DO NOT DETACH FROM BOOK.

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** (97)	44 **Ru** 101.1	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **†Ac** (227)	104 **Rf** (267)	105 **Db** (270)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (282)	112 **Cn** (285)	113 **Uut** (285)	114 **Fl** (289)	115 **Uup** (288)	116 **Lv** (293)	117 **Uus** (294)	118 **Uuo** (294)

*Lanthanoid Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
†Actinoid Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

AP® CHEMISTRY EQUATIONS AND CONSTANTS

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

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

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

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

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

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

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

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

$$= K_a \times K_b$$

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

$$14 = pH + pOH$$

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

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

Equilibrium Constants

K_c (molar concentrations)

K_p (gas pressures)

K_a (weak acid)

K_b (weak base)

K_w (water)

KINETICS

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

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

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

k = rate constant

t = time

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

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

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

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

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

$$K = °C + 273$$

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

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = 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

$$q = mc\Delta T$$

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

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

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

$$\Delta G° = \Delta H° - T\Delta S°$$
$$= -RT \ln K$$
$$= -nFE°$$

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

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)

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, and 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 ≤ x < 10)*
Examples:

Number	**3.0 x 10^{3}**	**5.998 x 10^{5}**	**6.00000 x 10^{-23}**	**0.5 x 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:

$2.3 \times 5.78 = \textit{Calculator returns } 13.294$

$\quad 2.3 \textit{ has 2 sig.fig}$

$\quad 5.78 \textit{ has 3 sig.fig.}$

$2.3 \times 5.78 = 13 \quad \textit{The answer must be rounded to show 2 sig.fig}$

Example 2.

$$\frac{1.67x10^5 \times 0.00045}{2x10^{-23}} = \textit{calculator returns } 2.505000000x10^{24}$$

$\quad 1.67x10^5 \textit{ has 3 sig.figs}$

$\quad \textit{0.00045 has 2 sig.figs}$

$\quad \textit{2x10}^{-23} \textit{ has 1 sig.fig}$

$$\frac{1.67x10^5 \times 0.00045}{2x10^{-23}} = 3x10^{24} \ (\textit{rounded to } 1 \textit{ sig.fig})$$

Example 3

$\sqrt{2.3} = \textit{calculator returns } 1.516575089$

$\quad 2.3 \textit{ has } 2 \textit{ sig.figs}$

$\sqrt{2.3} = 1.5 \quad \textit{round answer to } 2 \textit{ sig.figs}$

(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. $123cm + 5.35cm = 128cm \ (\textit{rounded to } 10^0)$

b. $1.0001m + 0.0003m = 1.0004m \ (\textit{rounded to } 10^{-4})$

c. $1.002s - 0.998s = 0.004s \ (\textit{rounded to } 10^{-3})$

Example 5 : *Watch for numbers ending with zero!*

$10 + 0.0110 = \textit{calculator returns } 10.0110$

$\quad 10 : \textit{the uncertain digit appears in the } 10^1 \textit{ place}$

$\quad 0.0110 : \textit{the uncertain digit appears in the } 10^{-4} \textit{ place}$

$10 + 0.0110 = 10 \quad \textit{round answer to the } 10^1 \textit{ place}$

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

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. 450*nm* ____________________ *mm*

2. 34*km* ______________________*cm*

3. 43 000*mm* _____________________k*m*

4. 4.0×10^{6} *nm* __________________µ*m*

5. 3.98×10^{-3} km___________________µ*m*

6. 456*mm* ______________________*km*

7. 136 000*m* _____________________*km*

8. 4.89×10^{12} *mm* ________________*km*

9. 2.68×10^{6} *m* __________________*km*

10. 456 000 µ*m* _________________*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"!

12 ***inches*** in one ***foot***
365 ***days*** in one ***year***
7 days in one ***week***
1.0 x 10^9 ***bytes*** in one ***gigabyte***

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:

Area = (length -**m**) x (width -**m**) = $\mathbf{m^2}$

Volume = (length -**m**) x (width -**m**) x (height - **m**) = $\mathbf{m^3}$

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

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

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 final 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 km^2. 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/m^3?

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	^{14}C		14.00	6		
Pb-208						
		15	30.97			15
			55.845			23
Potassium Ion (cation)	K^{+}		39.10			18
Sulfur Ion (anion)	S^{2-}		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

Alpha particle $^{4}_{2}He$ (an alpha particle is a helium nucleus)

Proton $^{1}_{1}H$ (the most common hydrogen nucleus is a proton)

Neutron $^{1}_{0}n$

Electron $^{0}_{-1}e$ or $^{0}_{-1}\beta$ (also called a beta particle)

Positron $^{0}_{1}e$ or $^{0}_{1}\beta$

Gamma ray $^{0}_{0}\gamma$

Write out or complete the following nuclear reactions.

1) Phosphorus-32 decays by beta emission to form sulfur-32.

---

2) Francium-212 ($^{212}_{87}Fr$) decays by alpha emission.

---

3) Sodium-24 decays by beta emission.

---

4) Fluorine-18 decays to oxygen-18 by positron emission.

---

5) Krypton-76 absorbs a beta particle to form bromine-76.

---

6) Aluminum-27 absorbs an alpha particle to form phosphorus-30 and a particle.

---

7) Nitrogen-14 absorbs an alpha particle to form oxygen-17 and emits a particle.

---

8) When neptunium-239 decays, plutonium-239 is formed and a particle is emitted. (Be sure to include the correct particle in the equation.)

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?

Part IIIb: Electron Configuration & Orbital Diagrams

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.

$C = \lambda\nu$ **$E = h\nu$**

$C = 3.00 \times 10^{8}$ m/s **$h = 6.626 \times 10^{-34}$ 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 $3.76 \times 10^{14}\ s^{-1}$? 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

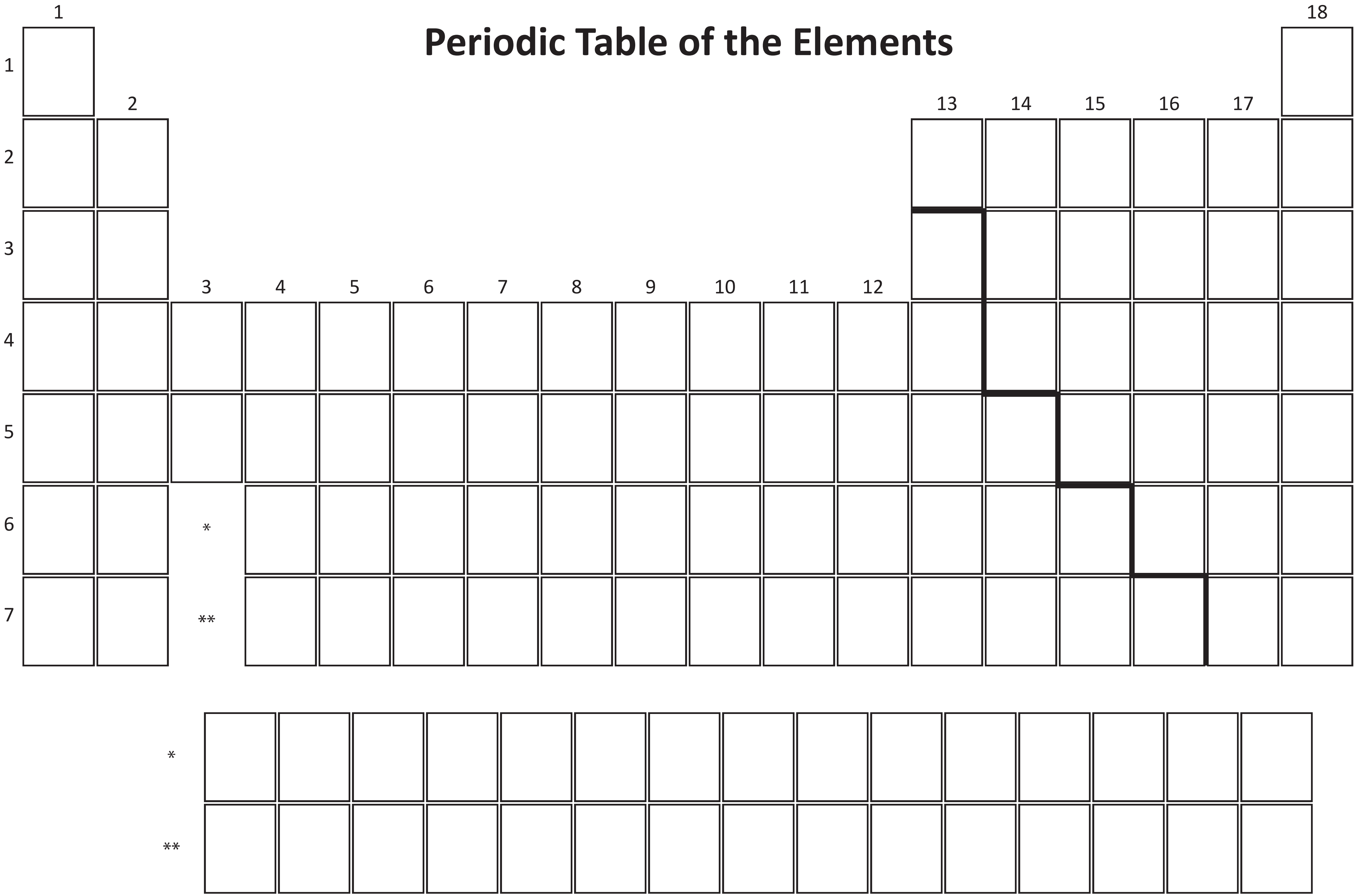
1. On the blank periodic table, color and label:
    a. alkali metals
    b. alkaline metals
    c. transition metals
    d. nonmetals
    e. metalloids
    f. halogens
    g. noble gases
    h. inner transition metals

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



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:

$\Delta EN > 2.0$ → Bond is ionic

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

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

$1.6 < \Delta EN \leq 2.0$ → 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**

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! → $Fe^{3+}{}_2 SO_4^{2-}{}_3$

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 = $(NH_4)_2CO_3$

No → **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, 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

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 = 6.02×10^{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:

$$\left(\frac{1\ mole}{6.02 \times 10^{23}\ particles}\right)\ or\ \left(\frac{6.02 \times 10^{23}\ particles}{1\ mole}\right)$$

$$\left(\frac{1\ mole}{molar\ mass\ in\ grams}\right)\ or\ \left(\frac{molar\ mass\ in\ grams}{1\ mole}\right)$$

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

Example Problems:

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

$$\mathbf{3.01 \times 10^{22}\ atoms} \left(\frac{1\ mole}{6.02x10^{23}\ atoms}\right) = 5 \times 10^{-2} \text{ moles}$$

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

$$\mathbf{4.0\ moles} \left(\frac{6.02x10^{23}\ molecules}{1\ mole}\right) = 2.41 \times 10^{24} \text{ molecules}$$

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

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

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

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

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

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

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

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

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

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

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

****

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?

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) →

___________________ b. _____ $Ba(NO_3)_2$(aq) + ______ Na_3PO_4(aq) →

___________________c. _____ Al(s) + ______ O_2(g) →

___________________d. ______ HBr(aq) + ______ KOH(aq) →

___________________e. ______ $Ca(NO_3)_2$ (aq) + ______ Na_2SO_4 (aq) →

3. In the following equations, label the oxidized element and the reduced element.

a. 2Na(s) + Cl_2(g) → 2NaCl(s)

b. 2NaBr(aq) + Cl_2(g) → 2NaCl(s) + Br_2(l)

Table 11.2

Activity Series of Metals

Decreasing reactivity ↓	Name	Symbol
	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

Name___________________________________ Period_____

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: $2KCl\ (aq) + Pb(NO_3)_2\ (aq) \text{ -> } 2KNO_3\ (aq) + PbCl_2\ (s)$

Complete Ionic Equation: $\cancel{2K^+}\ (aq) + 2Cl^-\ (aq) + Pb^{2+}\ \cancel{(aq) + 2}NO^{3-}\ \cancel{(aq) \text{ ->}}\ 2K^+\ \cancel{(aq) +}\ 2NO3^-\ (aq) + PbCl_2\ (s)$

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

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 + 2×1.01 + 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:

$$180.18 \text{ g/mol} \div 30.03 \text{ g/mol} = 6$$

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!

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	=	$\boxed{\text{46.3 g CO}}$
	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	=	$\boxed{\text{20.4 g } CO_2}$
	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 = $\boxed{\text{7.68 g } C_2H_2 \text{ excess (leftover)}}$

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 $\rightarrow$

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?

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

____$H_3PO_{4(aq)}$ + _____ $Ca(OH)_{2(aq)}$ $\rightarrow$

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