AP Chemistry Summer Assignment

Welcome to AP Chemistry. I am eagerly anticipating a great year of Chemistry. In order to ensure the best start for everyone next fall, I have prepared a summer assignment that reviews basic chemistry concepts.

There are a multitude of tremendous chemistry resources available via the internet. With ready access to hundreds of websites either in your home or at the local library, I am confident that you will have sufficient resources to prepare adequately.

Much of the material in this summer packet will be familiar to you. It will be important for everyone to come to class the first day prepared. While I will review, extensive remediation is not an option as we work towards our goal of being 100% prepared for the AP Exam in early May. There will be a test covering the basic concepts included in the summer packet during the first week of school.

You may contact me by email: astaub@upperdarbysd.org this summer. I will do my best to answer your questions asap.

This assignment is designed to be <u>six weekly assignments</u>. **<u>Please do not try to complete it all in the final week of the summer.</u>** Chemistry takes time to process and grasp at a level necessary for success in AP Chemistry. Remember, AP Chemistry is an equivalent course to 2 semesters of introductory Chemistry in college. Taking a college level course in high school is difficult, requires dedication, and is a great investment in your education. So prepare yourself and arrive ready to learn!

Please join our summer google classroom. Go to https://classroom.google.com and use the class code ***dnd6k7m***

Have a great summer!

Mr. Staub

<u>SUMMER ASSIGNMENT #1</u>

Significant Figures – Rules

Single Number

Significant figures are critical when reporting scientific data because they give the reader an idea of how well you could actually measure/report your data. Before looking at a few examples, let's summarize the rules for significant figures.

1) ALL non-zero numbers (1,2,3,4,5,6,7,8,9) are ALWAYS significant.
2) ALL zeroes between non-zero numbers are ALWAYS significant.
3) ALL zeroes which are SIMULTANEOUSLY to the right of the decimal point AND at the end of the number are ALWAYS significant.
4) ALL zeroes which are to the left of a written decimal point and are in a number >= 10 are ALWAYS significant.

A helpful way to check rules 3 and 4 is to write the number in scientific notation. If you can/must get rid of the zeroes, then they are NOT significant.

Examples: How many significant figures are present in the following numbers?

Number	# Significant Figures	Rule(s)
48,923	5	1
3.967	4	1
900.06	5	1,2,4
0.0004 (= 4 E-4)	1	1,4
8.1000	5	1,3
501.040	6	1,2,3,4
3,000,000 (= 3 E+6)	1	1
10.0 (= 1.00 E+1)	3	1,3,4

This gives you some idea of how to determine the number of significant figures in a single number.

S C I E N T I F I C   N O T A T I O N

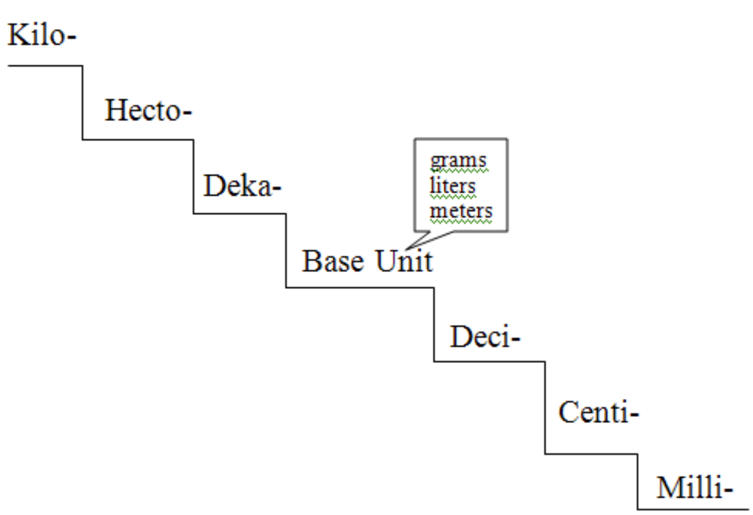
Change the following to Scientific Notation (maintain the number of significant figures):

1.	5.280 =	____________	11.	2,560 =	____________
2.	2,000 =	____________	12.	.0009 =	____________
3.	15 =	____________	13.	8,900,000 =	____________
4.	6,589,000 =	____________	14.	.0920 =	____________
5.	70,400,000,000 =	____________	15.	6,300 =	____________
6.	.00263 =	____________	16.	.90 =	____________
7.	.00589 =	____________	17.	250 =	____________
8.	.006 =	____________	18.	.006087 =	____________
9.	.400 =	____________	19.	500,000 =	____________
10.	.08060 =	____________	20.	.0000000105 =	____________

Metric Unit Conversions

<u>RULES</u>

- For 1 dimentional units each step on the staircase below represents one decimal place.
    - For example convert 3.5 mm to cm.
        - mm and cm are 1 dimentional units. Whereas mm^2 is a two dimentional unit.
        - Start at mm, the bottom of the staircase. cm are one step to the left. Move the decimal point one place to the left.
        - 3.5 becomes 0.35. Therefore 3.5 mm=0.35 cm
        - If no decimal point is present, assume there is a decimal point after the number. For example 3 cm = 30 mm
- For 2 dimentional units each step on the staircase below represents two decimal places.
    - For example convert 3.5 mm^2 to cm^2.
        - mm^2 and cm^2 are 2 dimentional units.
        - Start at mm, the bottom of the staircase. cm are one step to the left. Move the decimal point two place to the left.
        - 3.5 becomes 0.035. Therefore 3.5 mm^2=0.035 cm^2
        - If no decimal point is present, assume there is a decimal point after the number. For example 3 cm^2 = 300 mm^2
- For 3 dimentional units each step on the staircase below represents three decimal places.
    - For example convert 3.5 mm^3 to cm^3.
        - mm^3 and cm^3 are 3 dimentional units.
        - Start at mm, the bottom of the staircase. cm are one step to the left. Move the decimal point three place to the left.
        - 3.5 becomes 0.0035. Therefore 3.5 mm^3=0.0035 cm^3
        - If no decimal point is present, assume there is a decimal point after the number. For example 3 cm^3 = 3000 mm^3



1mm = 0.1 cm
1mm^2 = 0.01 cm^2
1 mm^3 = 0.001 cm^3

1cm = 10 mm
1 cm^2 = 100 mm^2
1 mm^3 = 1000 mm^3

1 km^3 = 1 x 10^{18} mm^3
1 mm^3 = 1 x 10^{-17} km^3

U N I T   A N A L Y S I S

Make the following Metric System conversions using “unit analysis” (you may use scientific notation):

		Standard Notation	Scientific Notation
1.	100 mg	____________g =	____________ g
2.	20 cm	____________ m =	____________ m
3.	50 L	____________ dm^3 =	____________ dm^3
4.	22 g	____________ cg =	____________ cg
5.	825 cm	____________ km =	____________ km
6.	2,350 kg	____________ g =	____________ g
7.	19 mL	____________ cm^3 =	____________ cm^3
8.	52 km	____________ m =	____________ m
9.	36 m^2	____________ cm^2 =	____________ cm^2
10.	18 cm^3	____________ mm^3 =	____________ mm^3
11.	6 g	____________ mg =	____________ mg
12.	4,259 mL	____________ dm^3 =	____________ dm^3

50L x 1000ml/1L = 50000cm3

Significant Figures – Rules

Math Operations

ADDITION AND SUBTRACTION:
When adding or subtracting numbers, count the NUMBER OF DECIMAL PLACES to determine the number of significant figures. The answer cannot CONTAIN MORE PLACES AFTER THE DECIMAL POINT THAN THE SMALLEST NUMBER OF DECIMAL PLACES in the numbers being added or subtracted.
Example:

23.112233	(6 places after the decimal point)
1.3324	(4 places after the decimal point)
+ 0.25	(2 places after the decimal point)
24.694633	(on calculator)
24.69	(rounded to 2 places in the answer)

Note: There are 4 significant figures in the answer.

MULTIPLICATION AND DIVISION:
When multiplying or dividing numbers, count the NUMBER OF SIGNIFICANT FIGURES. The answer cannot CONTAIN MORE SIGNIFICANT FIGURES THAN THE NUMBER BEING MULTIPLIED OR DIVIDED with the LEAST NUMBER OF SIGNIFICANT FIGURES.
Example:

23.123123	(8 significant figures)
x 1.3344	(5 significant figures)
30.855495	(on calculator)
30.855	(rounded to 5 significant figures)

Determining the number of significant figures in logarithms & antilogarithms:

- The logarithm (base 10) of x, log x = a, where x = 10a.
- The antilogarithm (base 10) of a, antilog a = x, where x = 10a.
- A logarithm is divided into two (2) parts by the decimal. The integer before the decimal is the characteristic and the numbers after the decimal are the mantissa.
- If a number is a logarithm, since the characteristic reflects the power of 10, i.e. the exponent, it is not considered to be part of the significant figures. Only the digits in the mantissa (after the decimal) are significant.

Antilog	Antilog Sig Figs	Log	Log Sig Figs
567 (5.67×10^{2})	3	$2.\underline{754}$ *The 2 is just a placeholder. The underlined digits are significant.*	3
0.0025 (2.5×10^{-3})	2	-2.60	2
205.203 (2.05203×10^{2})	6	2.312183	6
3.400×10^{20}	4	20.5315	4
0.0000002 (2×10^{-7})	1	-6.7	1

Significant Figures Calculations Practice

1. **Add-**

a) 16.5 + 8 + 4.37

b) 13.25 + 10.00 + 9.6

c) 2.36 + 3.38 + 0.355 + 1.06

d) 0.0853 + 0.0547 + 0.0370 + 0.00387

e) 25.37 + 6.850 + 15.07 + 8.056

2. **Subtract-**

a) 23.27 – 12.058

b) 13.57 – 6.3

c) 350.0 – 200

d) 27.68 – 14.369

3. **Multiply-**

a) 2.6 x 3.78

b) 6.54 x .037

c) 3.15 x 2.5 x 4.00

d) 0.085 x 0.050 x 0.655

e) 3.08 x 5.2

f) 0.0036 x 0.02

g) 4.35 x 2.74 x 3.008

h) 35.7 x 0.78 x 2.3

4. **Divide-**

a) 35 / 0.62

b) 39 / 24.2

c) 0.58 / 2.1

d) 40.8 / 5.05

e) 3.76 / 1.62

f) 0.075 / 0.030

5. **Logs-**

	answer	w/sig figs		answer	w/sig figs
a) log 5.89×10–3	4.701	3	b) ln 3.591	________	________
c) $10^{-2.22}$	________	________	d) $e^{5.61}$	________	________

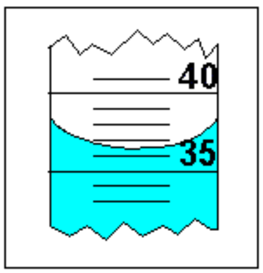
Uncertainty in Measurement

A. When taking **measurements** all certain digits plus the **first** uncertain number are significant.

Example: Your bathroom scale weighs in 10 Newton increments and when you step onto it, the pointer stops between 550 and 560. Your look at the scale and determine your weight to 557 N. You are certain of the first two places, 55, but not the last place 7. The last place is a guess and if it is your best guess it also is significant.

IN THE LABORATORY:
Masses should always be recorded to as many places after the decimal point as are read off the balance. Calculation of mass by difference using a tare should be reported to this same number of places.

Graduated cylinders should be read one decimal place past the smallest graduation.



EX. The Graduated cylinder to the left has a major scale of 5 mL. The minor scale is 1 mL. These values are determined by looking at the graduations (markings) and determining the valueof each line. Therefore this graduated cylinder should be red to the nearest 0.1 mL. One decimal place past the minor scal. The last decimal place should be estimated. I would read this cylinders volume as 36.4 mL. This reading has anuncertainty of +/- 0.1 mL. Thus the reading is really saying that the answer is between 36.3 mL and 36.5 mL.

U N C E R T A I N T I E S I N C A L C U L A T I O N S

1. When adding or subtracting numbers written with the ± notation, **always add the ± uncertainties** and then <u>round off the ± value to the largest significant digit. Round off the answer to match</u>.

   Example: (22.**<u>4</u>** ± .5) + (14.7**<u>6</u>** ± .25) = 37.**<u>16</u>** ±.75 = 37.**<u>2</u>** ± .8
   The uncertainty begins in the **<u>tenths</u>** place… it is the last significant digit.

2. When adding or subtracting numbers written in significant figures, show the uncertainty by rounding the <u>answer to match the largest place with uncertainty</u>.

   Example: 26**<u>7</u>** + 11.**<u>8</u>** = 27**<u>8.8</u>** = 27**<u>9</u>**
   The least accurate original measurement is only accurate to the **<u>ones</u>** place.

3. When multiplying or dividing measurements written in significant figures, show the uncertainty of your calculations by <u>rounding off your answer to **match the same number of significant figures as your *least* precise measurement (the measurement with the least number of significant figures).**</u>

   Example: 477.85 ÷ 32.6 = 14.657975 = <u>14.7</u>
   32.6 is the least accurate measurement with only 3 significant figures.

NOTE: There are two types of precision: "absolute precision" and "relative precision."

Example: 322.45 x 12.75 x 3.92 = 16116.051 = **161**00

All the measurements are accurate to the hundredth place (absolute precision) but the answer is rounded to 3 significant figures because 3.92 has only 3 significant figures (relative precision).

In Summary:

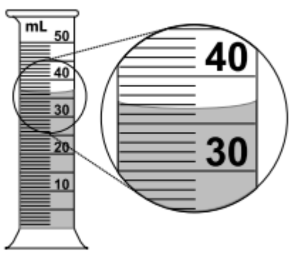
	Adding and Subtracting	**Multiplying and dividing**
#'s with ± notation	Rule 1	Don't Do This Case
#'s with significant figures	Rule 2	Rule 4

Uncertainties in Measurement

<u>Directions:</u> *For each measurement please record the major scale, minor scale and uncertainty. Than report your answer to the appropriate value with units. If the question has any extra calculations, please do them to the proper significant figures. If there are any extra questions, please answer them completely and in complete sentences.*

<u>Volume of a Liquid</u>

1.



a) What is the major scale? ________

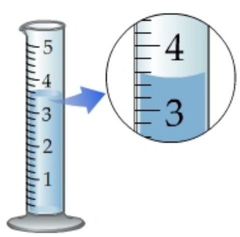
b) What is the minor scale? ________

c) What is the uncertainty of this device? ________

d) What is the volume of liquid? ________

e) Based on the uncertainty, what range can we assume the “actual volume” lies between? ________

2.



a) What is the major scale? ________

b) What is the minor scale? 0.2 mL

c) What is the uncertainty of this device? 0.02ml

d) What is the volume of liquid? 3.50ml

e) Based on the uncertainty, what range can we assume the “actual volume” lies between? 3.48ml - 3.52 ml

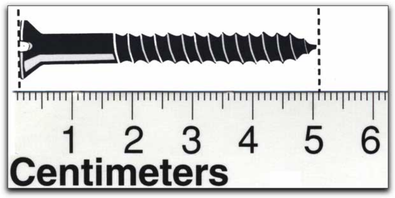
3. Add the volumes of the two graduated cylinders from above and record the value with proper significant figures?

4. What can you say about the precision of graduated cylinder 2? Explain.

5. How should our knowledge of uncertainty effect out selection of tools in the lab?

Measuring length

6.



a) What is the major scale? _________

b) What is the minor scale? _________

c) What is the uncertainty of this device? ________

d) What is the length of the screw? _________

e) Based on the uncertainty, what range can we assume the "actual length" lies between? ________

7. If a cube with a mass of 21.56g and sides measuring 2.04 cm in length, what is its density? Show all work.

8. What substance is the cube in number 7 made out of?

Material	Density (g/cm^3)
Aluminum	2.64
Brass	8.55
Brick (red, common)	1.92
Coal (anthracite)	1.51

9. Calculate the percent error. Is this an acceptable range of error? Explain.

SUMMER ASSIGNMENT #2

The first 4 pages are rules and guidelines. I would print these out and put them in page protectors and keep them in my chemistry binder. They will be helpful for the entire course.

Inorganic Chemical Nomenclature

***Binary* Compound Names for Non-Metal Ions**

H^-	hydride	N^{3-}	nitride	O^{2-}	oxide	F^-	fluoride
H^+	hydrogen	P^{3-}	phosphide	S^{2-}	sulfide	Cl^-	chloride
		As^{3-}	arsenide	Se^{2-}	selenide	Br^-	bromide
				Te^{2-}	telluride	I^-	iodide

Polyatomic Ion Names

NH_3	ammonia	ClO_4^-	perchlorate	PO_4^{3-}	phosphate
NH_4^+	ammonium	ClO_3^-	chlorate	HPO_4^{2-}	monohydrogen
H_3O^+	hydronium	ClO_2^-	chlorite		phosphate
CH_3COO^-	acetate	ClO^-	hypochlorite	$H_2PO_4^-$	dihydrogen
($C_2H_3O_2^-$)		CrO_4^{2-}	chromate		phosphate
AsO_4^{3-}	arsenate	$Cr_2O_7^{2-}$	dichromate	PO_3^{3-}	phosphite
BO_3^{3-}	borate	CN^-	cyanide	SeO_4^{2-}	selenate
$B_4O_7^{2-}$	tetraborate	OH^-	hydroxide	SiO_3^{2-}	silicate
BrO_3^-	bromate	IO_4^-	periodate	SiF_6^{2-}	hexafluorosilicate
BrO^-	hypobromite	IO_3^-	iodate	SO_4^{2-}	sulfate
CO_3^{2-}	carbonate	IO^-	hypoiodite	HSO_4^-	hydrogen sulfate
HCO_3^-	hydrogen carbonate	MnO_4^-	permanganate		(bisulfate)
	(bicarbonate)	NO_3^-	nitrate	SO_3^{2-}	sulfite
		NO_2^-	nitrite	HSO_3^-	hydrogen sulfite
		$C_2O_4^{2-}$	oxalate		(bisulfite)
		O_2^{2-}	peroxide	$C_4H_4O_6^{2-}$	tartrate
				$S_2O_3^{2-}$	thiosulfate

Common Acid Names

$HC_2H_3O_2$	acetic acid	HNO_3	nitric acid
CH_3COOH	acetic acid	H_3PO_4	phosphoric acid
H_2CO_3	carbonic acid	H_2SO_4	sulfuric acid
HCl	hydrochloric acid		

ION FORMULA CHART

1+	**2+**	**3+**
ammonium, NH_4	barium, Ba	aluminum, Al
cesium, Cs	beryllium, Be	chromium (III), Cr
copper(I) Cu	cadmium, Cd	cobalt(III), Co
gold(I) Au	calcium, Ca	gallium, Ga
hydrogen, H	chromium, Cr	gold(III) Au
lithium, Li	cobalt(II), Co	iron(III), Fe
potassium, K	copper(II), Cu	
rubidium, Rb	iron(II), Fe	
silver, Ag	lead(II), Pb	
sodium, Na	magnesium, Mg	
	mercury(I), Hg_2	**4+**
	mercury(II) Hg	lead(IV), Pb
	nickel, Ni	tin(IV), Sn
	strontium, Sr	
	tin(II), Sn	
	Zinc, Zn	

1-	**2-**	**3-**
acetate, $C_2H_3O_2$	carbonate, CO_3	borate, BO_3
bromate, BrO_3	chromate, CrO_4	nitride, N
bromide, Br	dichromate, Cr_2O_7	phosphate, PO_4
chlorate, ClO_3	hydrogen phosphate, HPO_4	phosphide, P
chloride, Cl	oxide, O	
chlorite, ClO_2	oxalate, C_2O_4	
cyanide, CN	peroxide, O_2	
Fluoride, F	selenide, Se	
dihydrogen phoshate, H_2PO_4	sulfate, SO_4	
bicarbonate, HCO_3	sulfide, S	
hydrogen sulfate, HSO_4	sulfite, SO_3	**4-**
hydroxide, OH	tartrate, $C_4H_4O_6$	carbon, C
iodate, IO_3	telluride, Te	silicon, Si
iodide, I	thiosulfate, S_2O_3	
nitrate, NO_3		
nitrite, NO_2		
permanganate, MnO_4		

COVALENT PREFIXES

1- mono	2 – di	3 – tri	4 – tetra	5 – penta
6 – hexa	7 – hepta	8 – oct	9 – nona	10 - deca

HYDROCARBONS

Prefixes 1 – meth 2 – eth 3 – prop 4 – but

Suffixes - ANE – C_nH_{2n+2} - ENE – C_nH_{2n} - YNE = C_nH_{2n-2}

STUFF I SHOULD KNOW FOR THE AP TEST BUT DO NOT KNOW YET

IONS LIST

acetate	$C_2H_3O_2^-$	ferric	Fe^{3+} (yellow)	oxalate	$C_2O_4^{2-}$
aluminum	Al^{3+}	ferrous	Fe^{2+} (green)	oxide	O^{2-}
ammonium	NH_4^+	fluoride	F^-	perbromate	BrO_4^-
barium	Ba^{2+}	hydrogen	H^+	perchlorate	ClO_4^-
bicarbonate	HCO_3^-	hydronium	H_3O^+	periodate	IO_4^-
bisulfate	HSO_4^-	hydroxide	OH^-	permanganate	MnO_4^- *(purple)*
bisulfide	HS^-	hypobromite	BrO^-	peroxide	O_2^{2-}
bisulfite	HSO_3^-	hypochlorite	ClO^-	phosphate	PO_4^{3-}
bromate	BrO_3^-	hypoiodite	IO^-	phosphide	P^{3-}
bromide	Br^-	iodate	IO_3^-	phosphite	PO_3^{3-}
bromite	BrO_2^-	iodide	I^-	potassium	K^+
calcium	Ca^{2+}	iodite	IO_2^-	silver	Ag^+
carbonate	CO_3^{2-}	lead	Pb^{2+}	sodium	Na^+
chlorate	ClO_3^-	lithium	Li^+	stannic	Sn^{4+}
chloride	Cl^-	magnesium	Mg^{2+}	stannous	Sn^{2+}
chlorite	ClO_2^-	manganese	Mn^{2+}	strontium	Sr^{2+}
chromate	CrO_4^{2-} (yellow)	mercuric	Hg^{2+}	sulfate	SO_4^{2-}
chromium	Cr^{3+}	mercurous	Hg_2^{2+}	sulfide	S^{2-}
cupric	Cu^{2+} (blue)	nickel	Ni^{2+} (green)	sulfite	SO_3^{2-}
cuprous	Cu^+ (green)	nitrate	NO_3^-	thiocyanate	SCN^-
cyanide	CN^-	nitride	N^{3-}	thiosulfate	$S_2O_3^{2-}$
dichromate	$Cr_2O_7^{2-}$ (orange)	nitrite	NO_2^-	zinc	Zn^{2+}

SOLUBILITY RULES

Always soluble:

alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+), NH_4^+,
NO_3^-, ClO_3^-, ClO_4^-, $C_2H_3O_2^-$, HCO_3^-

Generally soluble: (mnemonics)

Cl^-, Br^-, I^-	Soluble except Ag^+, Pb^{2+}, Hg_2^{2+}	(AP/H)
F^-	Soluble except Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Mg^{2+}	(CBS-PM)
SO_4^{2-}	Soluble except Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	(CBS/PBS)

Generally insoluble:

O^{2-}, OH^- Insoluble except alkali metal ions and NH_4^+
Ca^{2+}, Sr^{2+}, Ba^{2+} (CBS) somewhat soluble

CO_3^{2-}, PO_4^{3-}, S^{2-}, SO_3^{2-}, $C_2O_4^{2-}$, CrO_4^{2-}
Insoluble except alkali metals and NH_4^+

GASES THAT FORM

$\rightarrow \mathbf{H_2CO_3} \rightarrow \mathbf{CO_2} + H_2O$ $\rightarrow \mathbf{NH_4OH} \rightarrow \mathbf{NH_3} + H_2O$
$\rightarrow \mathbf{H_2SO_3} \rightarrow \mathbf{SO_2} + H_2O$ $\rightarrow \mathbf{H_2S}$
$\rightarrow 2\mathbf{HNO_2} \rightarrow \mathbf{NO} + \mathbf{NO_2} + H_2O$ $\rightarrow \mathbf{HCN}$

WEAK ELECTROLYTES

Weak Acids (*esp.* $HC_2H_3O_2$ and HF)
(Memorize the 8 strong acids… all others are weak)

HCl	hydrochloric acid	HNO_3	nitric acid
HBr	hydrobromic acid	HIO_4	periodic acid
HI	hydroiodic acid	H_2SO_4	sulfuric acid
$HClO_4$	perchloric acid	$HClO_3$	chloric acid

Ammonium Hydroxide ($NH_4OH \approx NH_3(aq)$) Water (H_2O)

DRIVING FORCES — Double Replacement

- Insoluble Solid (Precipitate)
- Weak Electrolyte (H_2O or Weak Acid)
- Gas Formation

STRONG OXIDIZERS (Oxidizing Agents)

MnO_4^- in acid solution	$\rightarrow Mn^{2+} + H_2O$
MnO_2 in acid solution	$\rightarrow Mn^{2+} + H_2O$
MnO_4^- in neutral or basic sol'n	$\rightarrow MnO_2$
$Cr_2O_7^{2-}$ in acid solution	$\rightarrow Cr^{3+} + H_2O$
$Cr_2O_7^{2-}$ with a base	$\rightarrow CrO_4^{2-} + H_2O$
CrO_4^{2-} in basic solution	$\rightarrow CrO_2^- + H_2O$
HNO_3, concentrated	$\rightarrow NO_2 + H_2O$
HNO_3, dilute (e.g. 6 M)	$\rightarrow NO + H_2O$
H_2SO_4, hot, concentrated	$\rightarrow SO_2 + H_2O$
Free halogens (e.g. Cl_2)	$\rightarrow$ halide ions (Cl^-)
H_2O_2 in acid solution	$\rightarrow H_2O$
Note: H_2O_2 decomposes	$\rightarrow H_2O + O_2$
Na_2O_2	$\rightarrow NaOH$
$HClO_4$	$\rightarrow Cl^- + H_2O$

Other Oxidizers

Metal-"ic" ions (e.g. Sn^{4+}, Fe^{3+})	$\rightarrow$ "-ous" ions (Sn^{2+}, Fe^{2+})
H_2O	$\rightarrow H_2 + OH^-$

STRONG REDUCERS (Reducing Agents)

Halide ions (e.g. Cl^-)	$\rightarrow$ Free halogen (Cl_2)
Free metals	$\rightarrow$ metal ions
"ites" SO_3^{2-} or SO_2, NO_2^-	$\rightarrow$ "ates" SO_4^{2-}, NO_3^-
Free halogens, dil. basic sol'n	$\rightarrow$ hypohalite ions (ClO^-)
Free halogens, conc. basic sol'n	$\rightarrow$ halate ions (ClO_3^-)
$S_2O_3^{2-}$	$\rightarrow S_4O_6^{2-}$

Other Reducers

Metal-"ous" ions (e.g. Sn^{2+})	$\rightarrow$ "-ic" ions (Sn^{4+})
H_2O	$\rightarrow O_2 + H^+$

Stuff I Should Know (Page 2)

Complex Ions & Common Ligands

Ligands	polar molecules & anions	NH_3, H_2O, OH^-, CN^-, Cl^-	**Odd example:** $Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$
Central Ions	transition metals and Al^{3+}	Ag^+, Cu^{2+}, Ni^{2+}, Zn^{2+}, etc. & Al^{3+}	
Examples	usually twice the number of ligands as the charge on the central ion. **Key Words**: "excess, concentrated"	$Ag(CN)_2^-$, $Cu(NH_3)_4^{2+}$, $Ni(OH)_4^{2-}$, $Zn(NH_3)_4^{2+}$, $Al(OH)_6^{3-}$	**Reaction with Acid:** $Cu(NH_3)_4^{2+} + H^+ \rightarrow Cu^{2+} + NH_4^+$

Organic Chemistry & Functional Groups

alkanes	alkenes	alkynes	aromatics (benzene)
C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}	C_6H_6
alcohol	**aldehyde**	**ketone**	**ether**
R — OH	O ‖ R — C — H	O ‖ R — C — R	R — O — R
carboxylic acid	**ester**	**amine**	**amide**
O ‖ R — C — OH	O ‖ R — C — O — R	R — NH_2	O ‖ R — C — NH_2
Substituted benzene:	**ortho = 1,2**	**meta = 1,3**	**para = 1,4**

nuclear chem
alpha
${}^{4}_{2}He$
beta/electron
${}^{0}_{-1}e$
neutron
${}^{1}_{0}n$
positron
${}^{0}_{+1}e$

ΔH	ΔS	Spont.?
–	+	at all temps
+	+	high temps
–	–	low temps
+	–	no temps

Note: ΔS in J
ΔG & ΔH in kJ

K_{sp} & Solubility, s

1:1 $K_{sp} = s^2$
1:2 $K_{sp} = 4s^3$
1:3 $K_{sp} = 27s^4$
2:3 $K_{sp} = 108s^5$

Lewis Acids & Bases

$BF_3 + NH_3 \rightarrow BF_3NH_3$

acid anhydrides (oxides of nonmetals, CO_2)
basic anhydrides (oxides of metals, MgO)

$MgO + CO_2 \rightarrow MgCO_3$

decomposition reactions: $MgCO_3 \rightarrow MgO + CO_2$
Strange Examples: $P_4O_{10} + H_2O \rightarrow H_3PO_4$

Strange Ions: (nitride, N^{3-}) (hydride, H^-)
$Li + N_2 \rightarrow Li_3N$ $LiH + H_2O \rightarrow H_2 + Li^+ + OH^-$

Flame Test Colors

Barium – green
Sodium – yellow
Copper – blue (w/ green)
Potassium – lavender
Strontium – red
Lithium – red
Calcium – orange

Quantum Numbers

n	1, 2, 3, …
l	0 … (**n**–1)
m	–**l** … +**l**
m_s	+½, –½

l	0 = s, 1 = p, 2 = d, 3 = f

Writing Lewis Structures
hint: use one valence electron to connect F's or Cl's then determine lone pairs (Ex: XeF_4)

Product–Favored (Spontaneous) Reactions

$\Delta G < 0$ $E° > 0$ $K_{eq} > 1$

Properties Indicate Strength of Intermolecular Forces (IMF's)

IMF	BP	FP	H_{vap}	H_{fus}	VP
IMF	BP	FP	H_{vap}	H_{fus}	VP

Orders of Reactions & Graphs That Give Straight Lines

0 Order	1st Order	2nd Order
[R] vs. Time	ln[R] vs. Time	1/[R] vs. Time
slope = -k	slope = -k	slope = k

Electrochemical Cells

anode	cathode
oxidation	reduction
– side	+ side
lower E°	higher E°
e^- leave	e^- enter

Bond Orders

bond	B.O.	
single	1	σ
double	2	$\sigma+\pi$
triple	3	$\sigma+\pi+\pi$

SN & hybridization & shape

Steric Number	hybridization	basic shape
1	s	—
2	sp	linear
3	sp^2	△ planar
4	sp^3	tetrahedral
5	sp^3d	△ bipyramidal
6	sp^3d^2	octahedral

IMF's

London	nonpolar molecules, ex: CH_4, He
dipole–dipole	polar molecules, ex: H_2S, SO_2
hydrogen bonding	H–F, H–O–, H–N–, NH_3, H_2O amines and alcohols
metallic	metals, Ag, Pb
ionic	salts, NaCl, $CaCO_3$ (Note: "ates" contain covalent bonds)
covalent network	C(graphite), C(diamond), SiO_2, WC, Si, SiC (Note: graphite = London, too)

Activity of Metals (Four Groups)

Metals	React with...
Groups I & II	H_2O ex: $Li + H_2O \rightarrow Li^+ + OH^- + H_2$
all others	Non-oxidizing Acid, ex: HCl $Zn + 2HCl \rightarrow H_2 + ZnCl_2$
Cu, Ag, Hg	Oxidizing Acid, HNO_3 or H_2SO_4 (conc.) $Cu + HNO_3 \rightarrow NO_2 + H_2O + Cu^{2+}$
Au, Pt, Ir	Aqua Regia (HNO_3 + HCl)

Molecules & Compounds

Writing Formulas and Naming Compounds

Introduction

Writing formulas and naming compounds can be confusing because there are different types of compounds that follow different rules. Additionally, some compounds (H_2O, NH_3, CH_4, etc.) simply have ***common names*** that must be memorized.

The two types of compounds we will focus on first are ***ionic compounds*** (formed from positive and negative ions) and ***binary nonmetal compounds*** (molecular compounds). Later we will add ***acids***. So… you must recognize the ***type*** of compound before you try to name it. [Note: + ion = "cation" and – ion = "anion".]

	Ionic	**Binary Nonmetal**
Formula	+ ion before – ion ex: NaCl $(NH_4)_2SO_4$ Al_2S_3	usually the less electronegative atom is first ex: CO CO_2 N_2O
Naming	Name of cation + name of anion sodium chloride ammonium sulfate aluminum sulfide	Indicate the number (mono, di, tri, and kind of atoms. First element is simply name of element. Second element name ends with "ide" carbon monoxide carbon dioxide dinitrogen monoxide

I. Writing Ionic Formulas

	Cl^-	NO_3^-	S^{2-}	CO_3^{2-}	N^{3-}	PO_4^{3-}	OH^-
Na^+							
NH_4^+							
Sn^{2+}							
Hg_2^{2+}							
Al^{3+}							
Sn^{4+}							

II. Naming Ionic Compounds

Cation	Anion	Formula	Name
Cu^{2+}	OH^-		
Ba^{2+}	SO_4^{2-}		
NH_4^+	$Cr_2O_7^{2-}$		
Ag^+	$C_2H_3O_2^-$		

mono	di	tri	tetra	penta	hexa	hepta	octa	nona	deca

III. Writing Formulas of Binary Nonmetal Compounds

Name	Formula	Name	Formula
nitrogen trifluoride		phosphorus trichloride	
nitrogen monoxide		phosphorus pentachloride	
nitrogen dioxide		sulfur hexafluoride	
dinitrogen tetroxide		disulfur decafluoride	
dinitrogen monoxide		xenon tetrafluoride	

IV. Naming Binary Nonmetal Compounds

Name	Formula	Name	Formula
	CCl_4		HBr
	P_4O_{10}		N_2F_4
	ClF_3		XeF_3
	BCl_3		PI_3
	SF_4		SCl_2

V. Practice for Both Types of Compounds

Formula	Name
HCl	
PCl_5	
K_2S	
$NiSO_4$	
ClF_3	
OF_2	
$Al(OH)_3$	
NCl_3	
$(NH_4)_3PO_4$	

Formula	Name
	carbon dioxide
	ammonium carbonate
	sulfur dichloride
	calcium iodide
	boron trifluoride
	phosphorus triiodide
	magnesium perchlorate
	potassium permanganate
	aluminum phosphate

Naming

Type of Compound	Ionic	Acids	Molecular
How To Recognize	Recognize + and - ion	H+ and - ion	Not Ionic
How To Name	names of + ion then - ion	"ides" → hydro---ic acid "ates" → ----ic acid "ites" → ---ous acid S (add "ur") P (add "or")	mono, di, tri, tetra, penta, hexa, hepta, octa, nona ,deca names ends with "ide" penta**ao**xide → pentoxide, etc.

Indicate the Type of Compound and then name the compound using the appropriate rules:

1.	NaF	I	Sodium fluoride	19.	MnO_2	___	___
2.	$FeCl_3$	___	___	20.	H_2S	___	___
3.	CO_2	___	___	21.	$CuCl_2$	___	___
4.	$MgCl_2$	___	___	22.	$AgNO_3$	___	___
5.	HF	___	___	23.	CO	___	___
6.	SF_4	___	___	24.	H_3PO_4	___	___
7.	$HC_2H_3O_2$	___	___	25.	NaCl	___	___
8.	H_2O	___	___	26.	N_2O_5	___	___
9.	NH_3	___	___	27.	NO_2	___	___
10.	CaO	___	___	28.	HNO_3	___	___
11.	NH_4NO_3	___	___	29.	NaOH	___	___
12.	NaI	___	___	30.	$SnCl_2$	___	___
13.	$PbCO_3$	___	___	31.	$CaSO_4$	___	___
14.	Na_2O	___	___	32.	HBr	___	___
15.	$Ba(NO_3)_2$	___	___	33.	$Cu(OH)_2$	___	___
16.	K_2CrO_4	___	___	34.	$Zn(OH)_2$	___	___
17.	NO	___	___	35.	$BaCl_2$	___	___
18.	HCl	___	___	36.	PCl_5	___	___

Naming Acids

Write the formula of the polyatomic ion. Then Write the formula of the acid. Finally, write the name of the acid.

"ate" becomes "_____ic acid"
"ite" becomes "_____ous acid"
"ide" becomes "hydro_____ic acid

in sulfur compounds, add "ur"
in phosphorus compounds, add "or"

bromate	BrO_3^-	$HBrO_3$	bromic acid	perchlorate	_____	_______	____________
periodate	_____	_______	____________	bisulfate*	_____	_______	____________
carbonate	_____	_______	____________	hypoiodite	_____	_______	____________
peroxide*	_____	_______	____________	bicarbonate*	_____	_______	____________
chloride	_____	_______	____________	sulfate	_____	_______	____________
chlorite	_____	_______	____________	iodite	_____	_______	____________
thiosulfate	_____	_______	____________	acetate	_____	_______	____________
sulfide	_____	_______	____________	iodide	_____	_______	____________
dichromate	_____	_______	____________	bromide	_____	_______	____________
hypobromite	_____	_______	____________	hydroxide*	_____	_______	____________
sulfite	_____	_______	____________	phosphate	_____	_______	____________
chromate	_____	_______	____________	hypochlorite	_____	_______	____________
permanganate	_____	_______	____________	phosphite	_____	_______	____________
iodate	_____	_______	____________	oxide*	_____	_______	____________
perbromate	_____	_______	____________	fluoride	_____	_______	____________
cyanide	_____	_______	____________	thiocyanate	_____	_______	____________
chlorate	_____	_______	____________	bromite	_____	_______	____________
nitrate	_____	_______	____________	nitrite	_____	_______	____________

* = be careful

Organic Naming

Organic Nomenclature - Alkanes, Alkenes, Alkynes

Naming organic compounds can be a challenge to any chemist at any level. Historically, chemists developed names for new compounds without any systematic guidelines. In this century, the need for standardization was recognized. For simple molecules, the nomenclature system worked out by the International Union of Pure and Applied Chemists (IUPAC) works well. For complex molecules, the IUPAC names are so long that no one in their right mind would use them. The net result is that a hodgepodge of IUPAC names and historic or common names is used. Any one compound may have five or six different names.

So, what we want to accomplish in this module is simply to establish the fundamentals of the IUPAC system and apply them to naming **alkanes, alkenes and alkynes**. These groups are **hydrocarbons**, compounds made of the elements carbon and hydrogen.

Numerical Prefixes = Number of Backbone Carbon Atoms

The prefix in the name of an organic molecule indicates the number of carbon atoms found in the longest continuous chain of carbon atoms in the molecule. You need to memorize the following prefixes:

Prefix	# C atoms
meth-	1
eth-	2
prop-	3
but-	4
pent-	5
hex-	6
hept-	7
oct-	8
non-	9
dec-	10

Alkanes = -ane ending

The alkanes are the least complex hydrocarbons. The alkane family uses the *prefix for the number of carbons* and an *-ane ending*. An alkane can be recognized by its general formula, C_nH_{2n+2}, where *n* is the number of carbon atoms in the compound. For example, C_5H_{12} has five carbon atoms pentane. Each member of the alkane family differs from the next by a —CH_2— group, and all the carbons are connected by single bonds.

Example 1:
Name the following compounds:

a. CH_4
b. C_2H_6 or CH_3CH_3
c. C_3H_8 or $CH_3CH_2CH_3$
d. C_4H_{10} or CH_3 $CH_2CH_2CH_3$

Solution 1:
All of the formulas fit into general formula, C_nH_{2n+2}, therefore the bonds in these compounds are single bonds; they are alkanes. Use the numerical prefix for the number of carbon atoms with the -ane ending.

a. one C atom = methane
b. two C atoms = ethane
c. three C atoms = propane
d. four C atoms = butane

Alkenes = -ene ending

Hydrocarbons that contain multiple bonds are called **unsaturated hydrocarbons**. If the hydrocarbon has **one double bond**, its general formula will be C_nH_{2n}, where *n* is the number of carbon atoms in the compound. The alkene family uses the *-ene ending.* The double bond is stronger than a single bond, and the bond length between the carbon atoms is shorter in the double bond. It is also more reactive than a single bond since the π bond (the second pair of electrons) is farther from the nuclei.

Naming is a little bit more complex for alkenes than alkanes. Since the double bond could appear at various sites in a typical molecule, we have to specify where it is. To do so, number the carbon backbone so that the **lowest possible number** is used to describe the double bond position. The lowest number of the two C atoms involved in the double bond is used in front of the name to indicate the C=C position. The number is place at the beginning of the name and is separated with a dash.

In the expanded structure formulas shown below, it is understood that since H only forms one bond, any double bonds are between carbon atoms. The expanded structures give a bit more information about how many H atoms are attached to each C atom.

Example 2:
Name the following compounds.

a. C_2H_4 or $H_2C{=}CH_2$
b. C_3H_6 or $CH_3CH{=}CH_2$
c. C_4H_8 or $H_2C{=}CHCH_2CH_3$
d. C_4H_8 or $CH_3CH_2{=}CH_2CH_3$
e. C_5H_{10} or $CH_3CH_2CH_2CH{=}CH_2$

Solution 2:
a. 2 C atoms = ethene (since there are no options for the position of the C=C, we do not need to specify the position, as in 1-ethene)
b. 3 C atoms = propene (again, since there are no options for the position of the C=C, we do not need to specify 1-propene. Convince yourself that 1-propene and 2-propene are really the same molecule.)
c. 4 C atoms with the C=C after the #1 C atom = 1-butene
d. 4 C atoms with the C=C after the #2 C atom = 2-butene
e. 5 C atoms with the C=C after the #1 C atom = 1-pentene (Did you say 4-pentene? Remember that we want to number the backbone of C atoms so that the lowest numbers are used in the name. In this case, you want to number the C backbone from right to left. This same molecule could also be written $H_2C{=}CHCH_2CH_2CH_3$).

Alkynes = -yne ending

The alkyne family contains a **triple bond** between two C atoms. If the hydrocarbon has one triple bond, its general formula will be C_nH_{2n-2}, where *n* is the number of carbon atoms in the compound. The alkyne family uses the -yne ending. The triple bond is stronger than either the double or single bond, therefore it is also shorter and more reactive than the single or double bond.

Just as in the alkene family, the position of the triple bond is specified with a number at the beginning of the name.

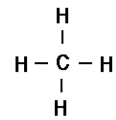
Example 3:
Name the following compounds.

a. $CH{\equiv}CH$
b. $CH{\equiv}CCH_2CH_2CH_2CH_3$
c. $CH_3C{\equiv}CCH_2CH_2CH_3$
d. $CH_3CH_2C{\equiv}CCH_2CH_3$
e. $CH_3CH_2CH_2C{\equiv}CCH_3$
f. $CH_3CH_2CH_2CH_2C{\equiv}CH$

Solution 3:

a. 2 C atoms = ethyne (this compound is commonly known as acetylene)
b. 6 C atoms, triple bond after the #1 C atom = 1-hexyne
c. 6 C atoms, triple bond after the #2 C atom = 2-hexyne
d. 6 C atoms, triple bond after the #3 C atom = 3-hexyne
e. 6 C atoms, triple bond after the #2 C atom = 2-hexyne (number the backbone from right to left)
f. 6 C atoms, triple bond after the #1 C atom = 1-hexyne (number the backbone from right to left)

NOMENCLATURE Worksheet
Draw the following organic molecules like the example.

Methane: 

1. Ethane

2. Propane

3. Decane

4. Propyne

5. 3-Octyne

6. 1-Propene

7. 2-Nonene

8. Nonane

9. 4-Nonyne

10. 3-Hexene

11. How many ways can you write butene? Draw them.

12. Why is 6-decene not possible? What would it be called?
Draw it.

Name the following compounds.

13. CH_3-CH_2-CH=CH-CH_2-CH_2-CH_3

14. CH_3-CH=CH-CH_2-CH_2-CH_3

15. CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3

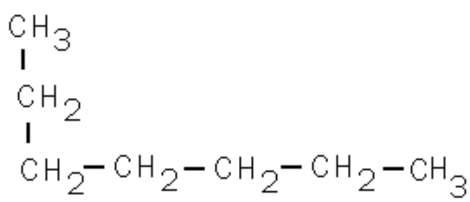
16. CH_3-CH_2-CH=CH-CH_2-CH_2-CH_2-CH_3

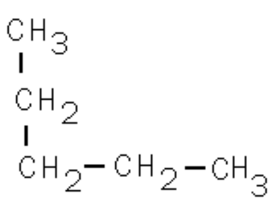
17. CH_3-CH_3

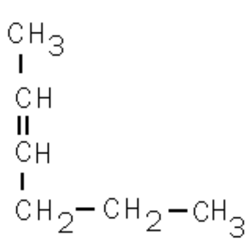
18. $CH_3-C\equiv C-CH_3$

19. $CH_3-C\equiv C-CH_2-CH_2-CH_3$

20. $CH_3-CH_2-CH_2-C\equiv C-CH_2-CH_2-CH_2-CH_2-CH_3$

21. 

22. 

23. 

<u>SUMMER ASSIGNMENT #3</u>

This assignment is about nuclear chemistry. What I am about to say will sound odd. Nuclear Chemistry is NOT part of the AP curricula. However, it is considered prior knowledge. So you will never have an AP question directly about Nuclear Chemistry, it is expected that an AP student understands the concept and may be asked to use this knowledge in conjunction with one of the six big ideas of AP Chemistry.

The 6 Big Ideas of AP Chemistry

- Big Idea 1: Structure or Matter
- Big Idea 2: Properties of Matter – Characteristics, States, and Forces of Attraction
- Big Idea 3: Chemical Reactions
- Big Idea 4: Rates of Chemical Reactions
- Big Idea 5: Thermodynamics
- Big Idea 6: Equilibrium

<u>You may want to refer to the two videos I linked in assignment #3. This assignment delves deeper into the nucleus and nuclear reaction; fission and fusion.</u>

Nuclear Chemistry: Crash Course Chemistry #38
https://www.youtube.com/watch?v=KWAsz59F8gA
Nuclear Chemistry Part 2: Fusion and Fission - Crash Course Chemistry #39
https://www.youtube.com/watch?v=FU6y1XIADdg

Study Guide for Nuclear Chemistry

1. <u>Positron</u>

- - particle of charge +1 and mass equal to that of an electron.
- $\boxed{{}^{0}_{1}\mathrm{e}}\ \boxed{{}^{0}_{+1}\beta}$

2. <u>Alpha particle</u>

- Emitted helium nucleus.
- ${}^{4}_{2}He^{+2}$

3. <u>Beta particle</u>

- Energetic electron from a decomposed neutron.
- ${}^{0}_{-1}e$ *or* β^{-}

4. <u>Transuranium elements</u>

- Element with atomic number greater than
  92

5. Gamma radiation

- High energy electromagnetic radiation
- $^{0}_{0}\gamma$

6. Transmutation

- . conversion of an atom of one element to an atom of another element

7. Fission

- Splitting of nucleus into two similar – sized pieces.

8. Fusion

- combination of two nuclei to form a large nucleus

9. Radioisotope

- Element with unstable nucleus

Nuclear Particles

Particle	Symbol	Charge
Proton	p+ $^{1}_{1}H$	+1
Neutron	n $^{1}_{0}n$	0
Electron/Beta	β- $^{0}_{-1}e$	-1
Positron	β+ $^{0}_{+1}e$	+1
Alpha particle	α $^{4}_{2}He$	+2
Gamma ray	γ	0

10. What is the charge on an alpha particle?

11. How many neutrons are there in an alpha particle?

12. What is the change in the atomic number when an atom emits an alpha particle?

13. What is the change in atomic mass when an atom emits an alpha particle.

14. What is the change in the atomic number when an atom emits a beta particle.

15. What is the change in the atomic number when an atom emits gamma radiation?

16. What particle is emitted in alpha radiation?

17. Which symbol is used for an alpha particle.

18. What is the minimum thickness needed to stop an alpha particle.

- A sheet of paper can stop an alpha particle.
- Alpha particles are the weakest form of radiation.

19. What symbol is used for beta radiation?

20. What is the minimum thickness needed to stop a beta particle.

- A sheet of aluminum foil.

21. What is the minimum thickness needed to stop gamma radiation

- Three inches of lead.

22. The most penetrating form of radiation is

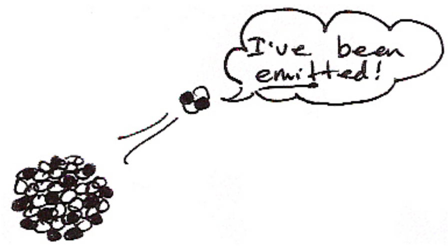
23. Which type of ionizing radiation can be blocked by clothing?

24. If the half life of a radioactive material is 8 years, how many years will it take for one half of the original amount to decay.

25. A piece of wood found in an ancient burial mound contains only half as much carbon-14 as a piece of woodcut from a living tree growing nearby. If the half-life for carbon-14 is 5730 years, what is the approximate age of the ancient wood

26. After 42 days, 2 g of phosphorous-32 has decayed to . 25 g. What is the half-life phosphorous-32.

27. Above what atomic number are all atoms radioactive?



The Nucleus – Radioactivity

NUCLEAR EQUATIONS

Write "isotopic symbols" for:

Example: $^{238}_{92}U$

alpha, α	beta, β^-	gamma, γ	positron, β^+	neutron, n°

Nuclear Change: ___ → ___ ___ → ___

Complete these nuclear reactions:

1. $^{238}_{92}U \rightarrow\ ^{234}_{90}Th + \square$ (_________ decay)

2. $^{234}_{90}Th \rightarrow\ ^{234}_{91}Pa + \square$ (_________ decay)

3. $^{234}_{91}Pa \rightarrow \square +\ ^{4}_{2}He$ (alpha decay)

4. $^{220}_{86}Rn \rightarrow \square +\ ^{4}_{2}He$ (alpha decay)

5. $^{216}_{84}Po \rightarrow \square +\ ^{0}_{-1}e$ (beta decay)

6. $^{14}_{6}C \rightarrow\ ^{14}_{7}N + \square$ (_________ decay)

7. $^{210}_{83}\text{Bi} \rightarrow \boxed{} + \boxed{}$ (beta decay)

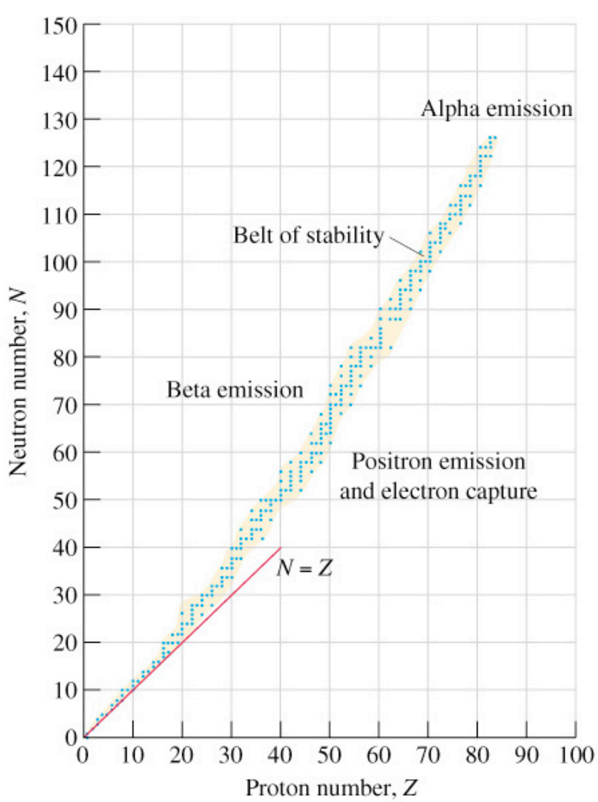
8. $^{1}_{0}\text{n} + {}^{10}_{5}\text{B} \rightarrow \boxed{} + {}^{4}_{2}\text{He}$ (neutron bombardment w/alpha emission)

9. $^{23}_{12}\text{Mg} \rightarrow \boxed{} + \boxed{} + {}^{0}_{0}\gamma$ (positron decay w/gamma emission)

Nuclear Chemistry

WHY ARE NUCLEI UNSTABLE?

Most of an atom's chemistry depends on it's ***electrons***. Gaining, losing, or sharing electrons is the basis of most chemistry In this chapter we are concentrating on the ***nucleus***.



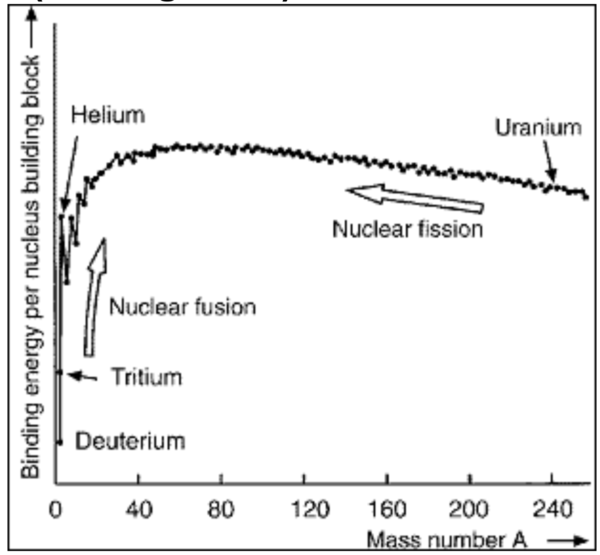
Radioactivity comes from unstable nuclei. What we know is that only a limited number of combinations of protons and neutrons form stable (un-radioactive) nuclei. There is a size limit (83 protons). So, if a nucleus has too many protons, too many neutrons, or is just too big, it will go through radioactive decay.

Guidelines:

- Given an isotope, compare it to the most common isotope (use the mass from the periodic table) and decide whether the isotope has too many p^+'s, too many n°'s, or is too big ($Z > 83$).
- Too many neutrons (left of the "belt of stability") will result in beta decay (neutron $\rightarrow$ proton + beta particle)
- Too many protons (right of the "belt of stability") will result in either positron decay or electron capture both of which involve a proton $\rightarrow$ neutron.
- If the atom is just too big, alpha decay allows the nucleus to lose two protons and two neutrons.

Another Idea: BINDING ENERGY (missing mass)

Calculate the mass of an atom of ^{56}Fe by adding up the masses of the particles. If you do and then compare it to the measured mass, you will find that there is some missing mass. This mass has been turned into energy ($E=mc^2$) and is called the "binding energy" which holds the nucleus together.



This graph is binding energy <u>per nucleon</u>

Binding energy increases as atomic number increases because more and more energy is needed to hold together the larger nuclei. However, if you divide a nucleus's binding energy by the number of protons and neutrons (nucleons) being held together, you get the graph to the right. The peak is at ^{56}Fe.

If you were a larger atom (like U) and broke into smaller pieces, there would be some missing mass. This is ***nuclear fission***.

If you are a small nucleus (like H) and join together with other atoms to form larger nuclei, there will also be missing mass (a lot more). This is ***nuclear fusion***.

<u>SUMMER ASSIGNMENT #4</u>

Assignment #4 deals with reactions and bonding. I have attached an activity series, solubility rules, electronegativity table and a VSPER Cheat Sheet. You WILL NOT be given these reference sheets on the AP test. However, you DO NOT have to memorize these rules. On the AP exam you will be given enough information to answer the questions. You will need to know your VSPER shapes for the AP exam. We will be learning some new VSPER shapes this year too.

TABLE 4.4 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li \longrightarrow Li^{+} + e^{-}$
Potassium	$K \longrightarrow K^{+} + e^{-}$
Barium	$Ba \longrightarrow Ba^{2+} + 2e^{-}$
Calcium	$Ca \longrightarrow Ca^{2+} + 2e^{-}$
Sodium	$Na \longrightarrow Na^{+} + e^{-}$
Magnesium	$Mg \longrightarrow Mg^{2+} + 2e^{-}$
Aluminum	$Al \longrightarrow Al^{3+} + 3e^{-}$
Manganese	$Mn \longrightarrow Mn^{2+} + 2e^{-}$
Zinc	$Zn \longrightarrow Zn^{2+} + 2e^{-}$
Chromium	$Cr \longrightarrow Cr^{3+} + 3e^{-}$
Iron	$Fe \longrightarrow Fe^{2+} + 2e^{-}$
Cobalt	$Co \longrightarrow Co^{2+} + 2e^{-}$
Nickel	$Ni \longrightarrow Ni^{2+} + 2e^{-}$
Tin	$Sn \longrightarrow Sn^{2+} + 2e^{-}$
Lead	$Pb \longrightarrow Pb^{2+} + 2e^{-}$
Hydrogen	$H_2 \longrightarrow 2H^{+} + 2e^{-}$
Copper	$Cu \longrightarrow Cu^{2+} + 2e^{-}$
Silver	$Ag \longrightarrow Ag^{+} + e^{-}$
Mercury	$Hg \longrightarrow Hg^{2+} + 2e^{-}$
Platinum	$Pt \longrightarrow Pt^{2+} + 2e^{-}$
Gold	$Au \longrightarrow Au^{3+} + 3e^{-}$

Ease of oxidation increases

Ion	General Solubility Rule
NO_3^-	All nitrates are soluble
$C_2H_3O_2^-$	All acetates are soluble (Ag $C_2H_3O_2$ only moderately)
Cl^-, Br^-, I^-	All chlorides, bromides and iodides are soluble except Ag^+, Pb^+ and Hg_2^{2+}. ($PbCl_2$ is slightly soluble in cold water and moderately soluble in hot water.)
SO_4^{2-}	All sulfates are soluble except those of Ba^{2+}, Pb^{2+}, Ca^{2+} and Sr^{2+}.
CO_3^{2-} and PO_4^{3-}	All carbonates and phosphates are insoluble except those of Na^+, K^+ and NH_4^+. (Many acid phosphates are soluble)
OH^-	All hydroxides are insoluble except those of Na^+ and K^+. Hydroxides of Ba^{2+} and Ca^{2+} are slightly soluble.
S^{2-}	All sulfides are insoluble except those of Na^+, K^+, NH_4^+ and those of the alkaline earths: Mg^{2+}, Ca^{2+}, Sr^{2+} and Ba^{2+}. (Sulfides of Al^{3+} and Cr^{3+} hydrolyze and precipitate as the corresponding hydroxides.
Na^+, K^+ and NH_4^+	All salts of sodium ion, potassium ion and ammonium ion are soluble except several uncommon ones.

Elements and Bonding

1) Classify each of the following elements as an alkali metal, an alkaline-earth metal, transition metal, metalloid, halogen, or noble gas based on its position in the periodic table:

- boron ______________________________________
- gold ______________________________________
- krypton ______________________________________
- calcium ______________________________________

2) How many valence electrons do each of the following elements have?

- carbon _____
- selenium _____
- xenon _____
- potassium _____

3) Which of the following ions are likely to be formed?

- N^{+5} ______
- He^{+} ______
- F^{-1} ______
- Al^{+2} ______
- P^{-3} ______
- Mg^{+2} _____

4) Explain why oxygen is a fairly reactive element while neon is not.

5) Explain why beryllium loses electrons when forming ionic bonds, while sulfur gains electrons.

6) Explain why fluorine and chlorine have similar reactivities (the word "valence" should be somewhere in your answer!)

Bonding Review

1) Barium iodide contains what type of bonding?

2) Carbon tetrachloride contains what type of bonding?

3) Molecular oxygen contains what type of bonding?

4) Liquid mercury contains what type of bonding?

5) Using electronegativity differences compare the bonding in carbon tetrachloride and molecular oxygen. (*See Table Below)*

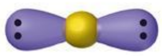
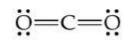
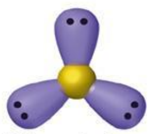
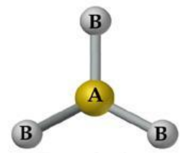
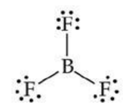
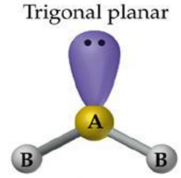
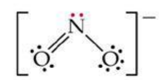
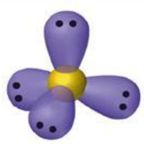
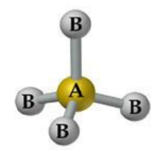
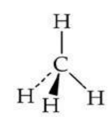
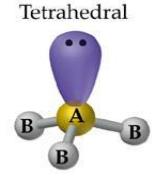
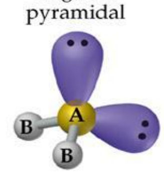
6) Draw Lewis structures, give molecular geometry name and indicate **molecular** polarity for the following species. (*VSPER Cheat Sheet on next page)*

A) CCl_4 B)CO_2 C) BCl_3 D) H_2O

H 2.20																	He n.a.
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne n.a.
Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar n.a.
K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.60	Mo 2.16	Tc 1.90	Ru 2.20	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.10	I 2.66	Xe 2.60
Cs 0.79	Ba 0.89	La 1.10	Hf 1.30	Ta 1.50	W 2.36	Re 1.90	Os 2.20	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.00	At 2.20	Rn n.a.
Fr 0.70	Ra 0.89	Ac 1.10	Rf n.a.	Db n.a.	Sg n.a.	Bh n.a.	Hs n.a.	Mt n.a.	Ds n.a.	Rg n.a.	Uub n.a.	—	Uuq n.a.	—	—	—	—

VSPER Cheat Sheet

TABLE 9.2 Electron-Domain Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains Around the Central Atom

Number of Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
2	 Linear	2	0	 Linear	
3	 Trigonal planar	3	0	 Trigonal planar	
		2	1	 Bent	
4	 Tetrahedral	4	0	 Tetrahedral	
		3	1	 Trigonal pyramidal	
		2	2	 Bent	

<u>Predicting Reaction Products</u>

Predict the products of each of the following chemical reactions, then balance the equation. If a reaction will not occur, explain why not:

1) ____ Ag_2SO_4 + ____ $NaNO_3$ $\rightarrow$

2) ____ NaI + ____ $CaSO_4$ $\rightarrow$

3) ____ HNO_3 + ____ $Ca(OH)_2$ $\rightarrow$

4) ____ $CaCO_3$ $\rightarrow$

5) ____ $AlCl_3$ + ____ $(NH_4)_3PO_4$ $\rightarrow$

6) ____ Pb + ____ $Fe(NO_3)_3$ $\rightarrow$

7) ____ C_3H_6 + ____ O_2 $\rightarrow$

8) ____ Na + ____ $CaSO_4$ $\rightarrow$

Balancing Equations and Simple Stoichiometry

Answers are provided on the second sheet. Please try to do the worksheet without referring to them, because you'll be expected to know this stuff the first day of school!

Balance the following equations:

1) ___ N_2 + ___ F_2 → ___ NF_3

2) ___ C_6H_{10} + ___ O_2 → ___ CO_2 + ___ H_2O

3) ___ HBr + ___ $KHCO_3$ → ___ H_2O + ___ KBr + ___ CO_2

4) ___ $GaBr_3$ + ___ Na_2SO_3 → ___ $Ga_2(SO_3)_3$ + ___ NaBr

5) ___ SnO + ___ NF_3 → ___ SnF_2 + ___ N_2O_3

Using the equation from problem 2 above, answer the following questions:

6) If I do this reaction with 35 grams of C_6H_{10} and 45 grams of oxygen, how many grams of carbon dioxide will be formed?

7) What is the limiting reagent for problem 6? ___________

8) How much of the excess reagent is left over after the reaction from problem 6 is finished?

9) If 35 grams of carbon dioxide are actually formed from the reaction in problem 6, what is the percent yield of this reaction?

Stoichiometry Practice

Solve the following stoichiometry grams-grams problems:

1) Using the following equation:

$$2\ NaOH + H_2SO_4 \rightarrow 2\ H_2O + Na_2SO_4$$

How many grams of sodium sulfate will be formed if you start with 200 grams of sodium hydroxide and you have an excess of sulfuric acid?

2) Using the following equation:

$$Pb(SO_4)_2 + 4\ LiNO_3 \rightarrow Pb(NO_3)_4 + 2\ Li_2SO_4$$

How many grams of lithium nitrate will be needed to make 250 grams of lithium sulfate, assuming that you have an adequate amount of lead (IV) sulfate to do the reaction?

SUMMER ASSIGNMENT #5

Assignment #5 deals with gas laws. I have also included some more stoichiometry.

Boyle's Law	Charles' Law	Guy-Lassac's Law	Combined Gas Law
For a given mass of gas at constant temperature, the volume of a gas varies inversely with pressure	The volume of a fixed mass of gas is directly proportional to its Kelvin temperature if the pressure is kept constant.	The pressure of a gas is directly proportional to the Kelvin temperature if the volume is kept constant.	Combines Boyle's, Charles', and the Temperature-Pressure relationship into one equation. Each of these laws can be derived from this law.
$PV = k$ $P_1V_1 = P_2V_2$	$\frac{V}{T} = k$ $V_1T_2 = V_2T_1$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P}{T} = k$ $P_1T_2 = P_2T_1$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{PV}{T} = k$ $V_1P_1T_2 = V_2P_2T_1$ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Dalton's Law	Ideal Gas Law
At constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the pressures exerted by each gas,	The Ideal Gas Law relates the pressure, temperature, volume, and mass of a gas through the gas constant "R".
$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$	$PV = nRT$

Abbreviations	Standard Conditions
atm = atmosphere mm Hg = millimeters of mercury torr = another name for mm Hg Pa = Pascal kPa = kilopascal K = Kelvin °C = degrees Celsius	0°C = 273 K 1.00 atm = 760.0 mm Hg = 76 cm Hg = 760 torr =101.325 kPa = 101, 325 Pa = 29.9 in Hg

Conversions	Gas Law's Equation Symbols
K = °C + 273 $F = 1.8°C + 32$ $°C = °F - 32/1.8$ 1 cm^3 (cubic centimeter) = 1 mL (milliliter) 1 dm^3 (cubic decimeter) = 1 L (liter) = 1000 mL	Subscript (1) = old condition or initial condition Subscript (2) = new condition or final condition Temperature must be in Kelvins n = number of moles = grams/Molar mass R = 8.31 L-kPa/ mol-K = 0.0821 L-atm/mol-K = 62.4 L-Torr/mol-K **You must have a common set of units in the problem**

Ideal Gas Law Problems

Background

The ideal gas law states that PV=nRT, where P is the pressure of a gas, V is the volume of the gas, n is the number of moles of gas present, R is the ideal gas constant, and T is the temperature of the gas in Kelvins.

Common mistakes:

- Make sure that T is expressed in Kelvins. Remember that Kelvins are degrees Celsius + 273.15.
- Using the wrong value for R. You need to make sure that you have the right value of R for the units you are using. In AP chemistry the value for R is 0.08206 L-Atm-mol^{-1}-K^{-1}. Make sure units agree!

1. How many moles of gas does it take to occupy 105 liters at a pressure of 1.3 atmospheres and a temperature of 240 K?

2. If you have a 60 liter container that holds 55 moles of gas at a temperature of 300°C, what is the pressure inside the container?

3. It is not safe to put aerosol canisters in a campfire, because the pressure inside the canisters gets very high and they can explode. If I have a 1.5 liter canister that holds 3 moles of gas, and the campfire temperature is 1400 °C, what is the pressure inside the canister?

4. How many moles of gas are in a 20 liter scuba canister if the temperature of the canister is 300 K and the pressure is 300 atmospheres?

5. I have a balloon that can hold 50 liters of air. If I blow up this balloon with 0.5 moles of oxygen gas at a pressure of 1 atmosphere, what is the temperature of the balloon?

Dalton's Law

6. A metal tank contains three gases: oxygen, helium, and nitrogen. If the partial pressures of the three gases in the tank are 35 atm of O_2, 5 atm of N_2, and 25 atm of He, what is the total pressure inside of the tank?

7. Blast furnaces give off many unpleasant and unhealthy gases. If the total air pressure is 0.99 atm, the partial pressure of carbon dioxide is 0.05 atm, and the partial pressure of hydrogen sulfide is 0.02 atm, what is the partial pressure of the remaining air?

8. If the air from problem 2 contains 22% oxygen, what is the partial pressure of oxygen near a blast furnace?

Gas Stoichiometry

9. For the reaction $2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$, how many liters of water can be made from 5 L of oxygen gas and an excess of hydrogen?

10. How many liters of water can be made from 55 grams of oxygen gas and an excess of hydrogen at STP?

11. How many liters of water can be made from 55 grams of oxygen gas and an excess of hydrogen at a pressure of 12.4 atm and a temperature of 85^0 C?

Dalton's Law of Partial Pressures

<u>Background Info</u>
When two or more gases are introduced into the same container, each gas individually expands to uniformly occupy that container. Thus each gas in the mixture has the same volume but depending on how many moles of each is present, exerts a different pressure, called its partial pressure. Dalton's Law of Partial Pressures states that in a mixture of gases, the total pressure is the sum of the individual pressures of each gas present in the mixture - i.e. the sum of the partial pressures. **THERE IS A DIRECT RELATIONSHIP BETWEEN PRESSURE AND MOLES!**

12. If 3 moles of N_2 and 4 moles of O_2 are placed in a 35L container at a temperature of 25°C, what will the pressure of the resulting mixture of gases be?

13. Two flasks are connected with a stopcock. The first flask has a volume of 5 liters and contains nitrogen gas at a pressure of 0.75 atm. The second flask has a volume of 8 L and contains oxygen gas at a pressure of 1.25 atm. When the stopcock between the flasks is opened and the gases are free to mix, what will the pressure be in the resulting mixture?

14. What's the partial pressure of carbon dioxide in a container that holds 5 moles of carbon dioxide, 3 moles of nitrogen, and 1 mole of hydrogen and has a total pressure of 1.05 atm?

Stoichiometry Review

15 .Determine the **formula weight** for the following:

a. N_2O_5

b. $CuSO_4$

c. $Ca(HCO_3)_2$

d. $CaSO_4 \cdot 2\ H_2O$

16. Calculate the percentage by mass of the following compounds:

a. SO_3

b. CH_3COOCH_3

c. Ammonium Nitrate.

17.Determine the empirical formula of the compounds with the following compositions by mass:

a.10. 4 % C, 27. 8% S , 61. 7 % Cl

b.21.7 % C, 9.6 % O, and 68.7 % F

18. Arsenic reacts with chlorine to form a chloride. If 1.587 g of arsenic reacts with 3.755 g of chlorine, what is the simplest formula of the chloride?

19. Washing soda is a hydrate of sodium carbonate. Its formula is Na_2CO_3. x H_2O. A 2.714 g Sample of washing soda is heated until a constant mass of 1.006 g of Na_2CO_3 is reached. What is x ?

20.What is the molecular formula of each of the following compounds?

a. empirical formula CH_2 , molar mass =84g/mol.

b. Empirical formula NH_2Cl, Molar mass = 51.5 g/ Mol

21. Sodium hydroxide reacts with carbon dioxide as follows:

$$2\ NaOH(s) + CO_2\ (g) \rightarrow Na_2CO_3\ (s) + H_2O(l)$$

Which reagent is the limiting reactant when 1.85 mol of sodium hydroxide and 1.00 mol carbon dioxide are allowed to react? How many moles of sodium carbonate can be produced? How many moles of the excess reactant remain after the completion of the reaction?

22. To prevent a condition called the “bends”, deep sea divers breathe a mixture containing, in mole percent, 10.0% O_2, 10.0% N_2 , and 80.0% He.

a. Calculate the molar mass of this mixture.

b. What is the ratio of the density of this gas to that of pure Oxygen?

23. A 2.0g sample of SX_6 (g) has a volume of 329.5 cm^3 at 1.00 atm and 20^oC. Identify the element ‘X’. Name the compound.

Summary Assignment #6

Assignment #6 deals with solutions and dilutions. I have also included some more stoichiometry.

Molarity and Dilutions Practice Problems

Molarity

Solution = solute + solvent

Solute – thing beings dissolved

Solvent – thing doing the dissolving (water is called the universal solvent)

Molarity = moles of solute/ liters of solution *M=n/V*

Units – mol/L or M

1) How many grams of potassium carbonate, K_2CO_3, are needed to make 250 mL of a 2.5 *M* solution? (*moles = molarity * volume. Make sure volume is in liters. Convert moles to grams)*

250 mL X 1L/1000mL -= 0.250 L

2.5M = n/0.250L

2.5 X 0.25 -= n

n=0.625 mole

2) How many liters of 4.0 *M* solution can be made using 125 grams of lithium bromide, LiBr?

3) What is the concentration of a solution that has a volume of 2.5 L and contains 660 grams of calcium phosphate, $Ca_3(PO_4)_2$? *(Molarity is a measure of concentration)*

4) How many grams of copper (II) fluoride, CuF_2, are needed to make 6.7 liters of a 1.2 *M* solution?

5) What is the concentration, in moles per liter, of a solution with a volume of 3.3 mL that contains 12 grams of ammonium sulfite, $(NH_4)_2SO_3$?

6) A 50 L baby pool has a chlorine concentration of 0.200 M. If 20 L of water evaporated on a hot summer day, what is the final concentration of the pool? *(Remember only water evaporated. Calculate moles of solute. Then divide by new volume)*

7) 233.76 g of NaCl is dissolved in 1 L of water. What is the concentration of this solution?

- What would happen to the concentration if we added another liter of water?

- What would happen to the concentration if we allow half of the water to evaporate?

Dilutions

Dilute solutions are often made by diluting concentrated solutions. The dilution formula is used for this. That formula is:

$$M_1V_1=M_2V_2$$

This formula can ONLY be used for dilution. **DO NOT USE IT FOR NEUTRALIZATION! YOU MUST USE STOICHIOMETRY FOR NEUTRALIZATION.**

8) What volume of concentrated hydrochloric acid (12 *M*) would be required to create a 500.0 mL solution with a concentration of 1.00 *M*?

9) What volume of the solution in question #10 would be required to create a 250 mL solution with a concentration of 0.100 *M*?

Chemical Equations and Stoichiometry

10. Oil paintings in which "white lead" has been used can be blackened by reaction with H_2S from air pollution or from the glaze over the painting itself. The blackening comes from the formation of lead sulfide, which may be cleaned off by washing with hydrogen peroxide, H_2O_2. The reaction for the cleaning process is

$$PbS\text{ (black solid)} + 4\ H_2O_2(aq) \rightarrow PbSO_4(s) + 4\ H_2O(l)$$

a. How many grams of H_2O_2 must be used to clean off 0.24 g of PbS?

b. If 0.072 g of H_2O form in the reaction, how many grams of $PbSO_4$ must also have been formed?

11. Methyl alcohol, CH_3OH, is a clean-burning, easily handled fuel. It can be made by the direct reaction of CO and H_2 (obtained from coal and water).

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

Assume you start with 12.0g of H_2 and 74.5g of CO;

a. Which of the reactants is in excess?

b. Which is the limiting reagent?

c. What mass (in grams) of the excess reagent is left after the reaction is complete?

d. How many grams of methyl alcohol can be obtained theoretically?

12. Zinc and chlorine react directly to give zinc chloride.

$$Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$$

If you begin with 1.00 mole of zinc and excess Cl_2, what is the theoretical yield of $ZnCl_2$ in grams? If you isolate 115g of $ZnCl_2$, what is the percent yield of the metal chloride?

13. Butane, which contains only C and H, is a commonly used fuel in camping stoves. To determine the formula of butane, assume you burn 0.580g of the gas and obtain 1.760g of CO_2 and 0.900g of H_2O. What is the empirical formula of butane?