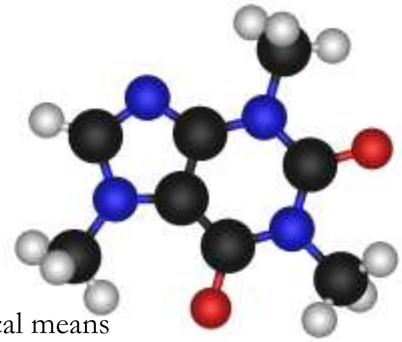
AP* Chemistry
Chemical Foundations and Stoichiometry: Review of Chemistry I



What is Chemistry?
Chemistry: the study of matter and change

Matter: has mass and occupies volume; exhibits inertia

Element: simplest form of matter that can't be broken into simpler substances by chemical means

Compound: two or more elements chemically combined; atoms bonded in a defined proportion

Pure Substance: material that is composed of only one type of particle; an element or compound

Mixture: material made of two or more pure substances physically combined; substances retain their properties

Making Measurements
The number of significant digits is determined by the precision of the measuring tool. The smaller the markings on the tool, the more precisely it can measure. In order to correctly use a measuring device, the operator must estimate one digit beyond the smallest marking on the tool. For digital balances, the balance "estimates" the last digit.

When a number is presented to you, it is necessary to understand how precisely the measurement was made. You must recognize how many significant figures are in the number.

1. Zeros at the beginning of a number are never significant *(serve only as place holders, do not communicate precision… "significance").*
2. Zeros at the end of a number are not significant **unless** there is a decimal point in the number. A decimal point **anywhere** in the number makes zeros at the end of a number significant.
3. Zeros that are between two nonzero numbers are always significant.

Significant figures in calculations:

- When multiplying and dividing numbers, count the number of significant figures in each number. Round your answer to the least number of significant figures.
- In adding and subtracting, look at the location of the last significant figure in each value. Round to the least specific decimal place of any number in the problem.

Accuracy: closeness of a measurement to the correct/accepted value; correctness
Precision: closeness of a set of measurements to one another; reproducibility

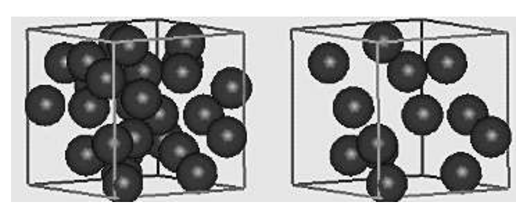
Dimensional Analysis
- Method for converting between units or equivalent values in calculations.
- Used in mole calculations and stoichiometry
- All work w/ units must be shown

Basic set-up for dimensional analysis problems:

$$\begin{array}{c|c|} \text{Given amount \& unit} & \text{Unit for answer w/ number} \\ \hline & \text{Given unit w/ number} \end{array} = \text{Answer w/ unit}$$

Density
Density: mass per unit volume

- A defined, intensive property
  - Pay attention to units and significant figures
  - Density of water = 1 g/mL



$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \qquad \text{For gases at STP, Density} = \frac{\text{Molar mass (from PT)}}{\text{Molar Volume (22.4 L/mol)}}$$

Significant Figure, Dimensional Analysis, and Density Sample Calculations:

1. How many centimeters tall is a 6.00 ft statue? (1 inch = 2.54 cm)

2. How many milliliters are in a gallon? (1 gallon = 3.785 L)

3. An empty vial has a mass of 99.85 g. If the vial has a mass of 244.03 g when completely filled with liquid mercury (d = 13.53 g/cm^3), what is the volume of the vial?

4. To measure the density of an unknown metal sample, a chemist performed the following set of measurements. He massed the sample and found its mass to be 69.966 g. Then he placed the sample in a graduated cylinder containing 50.000 mL of oil. The level of the oil rose to 74.403 mL. What is the density of the unknown metal?

Historical Perspective

Inquiry into the composition of matter dates back to early history. Around 400 BC, Greek philosopher Democritus is credited with describing the world as made up of tiny indivisible particles called *atomos* (indivisible). This idea lacked evidence and support. For hundreds of years chemistry went through significant changes until a focus on quantitative measurement and experimental evidence was able to lead to the development of scientific laws.

<u>Law of Conservation of Mass</u>: matter can neither be created nor destroyed

<u>Law of Definite Proportions</u>: a given compound always contains the same proportions of elements by mass

<u>Law of Multiple Proportions</u>: when two elements combine to form a series of compounds, the ratio of the masses of the second element that combine with 1 g of the first element can always be reduced to small, whole numbers

<u>Dalton's Atomic Theory</u>:

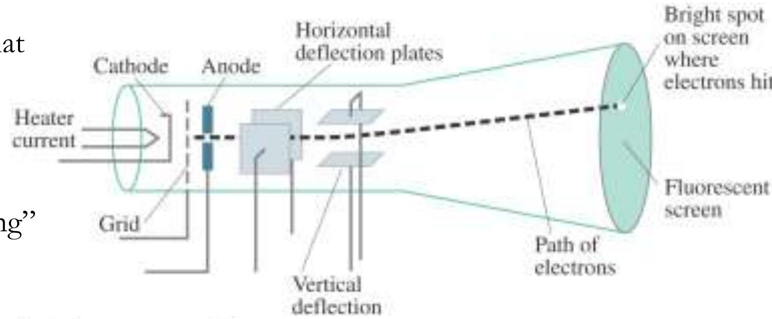
1. Each element is composed of extremely small particles* called atoms. *Thought to be indivisible, now known to be divisible
2. All atoms of a given element are identical*, but the atoms of one element are different from the atoms of all other elements. *Since disproved
3. Atoms of one element cannot be changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.
4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kinds of atoms.

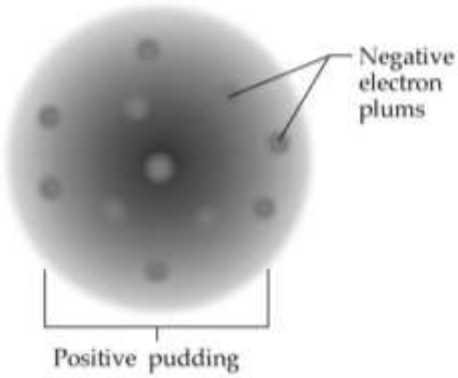
<u>Avogadro's Hypothesis</u>: Equal volumes of different gases, and the same temperature and pressure, contain the same number of particles. *(Led to concept of the mole)*

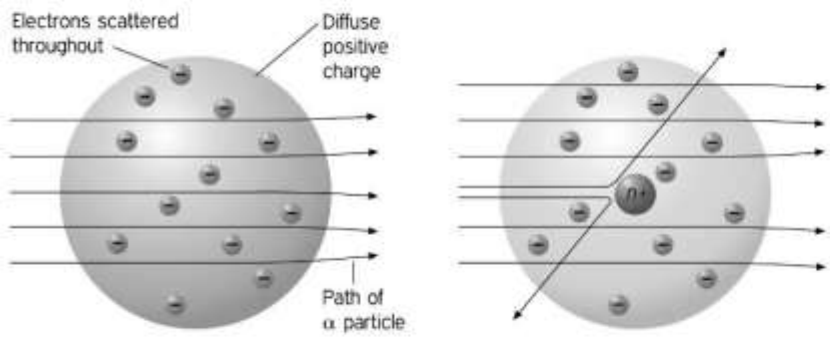
Discovery of Atomic Structure

Dalton's Atomic Theory was significant because it was based on experimental evidence. Many other advances were made during the Scientific/Industrial Revolution as equipment was developed that allowed precise measurement. Boyle, Gay-Lussac, Lavoisier, Proust, and others made discoveries that changed the course of chemistry. Understanding the atom and its structure was the next big challenge.

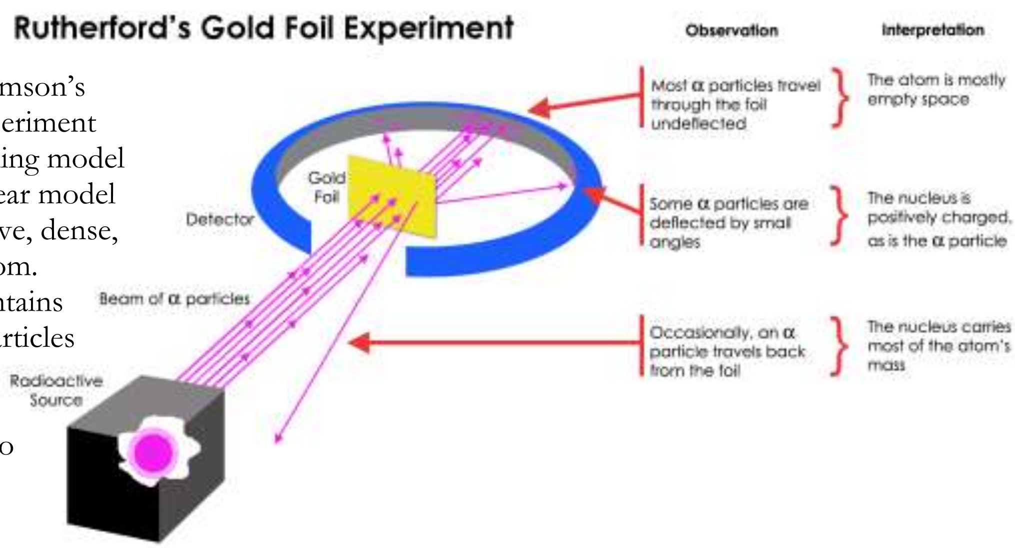
When working with cathode rays, J.J. Thomson proved that all atoms consisted of the same small negatively-charged particles. He determined the charge-to-mass ratio by looking at the deflection of the cathode beam in an electric field. This experiment proved the existence of electrons and helped Thomson develop the "Plum Pudding" model of the atom.



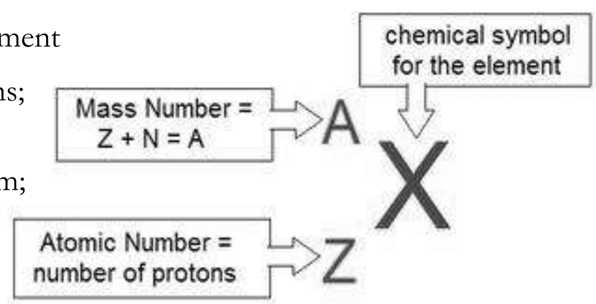




Ernest Rutherford used radioactive materials to attempt to support Thomson's atomic model. The result of his experiment disproved aspects of the Plum Pudding model and laid the ground work for a nuclear model of the atom. He discovered a massive, dense, positively-charged portion of the atom. This is called the nucleus, which contains the protons and neutrons. These particles account for most of the mass and all of the positive charge in an atom. The electrons were thought to orbit the nucleus.



<u>Key Aspects of Atomic Structure</u>:



Atom: smallest piece of an element that retains the properties of that element

Nucleus: dense center of the atom that contains the protons and neutrons;
>99% of mass but <1% of volume

Proton: positively charged subatomic particle; determines identity of atom;
found in nucleus

Neutron: neutral subatomic particle; found in nucleus;
contributes mass but no charge; changes nuclear properties

Electron: negatively charged subatomic particle; determines charge of atom/ion; found outside nucleus; mass ~ 0

Atomic Number (Z): number of protons in an atom

Mass Number: number of protons and neutrons in an atom; actual mass of atom is <u>not</u> a whole number
(actual mass of atom accounts for small mass of electrons AND mass defect- mass lost in nucleus formation)

Isotope: atoms of the same element (same Z, # of protons) with different masses (different number of neutrons)

Ion: charged atom; has gained or lost electrons

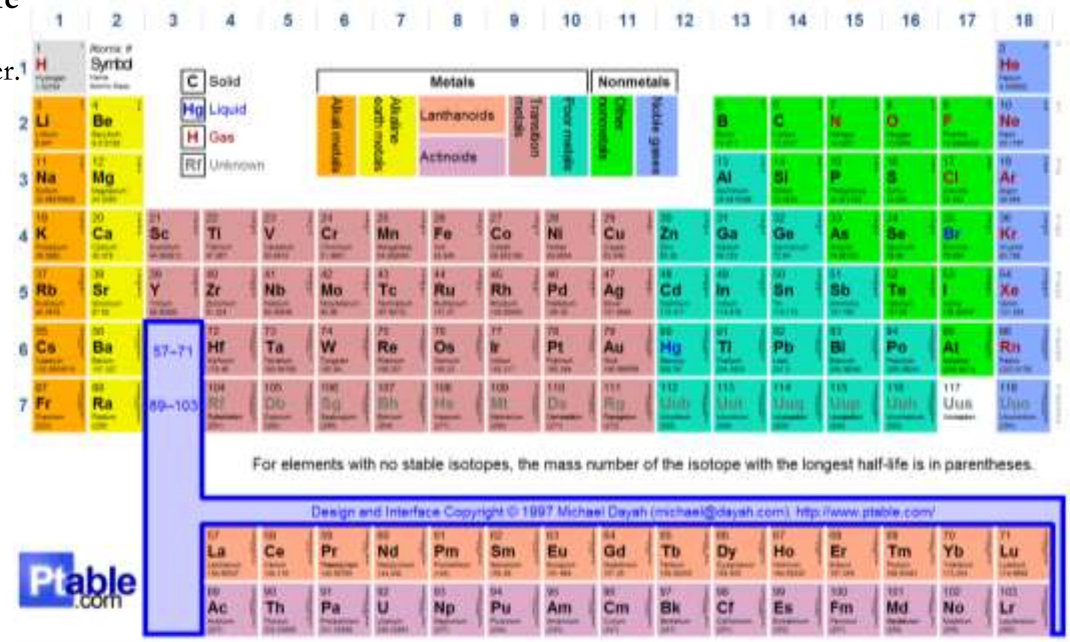
Periodic Table & Nomenclature

The periodic table is arranged in order of increasing atomic number.[1] When arranged by increasing atomic number, chemical and physical properties repeat in regular (periodic) intervals.

Rows on PT = period

Columns on PT = group/family
- "Tall" columns, 1-2, 13-18 = Representative elements
- "Short" columns, 3-12 = Transition elements

Important Family Names:
- 1 = alkali metals
- 2 = alkaline earth metals
- 3-12 = transition metals
- 17 = halogens
- 18 = noble gases



Metals = left side **Nonmetals = right side**

Metalloids/semiconductors = "stair steps"

<u>Nomenclature</u>

Ionic compounds:

Cation and anion are combined so that overall charge is neutral (cancelled)

Cation = metal/polyatomic with positive charge… Name of cation is same as name of element

Anion = nonmetal/polyatomic with negative charge… Name of anion is name of element with –ide ending (polyatomic ions don't change suffix)

Polyatomic ions to know:

–1 Charge				**–2 Charge**		**–3 Charge**	
$C_2H_3O_2^-$	Acetate	ClO_2^-	Chlorite	SO_4^{2-}	Sulfate	PO_4^{3-}	Phosphate
NO_3^-	Nitrate	ClO^-	Hypochlorite	SO_3^{2-}	Sulfite	PO_3^{3-}	Phosphite
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate	CO_3^{2-}	Carbonate		
CN^-	Cyanide	HSO_4^-	Hydrogen sulfate	CrO_4^{2-}	Chromate	**+ 1 Charge**	
OH^-	Hydroxide	HSO_3^-	Hydrogen sulfite	$Cr_2O_7^{2-}$	Dichromate	H_3O^+	Hydronium
MnO_4^-	Permanganate	$H_2PO_4^-$	Dihydrogen phosphate	$S_2O_3^{2-}$	Thiosulfate	NH_4^+	Ammonium
ClO_4^-	Perchlorate	IO_3^-	Iodate	$C_2O_4^{2-}$	Oxalate		
ClO_3^-	Chlorate	SCN^-	Thiocyanate	HPO_4^{2-}	Hydrogen phosphate		

Many transition metals need Roman numerals to communicate charge: Cu^{2+} = copper(II); Cu^+ = copper(I)

Silver = always +1; cadmium and zinc = always +2… Don't need a Roman numeral

Covalent compounds:

Prefixes are used to communicate number of each element in molecule

Mono-	1	Tri-	3	Penta-	5	Hepta-	7	Nona-	9
Di-	2	Tetra-	4	Hexa-	6	Octa-	8	Deca-	10

Special cases…

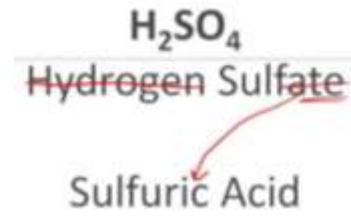
The "mono" prefix is never used on the first atom — FCl = fluorine monochloride

Double vowels are usually dropped (oo & ao)… — CO = carbon monoxide

but some are kept (io & ii) — SO_3 = sulfur trioxide

Acids:

Look at the non-hydrogen portion (anion)
The number of hydrogen (subscript) is based upon the charge of the anion



If anion ends in…

__________________-ide = hydro-__________________-ic acid

__________________-ate = __________________-ic acid

__________________-ite = __________________-ous acid

Naming & Formula Writing Practice:

1. Name the following compounds.
   - $CuSO_4$
   - S_2O_5
   - HClO
   - $HBrO_3$
   - $Na_2Cr_2O_7$
   - $Zn(NO_2)_2$
   - NO
   - $PbCrO_4$
   - CH_3COOH
2. Write the formulas for the following compounds.
   - Iron(III) oxide
   - Calcium hydroxide
   - Phosphoric acid
   - Hydrosulfuric acid
   - Lithium hydrogen sulfite
   - Tetraphosphorus decoxide
   - Carbon tetrachloride
   - Nickel(II) oxalate
   - Aluminum carbonate

Compound Stoichiometry

The concept of the counting the number of particles in a sample was first imagined by Avogadro when comparing the volumes of gas samples. Avogadro's number was specifically chosen to connect the number of particles in a sample to the measured average atomic mass. Molar volume is based upon the volume of that number of particles at standard temperature and pressure.

- A mole is a measure of gas volume at STP: **22.4 L/mol**
- A mole is a counting number: **6.02 x 10^{23} particles**
- A mole is related to mass: **molar mass is atomic mass from periodic table measured in g/mol**

Mole Sample Calculations:

1. How many moles contain 6.85 x 10^{24} atoms of Al?
2. How many atoms of each element are in 2.55 mol of P_4O_{10}?
3. How many liters would 10.00 grams of oxygen occupy at STP?

Analyzing Compounds

Percent Composition: The percent by mass of each element in a compound.

$$\text{Percent composition} = \frac{\text{Mass of element}}{\text{Mass of compound}} \times 100$$

Empirical formula: a formula with the lowest whole-number ratio of elements in a compound
Steps-

1. If given as percent composition, change the "%" to a "g" *(There are no steps, simply write the new unit)*
2. Find the number of moles of each element using the given mass and the molar mass.
3. Divide each number of moles by the smallest number in the ratio.
4. Round the number to the closest whole numbers to find the mole ratio (ratio of each type of atom in the formula) which gives the subscript of each element.
   *5. If the number in the mole ratio comes out to
      a) 0.5, multiply all numbers by 2 to find the actual ratio
      b) 0.33 or 0.67, multiply all numbers by 3…
      c) 0.25 or 0.75, multiply all numbers by 4…
      d) Other "even" decimal (1/5, 1/6, etc), multiply all numbers by appropriate number

Molecular formula: a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of a compound
Steps-

1. Find the empirical formula
2. Find the molar mass of the empirical formula. *(You will be given the molar mass of the molecular formula.)*
3. Divide the molecular mass by the empirical mass.
4. Multiply the subscripts in the empirical mass by the number just calculated to find the molecular formula.

Percent, Empirical, and Molecular Sample Calculations:

1. What percent of the mass of water comes from the oxygen?

2. If we react 6.03 grams of magnesium by heating it, the mass increases to 10.00 grams. Find the percent composition.

3. A compound is found to contain 40.9% carbon, 4.58% hydrogen, and 54.5% oxygen. Find the empirical formula.

4. Calculate the empirical formulas for the following three iron ores:
   a) Fe: 77.7%, O: 22.3%

   b) Fe: 72.4%, O: 27.6%

   c) Fe: 70.0%, O: 30.0 %

5. A sample of a compound with a molar mass of about 126 g/mol is found to consist of 28.57% carbon, 4.80% hydrogen, and the remainder nitrogen. Find the molecular formula.

Hydrate: an ionic compound with a specific number of water molecules bound to each formula unit

Naming Hydrates

- Name ionic portion using rules for naming ionic compounds
- Use same prefixes used when naming molecular compounds to indicate the # of water molecules

  $CaCl_2 \cdot 2H_2O$ = calcium chloride dehydrate

  $CuSO_4 \cdot 5H_2O$ = copper(II) sulfate pentahydrate

  $CoCl_2 \cdot H_2O$ = cobalt(II) chloride monohydrate

Analyzing Hydrates *(How much water?)*

To determine formulas

- Find mass of each and convert both masses to moles
- Make a ratio of moles water/moles solid

To determine percent composition

- Find mass of each
- Make a ratio of water to compound

Hydrate Sample Calculations:

1. What is the percent water in copper(II) sulfate pentahydrate?

2. If 11.75 grams of cobalt(II) chloride hydrate is heated, 9.25 g anhydrous $CoCl_2$ remain. What is formula and name for this hydrate?

3. How much does the mass of 12.65-g sample of copper(II) nitrate hexahydrate decrease when heated?

Combustion Analysis: process of determining the formula of an organic compound by looking at the products produced in a combustion reaction

- An unknown organic compound is reacted with excess oxygen to produce carbon dioxide and water.
- The empirical formula of the organic compound is determined by analyzing the mass of products.
  - Stoichiometric principle that allows analysis:
    - All of the carbon from the CO_2 came from the organic compound.
    - All of the hydrogen from the H_2O came from the organic compound.
    - Oxygen in products came from O_2 ***AND*** any oxygen in the organic compound.
      - Any mass in the organic compound that is not carbon or hydrogen must be oxygen unless otherwise specified (halogens and nitrogen can appear).

Combustion Analysis Sample Calculations:

1. An unknown hydrocarbon (C_xH_y) with a mass of 1.125 g is burned completely in excess oxygen. 3.447 g of carbon dioxide and 1.647 g of water vapor are collected. Determine the empirical formula of the hydrocarbon.

2. Methyl salicylate is made up of carbon, hydrogen, and oxygen atoms. When a sample of methyl salicylate weighing 5.287 g is burned in excess oxygen, 12.24 g of carbon dioxide and 2.522 g of water are formed. What is the empirical formula for methyl salicylate?

Reaction Stoichiometry

Stoichiometry is about the mathematical relationships between amounts of reactants and products. When we are given a certain amount of one substance, we can use conversion factors and the balanced equation to find how much of another substance is used/created. The relationship between different substances is expressed as a mole ratio based on the coefficient of the balanced equation.

All other stoichiometry problems use the mole ratio but some have additional steps before and/or after the mole ratio depending on the given quantity and the desired unit for the final answer. If the given quantity is not in moles, it must first be changed to moles.

amount given → moles given → moles unknown → amount unknown

Depending on the amount of each reactant present at the beginning of a reaction, one may run out before the others. The limiting reactant(s) are any reactant(s) that are completely used up in a chemical reaction. The limiting reactant determines the amount of product that can be formed in a reaction. The excess reactant(s) are the reactant(s) present in a quantity that is more than sufficient to react with a limiting reactant. Any reactant that remains after the limiting reactant is used up in a chemical reaction is an excess reactant.

Reactions do not always proceed perfectly when carried out in the lab. Often a different amount of product is collected than what should be collected. The ratio of the amount that is collected (actual yield) to the amount that should be collected (theoretical yield) is called the percent yield.

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \text{Percent yield}$$

Stoichiometry Sample Calculations:

1. a) For the following equation, calculate how many grams of each product would be produced by complete reaction of 10.0 g of the glucose.

   b) How many grams of O_2 were consumed in the reaction? *(What are the 2 ways we can find this?)*

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$$

2. For the following equation, calculate how many liters of nitrogen dioxide would be produced by complete reaction of 34.0 L of oxygen at STP.

$$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$$

3. For the following reaction: $2\ H_2(g) + CO(g) \rightarrow CH_3OH(l)$

   We start with 6.85×10^4 g of CO and 8.60×10^3 g of H_2. Calculate the following:

   a) The theoretical yield of methanol

   b) The amount of excess reactant remaining

   c) If 3.57×10^4 g of CH_3OH is actually produced, what is the percent yield of methanol?

BCA Tables

Multiple methods can be used to solve stoichiometry problems. As long as all work is correctly shown and clearly labeled, credit can be earned. This work may include conversions using molar mass, mole ratios, Avogadro's number, molar volume, ideal gas law, etc.

When looking at reactions in terms of limiting/excess, sometimes it is more efficient to look at the changes in amount of all substances at the same time. This can be achieved using BCA tables. BCA tables look at how the amount of each substance is changing relative to the other substances to determine what is limiting, how much is in excess, and the yield.

B = Before: the number of moles of each reactant before the reaction takes place

C = Change: the relative amount of moles used/created based upon the mole ratio

A = After: the number of moles of each reactant and product after the reaction takes place

How they work…

	$4\ NH_3$	+	$5\ O_2$	$\rightarrow$	$4\ NO$	+	$6\ H_2O$
B	#		#		0		0
C	–4x		–5x		+4x		+6x
A	# – 4x		# – 5x		0 + 4x		0 + 6x
	↑ One of these will end at 0 ↑						

BCA Stoichiometry Sample Calculations:

1. When 10.0 grams of ammonia reacts with 20.0 grams of oxygen according to the following reaction, calculate the following:
   a) Mass of water vapor created
   b) Mass of limiting reactant remaining
   c) Mass of nitrogen monoxide created

	$4\ NH_3$	+	$5\ O_2$	$\rightarrow$	$4\ NO$	+	$6\ H_2O$
B							
C							
A							