Dear AP Chemistry Students,

I am excited that you are thinking about taking AP Chemistry in the 22 – 23 school year with me and I am happy that I will have an opportunity to teach you for another year!

This is a hard class that requires dedication, tenacity, grit and fair amount of work and love for Chemistry. I know that with your passion and effort, we will have a great time learning this advanced material in Chemistry.

AP Chemistry is meant to be second-year course – which means there is a lot of new material to cover and very little time to go over the basics for the topics that we covered in Honors Chemistry. However because Chemistry is comprehensive, we can't forget about all the concepts learned in the last school year.

To help you be prepared I have compiled this notes and problems packet for you to complete during the summer. **DON'T TRY TO DO IT ALL AT ONCE** – practice a little at a time to keep those brain muscles flexed over the summer.

This packet will act as an initial review for the AP Chemistry course and at the end of the first full week of class, you will be tested on this material.

This packet will be graded on completion and we can review any questions you have on the packet. There are questions in the packet that will feel foreign – they are extension of the topics that we have covered in Honors Chemistry – the aim is to see and test your limits on such questions – if you get stummed on them, don't worry and remember they are there to challenge you. We will cover them and go over any other topic during the first week. Your aim should be to make a concerted effort to review these topics, I have no doubt in my mind that you will be able to master this Summer Packet!!!

Enjoy your summer break and I look forward to working and learning AP Chemistry with you!!!

Cheers,

Mr. Patolia

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 *	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 †	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

*Lanthanoids

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97

†Actinoids

89 Ac	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	= liter(s), milliliter(s)	mol	= mole(s)
g	= gram(s)	V	= volt(s)
nm	= nanometer(s)	J, kJ	= joule(s), kilojoule(s)
atm	= atmosphere(s)	mm Hg	= millimeters of mercury

ATOMIC STRUCTURE

$$E = h\nu$$

$$c = \lambda\nu$$

E = energy

ν = frequency

λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^8$ m s^{-1}

Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}

Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ where } a\,A + b\,B \rightleftarrows c\,C + d\,D$$

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ C$$
$$= K_a \times K_b$$

$$pH = -\log[H^+], \quad pOH = -\log[OH^-]$$

$$14 = pH + pOH$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$pK_a = -\log K_a, \quad pK_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)

K_p (gas pressures)

K_a (weak acid)

K_b (weak base)

K_w (water)

KINETICS

$$[A]_t - [A]_0 = -kt$$

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant

t = time

$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{total} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{M}$$

$$K = {}^\circ C + 273$$

$$D = \frac{m}{V}$$

$$KE_{molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = \varepsilon b c$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
M = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
ε = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1}K^{-1}
    = 0.08206 L atm mol^{-1} K^{-1}
    = 62.36 L torr mol^{-1} K^{-1}
1 atm = 760 mm Hg = 760 torr
STP = 273.15 K and 1.0 atm
Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nF E^\circ$$

$$I = \frac{q}{t}$$

$$E_{cell} = E_{cell}^\circ - \frac{RT}{nF} \ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient

Faraday's constant, F = 96,485 coulombs per mole
of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

Common Polyatomic Ions

acetate	$C_2H_3O_2^-$
ammonium	NH_4^+
arsenate	AsO_4^{3-}
arsenite	AsO_3^{3-}
azide	N_3^-
benzoate	$C_7H_5O_2^-$
borate	BO_3^{3-}
bromate	BrO_3^-
carbonate	CO_3^{2-}
chlorate	ClO_3^-
chlorite	ClO_2^-
chromate	CrO_4^{2-}
cyanide	CN^-
dichromate	$Cr_2O_7^{2-}$
dihydrogen phosphate	$H_2PO_4^-$
dihydrogen phosphite	$H_2PO_3^-$
hydrogen carbonate	HCO_3^-
hydrogen phosphate	HPO_4^{2-}
hydrogen phosphite	HPO_3^{2-}
hydrogen sulfate	HSO_4^-
hydrogen sulfide	HS^-
hydrogen sulfite	HSO_3^-
hydroxide	OH^-
hypochlorite	ClO^-
iodate	IO_3^-
manganate	MnO_4^{2-}
nitrate	NO_3^-
nitrite	NO_2^-
oxalate	$C_2O_4^{2-}$
perchlorate	ClO_4^-
permanganate	MnO_4^-
peroxide	O_2^{2-}
phosphate	PO_4^{3-}
phosphite	PO_3^{3-}
silicate	SiO_3^{2-}
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
tartrate	$C_4H_4O_6^{2-}$
thiocyanate	SCN^-
thiosulfate	$S_2O_3^{2-}$

AsO_3^{3-}	arsenite
AsO_4^{3-}	arsenate
BO_3^{3-}	borate
BrO_3^-	bromate
$C_2H_3O_2^-$	acetate
$C_2O_4^{2-}$	oxalate
$C_4H_4O_6^{2-}$	tartrate
$C_7H_5O_2^-$	benzoate
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
CN^-	cyanide
CO_3^{2-}	carbonate
$Cr_2O_7^{2-}$	dichromate
CrO_4^{2-}	chromate
$H_2PO_3^-$	dihydrogen phosphite
$H_2PO_4^-$	dihydrogen phosphate
HCO_3^-	hydrogen carbonate
HPO_3^{2-}	hydrogen phosphite
HPO_4^{2-}	hydrogen phosphate
HS^-	hydrogen sulfide
HSO_3^-	hydrogen sulfite
HSO_4^-	hydrogen sulfate
IO_3^-	iodate
MnO_4^-	permanganate
MnO_4^{2-}	manganate
N_3^-	azide
NH_4^+	ammonium
NO_2^-	nitrite
NO_3^-	nitrate
O_2^{2-}	peroxide
OH^-	hydroxide
PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
$S_2O_3^{2-}$	thiosulfate
SCN^-	thiocyanate
SiO_3^{2-}	silicate
SO_3^{2-}	sulfite
SO_4^{2-}	sulfate

AP Chemistry Summer Review Part I:
Physical & Chemical Changes, Matter & Energy

1. Label each as either physical or chemical change.

    a. corrosion of aluminum metal by hydrochloric acid

    b. melting wax

    c. pulverizing an aspirin tablet

    d. digesting a Three Musketeers® bar

    e. explosion of nitroglycerin

    f. a burning match

    g. metal warming up, due to the burning match

    h. water vapor condensing on the metal

    i. the metal oxidizes, becoming dull and brittle

    j. salt being dissolved by water

2. For each process described, state whether the material being discussed (**in bold**) is a mixture or compound, <u>and</u> state whether the change is physical or chemical.

    a. An **orange liquid** is distilled (boiled to separate components with different boiling points), resulting in the collection of a red solid and a yellow liquid.

    b. A **colorless, crystalline solid** is decomposed, leaving a pale yellow-green gas and and a soft, shiny metal.

    c. A **cup of tea** becomes sweeter as sugar is added to it.

3. Classify each as mixture (homogeneous or heterogeneous) or pure substance (elements or compounds).

    a. water

    b. blood

    c. the oceans

d. iron

e. brass (an alloy of zinc and copper)

f. wine

g. sodium bicarbonate (baking soda)

4. Explain how the five states of matter and energy are related. (HINT: Think of the motion of the particles!)

5. Consider the burning of gasoline and the evaporation of gasoline. Which represents a physical change and represents a chemical change? Give the reason for your answer.

6. A) Label the arrows on the diagram below with the correct phase change processes. B) Draw a particle diagram representing each phase.

Solid   Liquid   Gas

7. Describe the three main intermolecular forces and explain how their relationship is important in determining a compound's state of matter at a particular temperature. → This is a major concept on the AP Chem Exam!

AP Chemistry Summer Review Part II: Uncertainty in Measurement and Calculations:

1. Exact Numbers:

Counted numbers and _definitions_ do not involve any measurement and are considered as exact numbers with an infinite number of significant figures. Do not consider them when determining significant figures for your final answer.

> _Definitions_: 1 week = 7 days.
> 1 mile = 5,280 feet
> 1 yard = 3 feet

> _Counted_: 5 Players on the basketball court.
> 23 students in a room
> 25 pennies used by a class in an experiment.

2. Measured Numbers:

All **measured numbers** have some degree of uncertainty.

When recording measurements, **record only the significant figures**. Record measurements to include one decimal estimate beyond the smallest increment on the measuring device.

Examples (consider a measuring instrument like a ruler):

➤ If smallest increment = 1m, then record measurement o 0.1m (i.e. 3.1_m_)

➤ If smallest increment = 0.1m, then record measurement to 0.01m (i.e. 5.67 _m_)

➤ If smallest increment = 0.01m, then record measurement to 0.001m (i.e.12.675 _m_)

c. Unless otherwise stated the uncertainty in the last significant figure _(the uncertain digit)_ is assumed to be ±1 unit. Modern digital instruments and many types of volumetric glassware will state the level of uncertainty.

3. Rules for counting Significant Figures.

a.**Non-Zero Numbers**: Non-zero numbers are always significant.
b. **Zeros:**

1: Leading zeros that come before the first non-zero number are **never** significant
2. Captive zeros (_sandwich zeros_) that fall between two non-zero digits are **always** significant.
3. Ending zeros that appear after the last non-zero digit are significant only when a decimal point appears somewhere in the number.

Examples:

Number	0.005	5005	5005.00	500.	0.0050
Sig Figs	1	4	6	3	2

c. Scientific Notation: Significant figures are recorded in the mantissa (*number $1 \le x < 10$*)
Examples:

Number	3.0×10^3	5.998×10^5	6.00000×10^{-23}	0.5×10^4
Sig Figs	2	4	6	1

4. Rules for Using Significant Figures in Calculations

(a) Multiplication, Division, Powers and Roots:-"LEAST SIG.FIG RULE"

1. The result should be reported to the same number of significant figures as the measured number having the **least number of significant figures.**

2. Only consider the number of significant figures in each of the **measured numbers! (not constants)**

> *Example 1:*
>
> $2.3 \times 5.78 = Calculator\ returns\ 13.294$
>
> $2.3\ has\ 2\ sig.fig$
>
> $5.78\ has\ 3\ sig.fig.$
>
> $2.3 \times 5.78 = 13$ *The answer must be rounded to show 2 sig.fig*

> *Example 2.*
>
> $\dfrac{1.67x10^5 \times 0.00045}{2x10^{-23}} = calculator\ returns\ 2.505000000x10^{24}$
>
> $1.67x10^5\ has\ 3\ sig.figs$
>
> $0.00045\ has\ 2\ sig.figs$
>
> $2x10^{-23}\ has\ 1\ sig.fig$
>
> $\dfrac{1.67x10^5 \times 0.00045}{2x10^{-23}} = 3x10^{24}\ (rounded\ to\ 1\ sig.fig)$

> *Example 3*
>
> $\sqrt{2.3} = calculator\ returns\ 1.516575089$
>
> $2.3\ has\ 2\ sig.figs$
>
> $\sqrt{2.3} = 1.5$ *round answer to 2 sig.figs*

(b) Addition and Subtraction: "LEAST PRECISE DECIMAL RULE"

1. The result should be reported with the same decimal precision as the measured number having the uncertain digit in the **least precise decimal place.**

2. Only consider the decimal precision in each of the **measured numbers! (not constants)**

> *Example 4: a - c*
>
> a. $123cm + 5.35cm = 128cm\ (rounded\ to\ 10^0)$
>
> b. $1.0001m + 0.0003m = 1.0004m\ (rounded\ to\ 10^{-4})$
>
> c. $1.002s - 0.998s = 0.004s\ (rounded\ to\ 10^{-3})$

> *Example 5: Watch for numbers ending with zero!*
>
> $10 + 0.0110 = calculator\ returns\ 10.0110$
>
> 10: *the uncertain digit appears in the* 10^1 *place*
>
> 0.0110: *the uncertain digit appears in the* 10^{-4} *place*
>
> $10 + 0.0110 = 10$ *round answer to the* 10^1 *place*
>
> *Rationale*: The uncertainty in the measured number 10 is ±1. The uncertainty alone in the first number (10) is greater than the entire second number (0.0110).

Problems

How many significant figures in the following numbers:

1. _____ 1,245m

2. _____ 0.030m

3. _____ 10,000m

4. _____ 1.340×10^{23} m

5. _____ 3.02003×10^{14} m

6. _____ 0.0000001m

7. _____ 1,000.

8. _____ 0.10000010

9: Convert the following numbers into standard scientific notation:

a. 96.3×10^4 g _____

b. 0.05×10^{23} s _____

c. 123×10^{-7} m _____

Problems 10 – 18: Perform the following Calculations and record your answers in the proper number of significant figures and units.

10. $0.6030s + 0.82s =$

11. $4.1m + 0.3789m - 153.22m =$

12. $\dfrac{0.307g}{(1.0 \times 10^{-3})ml} =$

13. $\sqrt[3]{5.33 \times 10^5 m} =$

Name: _____

AP Chemistry Worksheet 1: *Significant Figures and Dimensional Analysis*

For each problem below, write the equation and show your work. Always use units and box your final answer.

1. Round each of the following numbers to four significant figures, and express the result in scientific notation:

   a. 300.235800   _____

   b. 456,500   _____

   c. 0.006543210   _____

   d. 0.000957830   _____

   e. − 0.035000   _____

2. Carry out the following operations, and express the answers with the appropriate number of significant figures:
   a. 1.24056 + 75.80   _____

   b. 23/67 - 75   _____

   c. 890,000 x 112.3   _____

   d. 78,132 / 2.50   _____

3. Perform the following conversions: (You need to go online to look up some conversion factors between metric and English units.)
   a. 8.60 mi to m   _____

   b. 3.00 days to s   _____

   c. $1.55/gal to dollars per liter   _____

   d. 75.00 mi/hr to m/s   _____

   e. 55.35 ft^3 to cm^3 _____

5

Part II: Simple Metric Conversions and Consistent Units

Section 1: Metric Conversions

Fill in the chart below with the metric conversion units. Memorize the ones in bold type! An example is given:

Prefix	Symbol	Power of 10	Meaning
deci-	d	10^{-1}	10 times smaller than base unit
centi-			
milli-			
micro-			
nano-			
kilo-			

Make the following conversions – preserve the number of significant figures in the answer!

1. 450nm _____ mm

2. 34km _____ cm

3. 43 000mm _____ km

4. 4.0×10^{6} nm _____ µm

5. 3.98×10^{-3} km_____ µm

6. 456mm _____ km

7. 136 000m _____ km

8. 4.89×10^{12} mm _____ km

9. 2.68×10^{6} m _____ km

10. 456 000 µm _____ mm

Unit Multiplication – Dimensional Analysis – Factor Labeling

Units:

In the world of mathematics numbers often exist as abstract and unit-less entities. However, in the world of physics and chemistry where numbers are based upon experimentation and measurement all numbers are based in a physical reality. ***As a result, every number consists of two important parts.*** The first is a **magnitude** and the second equally important part is a **unit**. It is the unit that gives physical, real-world meaning to the number. We never write one without the other!

Examples: Note that these are all "equivalence statements"!

<div align="center">

12 ***inches*** in one ***foot***

365 ***days*** in one ***year***

7 days in one ***week***

1.0×10^9 ***bytes*** in one ***gigabyte***

</div>

Derived Units and Calculations

Many of the common units we use are actually derived units that result from performing mathematical operations on the basic units. ***When performing mathematical operations the units are treated and manipulated as if they were algebraic variables.*** Here are a few examples:

<div align="center">

Area = (length -**m**) x (width -**m**) = **m^2**

Volume = (length -**m**) x (width -**m**) x (height - **m**) = **m^3**

Velocity = (distance traveled – **m**)/(time -**s**) = **m/s**

Density = (mass – **g**)/(volume - mL) = **g/mL**

</div>

Unit Conversions

It is often necessary to convert from one system of units to another. The most efficient way to do this is using a process known as "*unit multiplication*", "*factor labeling*" or "*dimensional analysis*".

"goal posting"

One useful version of this method is called "goal posting". ***Step 1:*** Draw a "*goal post*"with the horizontal bar extending on each side. ***Step 2:*** Place the original number and unit to the left. Place the final unit on the right. ***Step 3:*** Move the original unit (cm) from the top left (*numerator*) to the bottom of the conversion factor (*denominator*). Now there is no confusion about which form of the conversion factor you will use. If you have done this correctly the original units on the top (*cm*) will be cancelled by the same unit in the denominator of the conversion factor.

Example: Consider a car traveling at ***35 m/s*** in the metric system. What would be the corresponding length in the English system (***miles / hour***)?

Solution: Note that velocity is a derived unit and has two units that must be converted: Length (Meters → miles) and Time (seconds → Hours).

Step 1: The derived unit has consists of two different units – one in the numerator and one in the denominator. Place the numerator unit *together with the number* on the "top" of the goalpost. Place the denominator units on the "bottom" of the goal post.

Step 2: The top unit will be moved down and to the right, the bottom unit will be moved up and to the right.

35 m	1.094 *yds*	1 mile	60 s	60 minute	78 miles
s	1 m	1760 *yds*	1 minute	1 hour	hour

Note that the only unit not cancelled in the numerator is miles. The only unit not cancelled in the **denominator** is hours. This gives us the finál unit of miles/hour which the correct unit for the result.

Dimensional Analysis Practice Problems

1. I have 470 milligrams of table salt, which is the chemical compound NaCl. How many liters of NaCl solution can I make if I want the solution to be 0.90% NaCl? (9 grams of salt per 1000 grams of solution).

The density of the NaCl solution is 1.0 g solution/mL solution.

2. I have a bar of gold that is 7.0 in × 4.0 in × 3.0 in. The density of gold is 19.3 g/cm^3. The price of gold currently is $1,945.94 per ounce. How much is my gold bar worth?

3. The roof of a building is 0.2 km². During a rainstorm, 5.5 cm of rain was measured to be sitting on the roof. What is the mass in kg of the water on the roof after the rainstorm? (Density of rainwater = 1g/mL).

4. The bromine content of the ocean is about 65 g of bromine per million g of sea water. How many mL of ocean must be processed to recover 500. mg of bromine, if the density of sea water is 1.0×10^3 kg/m³?

5. Light travels 186 000 miles / second. How long is a light year in meters? (1 light year is the distance light travels in one year)

Part IIIa: Subatomic Particles, Isotopes and Ions

Element or Ion	Abbreviation	Atomic Number (Z)	Average Atomic Mass (A)	Protons*	Neutrons* (for most common isotope unless otherwise noted)	Electrons*
Oxygen	O	8	16.00			
Bismuth	Bi		209.0			
	F-					
Carbon	C	6	12.01			
Carbon-14	14C		14.00	6		
Pb-208						
		15	30.97			15
			55.845			23
Potassium Ion (cation)	K+		39.10			18
Sulfur Ion (anion)	S2-		32.07			

*- Calculate the number of protons, neutrons, and electrons for the most prevalent isotope

Average Atomic Masses:

Silver has two isotopes, one with 60 neutrons and the other with 62 neutrons. Give the chemical notation for each of these isotopes and calculate the relative abundance for each isotope given that the average atomic mass for silver is 107.87 amu.

Potassium has three isotopes. The number of neutrons and the natural abundance of these are: 20 neutron (93.23%); 21 neutrons (0.012%); and 22 neutrons (6.73%). Give the chemical notation for each of these isotopes and calculate the average atomic mass for potassium.

PART IIIB: NUCLEAR CHEMISTRY — HALF LIFE PROBLEMS

1. A 2.5 gram sample of an isotope of strontium-90 was formed in a 1960 explosion of an atomic bomb at Johnson Island in the Pacific Test Site. The half-life of strontium-90 is 28 years. In what year will only 0.625 grams of this strontium-90 remain?

2. Actinium-226 has a half-life of 29 hours. If 100 mg of actinium-226 disintegrates over a period of 58 hours, how many mg of actinium-226 will remain?

3. The half-life of isotope X is 2.0 years. How many years would it take for a 4.0 mg sample of X to decay and have only 0.50 mg of it remain?

4. The half-life of Po-218 is three minutes. How much of a 2.0 gram sample remains after 15 minutes? Suppose you wanted to buy some of this isotope, and it required half an hour for it reach you. How much should you order if you need to use 0.10 gram of this material?

In the space below, write the full electron configurations of the following elements:

1. Chlorine   _____

2. Scandium   _____

3. Bromine   _____

In the space below, write the noble gas shorthand electron configurations of the following elements:

1. Chlorine   _____

2. Scandium   _____

3. Bromine   _____

4. Barium   _____

5. Cadmium   _____

Determine what elements are denoted by the following electron configurations:

1) $1s^2 2s^2 2p^6 3s^2 3p^5$ _____

2) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^1$ _____

3) $1s^2 2s^2 2p^3$ _____

4) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$ _____

5) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$ _____

Draw the complete orbital (arrow) diagrams for the flowing elements:

6)   Phosphorus

7)   Iron

Honors Chemistry Worksheet – Wavelength, frequency, & energy of electromagnetic waves.

Show ALL equations, work, units, and significant figures in performing the following calculations.

$C = \lambda v$

$C = 3.00 \times 10^8$ m/s

$E = hv$

$h = 6.626 \times 10^{-34}$ J•s (or J/Hz)

$J = $ Joule $Hz = $ hertz or s^{-1} or $1/s$

1. What is the wavelength of a wave having a frequency of 3.76×10^{14} s^{-1}? What is its energy?

2. What is the frequency of a 6.9×10^{-10} cm wave? What is its energy?

3. What is the frequency of a 7.43×10^{-5} mm wave? What is its energy?

4. What is the wavelength of a wave carrying 8.35×10^{-18} J of energy?

Part IV: Periodic Trends

1. On the blank periodic table, color and label:
   a. alkali metals
   b. alkaline metals
   c. transition metals
   d. nonmetals
   e. metalloids
   f. halogens
   g. noble gases
   h. inner transition metals

2. On the blank periodic table, color and label.
   a. the s block
   b. the p block
   c. the d block
   f. the f block

3. On the blank periodic table, draw arrows to show the following periodic trends across each period and down each group. Be sure to label which way the trend is increasing and which way it is decreasing.

   a. Atomic radius
   b. Ionization energy
   c. Electronegativity

Periodic Table of the Elements

(empty periodic table grid)

Part IV: Periodic Trends Worksheet

Directions: Use your notes to answer the following questions.

1. Rank the following elements by increasing atomic radius: carbon, aluminum, oxygen, potassium.

2. Rank the following elements by increasing electronegativity: sulfur, oxygen, boron, aluminum.

3. Why does fluorine have a higher ionization energy than iodine?

4. Why do elements in the same family generally have similar properties?

5. Indicate whether the following properties increase or decrease from left to right across the periodic table.
   a. atomic radius (excluding noble gases)

   b. first ionization energy

   c. electronegativity

6. What trend in atomic radius occurs down a group on the periodic table? What causes this trend?

7. What trend in ionization energy occurs across a period on the periodic table? What causes this trend?

8. Circle the atom in each pair that has the largest atomic radius.
   a. Al or B   c. Na or Al   e. S or O

   b. O or F   d. Br or Cl   f. Mg or Ca

9. Circle the atom in each pair that has the greater ionization energy.
   a. Li or Be   c. Ca or Ba   e. Na or K

   b. P or Ar   d. Cl or Si   f. Li or K

10. Define electronegativity.

11. Circle the atom in each pair that has the greater electronegativity.
    a. Ca or Ga   c. Br or As   e. Li or O

    b. Ba or Sr   d. Cl or S   c. O or S

Part V: Chemical Bonding

Section 1: Ionic Bonding

Ionic bonds involve a transfer of electrons from one atom (or atomic group) to another. **Cations** are positive ions resulting from the loss of electrons. **Anions** are negative ions resulting from the gain of electrons. Atoms generally lose or gain electrons to achieve a "stable octet" or set of 8 electrons in the valence shell (although there are exceptions!)

Metals tend to have low electronegativity and ionization energy and tend to form cations.

Nonmetals tend to have high electronegativity and tend to form anions.

Things to know:

1. Placement of metals and nonmetals on Periodic Table.
2. The charges/oxidation states taken by elements in different groups of Periodic Table.
3. Common Polyatomic Ions (memorize sulfate/sulfite, carbonate, phosphate/phosphite, permanganate, hydroxide, ammonium, nitrate/nitrite, hypochlorite/chlorite/chlorate/perchlorate – both names and formulas with charges!).

Section 2: Covalent Bonding

Covalent bonds involve a sharing of electrons between atoms. Usually both elements in a covalent bond are nonmetals.

Equal sharing of electrons produces a **nonpolar covalent bond** and occurs when the bonding atoms have equal or very similar electronegativity. Unequal sharing of electrons occurs when atoms have significantly different electronegativities and results in a **polar covalent bond** in which one atom has a partial negative charge and the other a partial positive charge.

Things to know:

1. Be able to determine whether a bond is ionic, polar covalent or nonpolar covalent based on the elements bonding and electronegativity chart.
2. Draw a basic Lewis Dot structure showing the placement of all electrons.

Bonding occurs on a spectrum based on the _difference in electronegativity_ between the two atoms involved in the bond. *Memorize the rules below and have a general sense of the electronegativities of common elements (& how the trend runs along the periodic table)!*

Difference in electronegativity

0	0.5	1.0	2.0	4.0
Nonpolar Covalent	Moderately Polar Covalent	Very Polar-covalent bond	Ionic bond	

Rules of thumb:

$\Delta EN > 2.0 \rightarrow$ Bond is ionic

$\Delta EN < 0.5 \rightarrow$ Bond is nonpolar covalent

$0.5 \leq \Delta EN \leq 1.6 \rightarrow$ Bond is polar covalent

$1.6 < \Delta EN \leq 2.0 \rightarrow$ Bond is polar covalent IF it involves two nonmetals, otherwise ionic.

H 2.1

Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.1	Se 2.5	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9															

Problems!

Bonding between	More electronegative element and value	Less electronegative element and value	Difference in electronegativity	Bond Type
Sulfur & Hydrogen				
Sulfur and cesium				
Chlorine and bromine				
Calcium and chlorine				
Oxygen and hydrogen				
Nitrogen & hydrogen				
Hydrogen & Fluorine				
Carbon and Oxygen				

Electron Dot Structures

The Electron Dot structure gives a *two-dimensional representation* of the molecular structure. The key consideration in drawing a Electron Dot structure is the application of the *octet rule*, which states that a molecule's atoms share electrons so that each is surrounded by eight valence electrons.

The first step in drawing a Electron Dot structure is to determine the *skeletal structure* of the molecule. The skeletal structure shows which atoms are bonded to a <u>central atom</u> using <u>at least</u> a single bond (represented by a dash). The central atom is usually the first atom in the chemical formula for the molecule.

The following rules give an organized method for drawing a valid Electron Dot Structure:

1. Using the column headings in the periodic table, determine the total number of valence electrons in the molecule by adding the valence electrons contributed by each atom (Ex: in CO_2 there should be 16 valence electrons – 4 from the C atom and 12 from the two O atoms). For a polyatomic ion, subtract one electron for each positive charge, and add one electron for each negative charge.

2. Identify the central atom (if two or more elements, central atom is usually the least electronegative). Draw a line-bond structure of the molecule bonding each outer atom to the central atom with a single bond. The line represents two shared electrons (1 covalent bond). You can use two dots to replace a line if that is easier for you.

3. Using a single dot to represent 1 electron, place dots around each atom until each atom has an octet of electrons. Remember that a line-bond represents **2** electrons. **Beware:** Hydrogen only gets 2 electrons (not an octet!).

4. Count the electrons in your Electron Dot structure. If the number of electrons in your diagram matches the total number of electrons from Step 1 → Congrats! You are done.

5. IF your Electron Dot diagram has MORE electrons then your total count in Step 1: Remove electron lone pairs (the dots) and add multiple bonds (the lines) between the central and peripheral atoms until the number of electrons in your diagram matches the number in Step 1. (Removing 1 lone pair from EACH of 2 atoms bonded together can be replaced by one line bond!)

6. IF your Electron Dot diagram has LESS electrons than your total count in Step 1: Add the extra electrons as dots onto the central atom.

7. Exceptions to the octet rule—*Central atoms that are in period 3 or higher can have more than eight valence electrons (a violation of the octet rule). Molecules with B or Al as a central atom (group III) may have a central atom with six valence shell electrons. Molecules with beryllium (Be) as a central atom (group II) may have a central atom with four valence shell electrons.*

Draw Electron Dot structures for the following compounds or polyatomic ions. Draw resonance structures if they exist.

Molecular Formula	Electron Dot Diagram	Molecular Formula	Electron Dot Diagram
CO_3^{2-}		SCl_6	
CS_2		SO_3	
NO_2^-		C_2Cl_4	
SF_4		ICl_5	
XeF_4		BBr_3	

Part VI: Nomenclature of Binary Compounds

** Before you start naming compounds or writing formulas from names be sure to review which elements are metals, transition metals & nonmetals and the charges they take as well as common polyatomic ions with their charges (makes this much easier!)

Part 1: Determine if the compound is ionic or covalent to decide which set of naming rules to apply:

    A. Ionic compound:
        i.   Compound contains a polyatomic ion
        ii.   Compound contains a metal and a nonmetal
    B. Covalent compound:
        i.   Compound contains only nonmetal elements

Part 2: Ionic Compound Nomenclature

    A. Name the cation
        i.   Univalent metal cations = same name as the element
            a. Na^+ = sodium, Ba^{2+} = barium, Al^{3+} = aluminium etc.
            b. These are usually Group 1, 2 and 13 elements

        ii.   Multivalent metal cations = same name as element + charge denoted by Roman Numeral in parenthesis
            a. Fe^{2+} = Iron (II), Fe^{3+} = Iron (III)
            b. Multivalent metal cation are usually in the transition metal block (Iron, Copper, Nickel, Chromium etc.)
            c. Silver is always 1+ (Ag^+) so it has no Roman Numeral
            d. Zinc is always 2+ (Zn^{2+}) so it has no Roman Numeral
            e. An easy way to remember charges for Al, Zn and Ag is noting that they form a diagonal step down starting with Al going down to the left (3+, 2+ and 1+)
            f. Pb and Sn are two metals not in the transition block that can take either the charge 2+ or 4+. As such, Pb and Sn always have a Roman Numeral when being named in a compound.

        iii.   If the cation is a polyatomic ion – it takes the same name as the ion. I.e. NH_4^+ is ammonium.

    B. Name the anion
        i.   Anion that is based on a nonmetal element:
            a. Use the root of the elemental name
            b. Change the suffix to -ide
            c. Cl^- = chloride, O^{2-} = oxide, P^{3-} = phosphide, N^{3-} = nitride etc.
        ii.   Anion that is a polyatomic ion:
            a. Use the name of the polyatomic ion
            b. SO_4^{2-} = sulfate, PO_3^{3-} = phosphite, CrO_4^{2-} = chromate etc.

C. *Examples:*
   $MgCl_2$ = magnesium chlorid
   $FeCl_3$ = iron (III) chloride
   NH_4Cl = ammonium chloride
   $Sn_3(PO_4)_2$ = Tin (II) phosphate
   $(NH_4)_2SO_4$ = ammonium sulfate

Part 3: Covalent Compound Nomenclature

   A. *Name the first element – use Greek Prefixes (except mono)*
      i.   Select the appropriate Greek prefix using subscript of the element
            a. Mono = one
            b. Di = two
            c. Tri = three
            d. Tetra = four
            e. Penta = five
            f. Hexa = six
            g. Hepta = seven
            h. Octa = eight
            i. Nona = nine
            j. Deca = ten
      ii.   Name the first element using the prefix and the element name:
            a. Do not use the prefix mono- for the first element. If there is only one
                atom of the first element in the compound "mono" is implied
   B. *Name the second element*
      i.   Select the appropriate Greek prefix using the subscript of the element
      ii.   Use the root of the element name for the second element
      iii.   Convert the suffix of the elemental name to -ide.

   C. *Examples:*
      H_2O = dihydrogen monoxide (the o from mono- gets dropped in monoxide)
      CO_2 = carbon dioxide
      CO = carbon monoxide
      PCl_5 = phosphorus pentachloride
      S_2O_3 = disulfur trioxide

Names to Formulas of Chemical Compounds

Metals or Polyatomic Ions Involved?

Yes → **Ionic**

No → **Covalent**

Ionic

Example – iron (III) sulfate

1. Use the name to determine the two ions in the compound → Fe and SO_4^{2-}

2. Write the cation first (remember Roman Numeral = charge on metal cation). Then write the anion. Include charges (for now) → $Fe^{3+} SO_4^{2-}$

3. Balance the charges on the two ions to obtain a neutral formula unit. The easy way is to "criss-cross" so that the charge on the cation becomes the subscript of the anion. The charge of the anion becomes the subscript on the cation. Use the lowest whole number ratio of subscripts! →

$$Fe^{3+} SO_4^{2-}$$

(criss-cross shown: $Fe_2^{3+} (SO_4)_3^{2-}$)

4. If the subscript of a polyatomic ion is greater than 1, put the whole polyatomic ion symbol in parentheses and the subscript outside the parenthesis. → $Fe^{3+}_2 (SO_4^{2-})_3$

5. Erase any ion charges in the formula → $Fe_2(SO_4)_3$

Examples: Cation + Monoatomic Anion
sodium fluoride = NaF, calcium bromide = $CaBr_2$, ammonium chloride = $AlCl_3$, iron (II) oxide = FeO, iron (III) oxide = Fe_2O_3

Examples: Cation + Polyatomic Anion
Copper (ii) phosphate = $Cu_3(PO_4)_2$, ammonium carbonate = $(NH_4)_2CO_3$

Covalent

1. 1st Greek prefix denotes subscript of first element

2. Write element symbol and subscript

3. 2nd Greek prefix denotes subscript of second element

4. Write symbol and subscript for second element

Examples:
carbon monoxide = CO
dinitrogen tetraoxide = N_2O_4
sulfur hexafluoride = SF_6
dihydrogen monoxide = H_2O
dihydrogen dioxide = H_2O_2
carbon tetrahydride = CH_4

Naming Binary Chemical Compounds

Metals or Polyatomic Ions Involved?

Yes → **Ionic**

No → **Covalent**

Ionic

Monovalent Cation
1. Name cation by element name
Ex. Na⁺ = Sodium,
Ca²⁺ = Calcium,
Ag⁺ = Silver

Na^+ = Sodium, Ca^{2+} = Calcium, Ag^+ = Silver

Multivalent Cation
1. Name cation by element name
→ Use Roman Numeral in parentheses to denote charge
Ex. Fe^{2+} = Iron (II),
Fe^{3+} = Iron (III)

Polyatomic Cation
1. Name cation by name of polyatomic cation
Ex. Ammonium

Monoatomic Anion
2. Name anion by element root.
3. Change suffix to -ide

Polyatomic Anion
2. Name anion by polyatomic anion name

Examples: Cation + Monoatomic Anion
sodium fluoride, calcium bromide, ammonium chloride, iron (II) oxide
Examples: Cation + Polyatomic Anion
sodium phosphate, ammonium carbonate, copper (II) sulfate

Covalent

1. 1ˢᵗ Greek prefix (don't use mono)
2. Name first element
3. 2ⁿᵈ Greek prefix
4. Root of 2ⁿᵈ element
5. Change suffix to –ide

Examples:
carbon monoxide
dinitrogen tetraoxide
phosphorus pentachloride
sulfur hexafluoride
dihydrogen monoxide
dihydrogen dioxide

AP Chemistry Worksheet 3: *Naming Inorganic Compounds*

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. Give the name for each of the following ionic compounds:

    a. AlF_3

    b. $Fe(OH)_2$

    c. $Cu(NO_3)_2$

    d. $Ba(ClO_4)_2$

    e. Li_3PO_4

    f. Hg_2S

    g. $Ca(C_2H_3O_2)_2$

    h. $Cr_2(CO_3)_3$

    i. K_2CrO_4

    j. $(NH_4)_2SO_4$

2. Write the chemical formula for each of the following compounds:

    a. copper (I) oxide

    b. potassium peroxide

    c. aluminum hydroxide

    d. zinc nitrate

    e. mercury (I) bromide

    f. iron (III) carbonate

    g. sodium hypobromite

3. Give the name or chemical formula, as appropriate, for each of the following acids:

    a. $HBrO_3$

    b. HBr

    c. H_3PO_4

    d. hypochlorous acid

    e. iodic acid

    f. sulfurous acid

Part VI: Problems - More Naming Practice!   **Acid, Ionic or Covalent?**

vanadium (V) phosphate _____   _____

sodium permanganate _____   _____

MnF_2 _____   _____

$Ni(SO_3)_2$ _____   _____

phosphorus triiodide _____   _____

H_3PO_4 _____   _____

HI _____   _____

Pb_3N_4 _____   _____

$Sn(OH)_2$ _____   _____

$SiCl_4$ _____   _____

$HClO_2$ _____   _____

Sodium sulfate _____   _____

Hydrosulfuric acid _____   _____

Nitrogen trifluoride _____   _____

Calcium phosphide _____   _____

B_2Si _____   _____

PCl_5 _____   _____

Percloric acid _____   _____

Manganese (IV) carbonate _____   _____

C_6H_{10} _____   _____

Carbon disulfide _____   _____

Iron (III) nitrate _____   _____

Copper (II) phosphite _____   _____

Sulfur hexachloride _____   _____

Part VII: Mole Conversions Notes & Practice Worksheet

There are three mole equalities. They are:

1 mol = 6.02 x 10^{23} particles

1 mol = molar mass in grams (periodic table)

1 mol = 22.4 L for a gas at STP

Each equality can be written as a set of two conversion factors. They are:

$$\left(\frac{1\ mole}{6.02\times 10^{23}\ particles}\right) or \left(\frac{6.02\times 10^{23}\ particles}{1\ mole}\right) \qquad \left(\frac{1\ mole}{molar\ mass\ in\ grams}\right) or \left(\frac{molar\ mass\ in\ grams}{1\ mole}\right)$$

$$\left(\frac{22.4\ L}{1\ mole}\right) or \left(\frac{1\ mole}{22.4\ L}\right)$$ at Standard Temperature and Pressure (0°C and 1 atm)

Example Problems:

1. How many moles of magnesium is 3.01 x 10^{22} atoms of magnesium?

$$3.01\ x\ 10^{22}\ atoms\left(\frac{1\ mole}{6.02 x 10^{23}\ atoms}\right) = 5\ x\ 10^{-2}\ moles$$

2. How many molecules are there in 4.00 moles of glucose, $C_6H_{12}O_6$?

$$4.0\ moles\left(\frac{6.02 x 10^{23}\ molecules}{1\ mole}\right) = 2.41\ x\ 10^{24}\ molecules$$

3. How many moles in 28 grams of CO_2 ?

Molar mass of CO_2   1 C = 1 x 12.01 g = 12.01 g

2 O = 2 x 16.00 g = <u>32.00 g</u>

44.00 g/mol

$$28\ g\ CO_2\left(\frac{1\ mole}{44.00\ g}\right) = 0.64\ moles\ CO_2$$

4. What is the mass of 5 moles of Fe_2O_3 ?

Molar mass Fe_2O_3   2 Fe = 2 x 55.6 g = 111.2 g

3 O = 3 x 16.0 g = <u>48.0 g</u>

159.2 g/mol

$$5\ moles\ Fe_2O_3\left(\frac{159.2\ g}{1\ mole}\right) = 800\ grams\ Fe_2O_3$$

5. Determine the volume, in liters, occupied by 0.030 moles of a gas at STP.

$$0.030\ mol\left(\frac{22.4\ L}{1\ mole}\right) = 0.67\ L$$

Mixed Mole Conversion Examples: **Given unit → Moles → Desired unit**

7. How many oxygen molecules are in 3.36 L of oxygen gas at STP?

$$3.36 \, L\left(\frac{1 \, mole}{22.4 \, L}\right)\left(\frac{6.02x10^{23} \, molecules}{1 \, mole}\right) = 9.03 \times 10^{22} \text{ molecules}$$

8. Find the mass in grams of 2.00×10^{23} molecules of F_2

Molar mass   2 F = 2 x 19 g = 38 g/mol

$$2.00 \times 10^{23} \text{ molecules}\left(\frac{1 \, mole}{6.02x10^{23} \, particles}\right)\left(\frac{38 \, g}{1 \, mole}\right) = \textbf{12.6 g}$$

****

Problems I: Mole Conversions Practice – Show Work

1. How many moles are 1.20×10^{25} atoms of phosphorous?

2. How many molecules are in 4.50 grams of N_2O_5?

3. What is the volume of 42.8 grams of water vapor at STP?

4. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (table sugar) when dissolved in water. It is marketed by G.D. Searle as *Nutra Sweet*. The molecular formula of aspartame is $C_{14}H_{18}N_2O_5$.

   a) Calculate the gram molar mass of aspartame.

   b) How many moles of molecules are in 10 g of aspartame?

   c) How many molecules are in 5 **mg** of aspartame?

   d) How many atoms of nitrogen are in 1.2 grams of aspartame?

AP Chemistry Worksheet 6: *The Mole*

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. The molecular formula of aspartame, the artificial sweetener marketed as NutraSweet, is $C_{14}H_{18}N_2O_5$.
   a. What is the molar mass of aspartame?

   b. How many moles of aspartame are present in 1.00 mg of aspartame? (1000 mg = 1g)

   c. How many molecules of aspartame are present in 1.00 mg of aspartame?

   d. How many hydrogen atoms are present in 1.00 mg of aspartame?

2. A sample of glucose, $C_6H_{12}O_6$, contains 2.03×10^{21} atoms of carbon.

   a. How many atoms of hydrogen does it contain?

   b. How many molecules of glucose does it contain?

   c. How many moles of glucose does it contain?

d. What is the mass of the sample in grams?

3. Calculate the following amounts:

   a. How many moles of chloride ions are in 0.0750 g of magnesium chloride?

   b. What is the mass, in grams, of 3.50×10^{-3} mol of aluminum sulfate?

   c. What is the mass, in grams, of 1.75×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$?

   d. What is the molar mass of cholesterol if 0.00105 mol weigh 0.406 g?

4. Calculate the number of molecules in:
   a. 0.0666 mol propane, C_3H_8, a hydrocarbon fuel

   b. A 50.0 mg tablet of acetaminophen, $C_8H_9O_2N$, an analgesic solid under the name of Tylenol

c. A tablespoon of table sugar, $C_{12}H_{22}O_{11}$, weighing 10.5 g

5. The allowable concentration level of vinyl chloride, C_2H_3Cl, in the atmosphere in a chemical plant is 2.0×10^{-6} g/L.
   a. How many moles of vinyl chloride in each liter does this represent?

   b. How many molecules per liter is this?

Chemical Reactions Review Sheet

Types of Chemical Reactions:

Combination or Synthesis $A + B \rightarrow AB$
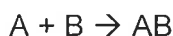
Decomposition $AB \rightarrow A + B$

Single Replacement $A + BC \rightarrow B + AC$
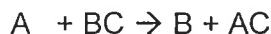
Double Replacement $AB + CD \rightarrow AD + CB$
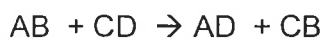
      Can be   a) acid-base if the reactants are acid & base and products are salt & water.

                  b) can be precipitation if a solid product forms

Hydrocarbon Combustion $C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$
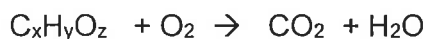
Oxidation-Reduction - Involve a transfer of electrons. Occurs during combustion, single replacement and can occur during synthesis and decomposition.

Problems:

1. A reaction occurs when aqueous lead (II) nitrate is mixed with an aqueous solution of potassium hydroxide. Write an overall, balanced equation for the reaction, including state designations.

2. For the following three reactions, label the type, predict the products (make sure formulas are correct), and balance the equation.

_____ a. _____C_3H_4 (g) + _____O_2 (g) $\rightarrow$

_____ b. _____ $Ba(NO_3)_2$(aq) + _____ Na_3PO_4(aq) $\rightarrow$

_____ c. _____ Al(s) + _____ O_2(g) $\rightarrow$

_____ d. _____ HBr(aq) + _____ KOH(aq) $\rightarrow$

_____ e. _____ $Ca(NO_3)_2$ (aq) + _____ Na_2SO_4 (aq) $\rightarrow$
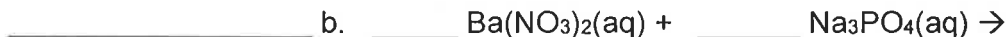
3. In the following equations, label the oxidized element and the reduced element.

    a. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

    b. $2NaBr(aq) + Cl_2(g) \rightarrow 2NaCl(s) + Br_2(l)$
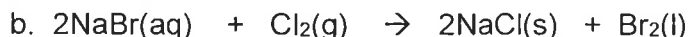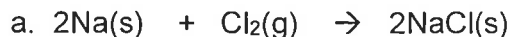
Table 11.2
Activity Series of Metals

	Name	Symbol
Decreasing reactivity ↓	Lithium	Li
	Potassium	K
	Calcium	Ca
	Sodium	Na
	Magnesium	Mg
	Aluminum	Al
	Zinc	Zn
	Iron	Fe
	Lead	Pb
	(Hydrogen)	(H)*
	Copper	Cu
	Mercury	Hg
	Silver	Ag

*Metals from Li to Na will replace H from acids and water; from Mg to Pb they will replace H from acids only.

Table 11.3
Solubility Rules for Ionic Compounds

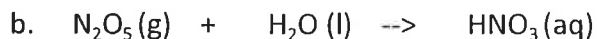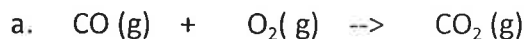| Compounds	Solubility	Exceptions
Salts of alkali metals and ammonia	Soluble	Some lithium compounds
Nitrate salts and chlorate salts	Soluble	Few exceptions
Sulfate salts	Soluble	Compounds of Pb, Ag, Hg, Ba, Sr, and Ca
Chloride salts	Soluble	Compounds of Ag and some compounds of Hg and Pb
Carbonates, phosphates, chromates, sulfides, and hydroxides	Most are insoluble	Compounds of the alkali metals and of ammonia

Reaction Review

1. What are 4 signs that a reaction is taking place? Think back to the lab:

2. What is does it mean when a substance is reduced? When it is oxidized? How is a single replacement reaction an oxidation-reduction reaction?

3. What are the 5 main types of chemical reactions? What type of reaction is an acid-base neutralization?

4. What does (s), (g), (l) and (aq) mean when placed near a chemical formula in an equation?

A) WRITE THE FORMULA FOR EACH MATERIAL CORRECTLY.
B) BALANCE THE EQUATION. SOME REACTIONS REQUIRE COMPLETION.
C) FOR EACH REACTION TELL WHAT TYPE OF REACTION IT IS.
D) For double and single replacement reactions – write the net ionic equations.

1. lead II nitrate and sodium iodide react to make lead iodide and sodium nitrate.

2. calcium carbonate decomposes when you heat it to leave calcium oxide and carbon dioxide.

3. ammonia gas when it is pressurized into water will make ammonium hydroxide.

4. aluminum hydroxide and sulfuric acid neutralize to make water and aluminum sulfate.

5. tetracarbon octahydride is burned in oxygen

6. sulfuric acid reacts with zinc

AP Chemistry Worksheet 5: *Balancing Equations & Patterns of Reactivity*

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. Balance the following equations:

    a. $CO (g) + O_2 (g) \rightarrow CO_2 (g)$

    b. $N_2O_5 (g) + H_2O (l) \rightarrow HNO_3 (aq)$

    c. $PCl_5 (l) + H_2O (l) \rightarrow H_3PO_4 (aq) + HCl (aq)$

    d. $CH_4 (g) + Br_2 (g) \rightarrow CBr_4 (l) + HBr (g)$

    e. $C_5H_{10}O_2 (l) + O_2 (g) \rightarrow CO_2 (g) + H_2O (l)$

    f. $Cr(OH)_3 (s) + HClO_4 (aq) \rightarrow Cr(ClO_4)_3 (aq) + H_2O (l)$
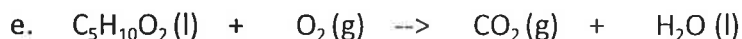
2. Write balanced chemical equations to correspond to each of the following descriptions:

    a. Solid calcium carbide, CaC_2, reacts with water to form an aqueous solution of calcium hydroxide and acetylene gas, C_2H_2.

    b. When solid potassium chlorate is heated, it decomposes to form solid potassium chloride and oxygen gas.

    c. Solid zinc metal reacts with sulfuric acid to form hydrogen gas and an aqueous solution of zinc sulfate.

    d. When liquid phosphorous trichloride is added to water, it reacts to form a solution of phosphorous acid and hydrochloric acid.

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e. When hydrogen sulfide gas is passes over solid hot iron (III) hydroxide, the reaction produces solid iron (III) sulfide and gaseous water.

3. a. What products form when a hydrocarbon is completely combusted in air?

   b. Write a balanced chemical equation for the combustion of octane, C_8H_{18} (l), in air.

   c. How can you determine the chemical formula of the product formed when the metallic element calcium combines with the nonmetallic element oxygen, O_2?

   d. Write the balanced chemical equation for the reaction described in (c).

4. Write a balanced chemical equation for the reaction that occurs when

   a. the hydrocarbon heptane, C_7H_{16} (l), is combusted in air

   b. the gasoline additive MTBE (methyl tertiary-butyl ether), $C_5H_{12}O$ (l), burns in air

   c. Rb (s) reacts with water

   d. Mg(s) reacts with Cl_2 (g)

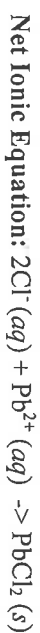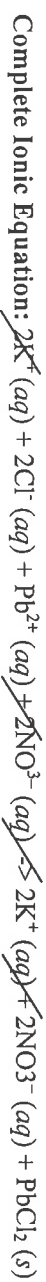5. Balance the following equations, and indicate what type of reaction each one is:

   a. $Al (s) + Cl_2 (g) \rightarrow AlCl_3 (s)$ _____

   b. $C_2H_4 (g) + O_2 (g) \rightarrow CO_2 (g) + H_2O (l)$ _____

   c. $Li(s) + N_2 (g) \rightarrow Li_3N (s)$ _____

   d. $PbCO_3 (s) \rightarrow PbO (s) + CO_2 (g)$ _____

   e. ___ $C_7H_8O_2 (l) + O_2 (g) \rightarrow CO_2 (g) + H_2O (l)$ _____
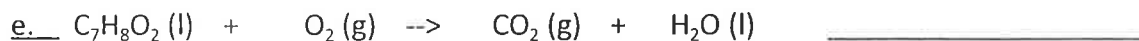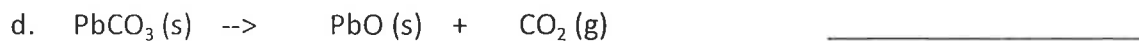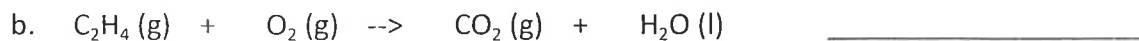
Net Ionic Equation Worksheet

READ THIS: When two solutions of ionic compounds are mixed, a solid may form. This type of reaction is called a **precipitation reaction**, and the solid produced in the reaction is known as the **precipitate**. You can predict whether a precipitate will form using a list of solubility rules such as those found in the table below. When a combination of ions is described as insoluble, a precipitate forms. There are three types of equations that are commonly written to describe a precipitation reaction. The **molecular equation** shows each of the substances in the reaction as compounds with physical states written next to the chemical formulas. The **complete ionic equation** shows each of the **aqueous** compounds as separate ions. Insoluble substances are not separated and these have the symbol (s) written next to them. Water is also not separated and it has a (l) written next to it. Notice that there are ions that are present on both sides of the reaction arrow -> that is, they do not react. These ions are known as **spectator ions** and they are eliminated from complete ionic equation by crossing them out. The remaining equation is known as the **net ionic equation.**

For example: The reaction of potassium chloride and lead II nitrate

Molecular Equation: $2KCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow 2KNO_3\ (aq) + PbCl_2\ (s)$

Complete Ionic Equation: $\cancel{2K^+\ (aq)} + 2Cl^-\ (aq) + Pb^{2+}\ (aq) \cancel{+ 2NO_3^-\ (aq)} \rightarrow \cancel{2K^+\ (aq)} \cancel{+ 2NO3^-\ (aq)} + PbCl_2\ (s)$

Net Ionic Equation: $2Cl^-\ (aq) + Pb^{2+}\ (aq) \rightarrow PbCl_2\ (s)$

Directions: Write balanced molecular, ionic, and net ionic equations for each of the following reactions. Assume all reactions occur in aqueous solution. Include states of matter in your balanced equation.

1. Sodium chloride and lead II nitrate

   Molecular Equation:

   Net Ionic Equation:

2. Sodium carbonate and Iron II chloride

   Molecular Equation:

   Net Ionic Equation:

3. Ammonium phosphate and zinc nitrate

   Molecular Equation:

Net Ionic Equation:

4. Iron III chloride and magnesium metal

Molecular Equation:

Net Ionic Equation:

5. Silver nitrate and magnesium iodide

Molecular Equation:

Net Ionic Equation:

6. Aluminum and copper (II) perchlorate

Molecular Equation:

Net Ionic Equation:

7. Sodium and water

Molecular Equation:

Net Ionic Equation:

8. Zinc and hydrochloric acid

Molecular Equation:

Net Ionic Equation:

Steps to Find Empirical & Molecular Formulas

Remember this:

"Percent to mass, Mass to mole,

Divide by small, Make it whole"

1. Determine the mass in grams of each element present in the sample. **"Percent to mass"**

    If the information in the problem is in terms of percent composition of each element →
    > a) assume you have 100 g of the sample to start with

    > b) The grams of each element (out of the 100 g sample) will just be the numerical value of its percent composition.

    EXAMPLE: You have a sample that is 40.0% carbon, 6.73% hydrogen and the rest oxygen. Find the empirical and molecular formulas.

    Step 1: $40.0\% + 6.73\% = 46.73\%$. The percentage of oxygen is $100\% - 46.73\% = 53.27\%$

    If I have 100 g of sample to start with, I have:

    > 40.0 grams Carbon, 6.73 grams Hydrogen and 53.27 grams Oxygen

2. Calculate the number of *moles* of each element. **"Mass to mole"**

    Step 2: Moles of Carbon = 40.0g C x 1 mol C/12.01g C = 3.331 mol C

    Moles Hydrogen = 6.73g H x 1 mol H/1.01g = 6.663 mol H

    Mole Oxygen = 53.27 g O x 1 mol O/16.0 g = 3.33 mol O

    DO NOT ROUND THESE NUMBERS → KEEP SEVERAL DECIMAL PLACES

3. Divide each by the smallest number of moles to obtain the *simplest whole number ratio*.

    "Divide by small"

    Step 3: The molar ratio of the elements in my compound is $C_{3.331}H_{6.663}O_{3.33}$. I want a whole number ratio, so I will divide all the subscripts by the smallest number of moles (3.331) to get:

    > $C_1H_2O_1$ → so my empirical formula is CH_2O

    If your number after dividing are values like 2.07, 1.1 etc. then round to the nearest whole number. If they are values like 3.5, 2.333 etc., then go to step 4.

4. If whole numbers are not obtained* in step 3), multiply through by the smallest integer that will give all whole numbers

> **"Make it whole"**
>
> Let's say that my empirical formula turned out to be $C_{2.333}H_4O_2$. 2.333 is not close enough to 2 to round down to 2. But I can multiply my formula through by 3 to get this:
>
> $$C_7H_{12}O_6$$

5. **Finding molecular formula:** If the molar mass of your empirical formula matches the molar mass of the final compound (as stated in the problem) → Hooray! You are done: your empirical formula IS your molecular formula.

> Step 5: For my example in step 1, it says that the molecular weight (molar mass) of my compound is 180.18 g/mol
>
> My empirical formula is CH_2O from step 3 has a molar mass of $(12.01 + 2 \times 1.01 + 16)$ g/mol = 30.03 g/mol. *So my empirical formula is not my molecular formula.*
>
> Now, divide molar mass of compound/molar mass of empirical formula:
>
> > 180.18 g/mol ÷ 30.03 g/mol = 6
>
> The molar mass of my compound is 6 times the molar mass of my empirical formula.
>
> Multiply the empirical formula subscripts by 6 to get the final molecular formula:
>
> > $6(CH_2O) = \mathbf{C_6H_{12}O_6}$ → The compound in my sample is glucose!

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. Determine the empirical formula of each of the following compounds if a sample contains
    a. 0.104 mol K, 0.052 mol C, and 0.156 mol O

    b. 5.28 g Sn and 3.37 g F

    c. 87.5 percent N and 12.5 percent H by mass

2. Determine the empirical formulas of the compounds with the following compositions by mass
    a. 10.4 percent C, 27.8 percent S, and 61.7 percent Cl

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

3. What is the molecular formula of each of the following compounds?
   a. empirical formula CH_2, molar mass = 84 g/mol

   b. empirical formula NH_2Cl, molar mass = 51.5 g/mol

4. Determine the empirical and molecular formulas of each of the following substances:
   a. Ibuprofen, a headache remedy contains 75.69 percent C, 8.80 percent H, and 15.51 percent O by mass; molar mass about 206 g

   b. Benzene contains only carbon and hydrogen and is 7.74% hydrogen by mass. The molar mass of benzene is 78.1 g/mol.

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5. Many homes in rural America are heated by propane gas, a compound that contains only carbon and hydrogen. Complete combustion of a sample of propane produced 2.641 g of carbon dioxide and 1.442 g of water as the only products. Find the empirical formula of propane. (Hint: Figure out how many moles of C and H were produced. They all came from the fuel.)

6. (This is probably the hardest problem in the whole packet!) Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005 g sample of menthol is combusted, producing 0.2829 g of CO_2 and 0.1159 g of H_2O.
   a. What is the empirical formula for menthol?

   b. If the compound has a molar mass of 156 g/mol, what is its molecular formula?

Steps to Solving Limiting Reagent Problems

Suppose 13.7 g of C_2H_2 reacts with 18.5 g O_2 according to the reaction below. What is the mass of CO_2 produced? What is the limiting reagent?

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(\ell)$$
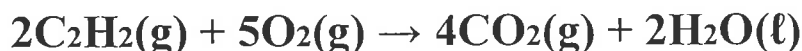
1. Find the mass of product yielded by the given amount of the first reactant. You can use either product (CO_2 or H_2O), but since the question asks about CO_2, it will be easier to use this product:

13.7 g C_2H_2	1 mole C_2H_2	4 mole CO_2	44.02 g CO_2	=	46.3 g CO
	26.04 g C_2H_2	2 mole C_2H_2	1 mole CO_2		

2. Find the mass of *the same product* (in this case CO_2) yielded by the given amount of the second reactant.

18.5 g O_2	1 mole O_2	4 mole CO_2	44.02 g CO_2	=	20.4 g CO_2
	32.00 g O_2	5 mole O_2	1 mole CO_2		

3. Since the 18.5 grams of O_2 produces *less CO_2*, it is the *limiting reagent* in this problem. This amount of O_2 gets used up first and "limits" how much CO_2 can be produced. The amount of CO_2 that can be produced is 20.4 grams (which you already calculated!)

4. You can repeat steps 1 and 2 for any number of reactants that you have a given mass for. The limiting reagent will ALWAYS be the *reactant that produces the least amount of product* (because it gets used up first).

5. **Finding the amount of excess reagent:** The excess reagent is the one that is NOT the limiting reagent. There will be some of this reagent leftover after the limiting reagent is completely used up.

   Figure out how much of the excess reagent must react completely with the given amount of the limiting reagent. Then subtract this amount from the given amount of the excess reagent.

18.5 g O_2	1 mole O_2	2 mole C_2H_2	26.02 g C_2H_2	= 6.02 g C_2H_2 used
	32.00 g O_2	5 mole O_2	1 mole C_2H_2	

13.7 g of C_2H_2 total – 6.02 g of C_2H_2 used = | 7.68 g C_2H_2 excess (leftover) |

Sharon Karackattu

Part VIII: Stoichiometry-Based Problems

1.   a) Nicotine is a stimulant and an addictive chemical found in tobacco. An analysis of nicotine produces the following percent composition: 74.03% carbon, 17.27% nitrogen, and 8.70% hydrogen. What is the empirical formula of nicotine?

     b) Further tests show that the molar mass of nicotine is 162.23 g/mol. Given this information, what is the molecular formula of nicotine?

2.   An ionic sample with a mass of 0.5000 g is determined to contain the elements indium and chlorine. If the sample has 0.2404 g of chlorine, what is the empirical formula of this ionic compound?

3. A 16.4 g sample of hydrated calcium sulfate is heated until all the water is driven off. The calcium sulfate that remains has a mass of 13.0 g. Find the formula and the chemical name of the hydrate.

4.   ____C_3H_8 + ____O_2 →

     a. What type of reaction is written above? _____

     b. Predict the products of the reaction and balance it.

     c. If I start with 5.00 grams of C_3H_8 and 5.00 grams of O_2, what is the limiting reagent? What is my theoretical yield of the carbon containing product?

d. I get a percent yield of 75%. How many grams of the carbon containing product did I make?

5. Magnesium undergoes a single replacement reaction with hydrochloric acid.
a) *Write the Balanced Equation:*

b) Which element is oxidized? _____ Which element is reduced? _____

c) How many grams of hydrogen gas can be produced from the reaction of 3.00 g of magnesium with 4.00 g of hydrochloric acid?

d) Identify the limiting and excess reactants. How many grams of the excess reagent are leftover?

e) If the hydrogen gas is produced at 48°C and 2.5 atm of pressure, what is the volume produced in liters?

6. Sulfur reacts with oxygen to produce sulfur trioxide gas.
a) *Write the Balanced Equation:*

b) If 6.3 g of sulfur reacts with 10.0 g of oxygen, what is the theoretical yield of sulfur trioxide gas in grams?

c) What is the limiting reagent? How many grams of the excess reagent is leftover?

d) The sulfur trioxide gas produced had a volume of 5.4 L and was produced at 98°C. What is the pressure of the gas in kPa?

AP Chemistry Worksheet 8: *Stoichiometry*

Prior to solving worksheet 8 and 9 - watch videos using the links below. Type the link (you will not be able to click on it) **in chrome or safari.**

Youtube links:
Part 1 - http://youtu.be/VHX9MauM8sM
Part 2 - http://youtu.be/Oaw9yREEs2Y
Part 2a - http://youtu.be/JxELYaiKHss
Part 3 - http://youtu.be/vJX1f1YTiCg
Part 3a - http://youtu.be/sGZ10SfzE3I
Part 4 - http://youtu.be/hnzu6kLNIpA
Part 4a - http://youtu.be/SC5_kvXPXAw
Part 5 - http://youtu.be/2vofw_6VsOk
Part 5a - http://youtu.be/x71ViHWvH9M

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. Why is it essential to use balanced chemical equations in solving stoichiometry problems?

2. The fermentation of glucose, $C_6H_{12}O_6$, produces ethyl alcohol, C_2H_5OH, and CO_2 as shown here:

$$C_6H_{12}O_6 \text{ (aq)} \rightarrow 2\, C_2H_5OH\text{(aq)} + 2\, CO_2\text{(g)}$$
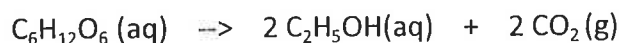
a. How many moles of CO_2 are produced when 0.300 mol of $C_6H_{12}O_6$ reacts in this fashion?

b. How many grams of $C_6H_{12}O_6$ are needed to form 2.00 g of C_2H_5OH?

c. How many molecules of CO_2 form when 2.00 g of C_2H_5OH are produced?

3. Aluminum sulfide reacts with water to form aluminum hydroxide and hydrogen sulfide.
   a. Write the balanced chemical equation for this reaction.

   b. How many grams of aluminum hydroxide are obtained from 10.5 g of aluminum sulfide?

4. Automotive air bags inflate when sodium azide, NaN_3, rapidly decomposes to its component elements:

$$2\ NaN_3\ (s)\ \text{-->}\ Na(s)\ +\ 3\ N_2\,(g)$$

   a. How many moles of N_2 are produced by the decomposition of 1.50 moles of NaN_3?

b. How many grams of NaN$_3$ are required to form 5.00 g of nitrogen gas?

c. How many grams of NaN$_3$ are required to produce 10.0 L of nitrogen gas if the gas has a density of 1.25 g/L?

5. A piece of aluminum foil 0.550 mm thick and 1.00 cm square is allowed to react with bromine to form aluminum bromide.
   a. How many moles of aluminum were used? (The density of aluminum is 2.699 g/cm^3.)

   b. How many grams of aluminum bromide form, assuming that the aluminum reacts completely?

AP Chemistry Worksheet 9: *Limiting Reactants & Theoretical Yield*

For each problem below, write the equation and show your work. Always use units and box in your final answer.

1. A manufacturer of bicycles has 50 wheels, 30 frames, and 24 seats.
   a. How many bicycles can be manufactured using these parts?

   b. How many parts of each kind are left over?

   c. Which part is like a limiting reactant in that it limits the production of bicycles?

2. The fizz produced when an Alka-Seltzer tablet is dissolved in water is due to the reaction between sodium
   bicarbonate, $NaHCO_3$, and citric acid, $H_3C_6H_5O_7$:

$$3\ NaHCO_3\ (aq)\ +\ H_3C_6H_5O_7\ (aq)\ -->\ 3\ CO_2\ (g)\ +\ 3\ H_2O(l)\ +\ Na_3C_6H_5O_7\ (aq)$$
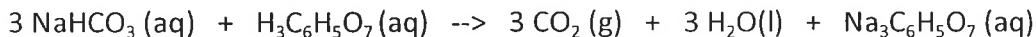
In a certain experiment 1.00 g of sodium bicarbonate and 1.00 g of citric acid are allowed to react.

    a. Which reactant is the limiting reactant? You must show work to support your answer.

    b. How many grams of carbon dioxide form?

    c. How much of the limiting reactant is left when the reaction is complete?

    d. How much of the excess reactant remains after the reaction is complete?

3. When hydrogen sulfide gas is bubbled into a solution of sodium hydroxide, the reaction forms sodium sulfide and water. How many grams of sodium sulfide are formed if 2.50 g of hydrogen sulfide is bubbled into a solution containing 1.85 g of sodium hydroxide, assuming that the limiting reagent is completely consumed?

4. Solutions of sulfuric acid and lead (II) acetate react to form solid lead (II) sulfate and a solution of acetic acid. If 10.0 g of sulfuric acid and 10.0 g of lead (II) acetate are mixed, calculate the number of grams of sulfuric acid, lead (II) acetate, lead (II) sulfate, and acetic acid present in the mixture after the reaction is complete.

5. A student reacts benzene, C_6H_6, with bromine, Br_2, to prepare bromobenzene, C_6H_5Br, and HBr.

   a. What is the theoretical yield of bromobenzene in this reaction when 30.0 g of benzene reacts with 65.0 g of bromine?

   b. If the actual yield of bromobenzene was 56.7 g, what was the percent yield?

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AP Chemistry Worksheet 10: Gases

1. What are the postulates of Kinetic Molecular Theory (KMT)

2. Convert following pressure units:

    a. 3800 torr to atm

    b. 800 mmHg to Pa

    c. 2.7 atm to mmHg and KPa

    d. 55.3 Psi to atm

    e. 103.74 KPa to mmHg

3. Explain the following terms STP and Absolute temperature. Provide significance of both.

4. A gas is stored in a closed container at 2.7 atm. The container occupies the volume of 289 ml. If the gas is released out of the container into a new container with a volume of 900 ml, calculate the new pressure of the gas in atm.

5. A gas has a volume of 350 ml at 45°C. If the volume changes to 400 ml, what is the new temperature?

6. Maybelline Cousteau's backup oxygen tank reads 900 mmHg while on her boat, where the temperature is 27°C. When she dives down to the bottom of an unexplored methane lake on a recently-discovered moon of Neptune, the temperature will drop down to –183°C. What will the pressure in her backup tank be at that temperature?

7. 400 ml of a gas is contained at 300 mm Hg and 0°C. What will its volume be in mL at 140 mm Hg and 10°C?

8. 100 g of oxygen(O_2) is added to the gas chamber. What is the volume of the gas at STP.

9. 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?

Part IX: Gas Laws, Molarity, pH and Putting it all Together

1. The following questions pertain to the reaction below:

    _____HBr +   _____ Ca →
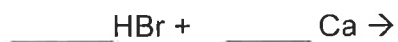
    a.   What type of reaction is shown above? _____

    b. Predict products and then balance the reaction.

    c. Name the ionic product of the reaction. _____

    d. Which element is oxidized? _____ Which element is reduced? _____

    e. 1.7 grams of Ca are mixed with 850.6 mL of 0.043 M HBr. What is the maximum theoretical yield of the gaseous product in grams?

    f. How many grams of the excess reagent are leftover?

    g. What is the pH of the HBr solution?

    h. What is the OH⁻ concentration of the HBr solution?

    i. If the gas is produced at 89°C and 1.7 atm of pressure, what is the volume of gaseous product in mL?

    j. The pressure of the gas is changed to 250 mmHg and the volume is changed to 1.54 L. What is the temperature of the gas now?

Question 2: The following questions pertain to the reaction below

_____ $H_3PO_{4(aq)}$ + _____ $Ca(OH)_{2(aq)}$ →

a) What type of reaction is shown above? _____ (HINT:
It could be two of the types we learned about because one product is insoluble –
which one? _____).

b) Predict the products and balance the reaction.

c) Write the net ionic reaction for the reaction above.

d) Name the reactants and products. Identify acid, base, conjugate acid and conjugate
base.

e) If I have 7.62 grams of $Ca(OH)_2$, what volume of 0.050 M H_3PO_4 would be required to
react with it completely?

f) In the reaction, only 6.89 grams of the solid product were produced. What is the
percent yield of the reaction?

g) How many grams of the $Ca(OH)_2$ remained unreacted?

Question 3:

It takes combustion of 58.8 mL of liquid propane (C_3H_8), which has a density of 0.493 g/cm^3, to cook my hamburger. If air is 21.0% by volume O_2, how many liters of air at 27.0°C and 105.0 kPa will it take to cook my burger? (NOTE: this is not happening at STP!)

a) Write and balance the combustion reaction for propane

b) Calculate the grams of propane used to cook the burger

c) Calculate the moles of oxygen used to cook the burger

d) Calculate the volume of O_2 used to cook the burger

e) Calculate the volume of air used to cook the burger