

AP Chemistry 2020-2021 at a Glance:

- The College Board recommends having successfully completed Chemistry and Algebra 2 for success in this course. A grade of a “B” or better in both classes is strongly suggested.
- To succeed in this course, it will also be important to take the initiative and be resilient.
- This course will be specifically focused on preparing for the AP exam by looking at the following Big Ideas: Atomic Theory, Bonding, Chemical Reactions, Reaction Rates, Thermochemistry, and Equilibrium.
- The exam consists of two sections, a multiple choice section and a free response section both equally weighted, and will be given on Friday, May 7, 2021 at 8:00AM, if nothing changes.
- This summer assignment and more Chemistry topics will be reviewed again throughout the school year but fairly quickly.

AP Chemistry 2020-2021 Summer Assignment:

- Review outline sections 1-3.
- Check your understanding using the section problem sets and answers. If you have any questions, e-mail me at: jlouie@nUSD.org. I will have limited access to my email during the summer so I will do my best to answer questions when I am able or they will be answered during the first few days of class.
- It is expected that you fully understand the material both in the outlines and the problem sets by the time the new school year begins. This information will be reviewed only somewhat during the course, and it is important that you understand it for the first test of the semester.
- There are often other ways of doing some of the things here and if what is presented in this outline does not work for you, I would encourage you to search the Internet to find something that does.
- The first exam on sections 1-3 will occur within the first weeks of class. Be prepared with any questions regarding the problem sets.

Section I Chemical Foundations

(adapted from sarrep.com)

Metric Conversions

All measurements in chemistry are made using the metric system. In using the metric system you must be able to convert between one value and another. You must memorize the factors, prefixes and symbols in the chart below.

There are numerous ways to do metric conversion. Here is one method below:

Example: Convert 1.83×10^{-1} kilograms to centigrams.

Step 1. Find the absolute difference between the power of ten values of the units.

In this problem you are given 1.83 kilograms and you are converting to centigrams. Kilograms have a power of ten value of 3; centigrams have a power of ten value of -2.

$$|3 - (-2)| = 5$$

Step 2. Either add the above value from the original power of ten value if you are converting from a larger unit to a smaller unit OR subtract the above value to the original power of ten value if you are converting from a smaller unit to a larger unit.

In this problem we are going from kilograms to centigrams (a larger unit to a smaller unit) so we take our original power of ten value of -1 and add 5. This gives us a total of 4 for the final power to ten value.

Step 3. Put your answer in proper scientific notation.

1.83×10^{-1} kg is converted to 1.83×10^4 cg.

Factor	Prefix	Symbol
1×10^{12}	tera-	T
1×10^9	giga-	G
1×10^6	mega-	M
1×10^3	kilo-	k
1×10^2	hecto-	h
1×10^1	deca-	da
1×10^0	---	---
1×10^{-1}	deci-	d
1×10^{-2}	centi-	c
1×10^{-3}	milli-	m
1×10^{-6}	micro-	μ
1×10^{-9}	nano-	n
1×10^{-10}	angstrom	Å
1×10^{-12}	pico-	p
1×10^{-15}	femto-	f

Uncertainty in Measurement

- Depends on the precision of the measuring device
 - For example a measurement of 1.682956 grams is a more precise measurement than 1.7 grams
- Reliability in Measurements
 - Accuracy – the closeness to the actual scientific value
 - Precision – getting repeated measurements in repeated trials
- Types of Errors
 - Random: error in measurement has equal probability of being high or low
 - Systematic: errors all occur in the same direction

General Rules for determining if a Number is Significant

1. Draw a box around all nonzero digits beginning with the leftmost nonzero digit and ending with the rightmost nonzero digit in the number.
2. If a decimal is present, draw a box around any trailing zeros to the right of the original box.
3. Consider any all boxed digits significant.

Example 1: 20406: 5 significant digits

Example 3: 4000: 1 significant digit

Example 5: 0.002500: 4 significant digits

Example 2: 0.0045: 2 significant digits

Example 4: 4000.: 4 significant digits

Example 6: 3.00: 3 significant digits

Addition or Subtraction using Significant Figures

The answer can only be as precise as the least precise measurement.

Example

2.8701	(precise four places to the right of the decimal)
0.0673	(precise four places to the right of the decimal)
<u>+ 301.520</u>	(precise three places to the right of the decimal)
304.4574	→ rounds off to 304.457 (answer must be precise three places after the decimal)

Multiplication or Division using Significant Figures

The answer can have no more total significant figures than there are in the measurement with the smallest total number of significant figures.

Example:

12.257	(5 total significant figures)
<u>x 1.162</u>	(4 total significant figures)
14.2426	→ rounds off to 14.24 (4 total significant figures)

As a general rule, if you are unsure how many significant figures to use on the AP exam, use 3 significant figures. This may not always work but it will work most times. However you should always pay close attention to using the correct number of significant figures in all calculations.

Scientific Notation

In chemistry, we often use numbers that are either very large (1 mole = 602 200 000 000 000 000 000 atoms) or very small (the mass of an electron = 0.000 000 000 000 000 000 000 000 000 910 939 kg). Writing numbers with so many digits would be tedious and difficult. To make writing very large and small numbers easier, scientists use an abbreviation method known as scientific notation. In scientific notation the numbers mentioned above would be written as 6.022×10^{23} and 9.10939×10^{-31} .

Converting a number to or from scientific notation

- If you move the decimal place to the left, the power of 10 value increases.
- If you move the decimal place to the right, the power of 10 value decreases.

Example 1: Look at the first number from above: 602 200 000 000 000 000 000

To put this number in scientific notation you would move your decimal place until there is one number to the left of the decimal. To do this, we must move our decimal 23 places to the left. When you move the decimal to the left, the power of 10 value increases. It increases from 0 to 23. Thus, the answer is **6.022×10^{23}** .

Look at the second number from above: 0.000 000 000 000 000 000 000 000 000 910 939

To put this number in scientific notation we must move our decimal 31 places to the right. **REMEMBER: You should always have one digit to the left of the decimal when writing numbers in scientific notation.** Since we are moving our decimal to the right, we must decrease our power of 10 value. It decreases from 0 to -31. The answer is **9.10939×10^{-31}** .

Rules for multiplying & dividing using scientific notation:

- **When multiplying two numbers in scientific notation, ADD their power of 10 values.**

For example: $(3.45 \times 10^6)(4.3 \times 10^5) = 14.835 \times 10^{11}$. But, we must also remember to express our answer in significant figures. Thus, the final answer is 1.5×10^{12}

- **When dividing numbers in scientific notation, SUBTRACT the denominator's power of 10 value from the numerator's power of 10 value.**

For example: $(2.898 \times 10^{12}) \div (3.45 \times 10^{15}) = 0.840 \times 10^{-3}$ (I had to add the zero at the end to get the three significant figures needed.) I got 10^{-3} because $12 - 15 = -3$. Make sure your answer is in proper scientific notation (one number to the left of the decimal). In this problem we have to move the decimal one place to the right. When we move our decimal to the right, we decrease our power of 10. -3 decreases by 1 to -4 . Our final answer is: **8.40×10^{-4}** .

Dimensional Analysis

- Used to convert a number from one system of units to another.
- Understanding dimensional analysis is crucial. This process will help you when performing difficult calculations later in the year.
- Conversion factors **do not** need to be memorized.

Example: Calculate how many kilometers there are in 5 miles.

Solution: (Needed Equivalents: 1 mile = 1760 yards, 1 meter = 1.094 yards)

$$\underline{5 \text{ miles}} \times \frac{1760 \text{ yards}}{1 \text{ mile}} \times \frac{1 \text{ meter}}{1.094 \text{ yards}} \times \frac{1 \text{ kilometer}}{1000 \text{ meters}} = 8.04 \text{ (rounded to) } 8 \text{ kilometers}$$

Example: Convert 55.0 miles/hour to meters/second

Solution: (Needed Equivalents: 1 mile = 1760 yards, 1 meter = 1.094 yards, 1 hour = 60.0 minutes, 1 minute = 60.0 seconds)

$$\frac{55.0 \text{ miles}}{\text{hour}} \times \frac{1760 \text{ yards}}{1 \text{ mile}} \times \frac{1 \text{ meter}}{1.094 \text{ yards}} \times \frac{1 \text{ hour}}{60.0 \text{ minutes}} \times \frac{1 \text{ minute}}{60.0 \text{ seconds}} = 24.5785 \text{ (rounded to) } 24.6 \text{ m/s}$$

General Rules for Rounding Numbers in Chemistry

- **Rule 1. If the digit following the last significant figure is less than 5, the last significant figure remains unchanged. The digits after the last significant figure are dropped.**

Example: Round 23.437 to three significant figures.

Answer: 23.4

Explanation: 4 is the last significant figure. The next number is 3. 3 is less than 5. Thus, 4 remains unchanged and 37 is dropped.

- **Rule 2. If the digit following the last significant figure is 5 or greater, then 1 is added to the last significant figure. The digits after the last significant figure are dropped.**

Example: Round 5.383 to two significant figures.

Answer: 5.4

Explanation: 3 is the last significant figure. The next number is 8. 8 is greater than 5. Thus, 1 is added to 3 making it 4. The 83 is dropped.

- As a rule, when performing a series of calculations, **wait until the very end to round off to the proper number of significant figures instead of rounding off each intermediate result. If you are changing from addition/subtraction to multiplication/division or vice versa, note the number of sig figs by underlining the significant digits.**

Example 1: $10.82 + 2.5 + 2.64 =$

WRONG: $10.82 + 2.5 = 13.32$ (rounded to 13.3) then $13.3 + 2.64 = 15.94$ (rounded to) 15.9

CORRECT: $10.82 + 2.5 + 2.64 = 15.96$ (rounded to) 16.0 (precise to 1 place after the decimal)

Example 2: $(12.00 - 10.00) \div 12.00 =$

In this case you would subtract and NOTE to the proper number of significant figures and then divide because you are changing between significant figure rules. It is best to wait to round until the very end.

CORRECT: $(12.00 - 10.00) = 2.00$ (precise to 2 places after the decimal)

$2.00 \div 12.00 = 0.167$ (rounded to 3 total significant figures)

Temperature Conversion

C: Celsius or Centigrade

K: Kelvin (named for Lord Kelvin, a.k.a. William Thomson)

Fahrenheit is not used in AP Chemistry so we will ignore it.

$K = ^\circ\text{C} + 273$ $^\circ\text{C} = K - 273$ (These formulas must be memorized)

Example 1: Convert 29.0 °C to Kelvin.

$$K = 29.0 + 273$$

$$K = 302 \text{ K}$$

Example 2: Convert 888 K to °C.

$$^{\circ}\text{C} = 888 - 273$$

$$^{\circ}\text{C} = 615 \text{ }^{\circ}\text{C}$$

Density

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

Density (d) is the ratio of the mass (m) of a substance to the volume (v) occupied by the substance. Pure water is used as the standard in measuring density. The density of pure water is 1.0 g/mL. In more precise calculations the actual density of water will differ slightly. If a substance has a density less than water, it will float; if a substance has a density greater than water, it will sink.

Mass is expressed in grams (g). Volume is expressed in liters (L), milliliters (mL) or cubic centimeters (cm³). A mL is the same as a cm³. Thus, density of a liquid or solid can be expressed as g/mL or g/cm³. The density of a gas is expressed in g/L.

Example: A piece of wood has a volume of 3350 cm³. If the density of the wood is 0.512 g/mL, what is its mass?

$$d = 0.512 \text{ g/mL}$$

$$0.512 \text{ g/mL} = \frac{m}{3350 \text{ cm}^3}$$

$$v = 3350 \text{ cm}^3$$

$$3350 \text{ cm}^3$$

$$m = 1715.2 \text{ g}$$

$$1.72 \times 10^3 \text{ g (3 significant digits)}$$

$$m = x$$

Percent Error

The accuracy of your measurements can be checked by calculating the percent error. In a percent error calculation you will compare your experimental value to the accepted scientific value (referred to as the theoretical value).

Since you are taking the absolute value of the subtraction, **your percent error will always be a positive number.**

Remember to use significant figures in all percent error calculations.

In some quantitative lab experiments you may be expected to calculate your percent error. In general, a percent error of 10% or less is often considered acceptable. Any percent error greater than 10% will usually mean that there may be some mistake that needs to be addressed.

$$\% \text{Error} = \frac{|\text{theoretical yield} - \text{experimental yield}|}{\text{theoretical yield}} \times 100.$$

Example: The theoretical density of aluminum is 2.70 g/mL. In an experiment a student measures the mass of an aluminum bar to be 12.52 grams and finds the volume of the bar to be 4.71 mL. Calculate the student's percent error.

First find the experimental density:

$$d = \frac{12.52 \text{ grams}}{4.71 \text{ mL}} \quad d = 2.66 \text{ g/mL}$$

Now use the experimental value to find % error. $\%E = \frac{|2.70 - 2.66|}{2.70} \times 100. \rightarrow \frac{0.04}{2.70} \times 100. \rightarrow \%E = 1\%$

Classification of Matter

Chemistry is the study of matter

- Matter – anything that has mass and takes up space
 - Matter exists in states
 - Solid – rigid, fixed shape and volume
 - Liquid – definite volume, takes the shape of the container
 - Gas – no fixed volume or shape, very compressible

- Most matter exists as mixtures
 - Example: wood, wine
 - Homogeneous mixtures
 - Visibly indistinguishable parts
 - Called solutions
 - Example: air, brass, salt water
 - Heterogeneous mixtures
 - Visibly distinguishable parts
 - Example: sand in water, oil & water
- Can be separated into pure substances through physical changes
 - Physical changes do not change the chemical composition of the matter
 - Ways to separate mixtures
 - Distillation – separation by boiling point
 - Filtration – separation method used with a solid & liquid mixture where a barrier blocks solid particles from passing through
 - Chromatography – a series of methods that employ a system with two phases (states of matter)
 - Mobile phase (liquid or gas)
 - Stationary phase (solid)
 - Types of Chromatography
 - Paper Chromatography
 - Column Chromatography
- Pure substances
 - Matter made up of only one type of element or compound
 - Compound – substance with constant composition that can be broken down into elements by chemical means
 - Elements – a substance that cannot be decomposed into simpler substances by chemical or physical means
 - Atoms – the most basic unit of matter
 - Proton – positively charged particle
 - Neutron – neutral particle
 - Electron – negatively charged particle

Section 2

Atoms, Molecules, & Ions

Fundamental Chemical Laws

- Antoine Lavoisier (1743-1794)
 - Father of Modern Chemistry
 - Explained the nature of combustion; showed combustion involves oxygen, not phlogiston
 - His experiments suggested that mass is not created or destroyed (Law of Conservation of Mass)
- Joseph Proust (1754-1826)
 - Showed that a given compound contains exactly the same proportion of elements by mass.
 - Example: Copper(II) Carbonate has a definite mass ratio of Cu: 5.3 parts to 4 parts oxygen to 1 part carbon; CuCO_3 , g.f.m = (Cu, 63.5 + C, 12.0 + O, 48.0 = 123.5), composition = (Cu, 63.5/123.5 = 0.514; C, 12.0/123.5 = 0.0972; O, 48.0/123.5 = 0.389), ratio = (Cu, 0.514/0.0972 = 5.29, C, 0.0972/0.0972 = 1.00, O = 0.389/0.0972 = 4.00)
 - Proust's Law, a.k.a. Law of Definite Proportion
- John Dalton (1766-1844)
 - Proust's work inspired Dalton to think about atoms as parts of the elements
 - His reasoning was that if elements were composed of tiny individual parts(atoms) then a given compound should always contain the same combination of these atoms
 - Discovered that when 2 elements form a series of compounds the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers – Law of Multiple Proportions

· Example 1:

	Mass of Oxygen that Combines with 1 gram of carbon	Chemical Formula
Compound 1	1.33 grams	CO
Compound 2	2.66 grams	CO ₂

Since we now know the mass of carbon is 12.01 amu and the mass of oxygen is 16.00 amu we can understand the formula of the first compound to be **CO** ($12.01 \times 1.33 = 16.0$). In the second compound there is twice as much mass of oxygen per gram of carbon so we can determine the formula to be **CO₂**.

Dalton's Atomic Theory

- 1808 published – A New System of Chemical Philosophy
 - Each element is made up of atoms
 - Atoms of a given element are identical; atoms of different elements are different
 - Compounds are formed when atoms of different elements combine with one another. A given compound always has the same relative numbers and types of atoms.
 - Atoms are not changed in chemical reactions.
- Designed first table of atomic masses, though most of his masses were wrong because of his incorrect assumptions about formulas
- Joseph Gay-Lussac (1778-1850)
 - In 1809 performed experiments in which he measured the volumes of gases that reacted with each other (at constant temperature and pressure)
 - Example: 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of water vapor ($2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$)

- Amadeo Avogadro (1776-1856)
 - In 1811 Avogadro interpreted these results by proposing that at the same temperature and pressure, equal volumes of different gasses contain the same number of particles (Avogadro's hypothesis)
 - If Avogadro is correct, the Gay-Lussac's result can be expressed as: 2 molecules of hydrogen react with 1 molecule of oxygen to form 2 molecules of water

Early Experiments to Characterize the Atom

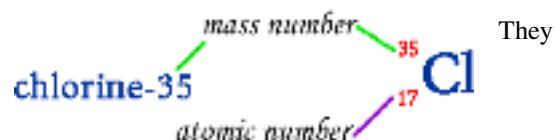
- The electron
 - J.J. Thomson
 - English physicist who studied electrical discharges in cathode-ray tubes
 - Cathode ray tubes produce a ray of light from the cathode end of the tube
 - Thomson postulated that the ray was a stream of negative charged particles, now called electrons
 - Charge-to-mass ratio = $-1.76 \times 10^8 \text{ C/g}$ (Coulombs per gram)
 - Since atoms are electrically neutral, Thomson assumed that atoms must contain a positively charged particle to balance the charge of the electron.
 - Plum Pudding Model (J.J. Thomson is given credit for this model but the idea was first suggested by English mathematician and physicist William Thomson (no relation) a.k.a. Lord Kelvin)
 - Robert Millikan (1868-1953)
 - Determined the magnitude of the electron charge.
 - With the magnitude of the electron charge and the charge to mass ratio, Millikan was able to calculate the mass of the electron to be $9.11 \times 10^{-31} \text{ kg}$
- Nuclear Atom
 - In 1911 Ernest Rutherford carried out experiments to test Thomson's plum pudding model
 - Rutherford shot alpha particles at a thin sheet of metal foil. He expected the alpha particles to travel through the gold foil
 - Most alpha particles passed straight through the gold foil but some were deflected at large angles and a few were deflected straight back
 - Rutherford said, "It's about as credible as shooting a 12" shell at a piece of tissue paper and having it come back and hit you."
 - Rutherford determined that the atom was mostly empty space and that at the core there is a small dense nucleus (Latin: little nut).

The Modern View of Atomic Structure

Subatomic Particle	Symbol	Location	Charge
proton	p^+	nucleus	+1
neutron	n^0	nucleus	0
electron	e^-	outside nucleus	-1

- **The chemistry of an atom results from its electrons.**
- The tiny nucleus accounts for almost all of the atom's mass (A nucleus the size of a pea would have a mass of 250 million tons (500 billion pounds).)
- Isotopes – atoms of an element with the same number of protons but different numbers of neutrons.
- Because the chemistry of an atom is due to its electrons, isotopes show almost identical chemical properties.

· There are different ways that isotopes can be represented. are shown to the right:



· Atomic number(Z) = # of protons = # electrons (in a neutral atom only)

· Mass number(A) = # protons + # neutrons

· **For example:** Two naturally occurring isotopes of chlorine are chlorine-35 & chlorine-37. Thirty-five and thirty-seven are the mass numbers for the two isotopes. Both isotopes have the same atomic number, number of protons and electrons.

Isotope name	Atomic number (z)	# protons	# neutrons	Mass number (A)	# electrons
chlorine-35	17	17	18	35	17
chlorine-37	17	17	20	37	17

Molecules & Ions

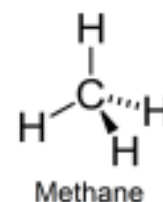
· The forces that hold atoms together in compounds is called chemical bonds.

· **Covalent Bonds** – atoms *share* electrons to form molecules.

· Molecules are represented by chemical formulas.



· **Structural formulas** give more information by indicating the individual bonds (shown with lines). The structural formula of methane, CH_4 , is shown to the right.



· A **space-filling model** shows the relative sizes of the atoms as well as their relative orientation in the molecule. The space filling model of methane is shown to the left.

· **Ball-and-stick models** are also used to represent ball-and-stick model of methane is represented to the right.



molecules. The

· Ionic Bonds occur between a cation and an anion. A cation is a positively charged ion that has lost an electron. An anion is a negatively charged ion that has gained an

charged ion that electron.

· Because anions and cations have opposite charges, they attract each other. as ionic bonding.

This is referred to

· A solid consisting of oppositely charged ions is called an ionic solid, or a salt.

An Introduction to the Periodic Table

· Most elements are metals. Metals are efficient conductors of heat and electricity, malleable, ductile and often lustrous. Metals tend to lose electrons when they form ions.

· There are few nonmetals. Nonmetals tend to gain electrons and acquire a negative charge when they form ions.

· Elements in the same vertical column are called groups or families and have similar chemical properties.

· Group 1 elements, (with the exception of hydrogen), are called the alkali metals and all have one valence electron and a 1+ charge.

· Group 2 elements are called alkaline-earth elements and all have 2 valence electrons and a 2+ charge.

· Groups 3-12 elements are called transition metals and most have 2 valence electrons. Their charges vary.

· Group 13 elements are part of the boron family and all have 3 valence electrons. Most have a 3+ charge.

· Group 14 elements are part of the carbon family (crystallogens) and all have 4 valence electrons. Most have a 4+ charge. Carbon can also be 4-. Some metals can also be 2+.

- Group 15 elements are part of the nitrogen family (pnictogens) and all have 5 valence electrons. Nonmetals have a 3- charge; metals vary in charge.
- Group 16 elements are part of the oxygen family and all have 6 valence electrons. Non-metals have a 2- charge; metals vary in charge.
- Group 17 elements are called halogens and all have 7 valence electrons and a 1- charge.
- Group 18, are called the noble gases. These elements have 8 valence electrons and are inert(non-reactive) and have no charge.
- The “charges” mentioned above refer to the most common ion formed when these elements form ionic compounds. Elements are neutral until they gain, lose or share their electrons and form compounds.
- Horizontal rows of elements on the periodic table are called periods.
- Vertical columns on the periodic table are called groups or families because they have the same number of valence electrons and thus similar properties.
- Be aware that the periodic table you will be able to use on the exam has no element names. As such you will want to be familiar with the names and symbols of elements 1-54, all of column 1,2,18, U, Th, Pt, Au, Ag, Cu, Pb, and I. If possible, be familiar with the atomic weights of very common elements like H, He, C, N, O, Na, and Cl.

Ionic Compounds

- An **ionic compound** is a compound that is formed between a **metal** and a **non-metal**. (Metalloids can also be used in ionic compounds, sometimes as cations and sometimes as anions, depending on the properties of the specific element.). In ionic compounds the **metal will always be a cation** and the **non-metal will always be an anion**. Please note, **the negative oxidation numbers we wrote on top of Groups 14, 15, 16 & 17 on our periodic tables refer only to the non-metals and the metalloids**. The metals in these columns have different oxidation numbers.
- When forming ionic compounds the goal is to balance the number of positive charges with the number of negative charges. More specifically, you want to ensure that the number of electrons that the cations are giving up is equal to the number of electrons the anions need so that both have full outer energy level.
- Binary ionic compounds contain only two elements, one is the cation and the other is the anion. It is important to remember that when writing binary ionic compounds **THE CATION MUST ALWAYS BE WRITTEN FIRST**. The rest of the rules will be outlined in the following example.

Part I. How to Write a Binary Ionic Compound Formula

Example: Write the formula for the compound between barium and sulfur.

Description of Action	Action
1. Write the symbol of the cation with its charge.	1. Ba ²⁺
2. To the right of the cation, write the anion and its charge.	2. Ba ²⁺ S ²⁻
3. Note if the charges equal zero. If they do, you are done. If not, add subscripts to ensure the charge equal zero. Remember, subscripts multiply the charge.	3. Ba ²⁺ S ²⁻ Result: Total charge = zero
4. Write out formula.	5. Ba and S do not need subscripts because the charge equals zero. So, the answer is: BaS

Sometimes groups of elements will combine such that they have an overall charge. These are known as polyatomic ions. The chart below lists polyatomic ions and their names. All will be used at different times in class, however, only the bolded ions and their charges (like hydroxide, nitrate, sulfate, carbonate, bicarbonate, etc.) need to be MEMORIZED for tests and quizzes. The names DO NOT need to be memorized. I suggest making flash cards and practicing the names and formulas.

Symbols of Common Polyatomic Ions					
(AsO ₃) ³⁻	arsenite	(C ₂ O ₄) ²⁻	oxalate	(N ₃) ⁻	azide
(AsO ₄) ³⁻	arsenate	(CrO ₄) ²⁻	chromate	(NH ₂) ⁻	amide
(BO ₃) ³⁻	borate	(Cr ₂ O ₇) ²⁻	dichromate	(NH ₄) ⁺	AMMONIUM
(B ₄ O ₇) ²⁻	tetraborate	(HCO ₃) ⁻	bicarbonate	(NO ₂) ⁻	nitrite
(BrO) ⁻	hypobromite	(HC ₂ O ₄) ⁻	bioxalate	(NO ₃) ⁻	nitrate
(BrO ₃) ⁻	bromate	(H ₃ O) ⁺	HYDRONIUM	(O ₂) ²⁻	peroxide
(CHO ₂) ⁻	formate	(HPO ₄) ²⁻	biphosphate	(OH) ⁻	hydroxide
(C ₂ H ₃ O ₂) ⁻	acetate	(H ₂ PO ₄) ⁻	dihydrogen phosphate	(PO ₃) ³⁻	phosphite
(C ₄ H ₄ O ₆) ⁻	tartrate	(HS) ⁻	bisulfide	(PO ₄) ³⁻	phosphate
(C ₆ H ₅ O ₇) ³⁻	citrate	(HSO ₃) ⁻	bisulfite	(SCN) ⁻	thiocyanate
(ClO) ⁻	hypochlorite	(HSO ₄) ⁻	bisulfate	(SO ₃) ²⁻	sulfite
(ClO ₂) ⁻	chlorite	(IO) ⁻	hypoiodite	(SO ₄) ²⁻	sulfate
(ClO ₃) ⁻	chlorate	(IO ₂) ⁻	iodite	(S ₂ O ₃) ²⁻	thiosulfate
(ClO ₄) ⁻	perchlorate	(IO ₃) ⁻	iodate	(SeO ₄) ²⁻	selenate
(CN) ⁻	cyanide	(IO ₄) ⁻	periodate	(SiF ₆) ²⁻	hexafluorosilicate
(CO ₃) ²⁻	carbonate	(MnO ₄) ⁻	permanganate	(SiO ₃) ²⁻	silicate

The word hydrogen can be substituted for the prefix bi- (i.e. hydrogen sulfide = bisulfide)

These polyatomic ions are treated just like any other ion when writing chemical formulas. As above, you would balance the charges with subscripts.

Part II: How to Write a Formula for an Ionic Compound that has a Polyatomic Ion

Example: Write the formula for the ionic compound formed between **aluminum** and **phosphite**.

Description of Action	Action
1. Write the symbol of the cation with its charge.	1. Al ³⁺
2. To the right of the cation, write the polyatomic anion and its charge.	2. Al ³⁺ (PO ₃) ³⁻
3. Note if the charges equal zero. If they do, you are done. If not, add subscripts to ensure the charge equal zero. Remember, subscripts multiply the charge.	3. Al ³⁺ (PO ₃) ³⁻ Result: Total charge = zero
4. Write out formula.	5. Al and PO ₃ do not need subscripts because the charge equals zero. Al (PO ₃)
5. Remember, do not touch anything inside the parenthesis . If there is no subscript outside the anion's parenthesis, remove the parenthesis.	6. Answer: Al PO ₃

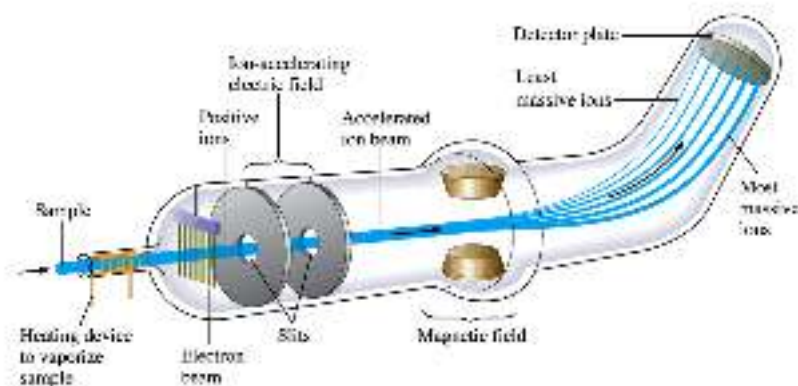
Many transition metals form ions with different charges. You should be at least familiar with the idea that charges of the following transition metals can be different. Others transition metals also have multiple charges but these are the ones you will be quizzed/tested on. Many can be figured out with a good understanding of the periodic table.

Metal	Charges	Metal	Charges
Sc	3+	Cu	1+, 2+
Ti	3+, 4+	Zn	2+
Cr	2+, 3+, 6+	Ag	1+
Mn	2+, 3+, 4+, 6+, 7+	Au	1+, 3+
Fe	2+, 3+	Hg	1+, 2+ *note mercury(I) is (Hg ₂) ²⁺
Co	2+, 3+	Sn	2+, 4+
Ni	2+, 3+	Pb	2+, 4+

Section 3 Stoichiometry

Atomic Masses

- The modern system of atomic masses, instituted in 1961, is based on carbon-12.
- Carbon-12 is assigned the mass of exactly 12 atomic mass units (amu)
- The most accurate method currently available for comparing the masses of atoms involves the use of the mass spectrometer.
- Atoms or molecules are passed through a beam of high speed electrons. This knocks electrons off the substance being analyzed and changes it to a positive ion. The electric field accelerates the positive ions into a magnetic field. Since an accelerated ion creates its own magnetic field, an interaction with the applied magnetic field causes a change in the path of the ion. The amount of deflection is a function of a substance's mass. The most massive ions are deflected the smallest amount, which causes the ions to separate. A comparison of the positions where the ions hit the detector plate gives very accurate values of their relative masses.
- For example: When isotopes carbon-12 and carbon-13 are analyzed in a mass spectrometer, the ratio of their masses is found to be: 1:1.0836129 To calculate the mass of carbon-13: $(1.0836129)(12 \text{ amu}) = 13.003355 \text{ amu}$.



- Since the modern system is based on carbon-12, it may seem surprising that carbon's atomic weight is 12.01 instead of 12.0. This is due to the fact that there are three carbon isotopes (^{12}C , ^{13}C & ^{14}C) and the atomic weight listed on the periodic table is an average value representing the mass and abundance of each isotope.

Average Atomic Weight

- The **atomic weight of an element is the weighted average of the masses of the isotopes of that element**. The weighted average is determined using the abundance and mass of each isotope. Most elements have more than one naturally occurring isotope.
- For example, there are two naturally occurring isotopes of copper, **copper-63** (62.93 amu) and **copper-65** (64.93 amu). The natural abundances of the isotopes are **69.17%** and **30.83%** respectively.
- **To determine the atomic weight:**

Step 1: Multiply the mass number and the relative abundance (as a decimal). The mass of the electron is insignificant in this calculation and is not used.

Isotope name	atomic mass	x	abundance (as a decimal)	=	result
copper-63	62.93 amu	x	0.6917	=	43.53
copper-65	64.93 amu	x	0.3083	=	20.02

- **Step 2:** Add up your results. **Atomic Weight = 63.55 amu**

· The mass of a proton is approx. 1.008 amu. The mass of a neutron is approximately 1.009 amu. The mass of the electron is insignificant. When elements are formed from the individual subatomic particles a large amount of energy in the form of heat is released. This loss of mass is called the **mass defect** and this conversion of mass into energy is seen in the equation: $E = mc^2$.

Gram Formula Mass / Molar Mass

· Gram formula mass (also known as molar mass) is defined as the atomic mass of one mole of an element, molecular compound or ionic compound. The answer must always be written with the unit g/mol (grams per mole).

Calculating a substance's gram formula mass (molar mass):

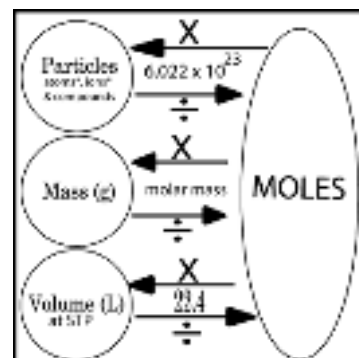
Calculate the gram formula mass (molar mass) of **ammonium phosphate**.

Description of Action	Action
1. Write the formula for the compound.	1. $(\text{NH}_4)_3\text{PO}_4$
2. Determine how many of each atom are in the compound. If there is a number outside the parenthesis, multiply each subscript by this number. If there is no subscript, assume it is one.	2. N: 3 H: 12 P: 1 O: 4
3. Multiply the number of atoms by its atomic weight.	3. N: $3 \times 14.01 = 42.03$ H: $12 \times 1.008 = 12.096$ P: $1 \times 30.97 = 30.97$ O: $4 \times 16.00 = 64.00$
4. Add your results and use the unit g/mol (grams per mole) on the end.	4. $42.03 + 12.096 + 30.97 + 64.00 = 149.10 \text{ g/mol}$.

Mole Conversions

- To convert between moles, atoms, molecules and liters you can use the following chart:
- One example is shown below:

Calculate the number of hydroxide ions in 23.5 grams of aluminum hydroxide.
 $23.5 \text{ g Al(OH)}_3 \div 78.0 \text{ g Al(OH)}_3 \times 6.022 \times 10^{23} \times 3 = 5.44 \times 10^{23} \text{ (OH)}^-$



Percent Composition

· When we calculate percent composition, we are determining the relative mass that each element contributes to the total mass of the compound. For example, if we calculate the gram formula mass of H_2O we will find it to be 18.0 g/mol. We arrived at this number by adding the mass of oxygen, 16.0, and the mass of two hydrogens, 2.0. Oxygen makes up 16.0 of the total 18.0 grams. Hydrogen is 2.0 of 18.0 grams. If we divide each element's total mass by the compound's total mass and multiply this result by 100, we get a percentage. This percentage is the element's percent composition.

Example (from above):

STEP 1 (find grams per mole)	STEP 2 (divide each element mass by grams per mole)	STEP 3 (multiply by 100)
H_2O : H: $2 \times 1.0 = 2.0$ O: $1 \times 16.0 = 16.0$ Total = 18.0 g/mol	\longrightarrow H: $2.0 \div 18.0 = 0.111$ O: $16.0 \div 18.0 = 0.889$ (rounded)	\longrightarrow H: $0.111 \times 100 = 11.1\%$ O: $0.889 \times 100 = 88.9\%$

Example: Calculate the percent composition of ammonium nitrate.

Description of Action	Action
1. Write the formula for the given compound.	1. NH_4NO_3
2. Record the amount of each element in the compound. (Note: We have 2 total nitrogen, so we record them together.) Multiply the amount of each element by its atomic weight (measured to the tenths place). Add the results to find the gram formula mass of the compound.	2. N: $2 \times 14.0 = 28.0$ H: $4 \times 1.0 = 4.0$ O: $3 \times 16.0 = 48.0$ 80.0 g/mol
3. Divide the total mass of each element by the gram formula mass. For these calculations your answer should have 3 places after the decimal (round if necessary).	3. N: $28.0 \div 80.0 = 0.350$ H: $4.0 \div 80.0 = 0.050$ O: $48.0 \div 80.0 = 0.600$
4. Multiply each result by 100. Add the % symbol to your new result. (If you were to add up your percentages the must equal 100.)	4. N: $0.350 \times 100 = 35.0\%$ H: $0.050 \times 100 = 5.0\%$ O: $0.600 \times 100 = 60.0\%$ 100.0%

Determining Empirical and Molecular Formula

Empirical Formula

Given: Percentage or mass of each element or compound.

1. Divide each percentage or mass given by the molar mass of the element or compound.
2. Divide each result by the smallest result.
3. Round and multiply (**if necessary**) each result by the SAME whole number to get a whole number result. (This step is not necessary for hydrates). Use the following chart:

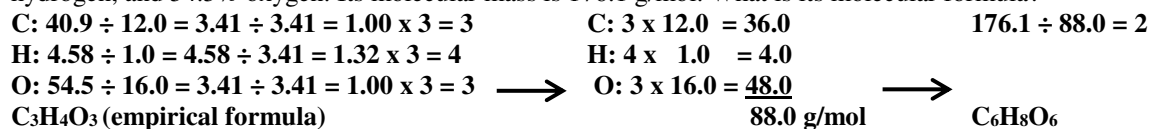
Round to:	Multiply by:
x.00	(use the other factor)
x.20	5
x.25	4
x.33	3
x.50	2
x.66	3
x.75	4
x.80	5

4. Write the formula using your results.

Molecular Formula (do all of the above)

5. Find the gram formula mass of the empirical formula.
6. Divide the molecular formula mass (given) by the empirical formula mass (calculated).
7. Multiply each subscript by the result. (The result MUST be a whole number.)

Example: Ascorbic acid, also known as vitamin C, has a percentage composition of 40.9% carbon, 4.58% hydrogen, and 54.5% oxygen. Its molecular mass is 176.1 g/mol. What is its molecular formula?



One question that has sometimes comes up on past AP exams is combustion analysis. To solve a combustion analysis problem one must know that all of the carbon in a compound is converted to carbon dioxide and all of the hydrogen in a compound is converted to water. Usually the percent analysis is given for any additional element, except oxygen. To find the mass of oxygen you must subtract the masses of the other elements from the total. Once

you know the mass of each element you can find the empirical and molecular formulas. Try the one below; the answers are given but the work is not.

Answer the following questions that relate to the analysis of chemical compounds.

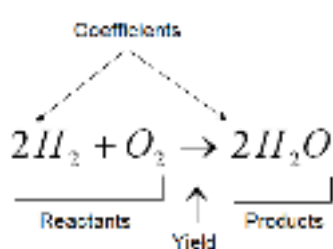
A compound containing the elements C, H, N, and O is analyzed. When a 1.2359 g sample is burned in excess oxygen, 2.241 g of $\text{CO}_2(g)$ is formed. The combustion analysis also showed that the sample contained 0.0648 g of H.

- Determine the mass, in grams, of C in the 1.2359 g sample of the compound. **0.6116 g C**
- When the compound is analyzed for N content only, the mass percent of N is found to be 28.84 percent. Determine the mass, in grams, of N in the original 1.2359 g sample of the compound. **0.3564 g N**
- Determine the mass, in grams, of O in the original 1.2359 g sample of the compound. **0.2031 g O**
- Determine the empirical formula of the compound. **$\text{C}_4\text{H}_5\text{N}_2\text{O}$**
- The molecular mass of the compound is 194.2 g/mol. Determine the molecular formula of the compound. **$\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$**

Chemical Equations

- Chemical changes are represented by chemical reactions.
- In a chemical reaction, the matter is not created or destroyed (Law of Conservation of Mass). Instead the atoms in the chemical reaction are rearranged. A balanced equation has equal numbers of each type of atom on each side of the equation.

Diatomic Elements
H₂ N₂ O₂
F₂ Cl₂ Br₂ I₂



Coefficients - the numbers written in front of the chemical formulas used in balancing chemical equations.

Reactants - the elements or compounds on the left hand side of the equation before the \rightarrow are known as the reactants.

Products - the elements or compounds on the right side of the equation after the \rightarrow are known as the products.

Yield(\rightarrow) - this arrow is called a yield sign. It separates the reactants from the products.

The word **diatomic** is used above when we read the equation. A **diatomic molecule** is made up of **two atoms of the same element**. These elements never exist as one naturally. When we are writing chemical equations you must always remember to write a subscript of two when these elements are by themselves. (When they are bonded to other elements they **do not** have to have a subscript of 2.) There are seven diatomic elements; they are listed in the chart to the right. You must **MEMORIZE** these seven diatomic elements. You can see that they kind of form a pattern on the periodic table.

TABLE 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane

Reactants		Products
$\text{CH}_4(g) + 2\text{O}_2(g)$	\longrightarrow	$\text{CO}_2(g) + 2\text{H}_2\text{O}(g)$
1 molecule + 2 molecules	\longrightarrow	1 molecule + 2 molecules
1 mole + 2 moles	\longrightarrow	1 mole + 2 moles
6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)	\longrightarrow	6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)
16 g + 2 (32 g)	\longrightarrow	44 g + 2 (18 g)
80 g reactants	\longrightarrow	80 g products

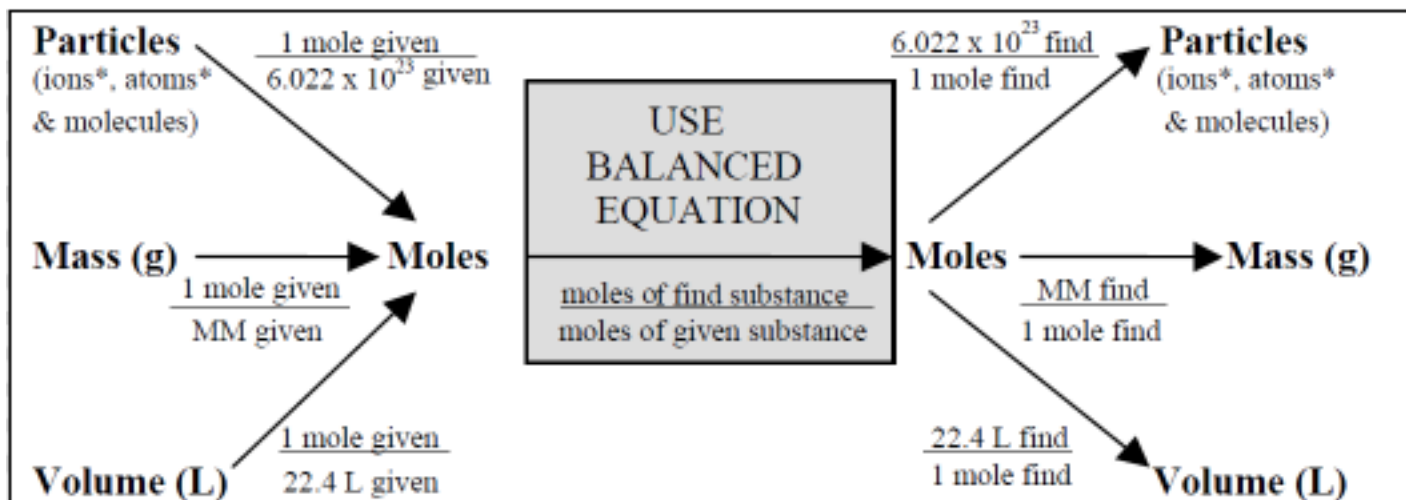
Balancing Chemical Equations

- There are three rules to follow when balancing chemical equations. They do not work every time, but they do work most times and most people find them helpful.
- If there is an element that is not bonded to any other atom or is diatomic, balance it last. For example in the equation: $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + \text{O}_2$, oxygen is diatomic. When balancing this equation, balance oxygen last.

- If you are balancing an equation with both hydrogen and oxygen and neither is a diatomic molecule, balance hydrogen last and balance oxygen second to last. For example in the equation: $\text{CO}_2 + \text{NH}_3 \rightarrow \text{OC}(\text{NH}_2)_2 + \text{H}_2\text{O}$, both hydrogen and oxygen are used and neither exist as a diatomic molecule. When balancing this equation, balance hydrogen last and oxygen second to last.
- You may come across symbols after every compound. (s) means the compound is in the solid phase, (l) means the compound is in a liquid phase, (g) means the compound is in the gas phase, and (aq) means the compound is dissolved in water (even if water is not one of the reactants or products)
- If you come to the point where you have an odd number of a certain element on one side of the equation and an even number of the same element on the opposite side on the equation, double the coefficient of the formula with the odd number of the element. If the coefficient is 1, change it to 2.
- Be aware of different kinds of reactions. Some are categorized as synthesis reactions (which are generalized to look like $\text{A} + \text{B} \rightarrow \text{AB}$), decomposition reactions (which are generalized to look like $\text{AB} \rightarrow \text{A} + \text{B}$), and combustion reactions (which are a type of *oxidation-reduction* reaction that will study further later, but generally look like $\text{C}_x\text{H}_y + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$)
- Ultimately it is a trial and error process. Keep at it until it works. Here are 10 practice problems. The answers are on the last page. Don't cheat. See if you can identify any synthesis, decomposition, or combustion reactions below.
 1. $\text{C}_{10}\text{H}_{16} + \text{Cl}_2 \rightarrow \text{C} + \text{HCl}$
 2. $\text{CO}_2 + \text{NH}_3 \rightarrow \text{OC}(\text{NH}_2)_2 + \text{H}_2\text{O}$
 3. $\text{Si}_2\text{H}_3 + \text{O}_2 \rightarrow \text{SiO}_2 + \text{H}_2\text{O}$
 4. $\text{Al}(\text{OH})_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{H}_2\text{O}$
 5. $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
 6. $\text{Fe}_2(\text{SO}_4)_3 + \text{KOH} \rightarrow \text{K}_2\text{SO}_4 + \text{Fe}(\text{OH})_3$
 7. $\text{C}_7\text{H}_6\text{O}_2 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
 8. $\text{H}_2\text{SO}_4 + \text{HI} \rightarrow \text{H}_2\text{S} + \text{I}_2 + \text{H}_2\text{O}$

Stoichiometry

- Here is a chart that provides a way for thinking about and doing stoichiometry:



START on THIS SIDE

FINISH on THIS SIDE

In a stoichiometry problem you always begin with a value of one of the units on the left and you always finish with a value of one of the units on the right. In every stoichiometry problem you must make a mole relationship comparison between substances being compared using their coefficients from the balanced chemical equation. This concept is used extensively.

Limiting Reactant

In a limiting reactant problem you will be given amounts of two of your reactants. The goal is to determine which you will run out of first. The only way to do this is to solve the stoichiometry problem using each reactant and its value. The reactant that produces **less** of the same product is the one you would run out of first, and thus your **limiting reactant**. The other reactant, the one you will have extra of when finished, is your **excess reactant**.

Percent Yield

The percent yield equation is used to determine the efficiency of a chemical procedure. When you perform a percent yield calculation you will be comparing how much product you **actually** produce in a lab situation with how much should theoretically be produced if the situation was ideal. The theoretical yield is what you determine in a stoichiometry problem. The percent yield equation is listed below.

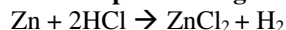
$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Example: Given the reaction $\text{C}_6\text{H}_6 + \text{HNO}_3 \rightarrow \text{C}_6\text{H}_5\text{NO}_2 + \text{H}_2\text{O}$, if a student began with 15.6 g of C_6H_6 in excess HNO_3 , then if all C_6H_6 were converted to product and isolated, the student would theoretically obtain 24.6 g of $\text{C}_6\text{H}_5\text{NO}_2$. If 18.0 g were actually produced, the percent yield could be calculated:

$$\text{percent yield} = \frac{18.0 \text{ g}}{24.6 \text{ g}} \times 100. \rightarrow \text{percent yield} = 73.2\%$$

Note: % error and % yield should add up to 100%

An example showing all three topics and the worked out solutions are included below.



Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas. 98.2 grams of zinc and 98.2 grams of hydrogen chloride react?

- Identify the limiting reactant. Support your answer with calculations. (3 points)
- How much of the excess reagent remains? (3 points)
- Calculate the volume of hydrogen gas produced. (3 points)
- If 12.5 liters of hydrogen gas are actually produced, what is the percent yield? (3 points)

$$\text{(a) } \frac{98.2 \text{ g Zn}}{1} \times \frac{1 \text{ mole Zn}}{65.4 \text{ g Zn}} \times \frac{1 \text{ mole H}_2}{1 \text{ mole Zn}} = 1.50 \text{ mol H}_2$$

$$\frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole H}_2}{2 \text{ mole HCl}} = 1.34 \text{ mol H}_2$$

Limiting Reagent: **HCl**

$$\text{(b) } \frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole Zn}}{2 \text{ mole HCl}} \times 65.4 \text{ g Zn} = 88.0 \text{ g Zn needed}$$

$$98.2 - 88.0 = 10.2 \text{ g of excess Zn}$$

$$\text{(c) } \frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole H}_2}{2 \text{ moles HCl}} \times 22.4 \text{ L H}_2 = 30.1 \text{ L H}_2$$

$$\text{(d) Percent Yield} = \frac{12.5}{30.1} \times 100 = 41.5\%$$

Balancing Answers:

- $\text{C}_{10}\text{H}_{16} + 8 \text{ Cl}_2 \rightarrow 10 \text{ C} + 16 \text{ HCl}$
- $\text{CO}_2 + 2 \text{ NH}_3 \rightarrow \text{OC}(\text{NH}_2)_2 + \text{H}_2\text{O}$
- $4 \text{ Si}_2\text{H}_3 + 11 \text{ O}_2 \rightarrow 8 \text{ SiO}_2 + 6 \text{ H}_2\text{O}$
- $2 \text{ Al}(\text{OH})_3 + 3 \text{ H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 6 \text{ H}_2\text{O}$
- $4 \text{ Fe} + 3 \text{ O}_2 \rightarrow 2 \text{ Fe}_2\text{O}_3$ (synthesis)
- $\text{Fe}_2(\text{SO}_4)_3 + 6 \text{ KOH} \rightarrow 3 \text{ K}_2\text{SO}_4 + 2 \text{ Fe}(\text{OH})_3$
- $2 \text{ C}_7\text{H}_6\text{O}_2 + 15 \text{ O}_2 \rightarrow 14 \text{ CO}_2 + 6 \text{ H}_2\text{O}$ (combustion)
- $\text{H}_2\text{SO}_4 + 8 \text{ HI} \rightarrow \text{H}_2\text{S} + 4 \text{ I}_2 + 4 \text{ H}_2\text{O}$

AP Chemistry Problem Set #1

1. Perform the following mathematical operations, and express the result to the correct number of significant figures.

- $(6.022 \times 10^{23}) \times (2.33 \times 10^3)$
- $1.00876 + 0.87206 - 0.0996$
- $(7.915 - 7.908) \div 7.915 \times 100.$
- $(3.000 \times 10^5) \div (4.00 \times 10^{-6})$
- $(2.38 \div 55.8) \times (6.022 \times 10^{23})$

2. Convert each of the following.

- 8.57 micrograms to centigrams
- 2.11×10^{-4} decaliters to milliliters
- 1.95×10^{11} nanometers to meters
- 2.27 kilograms to decigrams
- 6.19×10^{-8} megagrams to micrograms

3. Perform the following unit conversions. The unit conversions can be found online.

- 809 oz to kilograms
- 22.4 L to gallons
- 375 mL to quarts
- 221 pounds to grams
- 74°C to Kelvin

4. Solve the following using dimensional analysis. The unit conversions can be found online. Show all of your work.

a. A parsec is an astronomical unit of distance. 1 parsec = 3.26 light years (or the distance traveled by light in one year. Light speed = 186,282.397 miles per second. An object travels 9.6 parsecs. Calculate this distance in cm.

b. The front edge of the pitcher's mound is 60'6" from the rear point of home plate. If a power pitcher like Roger Clemens throws a fast ball at 95 miles/hour, how many seconds will it take for the ball to reach the catcher's mitt?

c. The current cost of gasoline is \$3.87/gallon. If my car gets 12.0 kilometers/liter, how many miles will I be able to travel if I put \$18.35 of gasoline in my car? If my house is 10.85 miles away from school, how many **complete round trips** can I make on \$18.35? At the above cost of gas, how much will I pay to make 194 **round trips** between home and work this year?

5. The density of pure platinum is 21.45 g/mL at 20°C . If 5.50 grams of pure platinum is added to 14.45 mL of water, to what volume will the level in the cylinder rise?

6. The amount of mercury in a polluted lake is $0.35 \mu\text{g Hg/mL}$, what is the total mass in kilograms of mercury in the lake? The lake has a surface area of $50. \text{mi}^2$ and an average depth of 30. ft.

7. A 20.00 gram sample of a solid is placed in a graduated cylinder and then filled to the 50.00 mL mark with benzene. The mass of the benzene and the solid together is 58.80 g. Assuming that the solid is insoluble in benzene and the density of benzene is 0.880 grams/cm³, calculate the density of the solid.

8. Cesium atoms are the largest naturally occurring atoms. The radius of a cesium atom is 2.62 Å (angstroms). How many cesium atoms would have to be laid side by side to give a row of cesium atoms 3.00 inches long? Assume that the atoms are spherical.

AP Chemistry Problem Set #2

1. _____ 2. _____ 3. _____ 4. _____ 5. _____
6. _____

1. Measurements indicate a charge of 0.444 C passes a point in 0.12 seconds. The current (i.e., the rate of charge flow, in C/s) is best expressed as:

- a. 0.27 C/s b. 0.270 C/s c. 3.7 C/s d. 3.70 C/s e. 3.700 C/s

2. A given sample contains 2.0 g of hydrogen, 33.1 g of sulfur, and 75.01 g of oxygen. What is the total mass of the sample?

- a. 110.12 g b. 110.1 g c. 110. g d. 1.1×10^2 g e. 1.1×10^{-2} g

3. The density of copper is 8.96 g/cm³. What is the mass of 18.88 cm³ of pure copper?

- a. 1.69 g b. 16.9 g c. 169 g d. 169.0 g e. 1690 g

4. The density of a piece of metal can be determined from mass and water displacement data. A piece of metal with a mass of 15.54 g is placed in a flask with a volume of 50.00 cm³. It is found that 40.54 g of water ($d=0.9971$ g/cm³) is needed to fill the flask with the metal in it. The density of the metal is most nearly (all answers in g/cm³):

- a. 1.66 b. 1.76 c. 9.46 d. 9.5 e. 40.7

5. Examples illustrating the Law of Multiple Proportions shown are:

- I. CO & CO₂
II. Ca & BaO
III. CaS & BaS
IV. Na₂CO₃ & Na₂SO₄
V. O₂ & O₃

- a. I only b. I & V c. III & IV d. III & V e. I & III

6. Which of the following elements is a transition metal?

- a. Ca b. S c. Fe d. N e. Cs

7. Complete the following table:

#p ⁺	#n ⁰	#e ⁻	mass #	atomic #	net charge	symbol
11	12	10				
			1	1	1+	
			79			Se ²⁻
		85	210	85		
	118			79	3+	
19			41		1+	

8. Identify each of the following elements:

- a. a member of the same family as oxygen whose most stable ion contains 54 electrons
b. a noble gas with 18 protons in the nucleus
c. a halogen with 85 protons and 85 electrons
d. a member of the alkali metal family whose most stable ion contains 18 electrons
e. a member of group 17 whose most stable ion contains 10 electrons

9. Suppose that a stable element, atomic number 119, symbol Pe, name Petrassium, is discovered.
- Would Pe be a metal or a non-metal? Explain/justify your answer.
 - What would be the most likely charge of the Pe ion in stable ionic compounds?
 - An isotope of Petrassium has a mass number of 291. How many neutrons does it have?
 - Write the formula for the compound formed between Pe and the carbonate (CO_3) ion.

AP Chemistry Problem Set #3

1. Convert each of the following. Assume that all substances are at standard temperature and pressure (STP):

- 43.0 grams of nitrogen dioxide to liters
- 56.0 milligrams of cesium iodide to molecules
- 2.6×10^{15} molecules of ammonium phosphate to grams
- 2.40 liters of carbon monoxide to moles
- 5.70 millimoles of sodium chlorate to grams

2. Sucrose, $C_{12}H_{22}O_{11}$, is the main compound in table sugar and is produced through the extraction and refining of cane or beet sugar. Show all of your work.

- Calculate the molar mass of sucrose.
- Calculate the percent composition of the elements in sucrose.
- What is the mass in grams of 8.88 moles of sucrose?
- How many molecules are there in 4.20 mg of sucrose?
- What is the mass of one molecule of sucrose?

3. Calculate the average atomic mass for each of the following groups of isotopes. The percentage listed with each isotope is the relative abundance.

- ^{28}Si (92.21%), ^{29}Si (4.70%), ^{30}Si (3.09%)
- ^{113}In (4.3%), ^{115}In (95.7%)
- ^{32}S (95.0%), ^{33}S (0.8%), ^{34}S (4.2%)
- ^{64}Zn (48.89%), ^{66}Zn (27.81%), ^{67}Zn (4.11%), ^{68}Zn (18.57%), ^{70}Zn (0.62%)
- ^{70}Ge (21.2%), ^{72}Ge (27.7%), ^{73}Ge (7.7%), ^{74}Ge (34.9%), ^{76}Ge (7.4%)

4. Solve each of the following empirical and molecular formulas:

a. A compound that contains only nitrogen and oxygen is 30.4% N by mass; the molar mass of the compound is 92 g/mol. What is the empirical formula of the compound? What is the molecular formula of the compound?

b. Maleic acid is an organic compound composed of 41.39% C, 3.47% H, and the rest oxygen. If 0.129 mol of maleic acid has a mass of 15.0 g, what are the empirical and molecular formulas of maleic acid?

c. Ascorbic acid, also known as vitamin C, has a percentage composition of 40.9% carbon, 4.58% hydrogen, and 54.5% oxygen. Its molecular mass is 176.1 g/mol. What is its molecular formula?

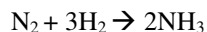
d. Aspirin contains 60.0% carbon, 4.48% hydrogen, and 35.5% oxygen. It has a molecular mass of 180.0 g/mol. What are its empirical and molecular formulas?

e. A substance has a percent composition of 39.6% carbon, 1.7% hydrogen, 58.7% chlorine. Determine the molecular formula of the substance if it has a molecular mass of 544.5 g/mol

5. Write complete balanced equations for each of the following:

- solid calcium oxide (CaO) reacts with aqueous hydrochloric acid (HCl) to yield aqueous calcium chloride (CaCl_2) and liquid water (H_2O)
- aqueous barium chloride (BaCl_2) reacts with aqueous sodium sulfite (Na_2SO_3) to produce solid barium sulfite (BaSO_3) and aqueous sodium chloride (NaCl)
- liquid benzene (C_6H_6) and oxygen gas (O_2) are reacted together which produces gaseous carbon dioxide (CO_2) and liquid water (H_2O)
- aqueous lead(II) nitrate ($\text{Pb}(\text{NO}_3)_2$) and aqueous potassium iodide (KI) are reacted together which yields solid lead(II) iodide (PbI_2) and aqueous potassium nitrate (KNO_3)
- fluorine gas (F_2) reacts with solid calcium chloride (CaCl_2) to yield solid calcium fluoride (CaF_2) and gaseous chlorine (Cl_2)

6. Consider the reaction:



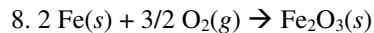
Identify the limiting reagent in each of the mixtures below. Assume all experiments occur at STP:

- 30 molecules of N_2 & 45 molecules of H_2
- 3 moles of N_2 & 1 mole of H_2
- 50.00 L of N_2 & 50.00 L of H_2
- 28.00 grams of N_2 & 8.00 grams of H_2
- 12 atoms of N_2 & 20 atoms of H_2

7. For each of the following identify the limiting reagent, theoretical yield and percent yield.

a. Aluminum (Al) reacts with aqueous chromium(II) oxide (CrO) to form aluminum oxide (Al_2O_3) and chromium (Cr). Determine the limiting reagent, theoretical yield and percent yield if 187.0 grams of chromium(II) oxide were used with 214.0 grams of aluminum and 88.0 grams of aluminum oxide were actually produced.

b. Zinc (Zn) reacts with hydrochloric acid (HCl) to form zinc chloride (ZnCl_2) and hydrogen gas (H_2). Calculate the limiting reagent, theoretical yield and percent yield if you are given 79.2 grams of zinc and 79.2 grams of hydrochloric acid and you actually produce 9.1 liters of hydrogen gas?



Iron (Fe) reacts with oxygen to produce iron(III) oxide (Fe_2O_3), as represented by the equation above. A 95.0 g sample of Fe(s) is mixed with 15.1 L of $\text{O}_2(g)$ at STP.

a. Calculate the number of moles of each of the following before the reaction begins.

- Fe(s)
- $\text{O}_2(g)$

b. Identify the limiting reactant when the mixture is heated to produce $\text{Fe}_2\text{O}_3(s)$. Support your answer with calculations.

c. Calculate the number of moles of $\text{Fe}_2\text{O}_3(s)$ produced when the reaction proceeds to completion.

AP Chemistry Problem Set #1 - ANSWERS

1. Perform the following mathematical operations, and express the result to the correct number of significant figures.

- $(6.022 \times 10^{23}) \times (2.33 \times 10^3) = \mathbf{1.40 \times 10^{27}}$
- $1.00876 + 0.87206 - 0.0996 = \mathbf{1.7812}$
- $(7.915 - 7.908) \div 7.915 \times 100. = \mathbf{0.09}$
- $(3.000 \times 10^5) \div (4.00 \times 10^{-6}) = \mathbf{7.50 \times 10^{10}}$
- $(2.38 \div 55.8) \times (6.022 \times 10^{23}) = \mathbf{2.57 \times 10^{22}}$

2. Convert each of the following.

- 8.57 micrograms to centigrams = $\mathbf{8.57 \times 10^{-4}}$ centigrams
- 2.11×10^{-4} dekaliters to milliliters = $\mathbf{2.11}$ milliliters
- 1.95×10^{11} nanometers to meters = $\mathbf{1.95 \times 10^2}$ meters
- 2.27 kilograms to decigrams = $\mathbf{2.27 \times 10^4}$ decigrams
- 6.19×10^{-8} megagrams to micrograms = $\mathbf{6.19 \times 10^4}$ micrograms

3. Perform the following unit conversions. The unit conversions can be found online.

- 809 oz to kilograms = $\mathbf{22.9}$ kilograms
- 22.4 L to gallons = $\mathbf{5.92}$ gallons
- 375 mL to quarts = $\mathbf{0.396}$ quarts
- 221 pounds to grams = $\mathbf{1.00 \times 10^5}$ grams
- 74° C to Kelvin = $\mathbf{347}$ Kelvin

4. Solve the following using dimensional analysis. The unit conversions can be found online. Show all of your work.

- A parsec is an astronomical unit of distance. 1 parsec = 3.26 light years (or the distance traveled by light in one year. Light speed = 186,282.397 miles per second. An object travels 9.6 parsecs. Calculate this distance in cm.

$$9.6 \text{ parsecs} \times \frac{3.26 \text{ yr}}{1.0 \text{ parsec}} \times \frac{365.25 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hr}}{1 \text{ day}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ sec}}{1 \text{ min}} \times \frac{186282.397 \text{ miles}}{1 \text{ sec}} \times \frac{5280 \text{ ft}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ foot}} \times \frac{2.54 \text{ cm}}{1 \text{ inch}}$$

$\mathbf{3.0 \times 10^{19} \text{ cm}}$

- The front edge of the pitcher's mound is 60'6" from the rear point of home plate. If a power pitcher like Roger Clemens throws a fast ball at 95 miles/hour, how many seconds will it take for the ball to reach the catcher's mitt?

$$\frac{60.5 \text{ ft}}{1} \times \frac{1 \text{ mile}}{5280 \text{ ft}} \times \frac{1 \text{ hour}}{95 \text{ miles}} \times \frac{60 \text{ min}}{1 \text{ hour}} \times \frac{60 \text{ sec}}{1 \text{ min}} = \mathbf{0.43 \text{ seconds}}$$

- The current cost of gasoline is \$3.87/gallon. If my car gets 12.0 kilometers/liter, how many miles will I be able to travel if I put \$18.35 of gasoline in my car? If my house is 10.85 miles away from school, how many **complete round trips** can I make on \$18.35? At the above cost of gas, how much will I pay to make 194 **round trips** between home and work this year?

$$1 \text{ gal} = 3.78541 \text{ L}$$

$$1 \text{ km} = 0.621371 \text{ mi}$$

$$\frac{18.35 \text{ dollars}}{1} \times \frac{1 \text{ gallon}}{3.87 \text{ dollars}} \times \frac{3.7854 \text{ L}}{1 \text{ gallon}} \times \frac{12.0 \text{ km}}{1 \text{ L}} \times \frac{0.62137 \text{ mi}}{1 \text{ km}} = \mathbf{134 \text{ miles}}$$

$$\frac{134 \text{ miles}}{1} \times \frac{1 \text{ way trip}}{10.85 \text{ miles}} \times \frac{1 \text{ round trip}}{2 \text{ way trip}} = 6.16 = \mathbf{6 \text{ complete round trips}}$$

$$\frac{194 \text{ trips}}{1} \times \frac{21.70 \text{ miles}}{1 \text{ round trip}} \times \frac{1 \text{ gallon}}{28.23 \text{ miles}} \times \frac{3.87 \text{ dollars}}{1 \text{ gallon}} = \mathbf{\$577}$$

5. The density of pure platinum is 21.45 g/mL at 20°C. If 5.50 grams of pure platinum is added to 14.45 mL of water, to what volume will the level in the cylinder rise?

$$\frac{21.45 \text{ g}}{1 \text{ mL}} = \frac{5.50 \text{ g}}{x \text{ mL}}$$

$$x = 0.256 \text{ mL}$$

$$0.256 \text{ mL} + 14.45 \text{ mL} = 14.706 \text{ mL}$$

$$14.706 \text{ mL (rounded to)} \mathbf{14.71 \text{ mL}}$$

6. The amount of mercury in a polluted lake is $0.35 \mu\text{g Hg/mL}$, what is the total mass in kilograms of mercury in the lake? The lake has a surface area of $50. \text{mi}^2$ and an average depth of $30. \text{ft}$.

$$1 \text{ ft} = 30.48 \text{ cm}$$

$$1 \text{ mi} = 160934 \text{ cm}$$

$$\frac{50 \text{ mi}^2 \times (5280 \text{ ft})^2}{1 \quad 1.00 \text{ mi}^2} \times 30 \text{ ft} = \frac{4.18176 \times 10^{10} \text{ ft}^3}{1} \times \frac{(12 \text{ in})^3}{1 \text{ ft}^3} \times \frac{(2.54 \text{ cm})^3}{1 \text{ in}^3} = 1.184142564 \times 10^{15} \text{ cm}^3$$

$$1.184142564 \times 10^{15} \text{ cm}^3 = \frac{1.184142564 \times 10^{15} \text{ mL}}{1} \times \frac{0.35 \mu\text{g}}{1 \text{ mL}} \times \frac{1 \text{ kg}}{10^9 \mu\text{g}} = \mathbf{4.1 \times 10^5 \text{ kg Hg}}$$

7. A 20.00 gram sample of a solid is placed in a graduated cylinder and then filled to the 50.00 mL mark with benzene. The mass of the benzene and the solid together is 58.80 g . Assuming that the solid is insoluble in benzene and the density of benzene is 0.880 grams/cm^3 , calculate the density of the solid.

$$\text{Benzene mass} = 58.80 \text{ g} - 20.00 \text{ grams} = 38.80 \text{ grams}$$

$$\text{Benzene density} = 0.880 \text{ g/mL (given)}$$

$$\text{Benzene volume} = 44.1 \text{ mL}$$

$$\text{Solid volume} = 50.00 - 44.0909 = 5.91 \text{ mL}$$

$$\text{Solid mass} = 20.00 \text{ grams (given)}$$

$$\text{Solid density} = \mathbf{3.38 \text{ g/mL}}$$

8. Cesium atoms are the largest naturally occurring atoms. The radius of a cesium atom is 2.62 \AA (angstroms). How many cesium atoms would have to be laid side by side to give a row of cesium atoms 3.00 inches long? Assume that the atoms are spherical.

$$1 \text{ \AA} = 10^{-10} \text{ m}$$

$$1 \text{ in} = 2.54 \text{ cm}$$

$$2.62 \text{ \AA} = \text{radius of 1.0 atom}$$

$$5.24 \text{ \AA} = \text{diameter of 1.0 atom}$$

$$\frac{3.00 \text{ in} \times 2.54 \text{ cm}}{1 \quad 1.00 \text{ in}} \times \frac{1.00 \text{ m}}{100 \text{ cm}} \times \frac{1.00 \text{ \AA}}{1 \times 10^{-10} \text{ m}} \times \frac{1 \text{ atom}}{5.24 \text{ \AA}} = \mathbf{1.45 \times 10^8 \text{ atoms}}$$

AP Chemistry Problem Set #2

1. C 2. B 3. C 4. A 5. A
6. C

1. Measurements indicate a charge of 0.444 C passes a point in 0.12 seconds. The current (i.e., the rate of charge flow, in C/s) is best expressed as:

- a. 0.27 C/s b. 0.270 C/s c. **3.7 C/s** d. 3.70 C/s e. 3.700 C/s

2. A given sample contains 2.0 g of hydrogen, 33.1 g of sulfur, and 75.01 g of oxygen. What is the total mass of the sample?

- a. 110.12 g b. **110.1 g** c. 110. g d. 1.1×10^2 g e. 1.1×10^{-2} g

3. The density of copper is 8.96 g/cm³. What is the mass of 18.88 cm³ of pure copper?

- a. 1.69 g b. 16.9 g c. **169 g** d. 169.0 g e. 1690 g

4. The density of a piece of metal can be determined from mass and water displacement data. A piece of metal with a mass of 15.54 g is placed in a flask with a volume of 50.00 cm³. It is found that 40.54 g of water (d=0.9971 g/cm³) is needed to fill the flask with the metal in it. The density of the metal is most nearly (all answers in g/cm³):

- a. **1.66** b. 1.76 c. 9.46 d. 9.5 e. 40.7

5. Examples illustrating the Law of Multiple Proportions shown are:

I. CO & CO₂

II. Ca & BaO

III. CaS & BaS

IV. Na₂CO₃ & Na₂SO₄

V. O₂ & O₃

- a. **I only** b. I & V c. III & IV d. III & V e. I & III

6. Which of the following elements is a transition metal?

- a. Ca b. S c. **Fe** d. N e. Cs

7. Complete the following table:

#p ⁺	#n ⁰	#e ⁻	mass #	atomic #	net charge	symbol
11	12	10	23	11	1+	Na⁺
1	0	0	1	1	1+	H⁺
34	45	36	79	34	2-	Se ²⁻
85	125	85	210	85	0	At
79	118	76	197	79	3+	Au³⁺
19	22	18	41	19	1+	K⁺

8. Identify each of the following elements:

- a. a member of the same family as oxygen whose most stable ion contains 54 electrons - **Te**
b. a noble gas with 18 protons in the nucleus - **Ar**
c. a halogen with 85 protons and 85 electrons - **At**
d. a member of the alkali metal family whose most stable ion contains 18 electrons - **K**
e. a member of group 17 whose most stable ion contains 10 electrons - **F**

9. Suppose that a stable element, atomic number 119, symbol Pe, name Petrassium, is discovered.

a. Would Pe be a metal or a non-metal? Explain/justify your answer. – **metal due to its location on the left side of the periodic table**

b. What would be the most likely charge of the Pe ion in stable ionic compounds? – **1+**

c. An isotope of Petrassium has a mass number of 291. How many neutrons does it have? - **172**

d. Write the formula for the compound formed between Pe and the carbonate (CO_3) ion. – **Pe_2CO_3**

AP Chemistry Problem Set #3

1. Convert each of the following. Assume that all substances are at standard temperature and pressure (STP)

- 43.0 grams of nitrogen dioxide to liters – **20.9 L**
- 56.0 milligrams of cesium iodide to molecules – **1.30×10^{20} molecules**
- 2.6×10^{15} molecules of ammonium phosphate to grams – **6.4×10^{-7} grams**
- 2.40 liters of carbon monoxide to moles – **0.107 moles**
- 5.70 millimoles of sodium chlorate to grams – **0.607 grams**

2. Sucrose, $C_{12}H_{22}O_{11}$, is commonly known as table sugar and is produced through the extraction and refining of cane or beet sugar. Show all of your work.

- Calculate the molar mass of sucrose. – **342.296 g/mol**
- Calculate the percent composition of the elements in sucrose. – **C (42.1%), H (6.5%), O (51.4%)**
- What is the mass in grams of 8.88 moles of sucrose? – **3040. g**
- How many molecules are there in 4.20 mg of sucrose? – **7.39×10^{18}**
- What is the mass of one molecule of sucrose? – **5.69×10^{-22} g**

3. Calculate the average atomic mass for each of the following groups of isotopes. The percentage listed with each isotope is the relative abundance.

- ^{28}Si (92.21%), ^{29}Si (4.70%), ^{30}Si (3.09%) – **28.1 g/mol**
- ^{113}In (4.3%), ^{115}In (95.7%) – **114.9 g/mol**
- ^{32}S (95.0%), ^{33}S (0.8%), ^{34}S (4.2%) – **32.1 g/mol**
- ^{64}Zn (48.89%), ^{66}Zn (27.81%), ^{67}Zn (4.11%), ^{68}Zn (18.57%), ^{70}Zn (0.62%) – **65.5 g/mol**
- ^{70}Ge (21.2%), ^{72}Ge (27.7%), ^{73}Ge (7.7%), ^{74}Ge (34.9%), ^{76}Ge (7.4%) – **72.0 g/mol**

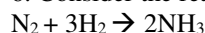
4. Solve each of the following empirical and molecular formulas:

- A compound that contains only nitrogen and oxygen is 30.4% N by mass; the molar mass of the compound is 92 g/mol. What is the empirical formula of the compound? What is the molecular formula of the compound? **NO_2 & N_2O_4**
- Maleic acid is an organic compound composed of 41.39% C, 3.47% H, and the rest oxygen. If 0.129 mol of maleic acid has a mass of 15.0 g, what are the empirical and molecular formulas of maleic acid? **CHO & $\text{C}_4\text{H}_4\text{O}_4$**
- Ascorbic acid, also known as vitamin C, has a percentage composition of 40.9% carbon, 4.58% hydrogen, and 54.5% oxygen. Its molecular mass is 176.1 g/mol. What is its molecular formula? – **$\text{C}_6\text{H}_8\text{O}_6$**
- Aspirin contains 60.0% carbon, 4.48% hydrogen, and 35.5% oxygen. It has a molecular mass of 180.0 g/mol. What are its empirical and molecular formulas? – **$\text{C}_9\text{H}_8\text{O}_4$ & $\text{C}_9\text{H}_8\text{O}_4$**
- A substance has a percent composition of 39.6% carbon, 1.7% hydrogen, 58.7% chlorine. Determine the molecular formula of the substance if it has a molecular mass of 544.5 g/mol – **$\text{C}_{18}\text{H}_9\text{Cl}_9$**

5. Write complete balanced equations for each of the following:

- solid calcium oxide reacts with aqueous hydrochloric acid to yield aqueous calcium chloride and liquid water – **$\text{CaO (s)} + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 \text{ (aq)} + \text{H}_2\text{O (l)}$**
- aqueous barium chloride reacts with aqueous sodium sulfite to produce solid barium sulfite and aqueous sodium chloride – **$\text{BaCl}_2 \text{ (aq)} + \text{Na}_2\text{SO}_3 \text{ (aq)} \rightarrow \text{BaSO}_3 \text{ (s)} + 2\text{NaCl (aq)}$**
- liquid benzene (C_6H_6) and oxygen gas are reacted together which produces gaseous carbon dioxide and liquid water – **$2\text{C}_6\text{H}_6 \text{ (l)} + 15\text{O}_2 \text{ (g)} \rightarrow 12\text{CO}_2 \text{ (g)} + 6\text{H}_2\text{O (l)}$**
- aqueous lead(II) nitrate and aqueous potassium iodide are reacted together which yields solid lead(II) iodide and aqueous potassium nitrate – **$\text{Pb(NO}_3)_2 \text{ (aq)} + 2\text{KI (aq)} \rightarrow \text{PbI}_2 \text{ (s)} + 2\text{KNO}_3 \text{ (aq)}$**
- fluorine gas reacts with solid calcium chloride to yield solid calcium fluoride and gaseous chlorine – **$\text{F}_2 \text{ (g)} + \text{CaCl}_2 \text{ (s)} \rightarrow \text{CaF}_2 \text{ (s)} + \text{Cl}_2 \text{ (g)}$**

6. Consider the reaction:



Identify the limiting reagent in each of the mixtures below. Assume all experiments occur at STP:

- 30 molecules of N_2 & 45 molecules of H_2 – **H_2**
- 3 moles of N_2 & 1 mole of H_2 – **H_2**
- 50.00 L of N_2 & 50.00 L of H_2 – **H_2**
- 28.00 grams of N_2 & 8.00 grams of H_2 – **N_2**
- 12 atoms of N_2 & 20 atoms of H_2 – **H_2**

7. For each of the following identify the limiting reagent, theoretical yield and percent yield.

a. Aluminum reacts with aqueous chromium(II) oxide to form aluminum oxide and chromium. Determine the limiting reagent, theoretical yield and percent yield if 187.0 grams of chromium(II) oxide were used with 214.0 grams of aluminum and 88.0 grams of aluminum oxide were actually produced.

L.R.: CrO

Theoretical: 93.46 g

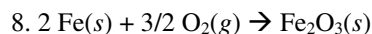
% Yield: 94.2%

b. Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas. Calculate the limiting reagent, theoretical yield and percent yield if you are given 79.2 grams of zinc and 79.2 grams of hydrochloric acid and you actually produce 9.1 liters of hydrogen gas?

L.R.: HCl

Theoretical: 24.32 L

% Yield: 37.4%



Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 95.0 g sample of $\text{Fe}(s)$ is mixed with 15.1 L of $\text{O}_2(g)$ at STP.

a. Calculate the number of moles of each of the following before the reaction begins.

(i) $\text{Fe}(s)$ – **1.70 moles**

(ii) $\text{O}_2(g)$ – **0.674 moles**

b. Identify the limiting reactant when the mixture is heated to produce $\text{Fe}_2\text{O}_3(s)$. Support your answer with calculations. - **O_2**

c. Calculate the number of moles of $\text{Fe}_2\text{O}_3(s)$ produced when the reaction proceeds to completion.

0.449 moles