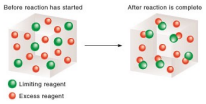
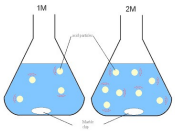
INTRO AND BACKGROUND:

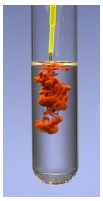
Reactions, Moles, Stoichiometry, and Solutions



$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

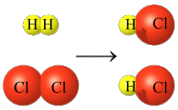
C=1		C=1
H=4	=	H=4
O=4		O=4





Chemical Reaction

Atoms are REARRANGED to form a different substance



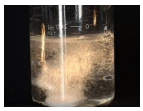
- Changes the way atoms are joined together
- Atoms CANNOT be created or destroyed!

Indicators of a Reaction

1) TEMPERATURE CHANGE

2) COLOR CHANGE

3) GAS FORMATION



4) PRECIPITATE

Symbols in Equations

+ separates the reactants or products

➔ separates reactants from products

$\rightleftharpoons$ indicates a reversible reaction

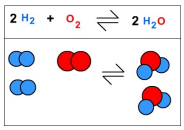
(s) solid

(g) gas

(l) liquid

(aq) aqueous or water solution

Δ indicates heat is supplied



CATALYST: speeds up reaction but is NOT a reactant or product

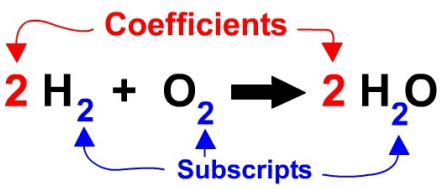
Balancing Chemical Equations

- Since we cannot break the Law of Conservation of Matter, equations MUST be balanced
- Balanced equations have the same number of each TYPE of atom on both sides of the equation
- COEFFICIENTS go in front of the formulas so the # of atoms of each element is the same on each side

WHAT GOES IN = WHAT COMES OUT!

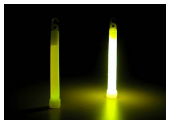
Balancing Chemical Equations

- Coefficients vs. Subscripts

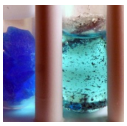


Types of Reactions

- Millions of reactions exist… but there are only several categories of reactions
- We will examine **FIVE** types:



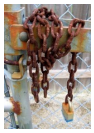
1) SYNTHESIS
2) DECOMPOSITION
3) SINGLE REPLACEMENT
4) DOUBLE REPLACEMENT
5) COMBUSTION



Synthesis Reaction

Two elements combine to make one compound

"Combination Reaction"



$$A + B \rightarrow AB$$

1) Two elements combine to form BINARY COMPOUNDS

EXAMPLE:

$$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$$

Synthesis Reaction

Two elements combine to make one compound

"Combination Reaction"

$$A + B \rightarrow AB$$

2) Metallic oxides and CO_2 combine to form CARBONATES

EXAMPLE:

$$CaO + CO_2 \rightarrow CaCO_3$$

Synthesis Reaction

Two elements combine to make one compound

"Combination Reaction"

$$A + B \rightarrow AB$$

3) Chloride salt and oxygen gas combine to form CHLORATES

EXAMPLE:

$$2\ KCl + 3\ O_2 \rightarrow 2\ KClO_3$$

Decomposition Reaction

One compound breaks down into two or more substances

$$AB \rightarrow A + B$$

1) BINARY COMPOUNDS will break down into their elements

EXAMPLE:

$$2\ NaCl \rightarrow 2\ Na + Cl_2$$

Decomposition Reaction

One compound breaks down into two or more substances

$$AB \rightarrow A + B$$

2) CARBONATES break down into the oxide and CO_2

EXAMPLE:

$$CaCO_3 \rightarrow CaO + CO_2$$

Decomposition Reaction

One compound breaks down into two or more substances

$$AB \rightarrow A + B$$

3) CHLORATES break down into the binary salt and oxygen gas

EXAMPLE:

$$2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$$

Single-Replacement Reaction

One element replaces another in a compound (element and compound become different element and compound)

$$A + BC \rightarrow B + AC$$

- **EXAMPLE:**

Predict the products for the reaction and balance.

$$Zn + 2\ HCl \rightarrow H_2 + ZnCl_2$$

USE ACTIVITY SERIES TO SEE IF THEY HAPPEN!

Activity Series

Metals from Li to Na will REPLACE H from acids and water… from Mg to Pb will REPLACE H from acids only!

Activity Series of Metals

	Name	Symbol
Decreasing reactivity ↓	Lithium	Li
	Potassium	K
	Calcium	Ca
	Sodium	Na
	Magnesium	Mg
	Aluminum	Al
	Zinc	Zn
	Iron	Fe
	Lead	Pb
	(Hydrogen)	(H)*
	Copper	Cu
	Mercury	Hg
	Silver	Ag

Higher metal (more active) can REPLACE any metal lower than it, otherwise NO REACTION WILL OCCUR!!

Double-Replacement Reaction

Two elements replace one another

$$AB + CD \rightarrow AD + CB$$

- **EXAMPLE:**

Predict the products for the reaction and balance.

$$3\ NaOH + FeCl_3 \rightarrow 3\ NaCl + Fe(OH)_3$$

USE SOLUBILITY RULES TO SEE IF THEY HAPPEN!

Combustion Reaction

Compound with only C and H (sometimes O as well) is reacted with O_2

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$

- **EXAMPLE:**

Predict the products for the reaction and balance.

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

IF INCOMPLETE THEN CO AND H_2O FORM!

Conversion

- Mole conversions are useful but not practical in a lab…

1 mole element = atomic mass (grams)

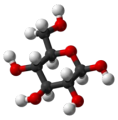
Get it right from the Periodic Table!!

- For example, 1 mole of arsenic has 74.9 g

Molar Mass

Mass of 1 mole of an element, molecule, or compound

- **How to Determine Molar Mass:**
  1) Determine the # of moles of the individual elements that make up the compound (just look at the formula)
  2) Look up the mass of each element
  3) Multiply the mass of each by the # of moles of each
  4) Add up the masses



Glucose
$C_6H_{12}O_6$

Molar Mass

Mass of 1 mole of an element, molecule, or compound

- **EXAMPLE:**

Find the molar mass of glucose ($C_6H_{12}O_6$).

6	C	x	12.01 g	=	72.06 g
12	H	x	1.008 g	=	12.096 g
6	O	x	16.00 g	=	+ 96.00 g

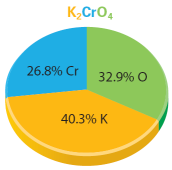
180.156 g/mol

Percent Composition

Amount or percent of an element in a compound by mass

- Determine the mass of each element and divide each by the total mass of the compound

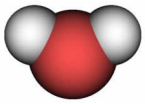
$$\frac{\text{Mass of element}}{\text{Mass of compound}} \times 100$$



Potassium chromate, K_2CrO_4

Mole

Amount of something



- When measuring molecules and atoms, we use moles
- Used to count very small items
- Helps convert from the sub-microscopic to the macroscopic

BUT, WHAT AMOUNT?

Mole

Amount of something

6.02×10^{23}

"Avogadro's Number"

THAT AMOUNT, BUT OF WHAT?

- REPRESENTATIVE PARTICLES: smallest piece of a chemical substance

  Ex: element = atom, ionic compound = formula unit, molecular compound = molecule

Conversions

1 mole = 6.02×10^{23} atoms

1 mole = 6.02×10^{23} molecules

1 mole = 6.02×10^{23} formula units

These can be used in DIMENSIONAL ANALYSIS problems!!

Atoms to Moles

- **EXAMPLE:**

  A sample of Mg has 1.25×10^{23} atoms of Mg. How many moles of Mg are contained in the sample?

$$1.25 \times 10^{23} \text{ atoms Mg} \times \frac{1 \text{ mole of Mg}}{6.02 \times 10^{23} \text{ atoms Mg}} =$$

0.208 mol Mg

Practice

- **EXAMPLES:**

  How many moles is 5.69 g NaOH?

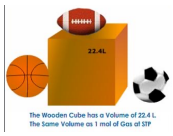
  How many grams are there in 0.107 moles of CO_2?

  How many atoms are there in 12.21 g of C?

Mole-Volume Relationship

- Many chemicals exist as gases but difficult to mass
- Moles of a gas can be related to volume (Liters), but temperature and pressure also play a role
- Standard Temp. and Pressure (STP): 0°C and 1 atm
- At STP:

1 mole gas = 22.4 liters



Empirical Formula

Lowest whole number ratio of elements in a compound (can NOT be reduced)

- **How to Determine:**
  1) Change the % to grams (if necessary)
  2) Convert grams to moles for each element
  3) Divide ALL of the mole answers by the SMALLEST # (mole ratio)
  4) If all whole #, then move on… if not then multiply to get whole # ***SEE NEXT SLIDE!**
  5) Use the whole # to represent the number of each element… write the formula

C_2H_2 Ethyne

C_8H_8 Styrene

Empirical Formula

Lowest whole number ratio of elements in a compound (can NOT be reduced)

- After dividing ALL by the lowest #… If any are not whole numbers you must multiply ALL of the numbers by a number to make them all whole:

IF YOU GET	MULTIPLY ALL BY
.5	2
.3 or .6	3
.25	4
.20	5

Appeared on an old AP Exam!

Empirical Formula

Lowest whole number ratio of elements in a compound (can NOT be reduced)

- **EXAMPLE:**

  Determine the empirical formula of the following compound: 12.12% C, 16.16% O, and 71.72% Cl.

$$12.12 \text{ g C} \times \frac{1 \text{ mole C}}{12 \text{ g C}} = \frac{1.01}{1.01} = 1$$

$$16.16 \text{ g O} \times \frac{1 \text{ mole O}}{16 \text{ g O}} = \frac{1.01}{1.01} = 1$$

$$71.72 \text{ g Cl} \times \frac{1 \text{ mole Cl}}{35 \text{ g Cl}} = \frac{2.05}{1.01} = 2$$

$COCl_2$

Phosgene gas (used in WWI)

Molecular Formula

Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

- **How to Determine:**
  1) Calculate the empirical formula (if needed)
  2) Calculate the molar mass of the empirical formula
  3) Divide the given **molecular molar mass** by the **empirical molar mass**
  4) Multiply subscripts of empirical formula by this #
  5) Write the molecular formula

Molecular Formula

Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

- **EXAMPLE:**

Determine the molecular formula of a compound composed of 85.7% C and 14.3% H with a molar mass of 70 g/mol.

$$85.7\text{ g C} \times \frac{1\text{ mole C}}{12\text{ g C}} = \frac{7.14}{7.14} = 1$$

$$14.3\text{ g H} \times \frac{1\text{ mole H}}{1\text{ g H}} = \frac{14.3}{7.14} = 2$$

CH_2

Molecular Formula

Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

- **EXAMPLE:**

Determine the molecular formula of a compound composed of 85.7% C and 14.3% H with a molar mass of 70 g/mol.

1	C	x	12.01 g	=	12.01 g
2	H	x	1.008 g	= +	2.016 g
					14.026 g

Molecular Formula

Formula indicating the exact numbers of atoms of each element in the formula (can be reduced)

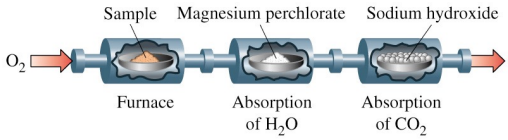
- **EXAMPLE:**

Determine the molecular formula of a compound composed of 85.7% C and 14.3% H with a molar mass of 70 g/mol.

$$\frac{\text{Molecular Mass}}{\text{Empirical Mass}} = \frac{70\text{ g}}{14.026\text{ g}} = 5$$

$CH_2 \Rightarrow C_5H_{10}$

Combustion Problems



Determine the amount of carbon, hydrogen, etc. present in the products formed which will tell you how much is present of each in the original sample (LAW OF CONSERVATION OF MASS) and then do the problems normally!

Practice

- **COMBUSTION EXAMPLE:**

Combustion of 10.68 g of Vitamin C (containing only C, H, and O) yields 16.01 g of CO_2 and 4.37 g of H_2O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of this compound?

$$16.01\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 4.37\text{ g C}$$

$$4.37\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.016\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.489\text{ g H}$$

4.37 g C + 0.489 g H = 4.859 g

10.68 g - 4.859 g = 5.821 g O

Practice

- **COMBUSTION EXAMPLE:**

Combustion of 10.68 g of Vitamin C (containing only C, H, and O) yields 16.01 g of CO_2 and 4.37 g of H_2O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of this compound?

Now you have the grams for each of the three elements in the compound... find the empirical and molecular as before!

$C_3H_4O_3$ = Empirical

$C_6H_8O_6$ = Molecular

Practice

- **COMBUSTION EXAMPLE:**

Phthalic anhydride contains only carbon, hydrogen, and oxygen and is used as a chemical intermediate in the production of plastics from vinyl chloride. When a 20.0 g sample of the compound was combusted in oxygen, 47.57 g of CO_2 and 4.86 g of H_2O were formed. Determine the empirical formula of the compound.

Balanced Equations

 # of 

 # of 

- Coefficients in a balanced chemical equation can represent a ratio of moles, molecules, liters (gases), or atoms... NOT GRAMS!
- Convert from an amount of one ingredient to another or to amounts of products
- Use MOLAR RATIO

Equations must be CORRECTLY WRITTEN and BALANCED in order to do these problems!!

Stoichiometry Problems

- **Always follow this same basic format...**

Grams (Reactant) ➡ Moles (Reactant) ➡ Moles (Product) ➡ Grams (Product)

*Liters can be SUBSTITUTED for Grams!!!

Use the MOLAR RATIO (from the Equation)

*Problems can start or end with Moles and SKIP Grams altogether!!

$$\text{Grams R} \times \frac{1 \text{ Mole R}}{\text{Grams R}} \times \frac{\text{Moles P}}{\text{Moles R}} \times \frac{\text{Grams P}}{1 \text{ Mole P}} =$$

Mass to Mass Conversions

- **EXAMPLE:**

If 10.1 g of Fe (3+) are added to a solution of copper (II) sulfate, how much solid copper would form?

$$2\ Fe + 3\ Cu(SO_4) \rightarrow 3\ Cu + Fe_2(SO_4)_3$$

$$10.1 \text{ g Fe} \times \frac{1 \text{ mole Fe}}{56 \text{ g Fe}} \times \frac{3 \text{ mole Cu}}{2 \text{ mole Fe}} \times \frac{64 \text{ g Cu}}{1 \text{ mole Cu}} =$$

17.3 g Cu

Mass to Volume Conversions

- **EXAMPLE:**

Potassium metal reacts with water to produce potassium hydroxide and hydrogen gas. If 12.1 g K is reacted completely, how many liters of H_2 gas can be produced at STP?

$$2\ K + 2\ H_2O \rightarrow 2\ KOH + H_2$$

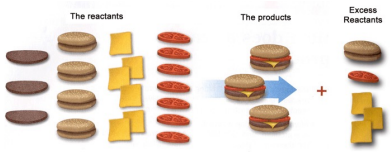
$$12.1 \text{ g K} \times \frac{1 \text{ mole K}}{39 \text{ g K}} \times \frac{1 \text{ mole } H_2}{2 \text{ mole K}} \times \frac{22.4 \text{ L } H_2}{1 \text{ mole } H_2} =$$

3.47 L H_2

Limiting Reagent

Reactant that runs out first in a chemical reaction

- Amount of reactants available for a reaction **LIMITS** the amount of product that can be made
- **EXCESS REAGENT**: reactant that is not used up completely



Limiting Reagent Problems

- **EXAMPLE:**

Copper reacts with sulfur to form copper (I) sulfide. If 10.6 g of Cu reacts with 3.83 g S, how much product will be formed?

$$2\ Cu + S \rightarrow Cu_2S$$

$$10.6\ g\ Cu \times \frac{1\ mole\ Cu}{64\ g\ Cu} \times \frac{1\ mole\ Cu_2S}{2\ mole\ Cu} \times \frac{160\ g\ Cu_2S}{1\ mole\ Cu_2S} = \boxed{13.2\ g}$$

Limiting Reagent

$$3.83\ g\ S \times \frac{1\ mole\ S}{32\ g\ S} \times \frac{1\ mole\ Cu_2S}{1\ mole\ S} \times \frac{160\ g\ Cu_2S}{1\ mole\ Cu_2S} = 19.2\ g$$

Limiting Reagent Problems

- **EXAMPLE:**

How much of the **excess reagent** will be left over from the previous problem?

$$2\ Cu + S \rightarrow Cu_2S$$

Use the Limiting Reagent!!
Sulfur is in EXCESS!

$$10.6\ g\ Cu \times \frac{1\ mole\ Cu}{64\ g\ Cu} \times \frac{1\ mole\ S}{2\ mole\ Cu} \times \frac{32\ g\ S}{1\ mole\ S} =$$

2.65 g NEEDED

3.83 g – 2.65 g = **1.18 g** EXCESS

Percent Yield

Amount of product formed compared to the amount that should have formed

$$\frac{\textbf{ACTUAL}}{\textbf{THEORETICAL}} \times 100$$

Should NEVER be greater than 100%... WHY?

- **How to Determine:**
  1) Actual is given or found in lab
  2) Calculate theoretical by dimensional analysis (may need limiting reagent)
  3) Use the equation

Multiple Choice Practice

- **EXAMPLE:**

When 8.0 g of N_2H_4 (32 g mol^{-1}) and 92 g of N_2O_4 (92 g mol^{-1}) are mixed together and react according to the equation below, what is the **maximum mass** of H_2O that can be produced?

$$2\ N_2H_4(g) + N_2O_4(g) \rightarrow 3\ N_2(g) + 4\ H_2O(g)$$

(A) 9.0 g
(B) 18 g
(C) 36 g
(D) 72 g
(E) 144 g

Multiple Choice Practice

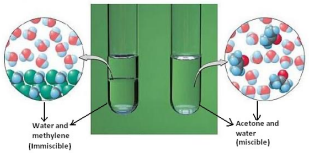
- **EXAMPLE:**

When hafnium metal is heated in an atmosphere of chlorine gas, the product of the reaction is found to contain 62.2% Hf by mass and 37.4% Cl by mass. What is the **empirical formula** for this compound?

(A) HfCl
(B) $HfCl_2$
(C) $HfCl_3$
(D) $HfCl_4$
(E) Hf_2Cl_3

Aqueous Solution

Solution in which H_2O is the solvent

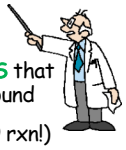


- Ionic and polar molecules dissolved best… **WHY?**
- **MISCIBLE**: liquids completely mix (alcohol and water)
- **IMMISCIBLE**: don't mix (oil and water)

Net Ionic Equations

Shows only the ions directly involved in a chemical reaction

- **SPECTATOR IONS**: appear in identical forms on both sides of the equation but are not directly involved in the reaction (cancel out)
- **Steps for Writing Net Ionic Equations:**



1) Write the balanced equation
2) Rewrite the equation showing the **IONS** that form in solution for each **SOLUBLE** compound
3) Cancel spectator ions (ALL cancel = NO rxn!)
4) Rewrite the final equation

Net Ionic Equations

Shows only the ions directly involved in a chemical reaction

- **EXAMPLE:**

$Pb(NO_3)_2$ (aq) + 2 KI (aq) ➔ PbI_2 (s) + 2 KNO_3 (aq)

Net Ionic Equations

Shows only the ions directly involved in a chemical reaction

- **PRACTICE:**

Predict the products (with solubility indicated), balance, and write the net ionic equation for each.

K_2CO_3 (aq) + $CaCl_2$ (aq) ➔

$Ba(NO_3)_2$ (aq) + NaCl (aq) ➔

Na_2S (aq) + $CuBr_2$ (aq) ➔

Molarity (M)

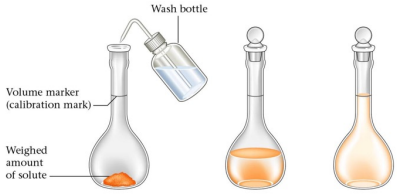
Moles of solute per liter of solution

$$M = \frac{\text{Moles of solute}}{\text{Liters of solution}}$$

- **EXAMPLE:**

Calculate the molarity of a solution when 11.5 g of NaOH is dissolved in enough water to make 1.5 liters of solution.

Preparing Solutions



- To make a certain concentration of solution, the **SOLID** should be weighed out first and then placed in a volumetric flask
- Dissolve the solid in **SOME** of the solvent then add the remaining solvent

Dilutions

$M_1V_1 = M_2V_2$

- **EXAMPLE:**

What volume of 16 M H_2SO_4 must be used to prepare 1.5 L of 0.10 M solution?

(16 M) (V_1) = (0.10 M) (1.5 L)

0.0094 L or 9.4 mL

Add 1490.6 mL of water to make 0.10M!

Dilutions

$M_1V_1 = M_2V_2$

- **PRACTICE:**

How many mL of water should be used to prepare 300.0 mL of 0.750 M NaBr solution using 2.00 M stock solution?