HONORS CHEM
1st Semester
Reference Sheets

Please be understanding if any of these reference sheets change, I am attempting to predict what we will cover and use for the entire school year…this is my best guess!

<u>Contains Reference Sheets for the following chapters.</u>

- Unit 1 – Chemistry Basics and Atomic Structure
- Unit 2 – Nuclear Chemistry
- Unit 3 – Electrons
- Unit 4 – Periodic Table
- Unit 5 – Bonding and Structure
- Unit 6 – Reactions
- Unit 7 – Stoichiometry

Units 8-14 are in a separate Reference Sheet photocopy packet.

Reference Sheets for
Unit #1 – Chemistry Basics and Atomic Structure

Dougherty Valley HS Chemistry
Chemistry Reference – Do Not Misplace!

Scientific Notation

Used to express a very large or very small number.
Move the decimal place to the right or to the left to produce a number between 1 and 10.
If you move the decimal to the right, your exponent will be negative.
If you move the decimal to the left, your exponent will be positive.
Adding and Subtracting numbers that are expressed in scientific notation require you to change the numbers so that they have the same exponents, you can do this by moving the decimal around a bit. You can also just use your calculator to add or subtract these numbers.
Multiplying numbers in scientific notation requires you to multiply the first factors then add the exponents.
Dividing numbers in scientific notation requires you to divide the first factors then subtract the exponents.

Dimensional Analysis

Dimensional analysis is a problem solving method that uses conversion factors.
A conversion factor is a ratio of equivalent values. For example; 1000m/1km
In solving dimensional analysis problems you always set the value you want over the value you already have. (What you want over what you got!)
You will cancel units and multiply to achieve your final value.

Accuracy and Precision

Accuracy refers to how close a measured value is to an accepted value.
Precision refers to how close a series of measurements are to one another.
Percent error is the ratio of an error to an accepted value.
Percent error = error/accepted value x 100 and should be expressed as a percentage.
It is irrelevant if the experimental value is larger or smaller than the accepted value.

Significant Figures

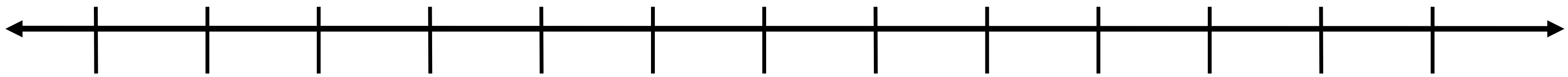
Significant figures include all known digits plus one estimated digit.
Non-zero numbers are always significant.
Zeros between non-zero numbers are always significant.
All final zeros to the right of the decimal place are significant.
Zeros that act, as placeholders are not significant.
Counting numbers and defined constants have an infinite number of significant figures.

Rounding Numbers

If the remainder *beyond the last digit* to be reported is less than 5, drop the last digit.
Rounding to one decimal place, the number 5.3467 becomes 5.3.
If the remainder is greater than 5, increase the final digit by 1. The number 5.798 becomes 5.8 if rounding to 1 digit.
To prevent rounding bias, if the remainder is exactly 5, then round the last digit to the closest even number. Thus the number 3.55 (rounded to 1 digit) would be 3.6 (rounding up) and the number 6.450 would round to 6.4 (rounding down) *if rounding to 1 decimal*.

Metric Units and Conversions

Value	1,000,000	100,000	10,000	1,000	100	10	1	1/10	1/100	1/1000	1/10000	1/100000	1/1000000
Exponent	10^6	10^5	10^4	10^3	10^2	10^1	10^0	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}



Prefix	mega			kilo	hecto	deka	base unit	deci	centi	milli			micro
Abbr.	M			k	h	da		d	c	m			μ

Working with quantities that are not in Scientific Notation

1. Find the prefix with which you are beginning. If the unit has no prefix attached, you are beginning with the "base unit" at 10^0.
2. Find the prefix for the answer you are seeking. If the unit has no prefix attached, you are converting to the "base unit" at 10^0.
3. Count the number of places on the number line to get from where you are starting to where you are finishing.
4. Now, move the decimal in the number you are converting that same number of places, and in the same direction that you moved on the number line above (if you moved left three spaces, you move the decimal left three spaces to complete the conversion).

Example: Convert 0.035 decimeters (dm) to millimeters (mm)

Solution: The prefix "milli" is two places (two powers of ten) to the right of the prefix "deci." Move the decimal two places to the right.

Answer: 0.035 dm = 3.5 mm

Working with numbers that are in Scientific Notation

1. Find the prefix with which you are beginning. If the unit has no prefix attached, you are beginning with the "base unit" at 10^0.
2. Find the prefix for the answer you are seeking. If the unit has no prefix attached, you are converting to the "base unit" at 10^0.
3. Count the number of places on the number line to get from where you are starting to where you are finishing.
4. If you moved to the right on the line, add the number of spaces to the exponent on 10.
5. If you moved to the left, subtract the number of spaces from the exponent on 10.

Example: Convert 1.35×10^2 centigrams (cg) to kilograms (kg)

Solution: The prefix "kilo" is five places (five powers of ten) to the left of the prefix "centi." Subtract five from the exponent.

Answer: 1.35×10^2 centigrams = $1.35 \times 10^{2-5}$ kilograms = 1.35×10^{-3} kg

Significant Figures in Measurement and Calculations

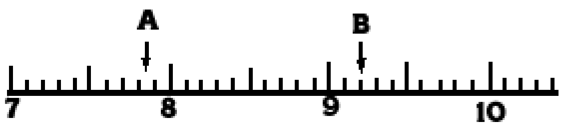
A successful chemistry student habitually labels all numbers, because the unit is important. Also of great importance is the number itself. Any number used in a calculation should contain only figures that are considered reliable; otherwise, time and effort are wasted. Figures that are considered reliable are called *significant figures*. Chemical calculations involve numbers representing actual measurements. In a measurement, significant figures in a number consist of:

Figures (digits) definitely known + One estimated figure (digit)

In class you will hear this expressed as "all of the digits known for certain plus one that is a guess."

Recording Measurements

When one reads an instrument (ruler, thermometer, graduate, buret, barometer, balance), he expresses the reading as one which is reasonably reliable. For example, in the accompanying illustration, note the reading marked *A*. This reading is definitely beyond the 7 cm mark and also beyond the 0.8 cm mark. We read the 7.8 with certainty. We further *estimate* that the reading is five-tenths the distance from the 7.8 mark to the 7.9 mark. So, we estimate the length as 0.05 cm more than 7.8 cm. All of these have meaning and are therefore significant. We express the reading as 7.85 cm, accurate to three significant figures. All of these figures, *7.85*, can be used in calculations. In reading B we see that 9.2 cm is definitely known. We can include one estimated digit in our reading, and we estimate the next digit to be zero. Our reading is reported as 9.20 cm. It is accurate to three significant figures.



Rules for Zeros

If a zero represents a measured quantity, it is a significant figure. If it merely locates the decimal point, it is not a significant figure.

Zero Within a Number. In reading the measurement 9.04 cm, the zero represents a measured quantity, just as 9 and 4, and is, therefore, a significant number. A zero between any of the other digits in a number is a significant figure.

Zero at the Front of a Number. In reading the measurement 0.46 cm, the zero does not represent a measured quantity, but merely locates the decimal point. It is not a significant figure. Also, in the measurement 0.07 kg, the zeros are used merely to locate the decimal point and are, therefore, not significant. Zeros at the first (left) of a number are not significant figures.

Zero at the End of a Number. In reading the measurement 11.30 cm, the zero is an estimate and represents a measured quantity. It is therefore significant. Another way to look at this: The zero is not needed as a placeholder, and yet it was included by the person recording the measurement. It must have been recorded as a part of the measurement, making it significant. Zeros to the right of the decimal point, and at the end of the number, are significant figures.

Zeros at the End of a Whole Number. Zeros at the end of a whole number may or may not be significant. If a distance is reported as 1600 feet, one assumes two sig figs. Reporting measurements in scientific notation removes all doubt, since all numbers written in scientific notation are considered significant.

significant.	1 600 feet	1.6×10^3 feet	Two significant figures
	1 600 feet	1.60×10^3 feet	Three significant figures
	1 600 feet	1.600×10^3 feet	Four significant figures

Sample Problem #1: Underline the significant figures in the following numbers.

(a) 0.0420 cm	answer = 0.0420 cm	(e) 2 403 ft.	answer = 2 403 ft.
(b) 5.320 in.	answer = 5.320 in.	(f) 80.5300 m	answer = 80.5300 m
(c) 10 lb.	answer = 10 lb.	(g) 200. g	answer = 200 g
(d) 0.020 ml	answer = 0.020 ml	(h) 2.4×10^3 kg	answer = 2.4×10^3 kg

Rounding Off Numbers

In reporting a numerical answer, one needs to know how to "round off" a number to include the correct number of significant figures. Even in a series of operations leading to the final answer, one must "round off" numbers. The rules are well accepted rules:

1. If the figure to be dropped is less than 5, simply eliminate it.
2. If the figure to be dropped is greater than 5, eliminate it and raise the preceding figure by 1.
3. If the figure is 5, followed by nonzero digits, raise the preceding figure by 1
4. If the figure is 5, not followed by nonzero digit(s), and preceded by an odd digit, raise the preceding digit by one
5. If the figure is 5, not followed by nonzero digit(s), and the preceding significant digit is even, the preceding digit remains unchanged

Sample Problem #2: Round off the following to three significant figures.

(a) 3.478 m	answer = 3.48 m	(c) 5.333 g	answer = 5.33 g
(b) 4.8055 cm	answer = 4.81 cm	(d) 7.999 in.	answer = 8.00 in.

Multiplication

In multiplying two numbers, when you wish to determine the number of significant figures you should have in your answer (the product), you should inspect the numbers multiplied and find which has the least number of significant figures. This is the number of significant figures you should have in your answer (the product). Thus the answer to 0.024 x 1244 would be rounded off to contain two significant figures since the factor with the lesser number of significant figures (0.024) has only *two* such figures.

Sample Problem #3: Find the area of a rectangle 2.1 cm by 3.24 cm.

Solution: Area = 2.1 cm x 3.24 cm = 6.804 cm^2

We note that 2.1 contains two significant figures, while 3.24 contains three significant figures. Our product should contain no more than *two* significant figures. Therefore, our answer would be recorded as 6.8 cm^2

Sample Problem #4: Find the volume of a rectangular solid 10.2 cm x 8.24 cm x 1.8 cm

Solution: Volume = 10.2 cm x 8.24 cm x 1.8 cm = 151.2864 cm^3

We observe that the factor having the least number of significant figures is 1.8 cm. It contains two significant figures. Therefore, the answer is rounded off to 150 cm^3.

Division

In dividing two numbers, the answer (quotient) should contain the same number of significant figures as are contained in the number (divisor or dividend) with the least number of significant figures. Thus the answer to 528 ÷ 0.14 would be rounded off to contain *two* significant figures. The answer to 0.340 ÷ 3242 would be rounded off to contain three significant figures.

Sample Problem #5: Calculate 20.45 ÷ 2.4

Solution: 20.45 ÷ 2.4 = 8.52083

We note that the 2.4 has fewer significant figures than the 20.45. It has only *two* significant figures. Therefore, our answer should have no more than two significant figures and should be reported as 8.5.

Addition and Subtraction

In adding (or subtracting), set down the numbers, being sure to keep like decimal places under each other, and add (or subtract). Next, note which column contains the first estimated figure. This column determines the last decimal place of the answer. After the answer is obtained, it should be rounded off in this column. In other words, round to the least number of decimal places in you data.

Sample Problem #6: Add 42.56 g + 39.460 g + 4.1g

Solution:

	42.56 g
	39.460 g
	4.1 g
Sum =	86.120 g

Since the number 4.1 only extends to the first decimal place, the answer must be rounded to the first decimal place, yielding the answer 86.1 g.

Average Readings

The average of a number of successive readings will have the same number of decimal places that are in their sum.

Sample Problem #7: A graduated cylinder was weighed three times and the recorded weighings were 12.523 g, 12.497 g, 12.515 g. Calculate the average weight.

Solution:

12.523 g
12.497 g
12.515 g
37.535 g

In order to find the average, the sum is divided by 3 to give an answer of 12.51167. Since each number extends to three decimal places, the final answer is rounded to three decimal places, yielding a final answer of 12.512 g. Notice that the divisor of 3 does not effect the rounding of the final answer. This is because 3 is an exact number - known to an infinite number of decimal places.

Common English and Metric Conversions Chart

American Linear Units	
12 inches (in)	1 foot (ft)
3 feet	1 yard (yd)
36 inches	1 yard
63,360 inches	1 mile (mi)
5,280 feet	1 mile
1,760 yards	1 mile

Weight and Mass	
1 Ton (T)	2,000 pounds
1 pound (lb)	16 ounces (oz)
1 Ton	32,000 ounces
1 metric ton (t)	1000 kg
60 grains	1 dram

Converting American Units	
Larger unit → smaller unit	*Multiply*
smaller unit → Larger unit	*Divide*

American to Metric Units	
1 inch	2.540 centimeters
1 foot	0.305 meters
1 yard	0.914 meters
1 mile	1.609 kilometers
1 gallon	3.78 Liters
1 quart	0.95 Liter
1 pound	0.45 kilogram
1 ounce	28.35 grams
1 fluid ounce	29.57 mL
1 grain	60 milligrams (mg)
1 teaspoon (tsp)	5 mL
1 fluid dram	4 mL
1 tablespoon (tbsp)	15 mL
1 pint (pt)	500 mL (approx)
1 quart (qt)	1000 mL (approx)
1 pound (lb)	453.6 g

American Capacity	
8 fluid ounces (fl oz)	1 cup
16 fluid ounces	2 cups
2 cups	1 pint (pt)
16 fluid ounces	1 pint
2 pints	1 quart (qt)
4 quarts	1 gallon
8 pints	1 gallon
32 fluid ounces	1 quart
8 fluid dram	1 fluid ounce
3 teaspoon (tsp)	1 tablespoon (tbsp)
6 teaspoon	1 fluid ounce
2 tablespoon	1 fluid ounce
1 drop (gtt)	1 minim
60 drop	1 fluid dram
60 drop	1 teaspoon
60 minims	1 fluid dram

Metric Units

mega (M)	*	*	kilo (k)	hector (h)	deka (da)	unit (m, g, L)	deci (d)	centi (c)	milli (m)	*	*	micro (mc) (u)

When going from larger unit to smaller unit move decimal to the right →

← **When going from smaller unit to larger unit move decimal to the left**

Time	
1 day	24 hours
1 hour (hr)	60 minutes (min)
1 minute	60 seconds (sec)
1 year (yr)	365.25 days
1 week	7 days
1 year	12 months (mon)
1440 minutes	1 day
3600 seconds	1 hour

Stones	
1 carat (karat)	200 mg

Metric to American Units	
1 km	0.621 miles
1 meter	1.094 yards
1 meter	3.281 feet
1 meter	39.370 inches
1 cm	0.3937 inch
1 Liter	0.26 gallon
1 Liter	1.06 quarts
1 kg	2.20 lbs
1 gram	0.035 oz
1 gram	15 grains
1 milliliter (mL)	15 minims

Temperature Formulas	
$C = \frac{(F - 32)}{1.8}$	$F = 1.8 \cdot C + 32$

Medical Application (Micrograms)	
1,000,000 micrograms (mcg)	1 gram
1,000,000 micrograms	1,000 mg
1 mL = 1 cc = 1 cm^3	
1 gram = 1 cm^3	
Nursing students 1fl oz = 30 mL	
Nursing students 1 in. = 2.5 cm	

Significant Figures with Scientific Notation Addition and Subtraction

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Speaking realistically, the problems discussed below can all be done on a calculator. However, you need to know how to enter values into the calculator, read your calculator screen, and round off to the proper number of significant figures. Your calculator will not do these things for you.

All exponents MUST BE THE SAME before you can add and subtract numbers in scientific notation. The actual addition or subtraction will take place with the numerical portion, NOT the exponent.

The student might wish to re-read the above two sentences with emphasis on the emphasized portions.

It might be advisable to point out again - DO NOT, under any circumstances, add the exponents.

Example #1: $1.00 \times 10^3 + 1.00 \times 10^2$

A good rule to follow is to express all numbers in the problem in the highest power of ten.

Convert 1.00×10^2 to 0.10×10^3, then add:

$$
\begin{aligned}
& \ 1.00 \times 10^3 \\
& +\ 0.10 \times 10^3 \\
& =\ 1.10 \times 10^3
\end{aligned}
$$

Example #2: The significant figure issue is sometimes obscured when numbers are in scientific notation. For example, add the following four numbers:

$$(4.56 \times 10^6) + (2.98 \times 10^5) + (3.65 \times 10^4) + (7.21 \times 10^3)$$

When the four numbers are written in the highest power, we get:

$$
\begin{aligned}
& \ 4.56 \times 10^6 \\
& \ 0.298 \times 10^6 \\
& \ 0.0365 \times 10^6 \\
& +\ 0.00721 \times 10^6 \\
& =\ 4.90171 \times 10^6
\end{aligned}
$$

The answer upon adding must be rounded to 2 significant figures to the right of the decimal point, thus giving 4.90×10^6 as the correct answer.

Generally speaking, you can simply enter the numbers into the calculator and let the calculator keep track of where the decimal portion is. However, you must then round off the answer to the correct number of significant figures.

Lastly, be warned about using the calculator. Students often push buttons without understanding the math behind what they are doing. Then, when the teacher questions their work, they say "Well, that's what the calculator said!" As if the calculator is to blame for the wrong answer. Remember, it is your brain that must be in charge and it is you that will get the points deducted for poor work, not the calculator.

Practice Problems

1) $(4.52 \times 10^{-5}) + (1.24 \times 10^{-2}) + (3.70 \times 10^{-4}) + (1.74 \times 10^{-3})$

2) $(2.71 \times 10^6) - (5.00 \times 10^4)$

Reminder: you must have the same exponent on each number of the problem.

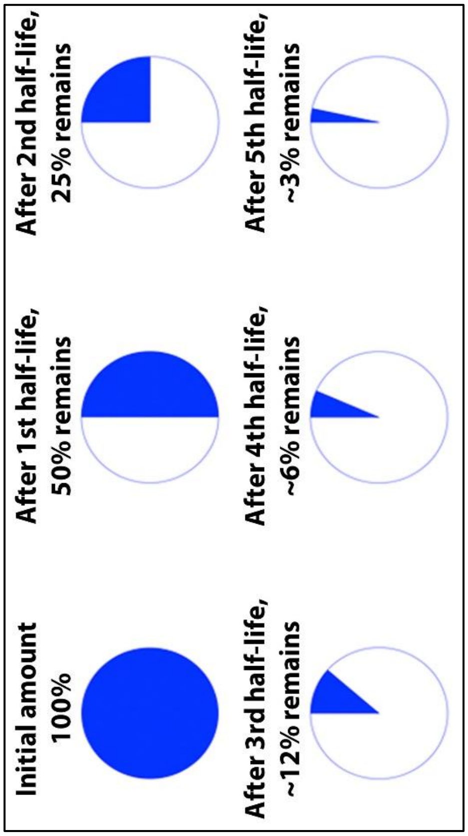
Reference Sheets for
Unit #2 – Nuclear Chemistry

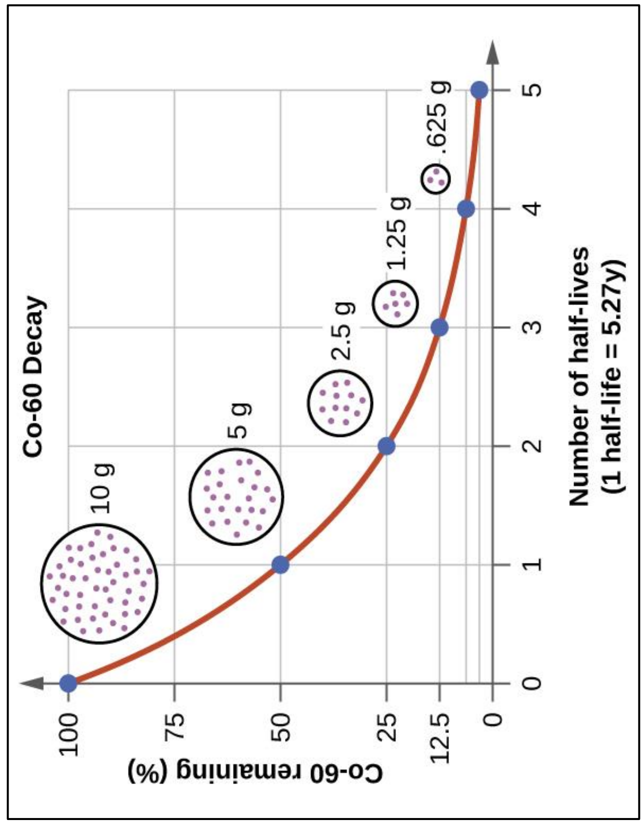
Nuclear Decay Organizer

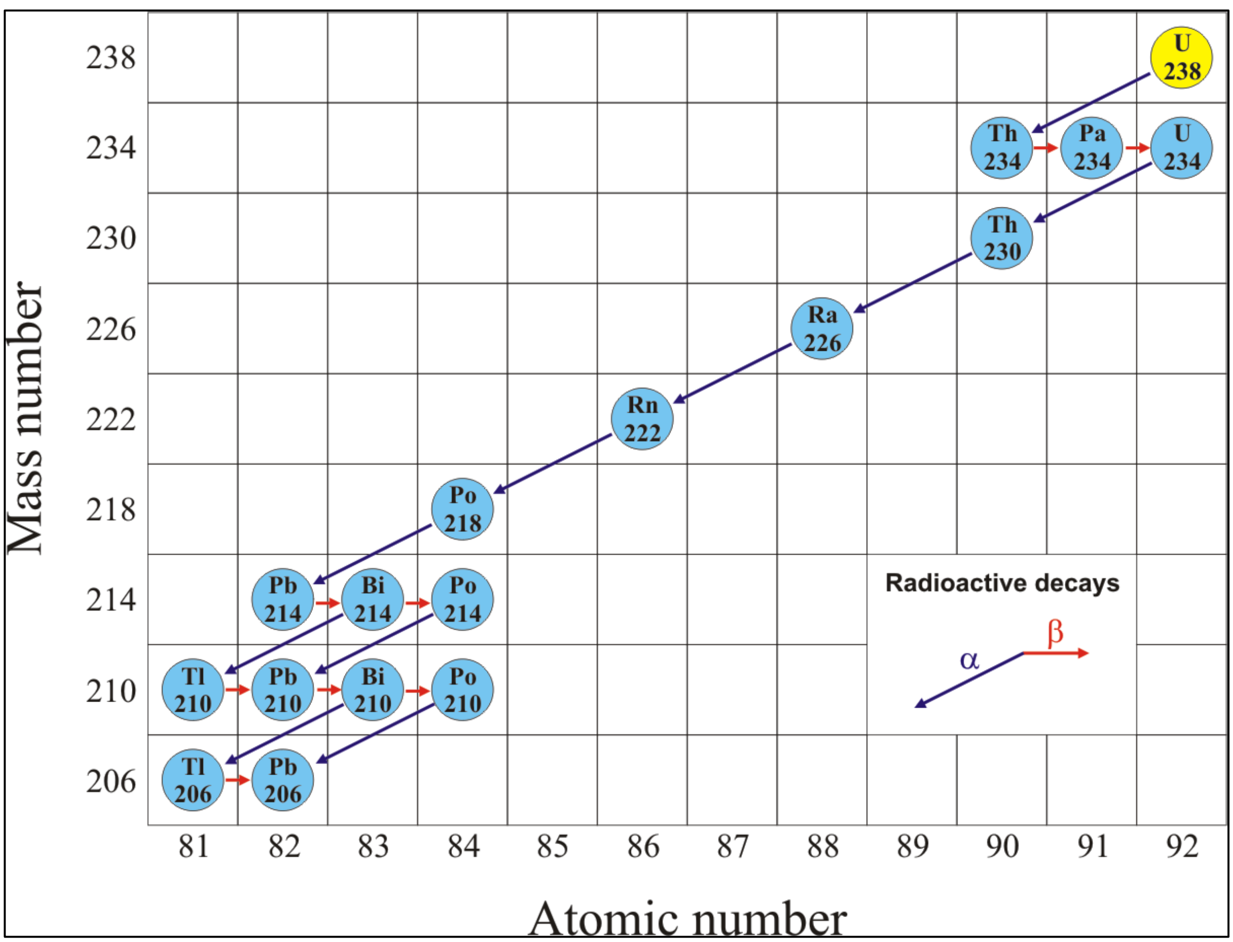
Students know the three most common forms of radioactive decay (alpha, beta, and gamma) and know how the nucleus changes in each type of decay.
Students know alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.
Students know some naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.

	Alpha Particle Emission	**Beta Particle Emission**	**Gamma Ray Emission**
Symbol	${}^{4}_{2}He^{2+}$ or ${}^{4}_{2}\alpha^{2+}$	${}^{0}_{-1}e$ or ${}^{0}_{-1}\beta$	${}^{0}_{0}\gamma$
Mass	Heavy	Light	No Mass
How it changes the nucleus	• Decreases the mass number by 4 • Decreases the atomic number by 2	• Converts a neutron into a proton • Increases atomic number by 1	No change to the nucleus
Penetration	Low	Medium	High
Protection provided by…	Skin	Paper, clothing	Lead
Danger	Low	Medium	High

Nuclear Info Sheet







Nuclear Info Sheet

Half Life Equation

$$A_E = A_s \times 0.5^n$$

A_E = amount ending
A_S = amount starting
n = number of half lives

of Half Lives

$$n = \frac{t}{h}$$

n = number of half-lives
t = time that has passed
h = length of a half life

% still radioactive or % decayed

$$\%_{still\ r.a.} = \frac{A_E}{A_S} \times 100$$

$$\%_{still\ r.a.} = 0.5^n \times 100$$

$$\%_{decayed} = 100 - \%_{still\ r.a.}$$

A_E = amount ending
A_S = amount starting
n = number of half lives

Solving for t, or h

$$\log\left(\frac{A_E}{A_S}\right) = n \times log(0.5)$$

Same as this version:

$$\log\left(\frac{A_E}{A_S}\right) = \frac{t}{h} \times log(0.5)$$

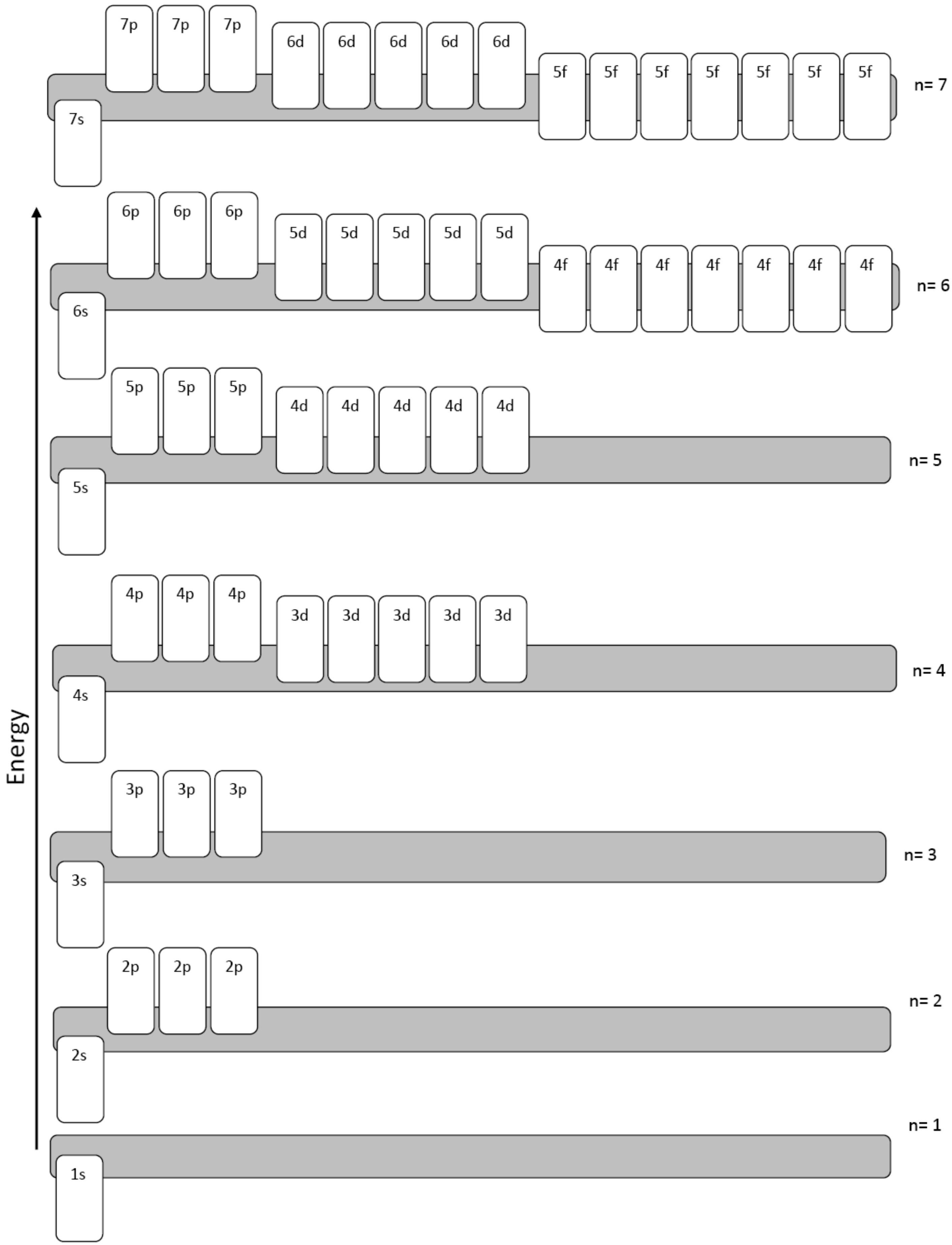
Simply isolate the variable you are trying to solve for

$$t = \frac{h \times \log\left(\frac{A_E}{A_S}\right)}{log(0.5)}$$

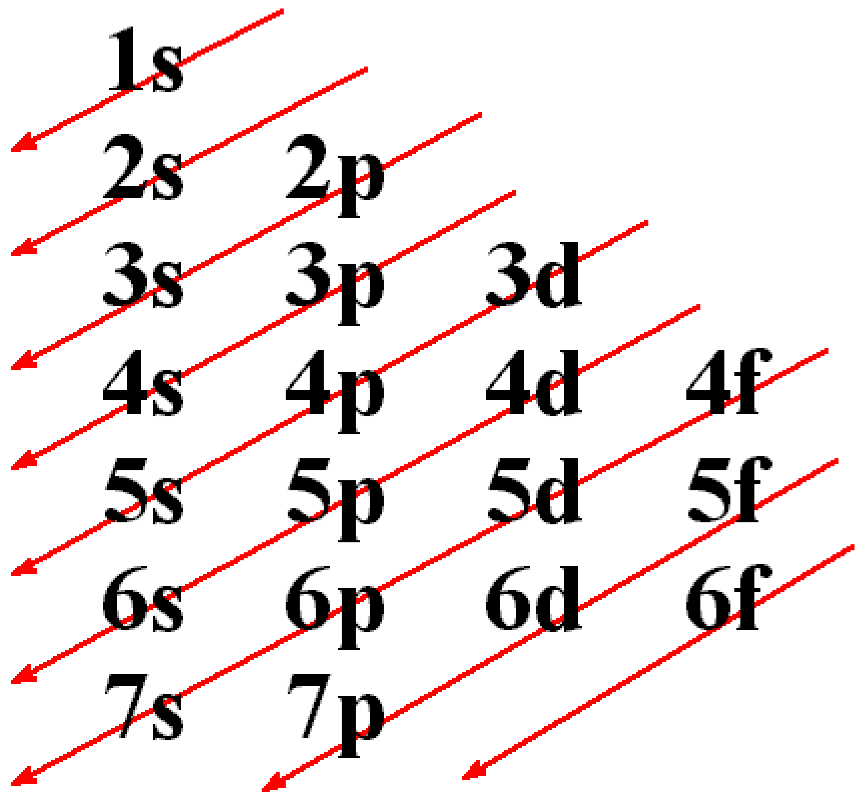
$$h = \frac{t \times \log(0.5)}{log\left(\frac{A_E}{A_S}\right)}$$

Reference Sheets for Unit #3 – Electrons

Orbital Diagram

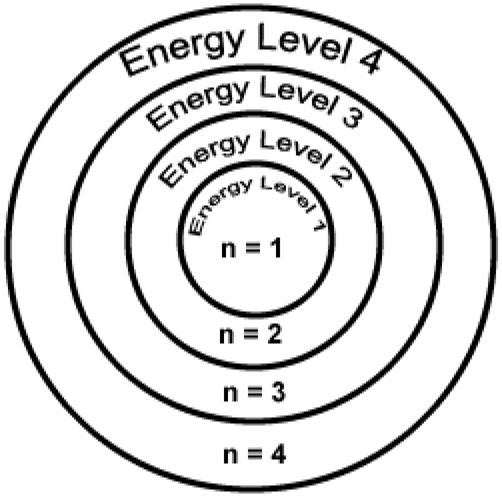


Order of Orbital Filling



Electron Orbitals

Quantum Numbers specify the properties of atomic orbitals and the properties of the electrons in orbitals
Orbitals are regions inside an energy level where the probability of finding an electron is very high.



Principal Quantum Number (*n*)	Sublevels in main energy level (*n* sublevels)	Number of orbitals (n^2)	Number of electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

A. Principal Quantum Number (*n*)
1. Indicates the main energy levels occupied by the electron
2. Values of n are positive integers
   a. n=1 is closest to the nucleus, and lowest in energy
3. The number of orbitals possible per energy level (or "shell") is equal to n^2

B. Angular Momentum Quantum Number
1. Indicates the shape of the orbital
2. Number of orbital shapes = *n*
   a. Shapes are designated *s, p, d, f*

C. Spin Quantum Number
1. Indicates the fundamental spin states of an electron in an orbital
2. Two possible values for spin, +1/2, -1/2
3. A single orbital can contain only two electrons, which must have opposite spin

<u>Electron Configurations</u>

1. Aufbau Principle
   a. An electron occupies the lowest-energy orbital that can receive it
2. Hund's Rule
   a. Orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin
3. Octet
   a. Highest energy level *s* and *p* electrons are filled (8 electrons)
   b. Characteristic of noble gases, Group 18
4. Noble gas configuration
   a. Outer main energy level fully occupied, usually (except for He) by eight electrons
   b. This configuration has extra stability

Element	**Configuration notation**	**Orbital notation**	**Noble gas notation**
Lithium	$1s^2 2s^1$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^1$
Beryllium	$1s^2 2s^2$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2$
Boron	$1s^2 2s^2 p^1$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^1$
Carbon	$1s^2 2s^2 p^2$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^2$
Nitrogen	$1s^2 2s^2 p^3$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^3$
Oxygen	$1s^2 2s^2 p^4$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^4$
Fluorine	$1s^2 2s^2 p^5$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^5$
Neon	$1s^2 2s^2 p^6$	____ ____ ____ ____ ____ 1s 2s 2p	$[He]2s^2 p^6$

Reference Sheets for
Unit #4 – Periodic Table

Periodic Table Structure Info Sheet

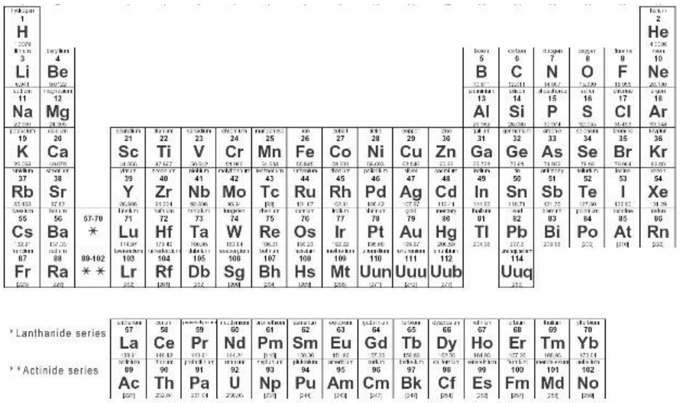
Periods (rows) →

Groups (columns) ↑

Mendeleev – Organized PT based on atomic masses & properties (almost right…)

Moseley – Organized PT based on atomic numbers (the way we do it now!)

Three classes of elements: Metals, non-metals, metalloids/semi-metals



Color code each class of element. Make a key here

Some videos about the structure & creation of the periodic table

- https://tinyurl.com/n4o9dns
- https://tinyurl.com/abq96op
- https://tinyurl.com/q2z47cl

Metal Properties:

Chemical Prop.	Physical Prop.
Few electrons in VALENCE shell (outer shell)	Ductile Malleable
Lose electrons easily	Good conductors
POSITIVE charge like Ca^{2+}	Shiny
Make Cations	Solid at room temp

Non-metal Properties:

Chemical Prop.	Physical Prop.
Almost full, or totally full valence shell	NOT Ductile NOT malleable
Tend to gain electrons	BAD conductors
NEGATIVE charge like N^{3-}	Mostly solid
Make ANIONS	Some are gas at room temp

Semi-metal Properties:

Chemical Prop.	Physical Prop.
Most have half full valence shell	Have properties of metals AND non-metals
Make anions OR cations depending on their environment	No way to know which properties of each

Things in the same period have:

Increasing atomic # and mass L→R

Same number of energy levels

Period 1 has 1 level

Period 2 has 2 levels etc…

Things in the same group have:

Increasing atomic # and mass ↓

Same number of valence electrons

Exceptions: d and f block

Similar physical and chemical properties

b/c they have same # of valence e^-s

Valence Electrons:

Outer electrons

Matches the "A" column number

1A has 1 v.e^-, 2A has 2v.e^-, etc.

d and f blocks don't follow rules

Shielding and Z_{eff}:

Outer electrons have trouble "seeing" the protons in the nucleus – the nucleus is "shielded" by the electrons. You can calculate how much "shielding" there is by calculating the "Effective Nuclear Charge"

$$Z_{eff} = Z - S$$

Z_{eff} = effective nuclear charge

Z = atomic #

S = all non-valence electrons

Periodic Table Structure Info Sheet

- ☐ Alkali metals
- ☐ Alkaline earth metals
- ☐ Transition metals
- ☐ Rare earth metals
- ☐ Metalloids (semi-metals)
- ☐ Other metals
- ☐ Non-metals
- ☐ Halogens
- ☐ Noble Gases

The Periodic Table of the Elements

Element name → Mercury
Atomic # → 80
Symbol → Hg
Avg. Mass → 200.59

1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Hydrogen 1 H 1.01																		Helium 2 He 4.00
Lithium 3 Li 6.94	Beryllium 4 Be 9.01												Boron 5 B 10.81	Carbon 6 C 12.01	Nitrogen 7 N 14.01	Oxygen 8 O 16.00	Fluorine 9 F 19.00	Neon 10 Ne 20.18
Sodium 11 Na 22.99	Magnesium 12 Mg 24.31												Aluminum 13 Al 26.98	Silicon 14 Si 28.09	Phosphorus 15 P 30.97	Sulfur 16 S 32.07	Chlorine 17 Cl 35.45	Argon 18 Ar 39.95
Potassium 19 K 39.10	Calcium 20 Ca 40.08		Scandium 21 Sc 44.96	Titanium 22 Ti 47.88	Vanadium 23 V 50.94	Chromium 24 Cr 52.00	Manganese 25 Mn 54.94	Iron 26 Fe 55.85	Cobalt 27 Co 58.93	Nickel 28 Ni 58.69	Copper 29 Cu 63.55	Zinc 30 Zn 65.39	Gallium 31 Ga 69.72	Germanium 32 Ge 72.61	Arsenic 33 As 74.92	Selenium 34 Se 78.96	Bromine 35 Br 79.90	Krypton 36 Kr 83.80
Rubidium 37 Rb 85.47	Strontium 38 Sr 87.62		Yttrium 39 Y 88.91	Zirconium 40 Zr 91.22	Niobium 41 Nb 92.91	Molybdenum 42 Mo 95.94	Technetium 43 Tc (98)	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.91	Palladium 46 Pd 106.42	Silver 47 Ag 107.87	Cadmium 48 Cd 112.41	Indium 49 In 114.82	Tin 50 Sn 118.71	Antimony 51 Sb 121.76	Tellurium 52 Te 127.60	Iodine 53 I 126.90	Xenon 54 Xe 131.29
Cesium 55 Cs 132.91	Barium 56 Ba 137.33	57-70 *	Lutetium 71 Lu 174.97	Hafnium 72 Hf 178.49	Tantalum 73 Ta 180.95	Tungsten 74 W 183.84	Rhenium 75 Re 186.21	Osmium 76 Os 190.23	Iridium 77 Ir 192.22	Platinum 78 Pt 195.08	Gold 79 Au 196.97	Mercury 80 Hg 200.59	Thallium 81 Tl 204.38	Lead 82 Pb 207.20	Bismuth 83 Bi 208.98	Polonium 84 Po (209)	Astatine 85 At (210)	Radon 86 Rn (222)
Francium 87 Fr (223)	Radium 88 Ra (226)	89-102 **	Lawrencium 103 Lr (262)	Rutherfordium 104 Rf (267)	Dubnium 105 Db (268)	Seaborgium 106 Sg (271)	Bohrium 107 Bh (272)	Hassium 108 Hs (270)	Meitnerium 109 Mt (276)	Darmstadtium 110 Ds (281)	Roentgenium 111 Rg (280)	Copernicium 112 Cn (285)	Ununtrium 113 Uut (284)	Ununquadium 114 Uuq (289)	Ununpentium 115 Uup (288)	Ununhexium 116 Uuh (293)	Ununseptium 117 Uus (294?)	Ununoctium 118 Uuo (294)

*lanthanides	Lanthanum 57 La 138.91	Cerium 58 Ce 140.12	Praseodymium 59 Pr 140.91	Neodymium 60 Nd 144.24	Promethium 61 Pm (145)	Samarium 62 Sm 150.36	Europium 63 Eu 151.97	Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.93	Dysprosium 66 Dy 162.50	Holmium 67 Ho 164.93	Erbium 68 Er 167.26	Thulium 69 Tm 168.93	Ytterbium 70 Yb 173.04
**actinides	Actinium 89 Ac (227)	Thorium 90 Th 232.04	Protactinium 91 Pa 231.04	Uranium 92 U 238.03	Neptunium 93 Np (237)	Plutonium 94 Pu (244)	Americium 95 Am (243)	Curium 96 Cm (247)	Berkelium 97 Bk (247)	Californium 98 Cf (251)	Einsteinium 99 Es (252)	Fermium 100 Fm (257)	Mendelevium 101 Md (258)	Nobelium 102 No (259)

Periodic Trends Info Sheet

IONIZATION ENERGY

The minimum amount of energy required to *remove* the most loosely bound electron, the valence electron, of an isolated neutral gaseous atom to form a cation. It is quantitatively expressed in symbols as:

$$X + \text{energy} \rightarrow X^{+} + e^{-}$$

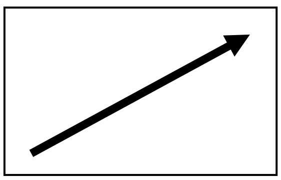
In other words…

How much energy does it take to take away an electron from an atom?

Trend

Increases →

Increases ↑

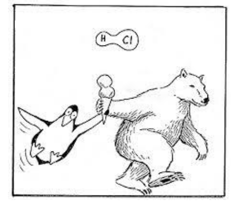


Highest = Fluorine only higher are Noble Gases (Rn)

Lowest = Francium

ELECTRONEGATIVITY

The tendency of an atom to attract a shared pair of electrons (or electron density) towards itself.



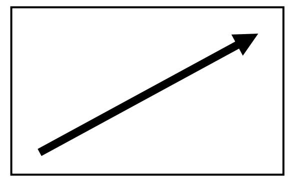
In other words…

How strong is the atom? How hard can it pull on electrons when sharing them with another atom?

Trend

Increases →

Increases ↑

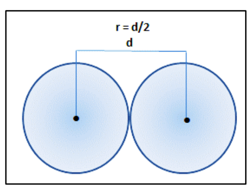


Highest = Fluorine

Lowest = Francium only lower are Noble Gases (He)

ATOMIC RADIUS

Usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding cloud of electrons. Since the boundary is not a well-defined physical entity, there are various non-equivalent definitions of atomic radius.



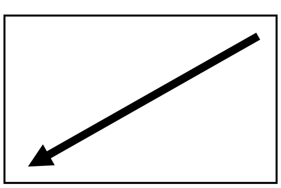
In other words…

Half the diameter of an atom. Hard to measure because atoms do not have actual tangible edges. Lots of ways to measure it.

Trend

Increases ←

Increases ↓



Highest = Francium

Lowest = Helium

ELECTRON AFFINITY

The amount of energy released* when an electron is *added* to one mole of a neutral atom, or molecule, in the gaseous state to form a negative ion. Usually written as a ΔE value.

$$X + e^{-} \rightarrow X^{-} + \text{energy}$$

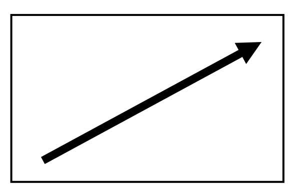
In other words…

How happy is the atom to gain that new electron? Really happy – much more stable, releases lots of energy. Not happy – not as much stability, doesn't release lots of energy.

Trend

Increases →

Increases ↑



Highest = Fluorine

Lowest = Francium only lower are Noble Gases (zero)

*Some electron affinities are actually positive ΔE – meaning energy is absorbed. They are not commonly talked about or used though.

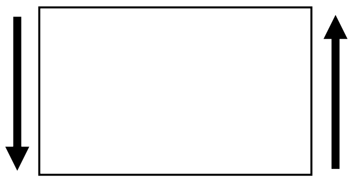
Periodic Trends Info Sheet

REACTIVITY

The impetus for which a chemical substance undergoes a chemical reaction, either by itself or with other materials, with an overall release of energy.

In other words…
How quickly, violently, readily, does an element undergo certain reactions. More reactivity means faster, more violent, easier reaction with lots of energy released.

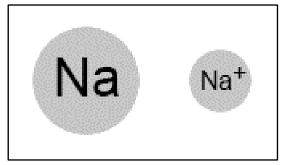
Trend
Metals
Increase ↓←
Non-metals
Increases ↑→



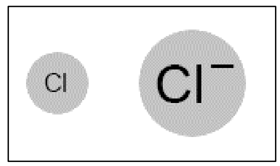
Highest metal = Francium
Highest non-metal = Fluorine
Lowest non-metal = Noble gases (He)

IONIC RADIUS

Cations:
Get smaller when they lose electrons



Anions:
Get larger when they gain electrons



EFFECTIVE NUCLEAR CHARGE - Zeff

The pull that the nucleus has on electrons

$$Zeff = Z - S$$

Z = # of protons
S = Core/Inner electrons (# of electrons in previous noble gas + any d or f electrons)

SUBSEQUENT IONIZATIONS

1st ionization energy
$X \rightarrow X^{+} + e^{-}$
2nd ionization energy
$X^{+} \rightarrow X^{2+} + e^{-}$
3rd ionization energy
$X^{2+} \rightarrow X^{3+} + e^{-}$

Ionization energy increases with each subsequent ionization because there is more attraction between the nucleus and the valence electrons each time you lower the number of valence electrons.

There is a huge leap in ionization energy once an atom loses all it's valence electrons because it now looks like a noble gas and really doesn't want to let go of any more electrons!

SHIELDING

When the inner/core electrons repel the valence electrons and prevents them from seeing the nucleus. Decreases how strongly the electrons are held onto by the nucleus.

BREAKS IN PATTERNS

There are *many* examples of elements that do not appear to follow the general trends typically described. This can be due to a variety of reasons. Here are two of many reasons why this can happen.

- Half-filled orbitals have slightly more stability than expected
  Example: p orbital set: ↑ ↑ ↑
- Unpairing an electron takes slightly more energy than expected
  Example: p orbital set: ↑↓ ↑ ↑

Warning…

- Don't over think this stuff.
- You can talk yourself into backwards answers.
- Focus on the fact that there are only a set number of trends to learn.
- Practice explaining each trend until you can do it in your sleep!
- There will ALWAYS be exceptions. Don't worry about that – focus on the pattern and answer questions based on the patterns.

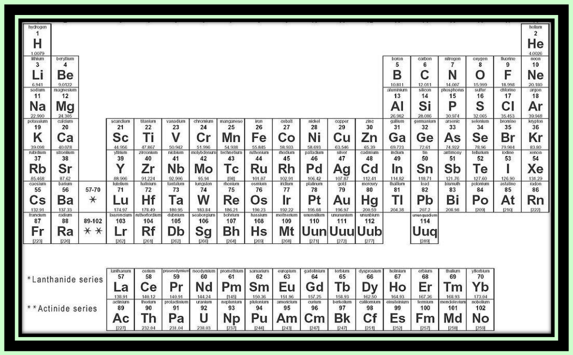
Warning…

- There is about to be a lot of notes because it takes a lot of words to explain
- You don't need EVERY word written down to understand it.
- Focus on listening and understanding.
- You can add to your notes at home.
- Capture enough to pay attention, leave space to come back and add/annotate.

Warning…

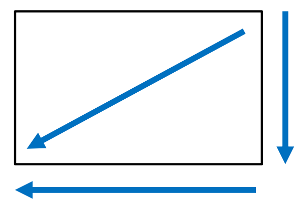
<u>Make sure you capture:</u>
What
How
Why
<u>Make sure you can tell me:</u>
What
How
Why

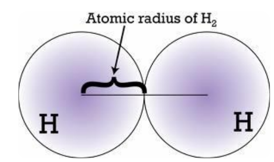
Periodic Trends



Atomic Radius

ATOMIC RADIUS

What	How
• ½ the distance between two bonded nuclei • Cant measure to the edge b/c orbitals aren't tangible!	

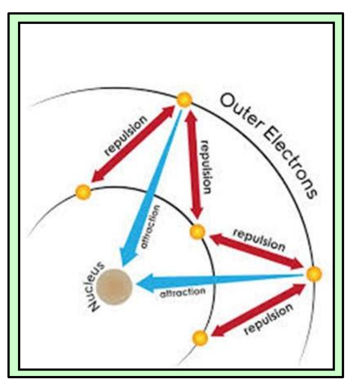


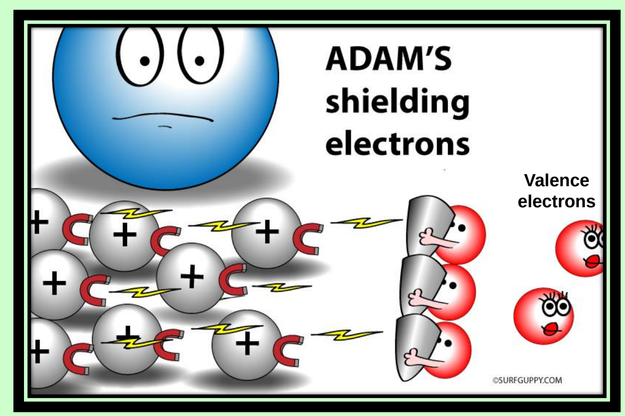
ATOMIC RADIUS

Why	
INCREASES DOWN	**DECREASES TO RIGHT**
• Adding energy levels Smaller "effective nuclear charge"	• Adding a proton = bigger change than adding an e-
• Inner e- keep valence e- from "feeling" the nucleus	• More protons pull the valence electrons in closer - "Greater Effective Nuclear Charge"
• Outer e-s are not as pulled in by the protons in the nucleus – there is more "shielding" by the inner electrons	• No increase in shielding b/c no new energy levels

Effective Nuclear Charge (Z_{eff})
The relative attraction the valence electrons have for the protons in the nucleus

Shielding Effect
The inner shell electrons repel the outer valence electrons – keeps the valence e- from "feeling" the nucleus. More energy levels - lowers the effective nuclear charge





Calculating Effective Nuclear Charge

The relative attraction the valence electrons have for the protons in the nucleus

$$Z_{eff} = Z - S$$

Z = nuclear attraction = # protons
S = the core/inner e- shielding the valence e-'s
= the total number of e- minus the e- in the highest occupied s and p energy levels
= *(# of e- in previous noble gas + any d or f e-'s past the noble gas in the element)*

Calculating Effective Nuclear Charge

$Z_{eff} = Z - S$

Magnesium
Z = 12 protons
S = Previous noble gas
= Neon = 10 electrons
Zeff = 12 - 10 = 2

Aluminum
Z = 13 protons
S = Previous noble gas
= Neon = 10 electrons
Zeff = 13 – 10 = 3

Aluminum is smaller – valence electrons are pulled in harder by the nucleus

IONIC RADIUS

What	How
The radius of an ion	Cation – always smaller
Cation – lost electrons Anion – gained electrons	Anion – always bigger

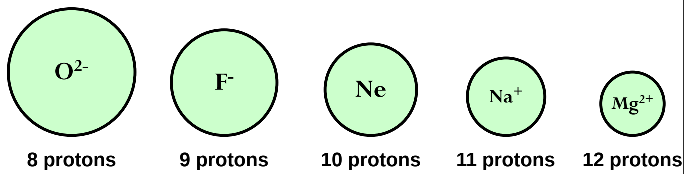
IONIC RADIUS

Why	
CATION SMALLER • Reduced repulsion between electrons • If you lose enough electrons you even drop down an energy level! Much smaller!	**ANION LARGER** • Extra valence electrons repel each other a bit more so it gets larger.

Isoelectric Species

Atoms/Ions that have the same number of e-

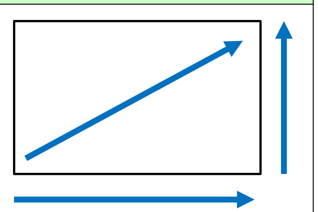
All these examples are $1s^2 2s^2 2p^6$

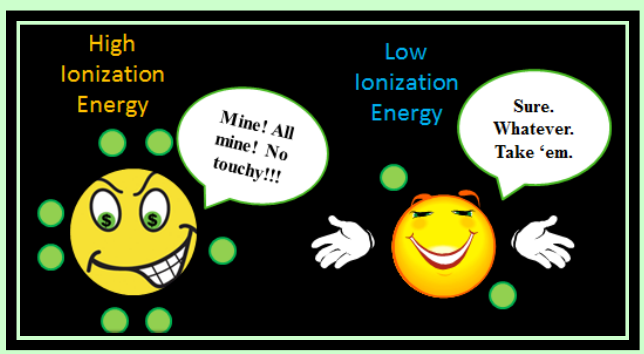


Increased protons can pull harder on the valence electrons – greater effective nuclear charge – so the radius is smaller even though they have the same number of electrons and energy levels

Ionization Energy

IONIZATION ENERGY

What	How
The energy required to remove on electron from a neutral atom of an element	 Noble Gases are HIGHEST! They REALLY don't want to let go of an e-



IONIZATION ENERGY

Why	
DECREASES DOWN • Electrons are further from nucleus in higher energy levels • Increased shielding from core e- causes nucleus to not pull as hard on valence e- • Lower effective nuclear charge –so they are more easily removed	**INCREASES TO RIGHT** • Closer to having a full stable valence shell • Increased effective nuclear charge means nucleus is pulling harder on the valence e- so they are harder to remove

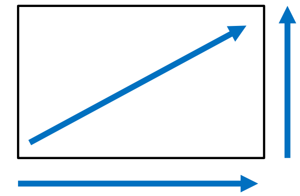
Subsequent Ionizations

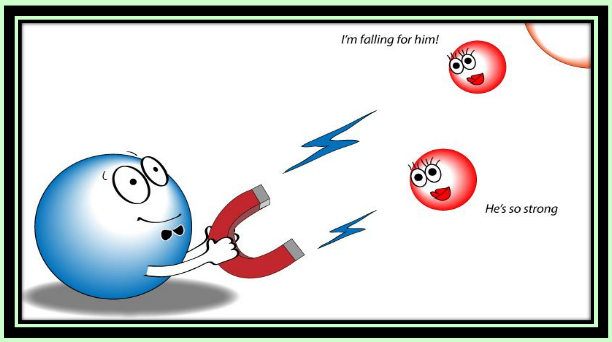
Every time you take an e- away it gets harder to take the next one. Radius is getting smaller, so nucleus can pull harder on the valence - harder to remove the next one. HUGE LEAP in I.E. once it's achieved noble gas configuration – why would it want to lose another one?!

Element	IE_1	IE_2	IE_3	IE_4
Na	496	4560		
Mg	738	1450	7730	
Al	578	1820	2750	11,600

Electronegatvity

ELECTRONEGATIVITY

What	How
A measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound *How strongly can one atom pull on the electrons being shared in a bond.*	 Noble Gases are LOWEST! They DON'T CARE about attracting electrons!



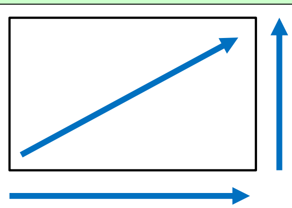
ELECTRONEGATIVITY

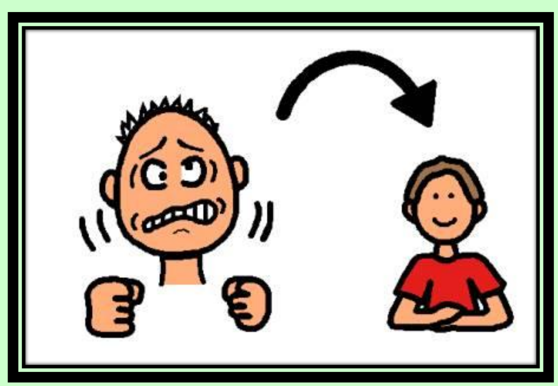
Why

DECREASES DOWN	INCREASES TO RIGHT
• e- are further from nucleus in higher energy levels	• Smaller radius, increased effective nuclear charge
• Increased shielding from core e- causes nucleus to not pull as hard on valence e-	• Nucleus is pulling harder on the valence electrons – which is where the bonding is occurring.
• So nucleus doesn't pull as hard on the bonding e-'s from another atom	

Electron Affinity

ELECTRON AFFINITY

What	How
How much energy is released when the atom gains an electron to make a negative ion. *How much stability does it gain once it is an anion. More energy released – more stable.*	 Noble Gases are LOWEST! They DON'T CARE about attracting electrons!



ELECTRON AFFINITY

Why

DECREASES DOWN	INCREASES TO RIGHT
• Electrons are further from nucleus in higher energy levels • Increased shielding from core e-'s causes the nucleus to not pull as hard on valence e-'s • So atom doesn't notice as much if it gains an electron – doesn't gain much stability	• Closer to filling valence shell – noble gas configuration is most stable

Reactivity

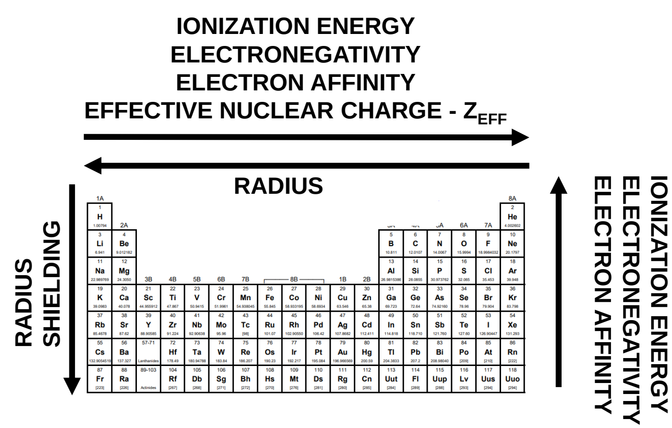
REACTIVITY

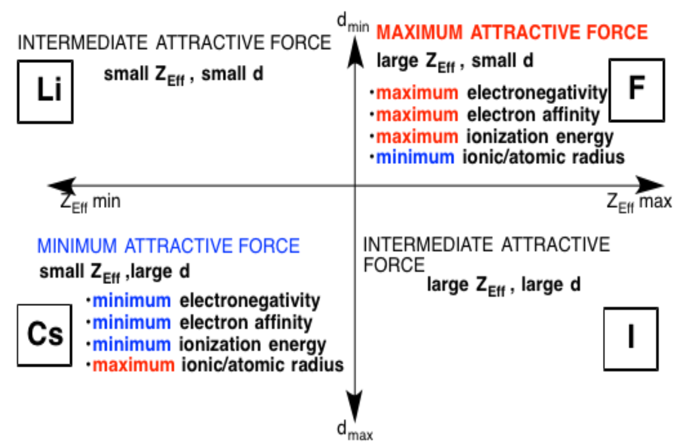
What	How
Elements in the same group have similar types of behaviors <u>because they have the same number of valence e-</u> BUT The MAGNITUDE of their reactions changes!	

REACTIVITY

Why	
<u>METALS INCREASE DOWN</u> • Larger radius and increased shielding means lower ionization energy so it is easier to remove electrons	**<u>NON-METALS INCREASE UP</u>** • Smaller radius and greater effective nuclear charge means higher electronegativity and electron affinity so it can attract an electron easier

Summary





<u>Brainiac Video</u> – *note: they augmented the reactions, but it is such a fun, silly, memorable video I think it is still worth watching* ☺

<u>Disposal of Sodium</u> – *old footage from WWII. Neat to see such old footage and how they actually disposed of the sodium after the war!*

Quick summary. Also has a quick but good explanation of some <u>exceptions</u> to the trends

https://www.youtube.com/watch?v=hePb00CqvP0

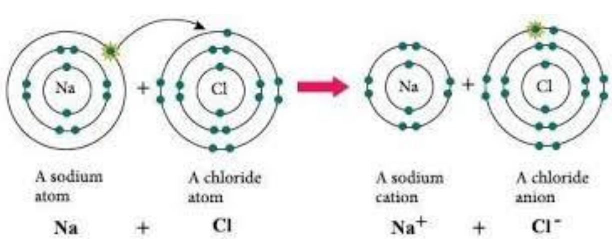
Reference Sheets for
Unit #5 – Bonding and Structure

DVHS Chemistry
Types of Bonds Reader

Directions:
Read this page and take notes and/or annotate it. We will not be doing a traditional lecture on this material because it is mostly review material. There is potentially information in here you may not be familiar with. If you come across anything you do not understand you need to ask about it! At the end there are questions to check that you were able to follow and grasp the material talked about here. These are selections of reading by various people, credit given when possible.

Types of Bonds – *by Janet Rae-Dupree, Pat DuPree. Edited to suit our purposes here.*
Atoms tend to arrange themselves in the most stable patterns possible, which means they have a tendency to complete or fill their valence shells. They join with other atoms in order to do that. The force that holds atoms together as "molecules" or "compounds" is referred to as a *chemical bond.* There are three main types of bonds, and they each may have some sub categories that are more specific.

Ionic Bond – *by Janet Rae-Dupree, Pat DuPree. Edited to suit our purposes here.*
Ionic bonding involves the transfer of electrons. One atom gains one or more electrons, while another atom loses one or more electrons. The atom that lost an electron carries a negative charge (anion), and the atom that lost an electron carries a positive charge (cation). Because opposite charges attract, the atoms bond together to form a compound. The electrostatic attraction is the "bond." Ionic bonds are formed when one atom has a low ionization energy (cation, metal) and another atom has a high electron affinity (anion, nonmetal). That essentially results in needing a metal and a non-metal to form an ionic bond.
Examples: NaBr – sodium is a metal, bromine is a nonmetal. Na turn into Na^+ and gives its electron to Br which turns into Br^-
MgF_2 – magnesium is a metal, fluorine is a nonmetal. Mg turns into Mg^{2+} and gives two electrons to the fluorine atoms. Each fluorine turns into F^- so you need two fluorine atoms in order to balance out the +2 charge of the magnesium.

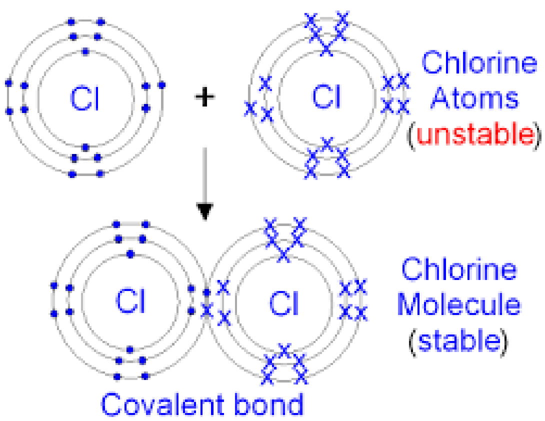


Covalent Bond – *by Janet Rae-Dupree, Pat DuPree. Edited to suit our purposes here.*

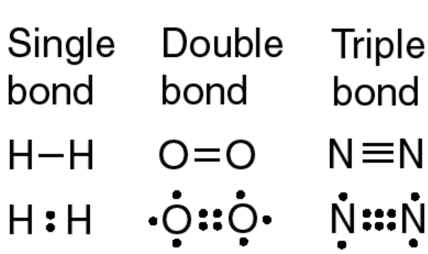
Covalent Bonds involve the sharing of electrons between two atoms. The pair of shared electrons holds the atoms close together and this sharing of electrons is the bond. Covalent bonds form when there is not an atom with a sufficiently low ionization energy to simply "give up" the electron, and the other atom does not have a sufficiently high electronegativity to "steal" the electron completely. That essentially results in needing two (or more) nonmetals to form the bond. One thing to note is that when a covalent bond is formed, neither atom truly has a complete valence shell. It "feels" like it has a full shell due to the sharing, but it doesn't completely "own" those electrons.

Examples: H_2O *– each hydrogen and each oxygen "donate" an electron to a shared bon. Both hydrogen and oxygen are nonmetals. They are not strong enough to steal the electron completely from the other atom.*

Cl_2 *– Each chlorine only has 7 valence electrons. If they each "donate" one to share then they are tricked into thinking they each have 8 valence electrons.*



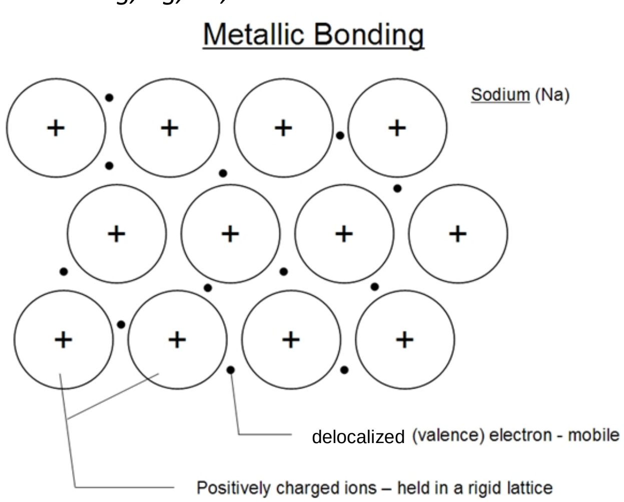
Nonmetals can donate one electron each to form a single bond (2 electrons being shared), or they can each donate two electrons to form a double bond (4 electrons being shared), or they can each donate three electrons to form a triple bond (6 electrons being shared).



Metallic Bond –

Metallic bonds form when one or more types of metals are together. The vacant p and/or d orbitals in the metal atoms' outer energy levels overlap, and allow outer valence electrons to move freely through the metal. The valence electrons basically "detach" and float around as a "sea of electrons." We say that the electrons are "delocalized" and no longer belong to any particular atom. This sea of electrons effect allows the electrons to flow through the material – this is how electricity can be conducted through something like a metal wire. All the delocalized electrons are flowing from one end of the wire to the other because they are not attached to their nuclei anymore. If you have more than one type of metal atom present it is called an "alloy." These metallic mixtures can form unique properties that can be useful.

Example: Cu – note: you do not need to include a subscript for metallic materials. It is not practical to tell someone how many atoms are present in a chunk of metal. Because the electrons essentially belong to every atom, it isn't really a molecule or compound in the traditional sense. If you just see a metallic element listed by itself you can assume there are lots present and it is a metallic bond – like Mg, Ag, Cu, Fe



DVHS Chemistry
Nomenclature Reader

Will be stamped and count for points towards your binder check grade!

Name: **Period:** **Seat#:**

Directions:
Read this page and take notes and/or annotate it. We will not be doing a traditional lecture on this material because it is mostly review material. There is potentially information in here you may not be familiar with. If you come across anything you do not understand you need to ask about it! At the end there are questions to check that you were able to follow and grasp the material talked about here. These are selections of reading by various people, credit given when possible.

Nomenclature – *by the International Union of Pure and Applied Chemistry Edited to suit our purposes here.*

A system of rules for naming chemical compounds. The International Union of Pure and Applied Chemistry (IUPAC) devised the system for naming compounds in order to ensure uniformity, consistency, and avoid ambiguity.

Ionic Compounds

For simple binary ionic compounds (ionic compounds composed of one kind of metal and one kind of nonmetal) the cation name comes first and then the anion. Simple *cations* take the name of their element, for example, the name for K^+ would be potassium, and the name for Zn^{2+} would be Zinc. Elements that form two or more ions need a roman numeral to denote their charge. For example, many transition metals can form ions with different charges, although they will always form cations. For example, copper can form Cu^+ which would be named copper (I), or it can form Cu^{2+}, which would be named copper (II).

Simple *anions* use the base name of the element, but end in *–ide.* For example, F^- would be named fluoride, and I^- would be named iodide.

Put the name of the cation and anion together to name the ionic compound. For example, NaCl is named sodium chloride. ZnS is named Zinc sulfide.

Compounds made with polyatomic ions do not change their anion name – keep the special polyatomic names for both cation and anion! NH_4I is named Ammonium Iodide, Li_2CO_3 is named Lithium carbonate.

Note: Some stable ions do not have noble gas configurations! Polyatomic Ions (ions with several types of atoms have names that will need to be memorized. Check your Common Ions table for the ions you will be required to memorize.

Naming Covalent Molecules – *by Eden Francis. Edited to suit our purposes here.*

Covalent molecules use a different system for nomenclature. Simple covalent molecules are generally named by using **Greek prefixes** to indicate how many atoms of each element are shown in the formula and the ending of the last element is changed to **-ide**.

The **mono-** prefix is usually not used for the first element in the formula. Some double vowels are omitted to help with ease of pronouncing the molecule name. The "o" and "a" endings of these prefixes commonly are dropped when they are attached to "oxide." See the table below for the Greek prefixes you will need to memorize.

number of atoms	prefix	example	
1	mono	NO	nitrogen monoxide
2	di	NO_2	nitrogen dioxide
3	tri	N_2O_3	dinitrogen trioxide
4	tetra	N_2O_4	dinitrogen tetraoxide
5	penta	N_2O_5	dinitrogen pentaoxide
6	hexa	SF_6	sulphur hexa fluoride
7	hepta	IF_7	iodine hepta fluoride
8	octa	P_4O_8	tetra phosphur decoxide
9	nona	P_4S_9	tetra phusphur nona sulphide
10	deca	AS_4O_{10}	tetra arsinic decoxide

Diatomic Molecules

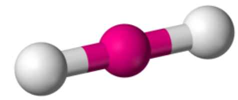
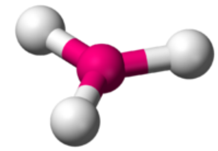
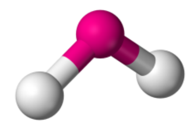
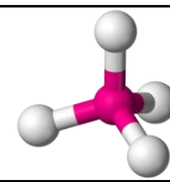
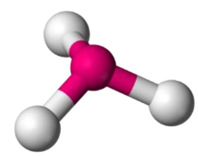
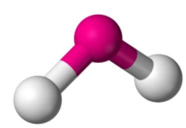
A molecule composed of only two atoms is said to be diatomic. There are several diatomic molecules made of the same element that you will need to memorize. Luckily, there is a mnemonic device for this. ***H***orses ***N***eed ***O***ats ***F***or ***Cl***ear ***Br***own Eyes ***(I)*** will help you remember that H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2 are all diatomic molecules. Another mnemonic is "H-7" which reminds you that there are seven diatomic elements, they make the shape of a seven on the periodic table, starting with N, and that Hydrogen is one of the diatomic elements.

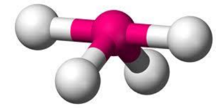
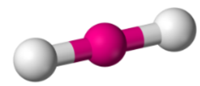
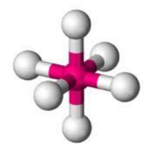
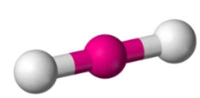
Practice – *We will review these names in class!*

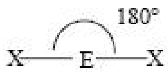
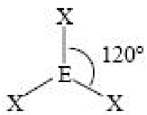
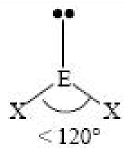
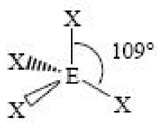
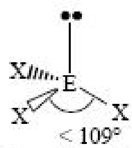
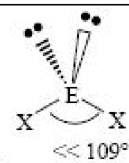
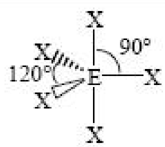
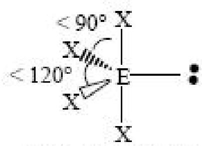
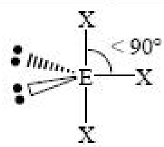
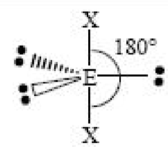
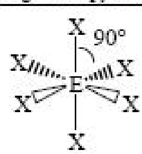
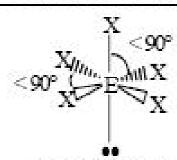
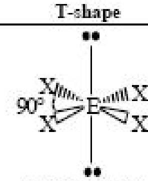
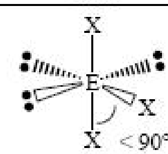
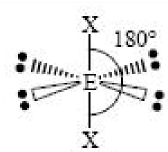
1. **CO**
2. **CO_2**
3. **S_4N_2**
4. **N_2O_6**
5. **PF_3**

VSEPR

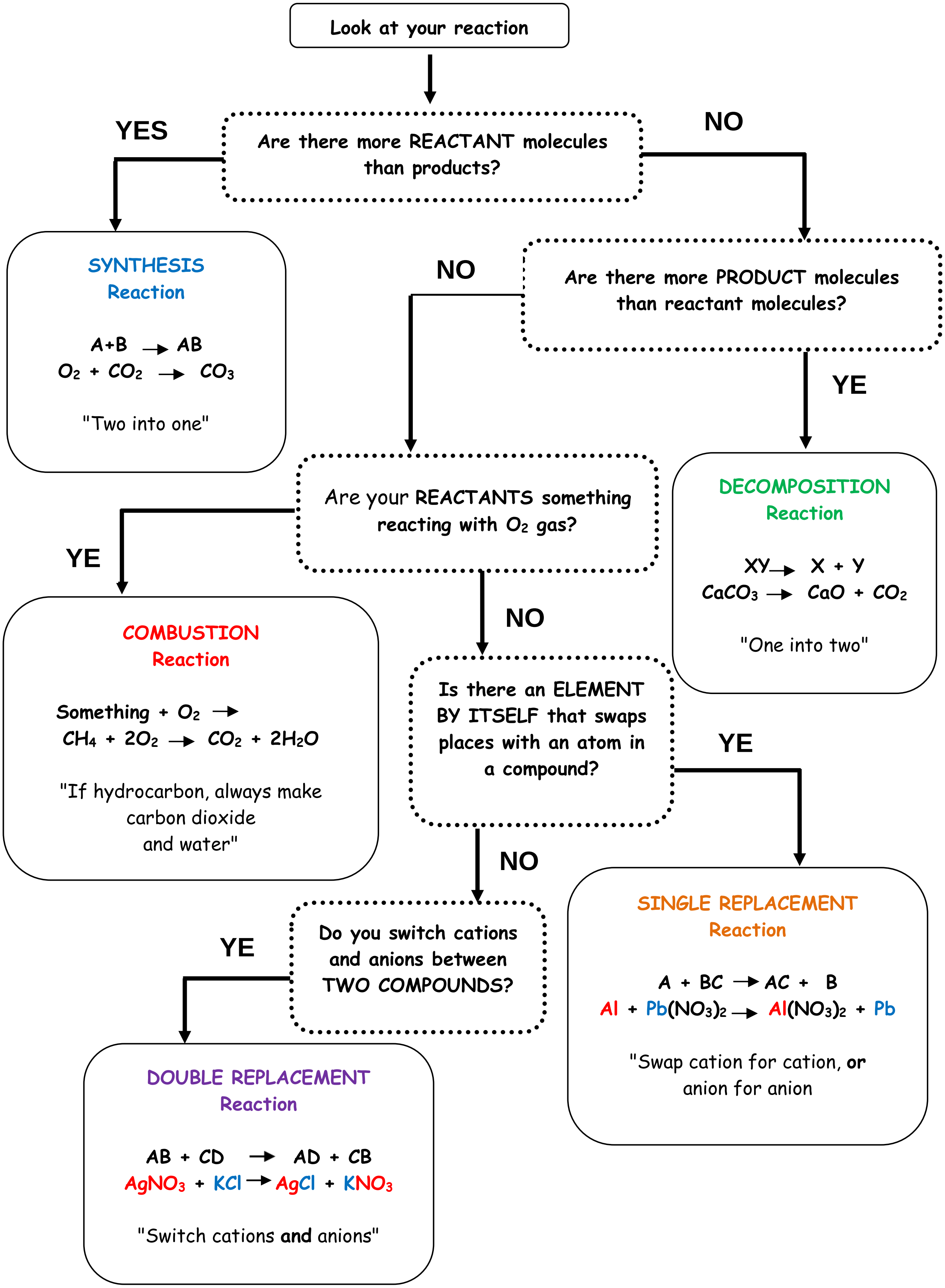
Predicting Molecular Geometry and Hybridization

Electron Groups	Bonding Groups	Lone Pairs	Electron Geometry *(Hybridization)*	Molecular Geometry *(VSEPR class)*	Approximate Bond Angles	Geometry Examples
2	2	0	Linear *(sp)*	Linear *(AX_2)*	180	
3	3	0	Trigonal Planar *(sp^2)*	Trigonal Planar *(AX_3)*	120	
	2	1		Bent *(AX_2E)*		
4	4	0	Tetrahedral *(sp^3)*	Tetrahedral *(AX_4)*	109.5	
	3	1		Trigonal Pyramidal *(AX_3E)*		
	2	2		Bent *(AX_2E_2)*		

Electron Groups	Bonding Groups	Lone Pairs	Electron Geometry *(Hybridization)*	Molecular Geometry *(VSEPR class)*	Approximate Bond Angles	Geometry Examples
5	5	0	Trigonal Bipyramidal *(sp^3d)*	Trigonal Bipyramidal *(AX_5)*	120 (in plane) 90 (above and below)	
	4	1		Seesaw *(AX_4E)*		
	3	2		T-Shaped *(AX_3E_2)*		
	2	3		Linear *(AX_2E_3)*	180	
6	6	0	Octahedral *(sp^3d^2)*	Octahedral *(AX_6)*	90	
	5	1		Square Pyrimidal *(AX_5E)*		
	4	2		Square Planar *(AX_4E_2)*		
	3	3		T-Shaped *(AX_3E_3)*		
	2	4		Linear *(AX_2E_4)*		

VSEPR Geometries					
Steric No.	<u>Basic Geometry</u> 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 **Linear**				
3	 **Trigonal Planar**	 **Bent or Angular**			
4	 **Tetrahedral**	 **Trigonal Pyramid**	 **Bent or Angular**		
5	 **Trigonal Bipyramid**	 **Sawhorse or Seesaw**	 **T-shape**	 **Linear**	
6	 **Octahedral**	 **Square Pyramid**	 **Square Planar**	 **T-shape**	 **Linear**

Reference Sheets for
Unit #6 – Reactions


Look at your reaction
Are there more REACTANT molecules than products?
YES
NO
SYNTHESIS Reaction
A+B → AB
O_2 + CO_2 → CO_3
"Two into one"
Are there more PRODUCT molecules than reactant molecules?
NO
YE
DECOMPOSITION Reaction
XY → X + Y
$CaCO_3$ → CaO + CO_2
"One into two"
Are your REACTANTS something reacting with O_2 gas?
YE
COMBUSTION Reaction
Something + O_2 →
CH_4 + 2O_2 → CO_2 + 2H_2O
"If hydrocarbon, always make carbon dioxide and water"
NO
Is there an ELEMENT BY ITSELF that swaps places with an atom in a compound?
YE
SINGLE REPLACEMENT Reaction
A + BC → AC + B
Al + Pb(NO_3)$_2$ → Al(NO_3)$_2$ + Pb
"Swap cation for cation, or anion for anion
NO
Do you switch cations and anions between TWO COMPOUNDS?
YE
DOUBLE REPLACEMENT Reaction
AB + CD → AD + CB
$AgNO_3$ + KCl → AgCl + KNO_3
"Switch cations and anions"

Activity Series Chart

Metals

Most Active ↑ … ↓ *Least Active*

Name	Symbol
Lithium	Li
Potassium	K
Barium	Ba
Strontium	Sr
Calcium	Ca
Sodium	Na
Magnesium	Mg
Aluminum	Al
Manganese	Mn
Zinc	Zn
Iron	Fe
Cadmium	Cd
Cobalt	Co
Nickel	Ni
Tin	Sn
Lead	Pb
Hydrogen	H
Copper	Cu
Silver	Ag
Mercury	Hg
Gold	Au

Lithium through Sodium can replace a Hydrogen in a water molecule

Magnesium through Lead can replace a Hydrogen in an acid molecule

Non-Metals

Name	Symbol
Fluorine	F
Chlorine	Cl
Bromine	Br
Iodine	I

***

Elements CANNOT replace anything ABOVE them.
The reaction DOES NOT OCCUR in this situation.

***

Examples: $ZnCl_2 + Mg \rightarrow MgCl_2$

Magnesium is above Zinc so the reaction happens

$ZnCl_2 + Cu \rightarrow$ No Reaction

Copper is below Zinc so no reaction happens

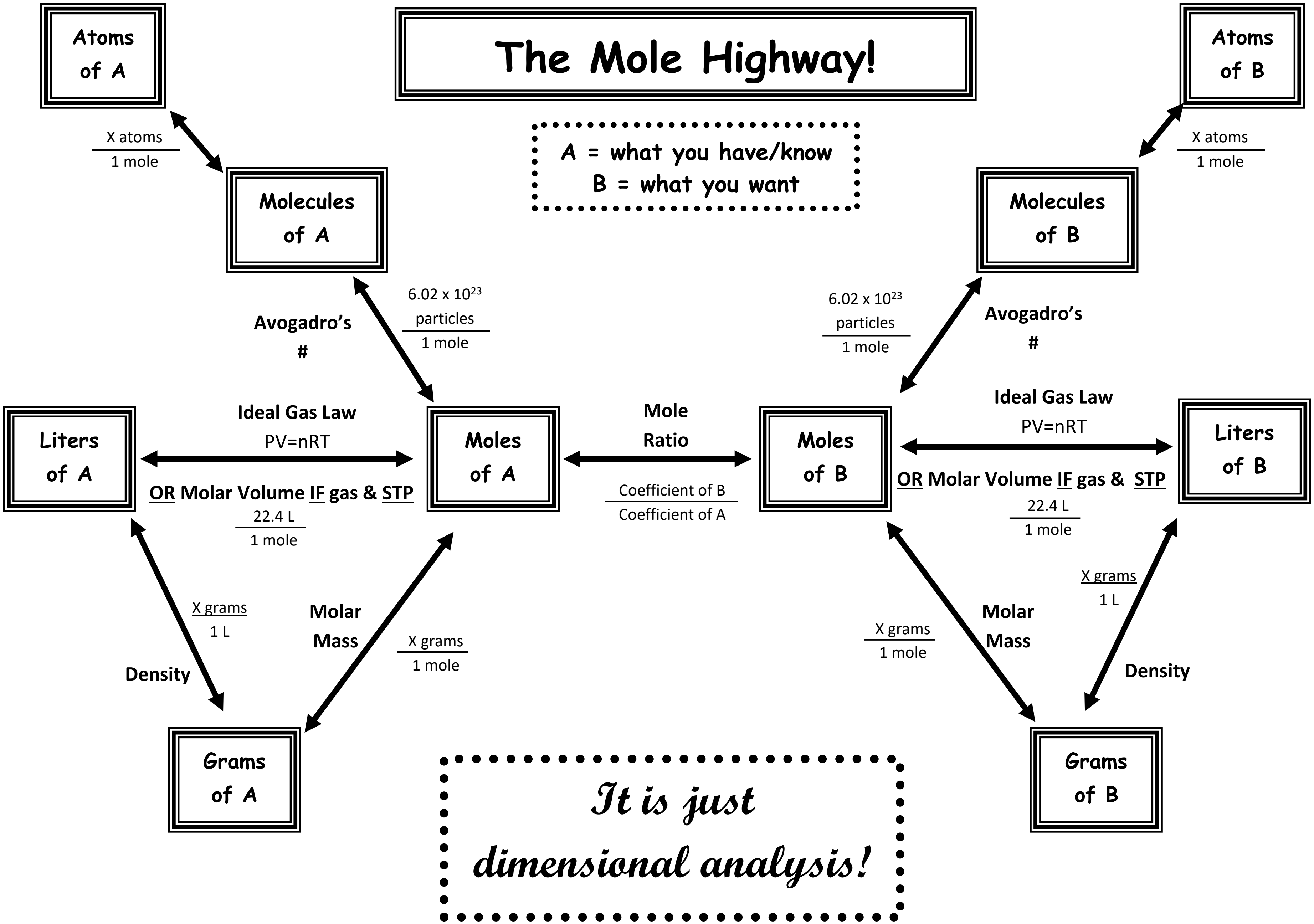
Solubility of Some Ionic Compounds in Water

<u>Always Soluble</u>		
Alkali metals =	Li^{+}, Na^{+}, K^{+}, Rb^{+}, Cs^{+}	AAA CNP
Ammonium =	NH_4^{+}	
Acetate =	$C_2H_3O_2^{-}$	
Chlorate =	ClO_3^{-}	
Nitrate =	NO_3^{-}	
Perchlorate =	ClO_4^{-}	
<u>Generally Soluble</u>		
Cl^{-}, Br^{--}, I^{-}	Soluble *<u>except</u>*: Ag^{+}, Pb^{2+}, Hg_2^{2+}	AP-H
F^{-}	Soluble *<u>except</u>*: Ca^{2+}, Ba^{2+}, Sr^{2+}, Pb^{2+}, Mg^{2+}	CBS-PM
Sulfate = SO_4^{2-}	Soluble *<u>except</u>*: Ca^{2+}, Ba^{2+}, Sr^{2+}, Pb^{2+}	CBS-P
<u>Generally Insoluble</u>		
O^{2-}, OH^{-}	Insoluble *<u>except</u>*: Alkali metals and NH_4^{+}	AA
	<u>Somewhat</u> soluble: Ca^{2+}, Ba^{2+}, Sr^{2+}	CBS
CO_3^{2-} S^{2-}, SO_3^{2-} PO_4^{3-} CrO_4^{2-}, $Cr_2O_4^{2-}$	Insoluble *<u>except</u>*: Alkali metals and NH_4^{+}	AA

Not Soluble = forms precipitate

Soluble = dissolves in water (aqueous)

Reference Sheets for
Unit #7 – Stoichiometry


The Mole Highway!
A = what you have/know
B = what you want
Atoms of A
X atoms / 1 mole
Molecules of A
Avogadro's #
6.02 x 10^23 particles / 1 mole
Liters of A
Ideal Gas Law
PV=nRT
OR Molar Volume IF gas & STP
22.4 L / 1 mole
Moles of A
Mole Ratio
Coefficient of B / Coefficient of A
Moles of B
Density
X grams / 1 L
Molar Mass
X grams / 1 mole
Grams of A
Atoms of B
X atoms / 1 mole
Molecules of B
6.02 x 10^23 particles / 1 mole
Avogadro's #
Ideal Gas Law
PV=nRT
Liters of B
OR Molar Volume IF gas & STP
22.4 L / 1 mole
X grams / 1 L
X grams / 1 mole
Molar Mass
Density
Grams of B
It is just dimensional analysis!

Stoichiometry Walk-Through

Some examples using the mole highway	
$N_2 + 3H_2 \rightarrow 2NH_3$ ***You start with 25.00g of N_2 - How much H_2 do you need?***	
Grams A → Moles A *Use molar mass A*	$\frac{25.00g\ N_2}{} \Big\vert \frac{1\ mol\ N_2}{28.01\ g\ N_2}$ $= 0.8925 mol\ N_2$
Moles A → Moles B *Use mole ratio B/A*	$\frac{0.8925 mol\ N_2}{} \Big\vert \frac{3\ mol\ H_2}{1\ mol\ N_2}$ $= 2.678\ mol\ H_2$
Grams A → Moles B *Use molar mass A, then mole ratio B/A*	$\frac{25.00g\ N_2}{} \Big\vert \frac{1\ mol\ N_2}{28.01\ g\ N_2} \Big\vert \frac{3\ mol\ H_2}{1\ mol\ N_2}$ $= 2.678\ mol\ H_2$
Grams A → Grams B *Use molar mass A, then mole ratio B/A, then molar mass B*	$\frac{25.00g\ N_2}{} \Big\vert \frac{1mol\ N_2}{28.01g\ N_2} \Big\vert \frac{3mol\ H_2}{1mol\ N_2} \Big\vert \frac{2.02g\ H_2}{1mol\ H_2}$ $= 5.409\ mol\ H_2$
Grams A → Molecules B *Use molar mass A, then mole ratio B/A, then Avogadro's # B*	$\frac{25.00g\ N_2}{} \Big\vert \frac{1mol\ N_2}{28.01g\ N_2} \Big\vert \frac{3mol\ H_2}{1mol\ N_2} \Big\vert \frac{6.02 \times 10^{23}\ molec.\ H_2}{1mol\ H_2}$ $= 1.612 \times 10^{24}$ molecules H_2

These are not all the combinations of routes on the mole highway, just some examples of possible routes

Winter Break Reminders and Suggestions:

When we return from Winter Break we will be starting a new chapter called “Advanced Chemical Ratios.” This chapter adds in a more complicated type of stoich called “limiting reagents.” It also adds in problems where you determine the formula of unknown compounds using stoichiometry and mole ratios.

There is no official homework over Winter Break, however, please make sure that you do not forget the following topics while on vacation! If you struggled with these topics during 1st semester please spend some time reviewing the topics. We want to make sure that everyone comes back from vacation ready to start 2nd semester off in a strong way!

Included in this handout is a list of topics to remember, a chart of notes where you learned the topics, starred (*) optional worksheets where you have practice problems, and a small practice test of some examples of the types of things we need to make sure we don’t forget how to do. Please realize that this practice test is not required, and it does not show every single possible thing you need to remember from 1st semester, it is just some examples to remind you.

We will keep using the same Composition Book and 3-ring binder 2nd semester so do not lose them or get new ones. You may take out your old rainbow packets BUT you need to keep them safe because they will be graded again at the end of 2nd semester. If you would rather leave them in your binder that is ok too. The gradebook starts over 2nd semester so everyone gets to start fresh and work towards completing all their work, doing well on quizzes and tests, etc.

If you have questions please email me. I will not be checking email daily, but I will check it occasionally over vacation. Thank you, and have a fabulous Winter Break!

Mrs. Farmer

Some Key Topics to Remember Over Vacation:

1. **Study your ions!**
   - ***There will be an ion quiz the week we return!***
   - The day is unannounced, but it will be during the first week.
   - Remember to know the ones on your ion sheet, but also any atoms from the periodic table s, p, d block that follow the pattern of the group numbers and those that use roman numerals.
2. **Types of bonds**
   - Identify if a molecule is ionic or covalent
3. **Writing formulas**
   - Crossing over to make neutral ionic compounds
   - Using prefixes to write covalent molecules
4. **Naming formulas**
   - Remember - two different ways to name things – one for ionic, one for covalent
5. **Type of reactions**
   - Be able to identify the type of reaction shown.
6. **Predicting products**
   - Use the main types of reactions to predict the products and write valid formulas for the products made – cross over if ionic, careful of diatomics, etc.
7. **Balancing equations**
   - Remember to balance AFTER predicting your products and writing valid formulas!
8. **Molar Conversions and Stoichiometry**
   - ***There will be a quiz on molar conversions and stoichiometry the week we return!***
   - The day is unannounced, but it will be during the first week.
   - Make sure you can do any type of problem given to you – don’t forget conversion factors like density, molar volume at STP, metric conversions thrown in, etc.

Where to Go to Refresh Your Memory Over Vacation:

Topic	Notes	* Optional Worksheets
Ions	N-17	R-2
Types of bonds	N-16	P5-WS16*
Writing/ naming formulas	N-16 N-17	P5-WS16*
Types of Reactions	N-23	P6-5*
Predicting products	N-24	P6-5*
Balancing equations	N-22	P6-5*
Molar conversions and stoich	N-25 N-26	P7-5*

***Remember –** You have your rainbow packets, reference pages, study materials, the class website has a “Resources” tab that has links to other websites and other practice, the “Notebook” tab has worksheets from my regular chem class that cover some similar basic level topics, and you have the entire internet at your fingertips too! ☺

Practice Test for Jogging Your Memory Before 2nd Semester:

1. Which of the following statements is **not** true of balancing a chemical equation?
   A) Subscripts in the reactants must be conserved in the products.
   B) Coefficients are used to balance the atoms on both sides.
   C) The law of conservation of matter must be followed.
   D) Phases are often shown for each compound but are not critical to balancing an equation.

2. The name for $Al(OH)_3$ is
   A) aluminum(III) hydroxide
   B) aluminum trihydroxide
   C) aluminum hydroxide
   D) monaluminum trihydroxide

3. Calculate the molecules of oxygen required to react with 35.9 g of sulfur.
   $2S + 3O_2 \rightarrow 2SO_3$
   A) 2.02×10^{24} molecules O_2
   B) 1.01×10^{24} molecules O_2
   C) 3.37×10^{23} molecules O_2
   D) 6.74×10^{23} molecules O_2

4. iron(III) phosphide is
   A) Fe_3P_2
   B) FeP
   C) Fe_3P
   D) FeP_3

5. Convert 9.51×10^{22} molecules NH_3
   A) 5.73×10^{46} mol
   B) 6.33 mol
   C) 2.69 mol
   D) 0.158 mol

6. The reaction $Pb(NO_3)_2 + Mg \rightarrow Pb + Mg(NO_3)_2$ is:
   A) synthesis
   B) acid-base
   C) single-replacement
   D) double-replacement

7. Sodium chloride and lead(II) nitrate react. Which is one of the products?
   A) $PbCl(s)$
   B) $Pb_2Cl(s)$
   C) $NaNO_3(aq)$
   D) $2NaNO_3(aq)$

8. The compound PI_3 is named
   A) potassium iodide
   B) monophosphorus iodide
   C) phosphorus iodide
   D) phosphorus triiodide

9. Which has covalent bond(s)?
   A) NaCl
   B) CaO
   C) CO_2
   D) Cs_2O

10. A 4.7-mol sample of $KClO_3$ was decomposed. How many moles of O_2 are formed? $2KClO_3 \rightarrow 2KCl + 3O_2$
    A) 7.1 mol
    B) 3.9 mol
    C) 4.7 mol
    D) 2.3 mol

11. The correct name for FeO is
    A) iron oxide
    B) iron(II) oxide
    C) iron(I) oxide
    D) iron monoxide

12. How many kilograms of silver can be produced when 40.3 g copper reacts with silver nitrate? Assume product has copper (II)
    A) 0.137 kg Ag
    B) 68.4 kg Ag
    C) 0.342 kg Ag
    D) 47.5 kg Ag

13. Ammonium sulfate is
    A) NH_4SO_3
    B) NH_4SO_4
    C) $(NH_4)_2SO_3$
    D) $(NH_4)_2SO_4$

14. What mass of oxygen gas is required to react completely with 18.8 g of C_6H_{14}?
    A) 5.72×10^3 g
    B) 33.2 g
    C) 6.98 g
    D) 66.3 g

15. How many atoms of calcium are present in 87.1 g of calcium?
    A) 3.61×10^{-24}
    B) 5.25×10^{25}
    C) 6.02×10^{23}
    D) 1.31×10^{24}

16. Convert: 45.0 g NaCl into mol
    A) 2.63×10^3
    B) 1.30
    C) 0.770
    D) 1.47×10^{23}

17. The charge on a barium ion is:
    A) +1
    B) +2
    C) +3
    D) -1

18. Convert: 2.64 g O_3 into molecules
    A) 1.59×10^{24}
    B) 7.63×10^{25}
    C) 3.31×10^{22}
    D) 9.13×10^{-26}

19. 2.85 moles of water weighs
    A) 1.58×10^{-1} g
    B) 51.3 g
    C) 6.32 g
    D) 21.0 g

20. Titanium(IV) oxide has the formula
    A) Ti_4O
    B) TiO_4
    C) Ti(IV)O
    D) TiO_2

21. The percent yield is a ratio of the ____________ yield to the _____________ yield, multiplied by 100%.

22. Which of the following formulas is **incorrect**?
    A) NaBr
    B) $AlCl_3$
    C) $CsCl_2$
    D) $Mg(OH)_2$

23. An aqueous solution of potassium chloride is mixed with an aqueous solution of sodium nitrate. Which is a product?
    A) $KCl(s)$
    B) $KNO_3(aq)$
    C) $KNa(aq)$
    D) $ClNO_3(aq)$

24. When they react, alkali metals:
    A) gain 1 electron
    B) gain 7 electrons
    C) gain or lose 7 electrons
    D) lose 1 electron

25. The molar mass of ammonium phosphate is
    A) 113.01 g/mol
    B) 131.05 g/mol
    C) 144.06 g/mol
    D) 149.10 g/mol

26. The balanced equation $P_4(s) + 6H_2(g) \longrightarrow 4PH_3(g)$ tells us that 5.0 mol H_2
    A) reacts with 2.5 mol P_4
    B) produces 10.0 mol PH_3
    C) cannot react with phosphorus
    D) produces 3.3 mol PH_3

27. Classify the following reaction: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
    A) Synthesis
    B) combustion
    C) double replacement
    D) single replacement

28. When the following equation is balanced using the smallest possible integers, what is the number in front of the substance in bold type? $Al + Fe_3O_4 \longrightarrow Al_2O_3 + \mathbf{Fe}$
    A) 1
    B) 3
    C) 6
    D) 9

29. True or false? Covalent bonding occurs when a metal reacts with a nonmetal.
    A) True
    B) False

30. Which of the following compounds contains an ionic bond?
    A) $HCl(g)$
    B) NaCl
    C) CCl_4
    D) SO_2

Answer Key

***Answer Key has not been checked! If you see typos please email me so I can fix them! ☺**

1. A
2. C
3. B
4. B
5. D
6. D
7. D
8. D
9. C
10. A
11. B
12. A
13. D
14. D
15. D
16. C
17. B
18. C
19. B
20. D
21. actual, theoretical
22. C
23. B
24. D
25. D
26. D
27. A
28. D
29. B
30. B