________

(l) 0.00881________________

(m)0.05500________________

(n) 0.0224________________

(o) 66.22________________

(3) Solve in the correct number of significant figures

(a) (15.470) (370) =

(b) (510.0) (32) =

(c) (325) (5.44562) =

(4) Express the following in the correct number of significant figures:

(a) 37.28 + 14.5 =

(b) 8000 – 3370 =

(c) 450.044 + 660 =

(d)12,310 + 23.5 =

(e) 4129 + 200 =

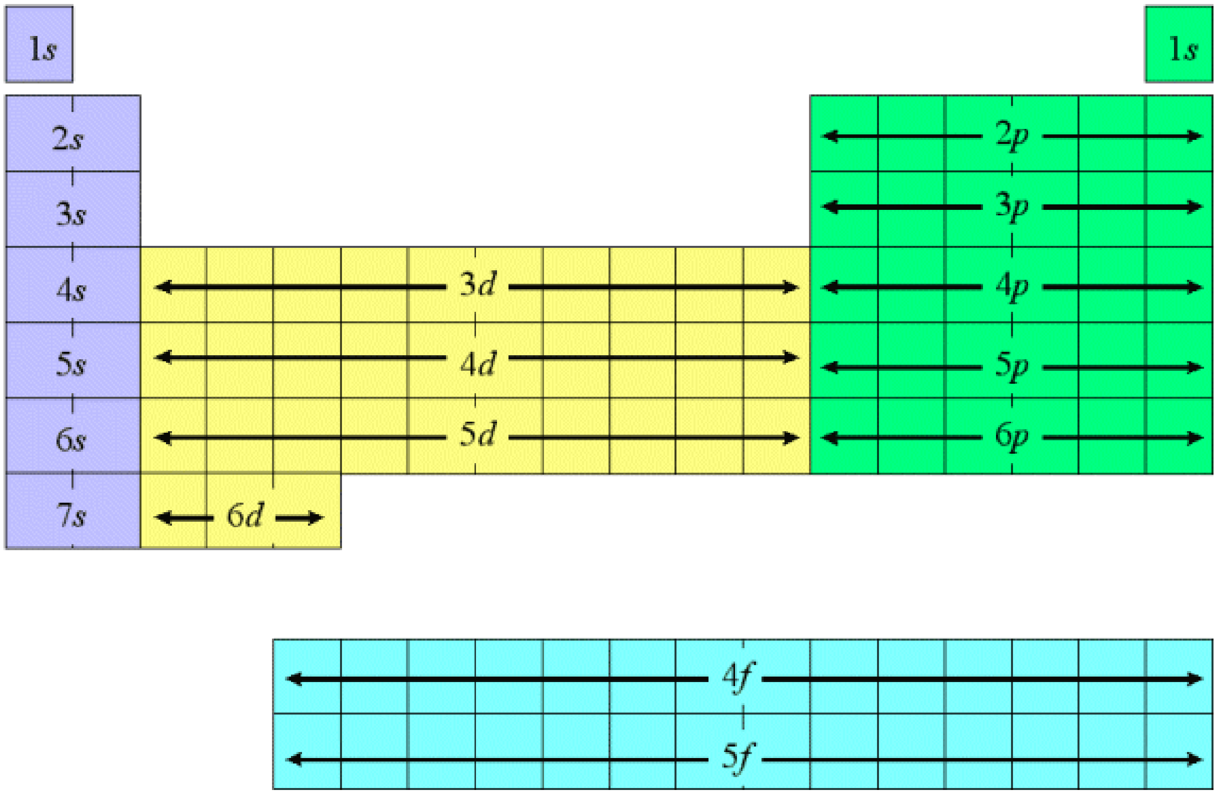
Exercise 8: Atomic Theory, Electron Configuration, Periodicity

Fill in the following table:

Element/Ion	# of protons	# of neutrons	# of electrons
Fe			
Na^+			
F			
	27		25
S^{2-}			
Cr^{+3}			

Exercise 9: Electron Configuration

Write the electron configuration for the following : Think of the periodic table in terms of s, p, d, and f orbitals and you will not go wrong



a) Li ____

b) Ca ____

c) Ca $^{2+}$ ____

d) F ____

e) F^{1-} ____

f) Al ____

Exercise 10 – Lewis Dot Diagrams and Valance Electrons

Fill in the table below:

	Li	Be	B	C	N	O	F	Ne
Valence Electrons								
Lewis Dot Diagram								

Exercise 11 – Periodic Trends

Place the following elements (S, Se, I, Ca, and Be) in order of

a) Increasing atomic radius

b) Decreasing ionization

c) Increasing electronegativity

d) Define ionization energy

e) Define electronegativity

Exercise 12 – Average Atomic Mass

Find the mass of an element and identify this element by symbol and name?

5% have a mass of 176

19% have a mass of 177

27% have a mass of 178

14% have a mass of 179

35% have a mass of 180

Exercise 13 – Bonding and Lewis Dot Structures

Use Table #6 from Part A and the table below to answer the following questions
WHEN ANSWERING THE QUESTIONS WATCH OUT FOR SOME RESONANCE STRUCTURES.

	Lewis Dot	Bonding Pairs	(Lone Pairs) Non-Bonding Pairs	Molecular Geometry	Hybridization	Angle
$BeCl_2$						
CO_2						
HCN						
BF_3						

	Lewis Dot	Bonding Pairs	(Lone Pairs) Non-Bonding Pairs	Molecular Geometry	Hybridizatio n	Angle
NO_3^{-1}						
SO_3						
CO_3^{2-}						
SO_2						

	Lewis Dot	Bonding Pairs	(Lone Pairs) Non-Bonding Pairs	Molecular Geometry	Hybridizatio n	Angle
O_3						
CH_4						
$SiCl_4$						
NH_4^+						

	Lewis Dot	Bonding Pairs	(Lone Pairs) Non-Bonding Pairs	Molecular Geometry	Hybridizatio n	Angle
NH_3						
PF_3						
H_2O						
OF_2						

Exercise 14 – Stoichiometry

1) Suppose a solution containing 4.50 grams of sodium phosphate is mixed with a solution containing 3.75 grams of barium nitrate.
   a. Write the balanced equation (Note this is a double replacement reaction).
   b. How many grams of barium phosphate can be produced?
   c. Which of the reactants is a limiting reagent?

2) Over the years, the mixing of solid iron (III) oxide with aluminum metal has been used for welding railroad rails. This is a single replacement reaction.
   a. Write the balanced equation representing this reaction.
   b. What masses of iron (III) oxide and aluminum must be used to produce 15.0 grams of iron?
   c. What is the maximum mass of aluminum oxide that can be produced?

Exercise 15 – Gas Laws

1. Convert the following to temperatures to Kelvin. (Kelvin = 273 + Celsius)
   a. 300 ˚C ___________

   b. 0 ˚C ___________

   c. -200 ˚C ___________

2. Pressure should be expressed in atm (atmospheres). Convert the following pressures to atm.   Note 1 atm =760 torr

   a. 380 torr ___________

   b. 1520 torr ___________

3. With what you learned about temperature and pressure above apply to the problem below and use the combined gas law ($\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$)

   A gas has a pressure of 345 torr at a temperature of -15 ˚C and a volume of 3.48 L. If conditions are changed so that the temperature is 36 ˚C and the pressure is 268 torr, what will be the volume of the sample?

4. Calculate the number of moles of an ideal gas if the gas is a 25 ˚C at a pressure of 750 torr at 2 liters. Use the ideal gas law to solve (PV = nRT). Remember that R = .0821).

5. Take the ideal gas law and manipulate the variables so that you can solve for molar mass and for density. Note that n = mass/molar mass

6. The mean molar mass of the atmosphere at the surface of Titan, Saturn's largest moon is 28.6 g/mol. Titan's surface temperature is 95 K and its pressure is 1.6 atm. Assuming ideal behavior, calculate the density of Titan's atmosphere under these conditions.

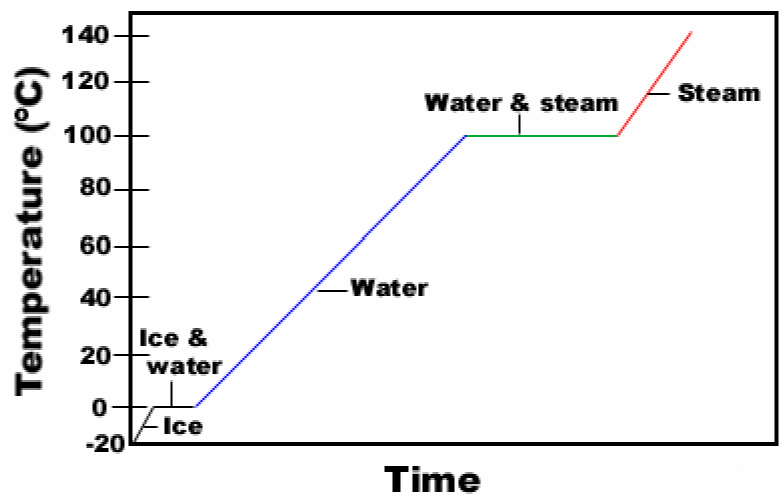
7. A 250 mL sample of oxygen is collected over water at 25°C and 760.0 torr pressure. Remember that $P_{gas} = P_{atmosphere} – P_{H2O}$
   (a) What is the pressure of the dry gas alone if the vapor pressure of water at 25°C = 23.8 torr).
   (b) How many moles of oxygen are collected?
   (c) How much Oxygen was collected?

8. Equal numbers of moles of He, Ar , and Ne are placed in a glass vessel at room temperature. If the vessel has a pinhole-sized leak, which of the following will be true regarding the relative values of the partial pressures of the gas remaining in the vessel after some of the gas mixture has effused?
   A) $P_{He} < P_{Ne} < P_{Ar}$
   B) $P_{He} < P_{Ar} < P_{Ne}$
   C) $P_{Ne} < P_{Ar} < P_{He}$
   D) $P_{Ar} < P_{He} < P_{Ne}$
   E) $P_{He} = P_{Ar} = P_{Ne}$

Exercise 16 – $Q=mc\Delta T$

Use the formulas, table, and heating curve below to solve the following questions.

Heating Curve



Important Formulas and Information

$Q = mc\Delta T$ Formula to solve for solid (ice in the heating curve above), liquid (water in the heating curve above), or gas (steam in the heating curve above)

$Q = m\Delta H_{vap}$ Formula to solve solids melting or liquids freezing

$Q = m\Delta H_{fusion}$ Formula to solve for liquids evaporating or gases condensing

Note :

m = mass in grams
C = specific heat in J/g°C
ΔT = Temperature (T_{final}- $T_{imitial}$)
Q= Heat in Joules
ΔH_{fusion} = use the table on the next page
$\Delta H_{vaporization}$ = use the table on the next page

Note:

- As one progresses on the heating curve from a solid to a gas, energy is absorbed and is expressed with a positive number
- As one progress on the heating curve from gas to solid, energy is released and is expressed with a negative number.

Some values for specific latent heats of fusion and vaporization:

Substance	latent heat of fusion J/g	°C Melting Temp.	latent heat of vaporization J/g	°C Boiling Temp.
Water	334	0	2260	100
Ethanol	109	-114	838	78
Ethanoic acid	192	17	395	118
Chlorofo	74	-64	254	62
Mercury	11	-39	294	357
Sulphur	54	115	1406	445
Hydrogen	60	-259	449	-253
Oxygen	14	-219	213	-183
Nitrogen	25	-210	199	-196

(1) Look at the table above – what are the units for:
  (a) Mass
  (b) Heat of Fusion
  (c) Heat of Vaporization

(2) Calculate the energy transferred when 4.6 grams of ice is melted at 0 ˚C. Is energy absorbed or released?

(3) Calculate the energy transferred when 9.8 g of water vapor condenses on a soda can at 100.0°C. Is energy absorbed or released?

(4) Mercury is a neurotoxin that when inhaled can be highly dangerous. Calculate the amount of energy required to change 14 g of liquid mercury into a gas at 357°C? See the values in the table. Is energy absorbed or released?

(5) Calculate the energy transferred in joules when 29. 5g of liquid water decreases from 14.0°C to 0.0°C. Is the energy absorbed or released?

Supplies for AP Chemistry:

1. **A sturdy three ring binder (3 inch)** to keep at home placing each unit's notes in **after** we have completed that unit in class. (Hence do not have to lug the entire curriculum to school each day)

2. **A sturdy 1 inch binder** to bring back and forth each day to keep the current units notes and assignments in

3. **Wide Ruled Loose-leaf paper** - easy to take notes on.

4. **Reinforcements** – the little circles that you place over the ripped holes on the loose-leaf paper.

5. **Tab Dividers** – you decide on how many to make your notebook work for you. Leave a section for vocabulary.

6. **Pencils with great erasers** - bring everyday

7. **Scientific Calculator** – bring everyday – you will lose credit if you do not bring every day. You need to be so comfortable with your calculator that its buttons and functions are second nature to you prior to taking the AP exam.

8. A bound **Graph Composition Book** (9.75 x 7.5 in) for your lab notebook. Please do not get the spiral bound graphing notebooks. There will be lots of labs in this course. Some will be require a full blown formal lab report and some will be activities to clarify material. We will fill the lab book with lots of hands on experiences.

9. A **valid student e-mail address** that is checked frequently. As I communicate frequently with students through email. When I send an email, please respond that you have received it. Communication is paramount.

Topic #1 – Scientists

- **Avogadro** – Determined the number of objects in one mole is 6.02×10^{23}
  - Determined that equal volumes of gases at the same temperature and pressure contain the same number of molecules.

- **Born** – Showed that the probability of the location of an electron can be determined.

- **Boyle** – Determined that pressure and volume of a gas vary inversely if amount and temperature is held constant. (Formula: $P_1V_1 = P_2V_2$)

- **Bronsted & Lowry –** Defined an acid as a proton (H^+) acceptor and a base as a proton donor.

- **Chadwick** – Discovered neutrons

- **Charles** –Determined that temperature and volume of a gas are directly related if amount and pressure is held constant. (Formula: $V_1/T_1 = V_2/T_2$)

- **Dalton** – Modern atomic theory. Determined $P_{total} = P_1 + P_2 + \ldots P_n$

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- **Dobereine**r –First attempt to classify elements by grouping elements with similar properties, the precursor to the periodic table.

- **Gay-Lussac**– Determined that temperature and pressure of a gas are directly related if amount and volume are held constant. (Formula: $P_1/T_1 = P_2/T_2$)

- **Graham** – Determined that the ratio of the rates of the diffusion of gases is equal to the square root of the inverse ratio of their molecular masses. $V_1/V_2 = \sqrt{m_2}/\sqrt{m_1}$

- **Heisenberg** – Uncertainty Principle – we can never know the exact position of an electron in an atom.

- **Hund** - Electrons will fill all empty orbitals before pairing up in one orbital.

- **Le Chatelier** – Determined that if stress is applied to a system at equilibrium that the system will shift so as to relieve the stress.

- **Lewis** – used dot diagrams to represent an atom and its outermost (valence) electrons. Defined an acid as an electron pair acceptor and a base as an electron pair donor.

- **Mendeleev** –Proposed that the properties of elements were a function of their atomic masses and thus formulated the periodic table.

- **Millikan** – Used oil drop experiment to determine charge on an electron.

- **Mosley** – Revised the periodic table. Properties of elements are based on atomic number rather than atomic mass.

- **Pauli** – Exclusion Principle – two electrons can never occupy the same space.

- **Planck** – Quantum Theory - energy is given off in packets called quanta.

- **Rutherford & Bohr** – Planetary model of atomic structure. Rutherford's gold foil experiment showed that the atom is made of mostly space and with dense nucleus.

- **Schrodinger** – Wave Equation – used to determine the position of an electron as a wave rather than a particle.

- **Thompson** – Discovered electrons using cathode ray tube. Determined charge/mass ratio.

Topic #2 - Matter & Properties

- **Matter** - A material that takes up space and has inertia
- **Mixture** -A material consisting of 2 or more substances that can be homogeneous or heterogeneous. Ex. Solutions, salt water, fruit salad, granite
- **Solution** - a homogeneous mixture (salt water, Kool-Aid, coffee)
- **Homogeneous** - A substance that is the same throughout. Could be mixture, compound, or element.
- **Heterogeneous** - A substance that is not the same throughout, so characteristics of different part of the substance vary. (Fruit salad)
- **Chemical property** - A property of a substance that deals with how the substance reacts with its surroundings (reacts with oxygen, flammable, decomposes spontaneously)
- **Physical property** - A property of a substance that does not deal with how it reacts. (ex. State of matter, color, density, boiling point, conductivity)
- **Chemical change** - A change that creates a new substance indicated by color change, change in mass, bubbles, precipitate, or energy changes.
- **Physical change** - A change in a substance that does not create a new substance. (cutting, melting, boiling, freezing, crushing, grinding)
- **Phase** - A distinct section of matter with uniform properties that is different from its surroundings (ice in water are two separate phases)
- **Precipitate** - A solid that forms out of mixing two solutions due to its insoluble nature. (Makes it cloudy.)
- **Compounds** - A substance composed or two or more atoms held together by chemical bonds
- **Elements** - A substance composed of one atom, and always has the same number of protons to determine the type of element. (see periodic table for the elements)

Topic #3 – Atomic Theory and Periodic Table

- **Atomic number** - The number of protons an element has that indicates what element it is.

- **Average atomic mass** - The average mass number of an element determined by considering the masses and relative proportions of the isotopes of that element.

- **Mass number** - The number of protons and neutrons in one atom that varies with isotopes as the neutrons change.

- **Anion** - The negative ion formed by a gain of 1 or more electrons

- **Cation** - A positive ion formed by a loss of 1 or more electrons

- **Cathode ray tube** - Tube used to determine the relative size and charge of an electron and proton "Crooke's tube"

- **Transition metal** - A metal found in the center of the periodic table that can have multiple oxidation numbers so it has a roman numeral with it when naming compounds.

- **Halogen** - An element found in group 17 with a -1 oxidation number.

- **Alkali metals** - An element found in Group 1 with +1 oxidation. They are very reactive.

- **Metal** – found to the left of the stair step line on the periodic table. Metals are ductile, malleable, lustrous, conductors and have + oxidation numbers

- **Nonmetal** - found to the right of the stair step line and do not have characteristics of metals. Often but NOT always gases.

- **Metalloid** - Elements having a common side with the stair step line excluding aluminum that have some characteristics of metals. Semiconductors are examples.

- **Ion** - An atom that has an unequal number of electrons and protons thus producing a charge. A positive ion (cation) is the result of too few electrons, while a negative ion (anion) is the result of too many electrons.

- **Atom** - The smallest piece of an element that retains the properties of an element. It is composed of protons, neutron, and electrons.

- **Isotope** - A form of an atom with a different number of neutrons thus having a different atomic mass than other atoms of this element.
  Example: Carbon –13 and Carbon –14.

- **Neutron** - A subatomic particle found in the nucleus with a mass number of 1 and no charge

- **Proton** - A subatomic particle found in the nucleus with a mass number of 1 and a positive charge

- **Electron** - A subatomic particle found outside the nucleus in a cloud with a mass number of 0 and a negative charge

- **Electron affinity** – The attraction an atom has for an electron

- **Electronegativity** – The attraction an atom has for a shared pair (bonding) of electrons

- **First ionization energy** - The energy required to remove the most loosely held electron from an atom.

- **Wave particle duality** – Waves have particle properties and particles have wave properties

- **Octet rule** – atoms are most stable with eight electrons in the outer energy level

- **Periodic table** – Arrangement of elements according to atomic number and electron configuration (first organized by Mendeleev)

- **Quantum numbers** – Numbers used to locate an electron in an atom.

- **Principal quantum number** – Designates the energy level and size of the electron cloud. Corresponds to periods on the periodic table.

- **Orbital quantum number** – Designates the energy sublevel and shape of the electron cloud (s,p,d, etc.).

- **Magnetic quantum number** – Designates the orientation of the electron cloud.

- **Spin quantum number** – Designates the direction of the electron. Allows us to distinguish between the two electrons in an orbital.

- **Atomic radius** – Distance from the nucleus to the outermost electron energy level.

- **<u>Sublevels</u>** –Differing electron orbitals within an energy level discernible by shape. S is a sphere. P has 2 lobes, and D has 4 lobes.

- **<u>Energy levels</u>** – Specific energy possessed by a group of electrons in an atom corresponding to distance from the nucleus.

- **<u>Electron cloud</u>** – Space that an electron has a high probability of occupying in an atom.

- **<u>+ charge</u>** – Charge of a proton (nucleus)

- **<u>- charge</u>** – charge of an electron (cloud surrounding nucleus)

- **<u>Orbitals</u>** – Space that can be occupied by electrons with the same energy level, sublevel, and orientation. Only two electrons can occupy an orbital.

- **<u>Noble Gases</u>** – Elements located in group 18 on the periodic table. They are Inert because they possess a full outer energy level of electrons.

- **<u>Diatomic molecules</u>** - The seven elements (gases) that when they are by themselves in elemental form are found in pairs. When combined with other elements or forming ions they need not be in pairs. (H, Cl, N, O, F, Br, I,)

Topic #4 - Units/Data/Lab

- **Quantitative** - Data concerning numerical measurement of a quantity
- **Qualitative** - Data concerning the general/subjective measurement of a quantity (color, “big”)
- **Mass** - The amount of matter (measured in grams using balance)
- **SI** - standard international unit for metrics (moles, kg, m, sec, K, ampere, candela) Liter is not included because it is a derived unit of m^3
- **Kelvin** - The standard international unit of temperature (amount of kinetic energy “heat”). 0K is absolute zero where motion is nonexistent and 273K = 0C freezing point of water
- **Accuracy** -How correct a measurement is
- **Precision** - How consistently the same a measurement is
- **Density** - mass divided by volume (g/ml)
- **Burette** - A calibrated glass tube that allows for careful release of liquids needed for titrations
- **Graduated cylinder** - calibrated container for measuring volume of a liquid
- **Erlenmeyer flask** - Flask with narrow neck and triangular shaped body
- **Florence flask** - Flask with narrow neck and round body
- **Percentage yield** - The percentage of actually product vs. expected product using calculations from stoichiometry. Actual product / theoretical product x 100
- **Volume** – Unit derived from length cubed.
- **ml = cm^3 = 1g** of water at 25 ^{0}C

Topic # 5 - Bonding & Intermolecular Forces

- **Single bond** – One shared pair of electrons between two atoms.
- **Double bond** – Two shared pair of electrons between two atoms.
- **Metallic bond** – Bond between two metal atoms held together by a "sea" of delocalized electrons.
- **VSEPR** – Valence Shell Electron Pair Repulsion theory for molecular shape. Electron pairs are as far apart as they can be on a single atom.
- **Tetrahedral** – Shape of an atom with 4 equally spaced bonds. Angles are 109.5^0 (CH4)
- **Bent** – Shape of an atom with 2 bonds and unshared pairs of electrons. (H2O)
- **Linear** – Shape of an atom with 180^o angles (straight line)
- **Polar** – Uneven charge distribution. A polar molecule has a more negative side and a more positive side. There are stronger intermolecular forces between polar molecules.
- **Nonpolar** – Even distribution of charge.
- **+ delta** –partial positive charge and **- delta**–partial negative charge
- **intermolecular forces**–Attractive forces between molecules.
- **dipole-dipole**–Intermolecular force between two polar molecules. The more positive side of one molecule is attracted to the more negative side of another molecule.
- **Ionic compound**-A compound formed by two ions held together by an attraction between the opposite charged ions. Bond due to transferring electrons causing the formation of ions. Formed when metal and nonmetal combine.
- **Molecular compound** - Compound formed by covalent bonds (sharing electrons.) Found when two nonmetals combine or in polyatomic ions.
- **Hydrogen bond** - A dipole/dipole attraction between molecules when hydrogen is part of the compounds. Water leads to hydrogen bonding. These are strong intermolecular attractions, but they are much weaker than covalent and ionic bonds.

- **Covalent bond** - Bond created when atoms share electrons found between two nonmetals.

- **Ionic bond** - Bond created when ions are attracted

- **Van der waals** - term for intermolecular forces (generally weak), including London dispersion forces, hydrogen bonding, and dipole -dipole attractions.

- **London dispersion** - An intermolecular force found in nonpolar molecules.

- **Lewis Structures**: drawing of covalent compound that represents bonds with dashes and unshared electrons with dots
  .
- **Lewis dot diagrams**: diagrams representing the valence electrons (outer level) of an element.

Topic # 6 – Energy& Thermochemistry

- **Energy** - The ability to do work. Many kinds, electrical, chemical, light, sound, potential, kinetic

- **Chemical energy** - energy that molecules contain by being bonded together

- **Law of conservation of mass** – energy (E=mc2) Einstein's law that says energy and matter are equivalent and can be converted from one to the other, while the total amount or energy+mass in the world is constant.

- **$E=mc^2$** Energy = mass x (speed of the light)

- **Exothermic** - A reaction that releases energy to its surrounding environment (burning a match)

- **Endothermic** -a reaction that absorbs energy from its surrounding environment

- **Activation energy** - the energy required for a reaction to get started. It is decreased by the use of catalysts.

- **Calorie** = unit for heat, nutritional value = 1000 (chemistry) calories

- **Calorie**= unit for heat

- **Joule**= J, SI unit for heat

- **Entropy**: S, measure of disorder, gases have the most S, solids the least

- **Enthalpy**: H, heat content, ΔH is negative for an exothermic reaction.

- **Rate law**: represents the effect of concentration on the speed of a reaction.

Topic # 7 – Solutions

- **Solute** - The dissolved material (often solid) salt in salt water, gas in soda
- **Solvent** - The dissolving material (often water) water in salt water, water in soda
- **Molarity** - A unit for concentration that is moles of solute per liter of solution. This can be used to determine molarity of gases or aqueous solutions.
- **Vapor pressure lowering (Raoult's law)**: P soln= X solv Psolv
  Vapor pressure is lowered by the addition of a solute.
- **Miscible**: two substances that will dissolve in each other
- **Alloy**: a solution made of two metals
- **Solubility**: measure of the amount of solute that can be dissolved in a specific amount of solvent (usually water).

Topic #8 – Stoichiometry

- **Stoichiometry**: A type of problem where we use mass and volume values to determine the amount of each reactant that will react and the amount of products that will form. It utilizes the mole/mole ratio determined by a balanced equation. 22.4L / mole of gas at STP can also be used.
- **Stoichiometry**: Relationship between the amounts of reactants and products in a chemical reaction.
- **Mole to mole ratio**: conversion from substance A to substance B: found by looking at the coefficients in the balanced equation.

Topic #9 – Mole

- **Formula mass** - The total mass of a formula measured in amu. For 1 atom/molecule and g for 1 mole.

- **Molecular formula** - Formula for the actual compound that exists.

- **Empirical formula** - Formula that indicates the ratio of atoms of each element present in a substance, but doesn't give the actual molecule that exists. (reduced form)

- **Formula** – A group of symbols to indicate the number and kind of atoms found in a compound.

- **Avogadro's number** - 6.02×10^{23}. Avogadro's number of something is 1 mole. 1 mole of atoms is equal to the formula mass in grams of that atom.

- **Mole** - the SI unit for measurement of the quantity of a substance = to Avogadro's number of atoms/molecules of that substance

- **Percent composition** - The percentage of each element found in a compound as calculated by comparing masses of the elements

- **Hydrate** - A molecule that has water molecules attached to it and is indicated by a dot between the number of water molecules and the actual substance

Topic #10 – Reactions &Compounds

- **Equation** - a representation of the number of moles, reacting substances, products, and the state that each substance is found.

- **Chemical reaction** - A reaction where a chemical change has taken place that can be represented with an equation.

- **Balanced equation** - A chemical equation that indicates the appropriate number of moles of each substance that will react and be formed to not break the law of conservation of mass - energy.

- **Reactants** -The substances that are present before a reaction occurs.

- **Products** - The substance(s) that are present after a reaction occurs.

- **Single displacement reactions** - A reaction where an element and an ionic compound trade “partners” to form a new ionic compound while the other element of the original compound is now by itself.

- **Double displacement reactions**- A chemical reaction where two compounds trade “partners” to form two new compounds.

- **Synthesis reaction** - A chemical reaction where two elements are combined to form a compound.

- **Decomposition reactions**- A chemical reaction where one compound is broken down into two or more elements.

- **Combustion reactions**-A chemical reaction where a hydrocarbon is combined with oxygen to form carbon dioxide and water.

- **Coefficients**-A number before a substance in an equation to indicate the number of moles of that substance that will react or be formed.

- **Subscripts**-Small number lowered after an atom within a compound to indicate how many atoms of that element are found in a compound. This cannot be changed to balance equations, because it determines what kind of substance is present.

- **Limiting reagent**-The reactant that is completely used up during a reaction if it goes to completion.

- **Enzymes** - catalysts found in the body
- **Catalysts** - chemicals that speed up a reaction, without being permanently changed in the reaction.
- **Reactants** -substances on the left side of a chemical equation, starting materials
- **Products**- substances on the right of a chemical equation, end materials
- **Reversible reaction**- a reaction that can proceed in both directions (forward and reverse), represented with double arrows,
- **Equilibrium** - a dynamic condition, where opposing reactions occur at equal rates.
- **Reaction rate**- the speed of a chemical reaction, dependent upon temperature, nature of the reactants, concentration and surface area.
- **Precipitate** -a solid formed from the mixing of two aqueous solutions
- **Net ionic equation** -reaction where all soluble compounds are broken down into ions and only actual participants are written
- **Spectator ions**- ions that do not participate in the chemical reaction.
- **Binary compound** - A compound consisting of two elements such as Calcium chloride
- **Polyatomic ion** - A group of atoms covalently bonded that have a + or –charge which will lead them to forming ionic bonds

Topic #11 – Redox

- **Anode** - The positive end of a cathode tube that attracts the anions
- **Cathode** - The negative end of a cathode ray tube that attracts the cations
- **Oxidation numbers** - A number that is assigned to an atom to indicate whether that atom wants to gain or lose electrons. A + (plus) means it wants to lose electrons, a – (negative) means it wants to gain electrons, and a 0 means it is stable.
- **Redox reaction** - reaction where electrons are transferred.
- **Reduction** - gaining electrons, oxidation number becomes more negative
- **Oxidation** - losing electrons, oxidation number becomes more positive

Topic #12 – Acids & Bases

- **pH**= -log[H_3O^+]= -log[H_+], measures acidity of a solution.
- **pH scale** from 0-14.
- **pOH**= -log[OH-]
- **Hydronium**= H_3O^+or H^+, characteristic of acids
- **Hydroxide**= OH-, characteristic of bases
- **Acetic Acid**:$HC_2H_3O_2$
- **Nitric acid**:HNO_3
- **Hydrochloric acid**: HCl
- **Sulfuric acid**:H_2SO_4
- **Acid**: below7 on pH scale, [H+] > [OH-]
- **Base**: above 7 on pH scale, [OH-] > [H+]
- **Neutral**: pH of 7, [H+] = [OH-]
- I**ndicator**: weak acid or base whose conjugate is a different color; changes colors in acids/bases. Examples: litmus, phenolphthalein
- **Titration**: addition of a solution of known concentration to a solution of unknown concentration to discover the unknown's concentration.
- **Endpoint:** end of titration, visual determination where there are equivalent concentrations of acid and base.
- **Equivalence pt.** mathematical determination where there are equivalent concentrations of acid and base.
- **Standard solution**: a solution with a known concentration
- **Concentration**: measurement of amount of solute relative to solvent or solution.
- **Dilute**: a "weak" solution, solution of low concentration, with large amounts of solvent.

- **Arrhenius Acids and Bases**: donate H^+(acid) or OH^-(base)
- **Bronsted-Lowry acid**: proton donor
- **Bronsted-Lowry base**: proton acceptor
- **Conjugate acid:** formed when a base gains a proton
- **Conjugate base**: formed when an acid donates a proton
- **Monoprotic**: acid containing one H
- **Diprotic:** acid containing 2 H's
- **Triprotic:** acid containing 3 H's
- **Polyprotic:** acid containing more than 1 ionizable H^+
- **Lewis acid:** electron acceptor
- **Lewis base**: electron donor
- **Ammonia**: NH_3 (weak base)