

This is an extra Summer Review Packet from another teacher.

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1	2	13	14	15	16	17	18
1 H 1.008	2 He 4.00	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3 Li 6.94	4 Be 9.01	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
11 Na 22.99	12 Mg 24.30	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
19 K 39.10	20 Ca 40.08	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (97)	44 Ru 101.1
37 Rb 85.47	38 Sr 87.62	57 *La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4
55 Cs 132.91	56 Ba 137.33	89 *Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)
87 Fr (223)	88 Ra (226)	101 Bi 208.98	102 Po (209)	103 At (210)	104 Rn (222)	105 Uuo (294)	106 Uus (294)
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
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
57 *La 138.91	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
89 *Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)
101 Bi 208.98	102 Po (209)	103 At (210)	104 Rn (222)	105 Uuo (294)	106 Uus (294)	107 Uuh (294)	108 Uuo (294)
111 Rg (282)	112 Cn (285)	113 Uut (285)	114 Fl (289)	115 Uup (288)	116 Lv (293)	117 Uus (294)	118 Uuo (294)
121 Sb 121.76	122 Te 127.60	123 I 126.90	124 Xe 131.29	125 Uuo (294)	126 Uus (294)	127 Uuh (294)	128 Uuo (294)
131 Bi 208.98	132 Po (209)	133 At (210)	134 Rn (222)	135 Uuo (294)	136 Uus (294)	137 Uuh (294)	138 Uuo (294)
141 Bi 208.98	142 Po (209)	143 At (210)	144 Rn (222)	145 Uuo (294)	146 Uus (294)	147 Uuh (294)	148 Uuo (294)
151 Bi 208.98	152 Po (209)	153 At (210)	154 Rn (222)	155 Uuo (294)	156 Uus (294)	157 Uuh (294)	158 Uuo (294)
161 Bi 208.98	162 Po (209)	163 At (210)	164 Rn (222)	165 Uuo (294)	166 Uus (294)	167 Uuh (294)	168 Uuo (294)
171 Bi 208.98	172 Po (209)	173 At (210)	174 Rn (222)	175 Uuo (294)	176 Uus (294)	177 Uuh (294)	178 Uuo (294)
181 Bi 208.98	182 Po (209)	183 At (210)	184 Rn (222)	185 Uuo (294)	186 Uus (294)	187 Uuh (294)	188 Uuo (294)
191 Bi 208.98	192 Po (209)	193 At (210)	194 Rn (222)	195 Uuo (294)	196 Uus (294)	197 Uuh (294)	198 Uuo (294)
201 Bi 208.98	202 Po (209)	203 At (210)	204 Rn (222)	205 Uuo (294)	206 Uus (294)	207 Uuh (294)	208 Uuo (294)
211 Bi 208.98	212 Po (209)	213 At (210)	214 Rn (222)	215 Uuo (294)	216 Uus (294)	217 Uuh (294)	218 Uuo (294)
221 Bi 208.98	222 Po (209)	223 At (210)	224 Rn (222)	225 Uuo (294)	226 Uus (294)	227 Uuh (294)	228 Uuo (294)
231 Bi 208.98	232 Po (209)	233 At (210)	234 Rn (222)	235 Uuo (294)	236 Uus (294)	237 Uuh (294)	238 Uuo (294)
241 Bi 208.98	242 Po (209)	243 At (210)	244 Rn (222)	245 Uuo (294)	246 Uus (294)	247 Uuh (294)	248 Uuo (294)
251 Bi 208.98	252 Po (209)	253 At (210)	254 Rn (222)	255 Uuo (294)	256 Uus (294)	257 Uuh (294)	258 Uuo (294)
261 Bi 208.98	262 Po (209)	263 At (210)	264 Rn (222)	265 Uuo (294)	266 Uus (294)	267 Uuh (294)	268 Uuo (294)
271 Bi 208.98	272 Po (209)	273 At (210)	274 Rn (222)	275 Uuo (294)	276 Uus (294)	277 Uuh (294)	278 Uuo (294)
281 Bi 208.98	282 Po (209)	283 At (210)	284 Rn (222)	285 Uuo (294)	286 Uus (294)	287 Uuh (294)	288 Uuo (294)
291 Bi 208.98	292 Po (209)	293 At (210)	294 Rn (222)	295 Uuo (294)	296 Uus (294)	297 Uuh (294)	298 Uuo (294)
301 Bi 208.98	302 Po (209)	303 At (210)	304 Rn (222)	305 Uuo (294)	306 Uus (294)	307 Uuh (294)	308 Uuo (294)
311 Bi 208.98	312 Po (209)	313 At (210)	314 Rn (222)	315 Uuo (294)	316 Uus (294)	317 Uuh (294)	318 Uuo (294)
321 Bi 208.98	322 Po (209)	323 At (210)	324 Rn (222)	325 Uuo (294)	326 Uus (294)	327 Uuh (294)	328 Uuo (294)
331 Bi 208.98	332 Po (209)	333 At (210)	334 Rn (222)	335 Uuo (294)	336 Uus (294)	337 Uuh (294)	338 Uuo (294)
341 Bi 208.98	342 Po (209)	343 At (210)	344 Rn (222)	345 Uuo (294)	346 Uus (294)	347 Uuh (294)	348 Uuo (294)
351 Bi 208.98	352 Po (209)	353 At (210)	354 Rn (222)	355 Uuo (294)	356 Uus (294)	357 Uuh (294)	358 Uuo (294)
361 Bi 208.98	362 Po (209)	363 At (210)	364 Rn (222)	365 Uuo (294)	366 Uus (294)	367 Uuh (294)	368 Uuo (294)
371 Bi 208.98	372 Po (209)	373 At (210)	374 Rn (222)	375 Uuo (294)	376 Uus (294)	377 Uuh (294)	378 Uuo (294)
381 Bi 208.98	382 Po (209)	383 At (210)	384 Rn (222)	385 Uuo (294)	386 Uus (294)	387 Uuh (294)	388 Uuo (294)
391 Bi 208.98	392 Po (209)	393 At (210)	394 Rn (222)	395 Uuo (294)	396 Uus (294)	397 Uuh (294)	398 Uuo (294)
401 Bi 208.98	402 Po (209)	403 At (210)	404 Rn (222)	405 Uuo (294)	406 Uus (294)	407 Uuh (294)	408 Uuo (294)
411 Bi 208.98	412 Po (209)	413 At (210)	414 Rn (222)	415 Uuo (294)	416 Uus (294)	417 Uuh (294)	418 Uuo (294)
421 Bi 208.98	422 Po (209)	423 At (210)	424 Rn (222)	425 Uuo (294)	426 Uus (294)	427 Uuh (294)	428 Uuo (294)
431 Bi 208.98	432 Po (209)	433 At (210)	434 Rn (222)	435 Uuo (294)	436 Uus (294)	437 Uuh (294)	438 Uuo (294)
441 Bi 208.98	442 Po (209)	443 At (210)	444 Rn (222)	445 Uuo (294)	446 Uus (294)	447 Uuh (294)	448 Uuo (294)
451 Bi 208.98	452 Po (209)	453 At (210)	454 Rn (222)	455 Uuo (294)	456 Uus (294)	457 Uuh (294)	458 Uuo (294)
461 Bi 208.98	462 Po (209)	463 At (210)	464 Rn (222)	465 Uuo (294)	466 Uus (294)	467 Uuh (294)	468 Uuo (294)
471 Bi 208.98	472 Po (209)	473 At (210)	474 Rn (222)	475 Uuo (294)	476 Uus (294)	477 Uuh (294)	478 Uuo (294)
481 Bi 208.98	482 Po (209)	483 At (210)	484 Rn (222)	485 Uuo (294)	486 Uus (294)	487 Uuh (294)	488 Uuo (294)
491 Bi 208.98	492 Po (209)	493 At (210)	494 Rn (222)	495 Uuo (294)	496 Uus (294)	497 Uuh (294)	498 Uuo (294)
501 Bi 208.98	502 Po (209)	503 At (210)	504 Rn (222)	505 Uuo (294)	506 Uus (294)	507 Uuh (294)	508 Uuo (294)
511 Bi 208.98	512 Po (209)	513 At (210)	514 Rn (222)	515 Uuo (294)	516 Uus (294)	517 Uuh (294)	518 Uuo (294)
521 Bi 208.98	522 Po (209)	523 At (210)	524 Rn (222)	525 Uuo (294)	526 Uus (294)	527 Uuh (294)	528 Uuo (294)
531 Bi 208.98	532 Po (209)	533 At (210)	534 Rn (222)	535 Uuo (294)	536 Uus (294)	537 Uuh (294)	538 Uuo (294)
541 Bi 208.98	542 Po (209)	543 At (210)	544 Rn (222)	545 Uuo (294)	546 Uus (294)	547 Uuh (294)	548 Uuo (294)
551 Bi 208.98	552 Po (209)	553 At (210)	554 Rn (222)	555 Uuo (294)	556 Uus (294)	557 Uuh (294)	558 Uuo (294)
561 Bi 208.98	562 Po (209)	563 At (210)	564 Rn (222)	565 Uuo (294)	566 Uus (294)	567 Uuh (294)	568 Uuo (294)
571 Bi 208.98	572 Po (209)	573 At (210)	574 Rn (222)	575 Uuo (294)	576 Uus (294)	577 Uuh (294)	578 Uuo (294)
581 Bi 208.98	582 Po (209)	583 At (210)	584 Rn (222)	585 Uuo (294)	586 Uus (294)	587 Uuh (294)	588 Uuo (294)
591 Bi 208.98	592 Po (209)	593 At (210)	594 Rn (222)	595 Uuo (294)	596 Uus (294)	597 Uuh (294)	598 Uuo (294)
601 Bi 208.98	602 Po (209)	603 At (210)	604 Rn (222)	605 Uuo (294)	606 Uus (294)	607 Uuh (294)	608 Uuo (294)
611 Bi 208.98	612 Po (209)	613 At (210)	614 Rn (222)	615 Uuo (294)	616 Uus (294)	617 Uuh (294)	618 Uuo (294)
621 Bi 208.98	622 Po (209)	623 At (210)	624 Rn (222)	625 Uuo (294)	626 Uus (294)	627 Uuh (294)	628 Uuo (294)
631 Bi 208.98	632 Po (209)	633 At (210)	634 Rn (222)	635 Uuo (294)	636 Uus (294)	637 Uuh (294)	638 Uuo (294)
641 Bi 208.98	642 Po (209)	643 At					

AP[®] CHEMISTRY EQUATIONS AND CONSTANTS

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

L, mL = liter(s), milliliter(s)
g = gram(s)
nm = nanometer(s)
atm = atmosphere(s)

mm Hg = millimeters of mercury
J, kJ = joule(s), kilojoule(s)
V = volt(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 = 6.022×10^{23} mol⁻¹

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

EQUILIBRIUM

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ where } aA + bB \rightleftharpoons cC + dD$$

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

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

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

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

$$= K_a \times K_b$$

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

$$14 = \text{pH} + \text{pOH}$$

$$\text{pH} = \text{p}K_a + \log \frac{[A^-]}{[HA]}$$

$$\text{p}K_a = -\log K_a, \text{ p}K_b = -\log K_b$$

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

$$PV = nRT$$

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

$$P_{\text{total}} = P_A + P_B + P_C + \dots$$

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

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

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

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

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure

V = volume

T = temperature

n = number of moles

m = mass

M = molar mass

D = density

KE = kinetic energy

v = velocity

A = absorbance

a = molar absorptivity

b = path length

c = concentration

Gas constant, R = $8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

$$= 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$$

$$= 62.36 \text{ L torr mol}^{-1} \text{ K}^{-1}$$

$$1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr}$$

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

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

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

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

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

$$= -RT \ln K$$

$$= -nFE^\circ$$

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

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	$\text{C}_2\text{H}_3\text{O}_2^-$
ammonium	NH_4^+
arsenate	AsO_4^{3-}
arsenite	AsO_3^{3-}
azide	N_3^-
benzoate	$\text{C}_7\text{H}_5\text{O}_2^-$
borate	BO_3^{3-}
bromate	BrO_3^-
carbonate	CO_3^{2-}
chlorate	ClO_3^-
chlorite	ClO_2^-
chromate	CrO_4^{2-}
cyanide	CN^-
dichromate	$\text{Cr}_2\text{O}_7^{2-}$
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	$\text{C}_2\text{O}_4^{2-}$
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	$\text{C}_4\text{H}_4\text{O}_6^{2-}$
thiocyanate	SCN^-
thiosulfate	$\text{S}_2\text{O}_3^{2-}$

AsO_3^{3-}	arsenite
AsO_4^{3-}	arsenate
BO_3^{3-}	borate
BrO_3^-	bromate
$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
$\text{C}_2\text{O}_4^{2-}$	oxalate
$\text{C}_4\text{H}_4\text{O}_6^{2-}$	tartrate
$\text{C}_7\text{H}_5\text{O}_2^-$	benzoate
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
CN^-	cyanide
CO_3^{2-}	carbonate
$\text{Cr}_2\text{O}_7^{2-}$	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
$\text{S}_2\text{O}_3^{2-}$	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 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 to 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 ($number\ 1 \leq x < 10$)

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:

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

2.3 has 2 sig.fig

5.78 has 3 sig.fig.

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

Example 2.

$$\frac{1.67 \times 10^5 \times 0.00045}{2 \times 10^{-23}} = \text{calculator returns } 2.505000000 \times 10^{24}$$

1.67x10⁵ has 3 sig.figs

0.00045 has 2 sig.figs

2x10⁻²³ has 1 sig.fig

$$\frac{1.67 \times 10^5 \times 0.00045}{2 \times 10^{-23}} = 3 \times 10^{24} \quad (\text{rounded to 1 sig.fig})$$

Example 3

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

2.3 has 2 sig.figs

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

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

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

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

Example 5: Watch for numbers ending with zero!

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

10: the uncertain digit appears in the 10^1 place

0.0110: the uncertain digit appears in the 10^{-4} place

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

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

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. 450nm _____ mm

2. 34km _____ cm

3. $43\,000\text{mm}$ _____ km

4. $4.0 \times 10^6 \text{ nm}$ _____ μm

5. $3.98 \times 10^{-3} \text{ km}$ _____ μm

6. 456mm _____ km

7. $136\,000\text{m}$ _____ km

8. $4.89 \times 10^{12} \text{ mm}$ _____ km

9. $2.68 \times 10^6 \text{ m}$ _____ km

10. $456\,000 \mu\text{m}$ _____ 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**

1.0×10^9 **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:

$$\underline{Area} = (\text{length} - \mathbf{m}) \times (\text{width} - \mathbf{m}) = \mathbf{m}^2$$

$$\underline{Volume} = (\text{length} - \mathbf{m}) \times (\text{width} - \mathbf{m}) \times (\text{height} - \mathbf{m}) = \mathbf{m}^3$$

$$\underline{Velocity} = (\text{distance traveled} - \mathbf{m}) / (\text{time} - \mathbf{s}) = \mathbf{m/s}$$

$$\underline{Density} = (\text{mass} - \mathbf{g}) / (\text{volume} - \mathbf{mL}) = \mathbf{g/mL}$$

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 <i>yds</i>	1 mile	60 s	60 minute	78 miles
s	1 m	1760 <i>yds</i>	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 \times 4.0 in \times 3.0 in. The density of gold is 19.3 g/cm³. 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 = 1 g/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 $1.0 \times 10^3 \text{ kg/m}^3$?

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	${}^4_2\text{He}$ (an alpha particle is a helium nucleus)
Proton	${}^1_1\text{H}$ (the most common hydrogen nucleus is a proton)
Neutron	${}^1_0\text{n}$
Electron	${}^0_{-1}\text{e}$ or ${}^0_{-1}\beta$ (also called a beta particle)
Positron	${}^0_1\text{e}$ 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 (${}^{212}_{87}\text{Fr}$) 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 following 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.

$$c = \lambda \nu$$

$$c = 3.00 \times 10^8 \text{ m/s}$$

$$E = h \nu$$

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

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

1. What is the wavelength of a wave having a frequency of $3.76 \times 10^{14} \text{ s}^{-1}$? What is its energy?
2. What is the frequency of a $6.9 \times 10^{-10} \text{ cm}$ wave? What is its energy?
3. What is the frequency of a $7.43 \times 10^{-5} \text{ mm}$ wave? What is its energy?
4. What is the wavelength of a wave carrying $8.35 \times 10^{-18} \text{ J}$ 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

1																	18
2																	
3																	
4																	
5																	
6			*														
7			**														

*														
**														

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.

[illegible]

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 central atom using at least 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 CO₂ 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 than 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

$\text{Sn}_3(\text{PO}_4)_2$ = Tin (II) phosphate

$(\text{NH}_4)_2\text{SO}_4$ = 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) → $\text{Fe}^{3+}\text{SO}_4^{2-}$

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! → $\text{Fe}^{3+}\text{SO}_4^{2-}$

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. → $\text{Fe}^{3+}_2(\text{SO}_4^{2-})_3$

5. Erase any ion charges in the formula → $\text{Fe}_2(\text{SO}_4)_3$

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 = $\text{Cu}_3(\text{PO}_4)_2$, ammonium carbonate = $(\text{NH}_4)_2\text{CO}_3$

No

Covalent

- 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 tetroxide = 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

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

No

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 _____

 $\text{Ni}(\text{SO}_3)_2$ _____

phosphorus triiodide _____

 H_3PO_4 _____

 HI _____

 Pb_3N_4 _____

 $\text{Sn}(\text{OH})_2$ _____

 SiCl_4 _____

 HClO_2 _____

Sodium sulfate _____

Hydrosulfuric acid _____

Nitrogen trifluoride _____

Calcium phosphide _____

 B_2Si _____

 PCl_5 _____

Perchloric 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:

$$1 \text{ mol} = 6.02 \times 10^{23} \text{ particles}$$

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:

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

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

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

Example Problems:

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

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

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

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

3. How many moles in 28 grams of CO_2 ?

Molar mass of CO₂

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

$$28 \text{ g CO}_2 \left(\frac{1 \text{ mole}}{44.00 \text{ g}} \right) = 0.64 \text{ moles CO}_2$$

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

Molar mass Fe_2O_3 $2 \text{ Fe} = 2 \times 55.6 \text{ g} = 111.2 \text{ g}$
 $3 \text{ O} = 3 \times 16.0 \text{ g} = \underline{48.0 \text{ g}}$
 159.2 g/mol

$$5 \text{ moles Fe}_2\text{O}_3 \left(\frac{159.2 \text{ g}}{1 \text{ mole}} \right) = 800 \text{ grams Fe}_2\text{O}_3$$

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

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

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

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

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

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

Molar mass $2 \text{ F} = 2 \times 19 \text{ g} = 38 \text{ g/mol}$

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

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 $\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$.

a) Calculate the gram molar mass of aspartame.

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

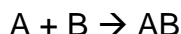
c) How many molecules are in 5 **mg** of aspartame?

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

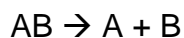
Chemical Reactions Review Sheet

Types of Chemical Reactions:

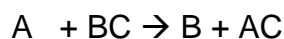
Combination or Synthesis



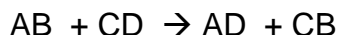
Decomposition



Single Replacement

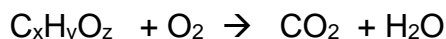


Double Replacement



- 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

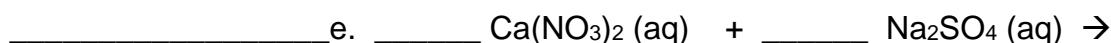
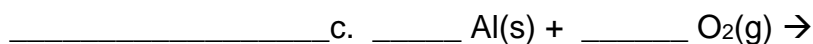
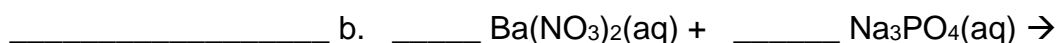
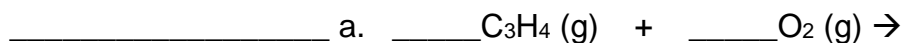


Oxidation-Reduction - Involve a transfer of electrons. Occurs during combustion, single replacement and can occur during synthesis and decomposition.

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.



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

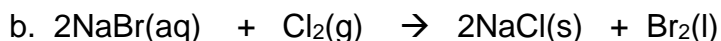
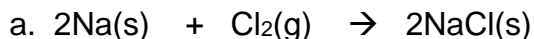


Table 11.2

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 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

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: $2\text{KCl} (aq) + \text{Pb}(\text{NO}_3)_2 (aq) \rightarrow 2\text{KNO}_3 (aq) + \text{PbCl}_2 (s)$

Complete Ionic Equation: $\cancel{2\text{K}^+ (aq)} + \cancel{2\text{Cl}^- (aq)} + \text{Pb}^{2+} (aq) + \cancel{2\text{NO}_3^- (aq)} \rightarrow \cancel{2\text{K}^+ (aq)} + \cancel{2\text{NO}_3^- (aq)} + \text{PbCl}_2 (s)$

Net Ionic Equation: $2\text{Cl}^- (aq) + \text{Pb}^{2+} (aq) \rightarrow \text{PbCl}_2 (s)$

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 = $40.0\text{g C} \times 1\text{ mol C}/12.01\text{g C} = 3.331\text{ mol C}$

Moles Hydrogen = $6.73\text{g H} \times 1\text{ mol H}/1.01\text{g} = 6.663\text{ mol H}$

Mole Oxygen = $53.27\text{ g O} \times 1\text{ mol O}/16.0\text{ g} = 3.33\text{ mol O}$

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 $\text{C}_{3.331}\text{H}_{6.663}\text{O}_{3.33}$. I want a whole number ratio, so I will divide all the subscripts by the smallest number of moles (3.331) to get:

$\text{C}_1\text{H}_2\text{O}_1 \rightarrow$ 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:



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:

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

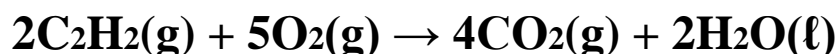
The molar mass of my compound is 6 times the molar mass of my empirical formula.

Multiply the empirical formula subscripts by 6 to get the final molecular formula:



Steps to Solving Limiting Reagent Problems

Suppose 13.7 g of C₂H₂ reacts with 18.5 g O₂ according to the reaction below. What is the mass of CO₂ produced? What is the limiting reagent?



- Find the mass of product yielded by the given amount of the first reactant. You can use either product (CO₂ or H₂O), but since the question asks about CO₂, it will be easier to use this product:

13.7 g C ₂ H ₂	1 mole C ₂ H ₂	4 mole CO ₂	44.02 g CO ₂	= 46.3 g CO
	26.04 g C ₂ H ₂	2 mole C ₂ H ₂	1 mole CO ₂	

- Find the mass of *the same product* (in this case CO₂) 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	

- Since the 18.5 grams of O₂ produces *less CO₂*, it is the **limiting reagent** in this problem. This amount of O₂ gets used up first and “limits” how much CO₂ can be produced. The amount of CO₂ that can be produced is 20.4 grams (which you already calculated!)
- 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).
- 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 ₂ H ₂	26.02 g C ₂ H ₂	= 6.02 g C ₂ H ₂ used
	32.00 g O ₂	5 mole O ₂	1 mole C ₂ H ₂	

13.7 g of C ₂ H ₂ total – 6.02 g of C ₂ H ₂ used = 7.68 g C ₂ H ₂ excess (leftover)

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. $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow$
 - 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:



- What type of reaction is shown above? _____
- Predict products and then balance the reaction.
- Name the ionic product of the reaction. _____
- Which element is oxidized? _____ Which element is reduced? _____
- 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?
- How many grams of the excess reagent are leftover?
- What is the pH of the HBr solution?
- What is the OH^- concentration of the HBr solution?
- If the gas is produced at 89°C and 1.7 atm of pressure, what is the volume of gaseous product in mL?
- 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



- 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