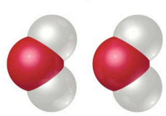
Chapter 3: Stoichiometry

Goal is to understand and become **proficient** at working with:

1. Chemical equations (Balancing REVIEW)
2. Some simple patterns of reactivity
3. Formula weights (REVIEW)
4. Avogadro's Number, molar mass and converting between mass and moles (REVIEW)
5. Empirical formulas from analysis
(Combustion analysis is challenging for some. You must work to understand and practice until you are proficient!)
6. Quantitative information from balanced equations
7. Limiting reactants and percent yield. You must work to understand and practice until you are proficient!

Subscripts and Coefficients Give Different Information

Chemical symbol	Meaning		Composition
H_2O	One molecule of water:		Two H atoms and one O atom
$2\ H_2O$	Two molecules of water:		Four H atoms and two O atoms

Subscripts tell the number of atoms of each element in a molecule or formula unit. (We can use the term “formula unit” for either molecular or ionic substances. However, when referring to ionic compounds, the term “formula unit” is correct, not “molecule”. Why is this?)

Coefficients tell the number of molecules or formula units (or the number of moles of molecules or formula units).

Balancing Chemical Equations - Steps

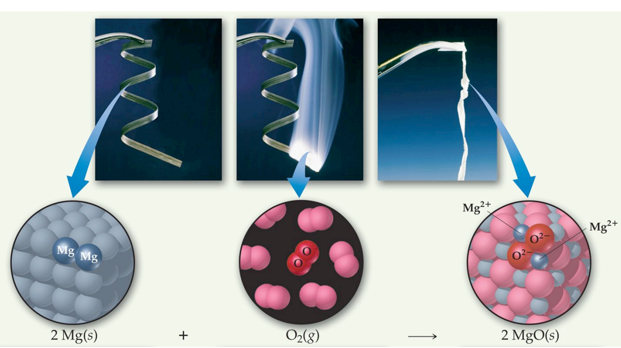
1. **Write correct formulas for reactants and products.**
2. Balance groups of atoms that are common to both reactants and products first, e.g. NO_3^-, SO_4^{2-}, etc.
3. Balance metals next.
4. Balance nonmetals last.
5. Check your answer.

Write the chemical formulas for reactants and products and then balance the following reactions, including phase symbols:

Hydrogen Gas Reacts with oxygen gas to produce water.

Aqueous barium chloride reacts with aqueous sodium phosphate to produce solid barium phosphate and a solution containing sodium chloride.

Reaction Types
1. Combination Reactions



Two or more substances react to form one product.

Examples:

$2\ Na(s) + Cl_2(g) \longrightarrow 2\ NaCl(s)$

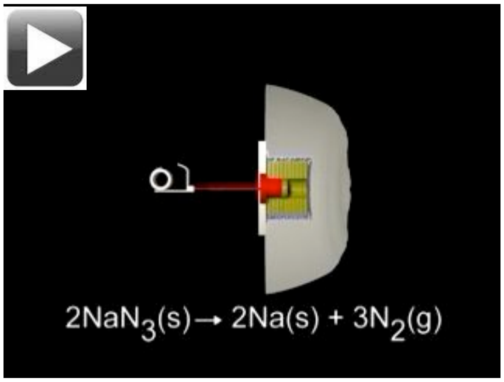
$N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$

$C_3H_6(g) + Br_2(l) \longrightarrow C_3H_6Br_2(l)$

$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g)$

Write the balanced chemical equation for the combination reaction between aluminum and fluorine gas. Include phase labels.

Reaction Types
2. Decomposition Reactions



One substance breaks down into two or more simpler substances.

Examples:

$CaCO_3(s)$ —> $CaO(s) + CO_2(g)$

2 $KClO_3(s)$ —> 2 $KCl(s)$ + 3 $O_2(g)$

2 $NaN_3(s)$ —> 2 $Na(s)$ + 3 $N_2(g)$

Chemical reaction to inflate an airbag with N_2 gas.

Write the balanced chemical equation for the following decomposition reaction:

When heated, solid sodium bicarbonate decomposes into sodium carbonate, water vapor and gaseous carbon dioxide. (Include phase labels.)

Reaction Types
3. Simple Organic Combustion Reactions

Organic combustion reactions involve an organic compound and oxygen gas as reactants. Simple organic compounds contain only C, H and O atoms.

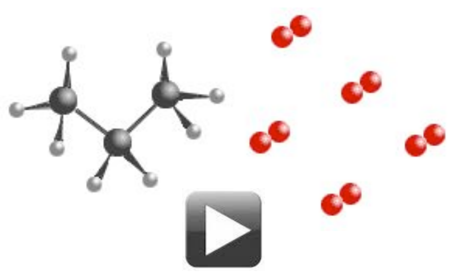
The products of a simple organic combustion are only CO_2 and H_2O. These reactions are always exothermic. You may have learned the term "exothermic" in a previous introductory course. If so, do you remember what it means?

To balance these reactions:

1. Write correct formulas for reactants and products.
2. Balance the C atoms with CO_2. (All the C is turned into CO_2)
3. Balance the H atoms with H_2O. (All the H is turned into H_2O)
4. Balance the O atoms with O_2.
5. Check your answer.

Balance the following combustion reaction:

$C_3H_8 + O_2$ —> ?



Reaction Types Review

Combination:

- **Easy to predict** product formulas and balance when a **metal and a nonmetal combine** to form a **binary ionic compound. You must be able to do this for exams!**
- **Can be hard to predict** products for reactions yielding **molecular compounds;** more familiarity with chemistry is needed; **you will not be ask to predict these on exams.**

Decomposition:

- Easy to predict product formulas and balance when a **binary ionic compound decomposes** into its elements. **You must be able to do this for for exams!**
- **Harder** to predict products for the **decomposition of ionic compounds containing polyatomic ions or the decomposition of molecular compounds**; more familiarity with chemistry is still needed; **you will not be ask to predict these on exams.**

Simple Organic Combustion:

- Easy to predict products and balance. **You must be able to do this for for exams!**

Formula Weight, Molecular Weight, and Molar Mass

Formula weight (FW):

Sum of atomic weights in **amu** for ionic compounds and elements.

This is the mass of a single formula unit.

- Example: determine the formula weight of aluminum oxide.

Molecular weight (MW):

Sum of atomic weights in **amu** for molecular compounds.

This is the mass of a single molecule.

- Example: determine the molecular weight of acetic acid.

Molar mass ($\mathcal{M}$):

Sum of atomic weights in **g** units for anything!

- This is the mass of one mole of a any compound/element.

Aluminum oxide molar mass:

Acetic acid molar mass:

Note: **Formula *mass*** or **molecular *mass*** are often used in place of weight. In fact, mass is “more correct”!

The Mole

The mole is the chemists counting unit. There are **6.022×10^{23}** items in a mole. We count in moles because atoms, molecules and ions are so incredibly small that even a µg sample of a substance contains a unimaginable number of atoms, molecules or ions!
There is a “National Mole Day” that starts at 6:02 AM on October 23. Really!

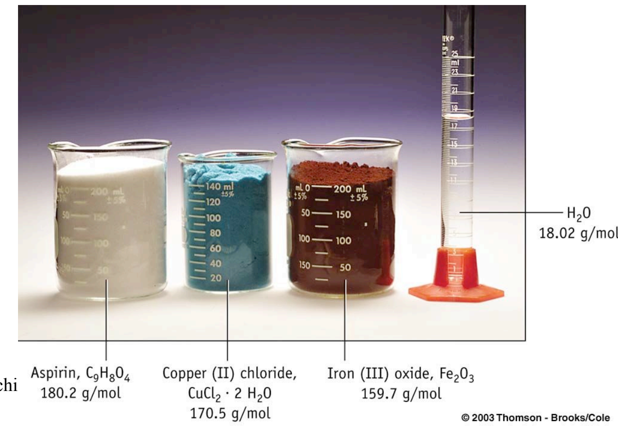
Examples of a mole:

1 mole of TNT, $CH_3C_6H_2(NO_2)_3$, molecules, contains:

- 7 moles of carbon
- 5 moles of hydrogen
- 3 moles of nitrogen
- 6 moles of oxygen

6.022×10^{23} molecules of TNT

How many oxygen atoms?

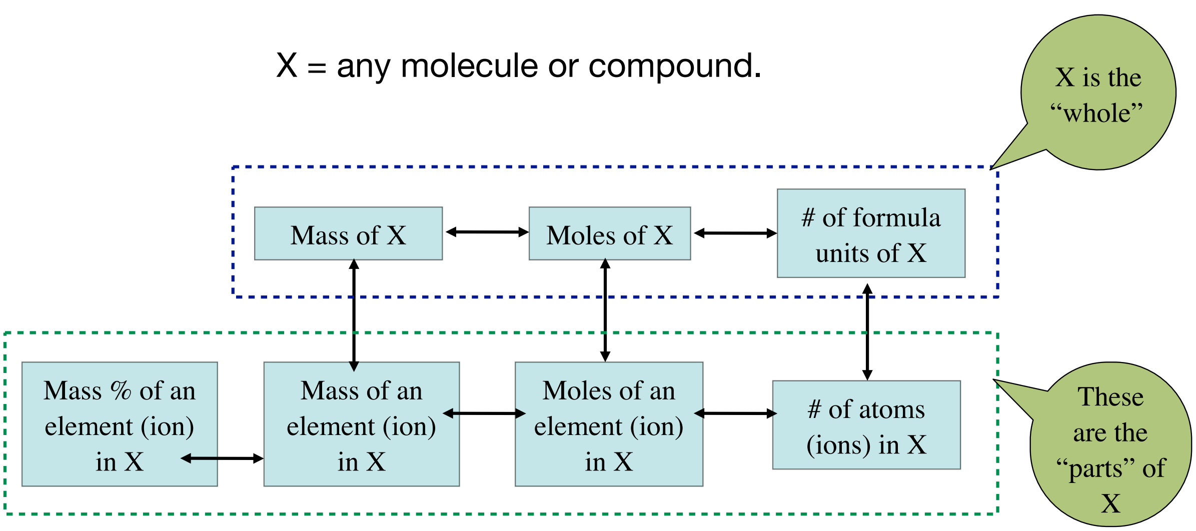


For TNT, $CH_3C_6H_2(NO_2)_3$, the molar mass is 227.14 g/mole

Conversions Between Mass, Moles, Formula Units and Atoms (Ions)

The molar mass is the bridge between mass (laboratory measurement) and amount in moles. We have no instruments in the lab that “count” moles.

X = any molecule or compound.



Molar Mass, formula units and atoms

Find the requested information.

1. The molar mass of zinc nitrate.
2. The number of grams in 2.75 mole of zinc nitrate.
3. The number of formula units in 72 mg of zinc nitrate.
4. A sample of zinc nitrate contains 3.6×10^{20} oxygen atoms. What mass of zinc nitrate is present in mg?

Percent Composition by Mass

1. The percent composition by mass of a compound is constant.
2. The percent composition by mass can be calculated for each element in a compound using the compound formula and the atomic weights of each element.
3. The sum of the percent composition of all elements in a compound must equal what?

Determine the percent composition of each element in zinc nitrate.

Empirical Formulas from Mass Data

If mass % or mass data is known - the empirical formula can be found for a compound.

1. Convert % composition to mass in g	2. Convert mass to moles	3. Find mole ratio(s)	4. Ratio(s) give empirical formula A_xB_y
% A —> g A —> *x* mol A % B —> g B —> *y* mol B		$\frac{x \text{ mol A}}{y \text{ mol B}}$	

In the reaction $Cr + O_2 \longrightarrow Cr_xO_y$, 0.345 g of Cr produces 0.504 g of Cr_xO_y. What is the empirical formula of Cr_xO_y?

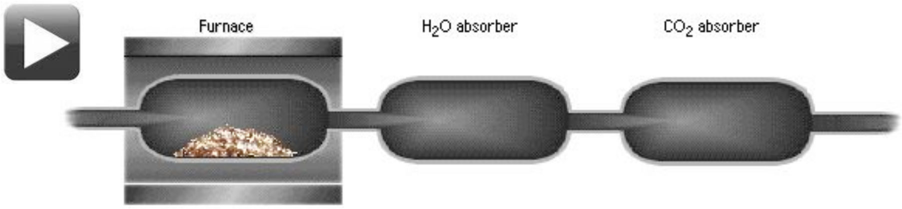
Molecular Formula From Empirical Formula

1. For molecular substances, the empirical formula may not be the molecular formula. However, they are related:
   (Molecular formula) = n(Empirical formula). **Where n is a whole number**.
2. To determine the molecular formula from the empirical formula, we need the molecular weight (or molar mass).

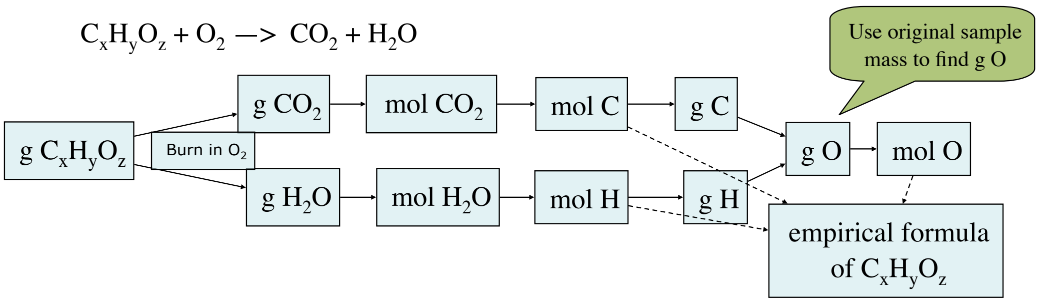
Adipic acid is used to make nylon. It contains 49.31% carbon, 6.90% hydrogen and the remainder oxygen by mass. The molar mass is 146.1g. What is the molecular formula of adipic acid?

Combustion Analysis to find Empirical Formula

The empirical formula of simple organic compounds can be determined by complete combustion and mass analysis of the products.



$C_xH_yO_z + O_2 \longrightarrow CO_2 + H_2O$



Practice Combustion Analysis Problem

Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005-g sample of menthol is combusted, producing 0.2829 g of CO_2 and 0.1159 g of H_2O. What is the empirical formula for menthol? If the compound has a molar mass of 156 g/mol, what is its molecular formula? Write the balanced combustion reaction.

Mass Relationships in Stoichiometry

The coefficients in a chemical equation can be interpreted in two ways:

1. As individual formula units
2. As **moles** of formula units

We will usually use the later - moles!

The coefficients allow us to convert from moles of one substance into moles of another using mole ratios (stoichiometric factors).

For the reaction 3A —> 2B the following mole ratios are indicated:

$$\frac{\text{3 moles reacted A}}{\text{2 moles B produced}} \quad \text{or} \quad \frac{\text{2 moles B produced}}{\text{3 moles A reacted}}$$

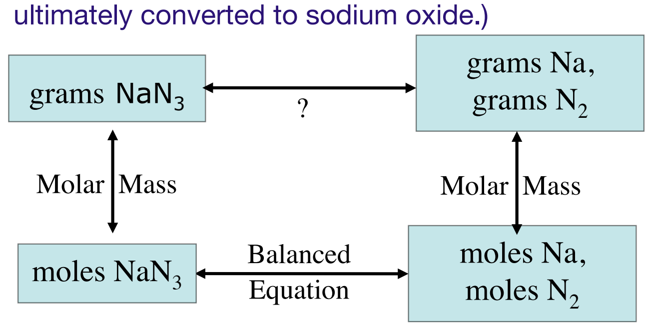
Examples of Stoichiometry

1. Automotive air bags inflate when sodium azide, NaN_3, rapidly decomposes to its component elements:

$2\ NaN_3(s) \longrightarrow 2\ Na(s) + 3\ N_2(g)$ (Remember from the air bag video that the Na is ultimately converted to sodium oxide.)

Molar Masses:

N_2	28.02 g/mol
NaN_3	65.02 g/mol
Na	22.99 g/mol



a) If 10.0 g of NaN_3 reacts, how many moles of Na form? What mass of sodium oxide can be obtained from this amount of Na?

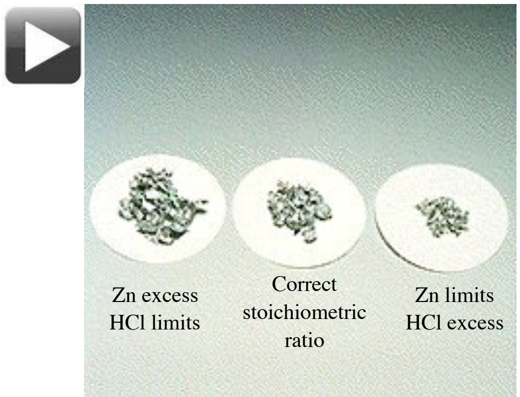
b) How many grams of NaN_3 are required to produce 10.0 ft^3 of nitrogen gas if the N_2 gas has a density of 1.25 g/L?

Limiting Reactant Calculations

In this type of calculation two (or more) reactant amounts are given. One of the reactants is completely used up, leaving an excess of the other reactant(s). We call the used up reactant the limiting reactant.

The limiting reactant determines the maximum amount of product(s) that can be formed.

The trick is to determine which reactant is limiting! For this we use stoichiometry.



$Zn(s) + 2\ HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

Practice Limiting Reactant Problem

Disulfur dichloride is used to vulcanize rubber. It can be made by treating molten sulfur with gaseous chlorine:

$S_8(l) + 4\ Cl_2(g) \longrightarrow 4\ S_2Cl_2(l)$

Starting with 63.0 g of sulfur and 63.0 g of chlorine, how many grams of disulfur dichloride can be produced? How many grams of each reactant remain?

Yields in Chemistry

The theoretical yield is the maximum amount of product that can be produced. This is **determined by a limiting reactant calculation.**

The actual yield is the amount of product actually **produced in an experiment**. This is **determined experimentally** and should be less then the theoretical yield. It should never be more than the theoretical yield!

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

Think about it.

- Why are percent yields often less that 100%?

- What would you think if you obtained a percent yield greater than 100%?

Yields in Chemistry

From our practice problem of limiting reactants, if only 99.6 g of S_2Cl_2 is actually produced in an experiment, what is the percent yield?

If this yield is typical, what mass of S_8 is needed to produce 50.0 g of S_2Cl_2?

Mixture Analysis

The composition of a mixture can be determined by chemical reaction and mass (gravimetric) analysis.

A 2.634-g mixture containing some $CuCl_2 \bullet (H_2O)_2$ is heated to eliminate the waters of hydration. After heating the mass remaining was 2.145 g. Assuming that the mass lost from the sample was only due to water loss from the $CuCl_2 \bullet (H_2O)_2$ in the sample, what is the mass percent of $CuCl_2 \bullet (H_2O)_2$ in the original mixture?

Problems from Text

3.90

a) One molecule of the antibiotic known as penicillin G has a mass of 5.342×10^{-21} g. What is the molar mass of penicillin G?

b) Hemoglobin, the oxygen-carrying protein in red blood cells, has four iron atoms per molecule and contains 0.340% iron by mass. Calculate the molar mass of hemoglobin.

Problems from Text

3.15

a) When the metallic element sodium combines with the nonmetallic element bromine, $Br_2(l)$, how can you determine the chemical formula of the product?

How do you know whether the product is a solid, liquid, or gas at room temperature?

Write the balanced chemical equation for the reaction.

b) When a hydrocarbon burns in air, what reactant besides the hydrocarbon is involved in the reaction?

What products are formed?

Write a balanced chemical equation for the combustion of benzene, $C_6H_6(l)$, in air.

Problems from Text

3.29

Without doing any detailed calculations (but using a periodic table to give atomic weights), rank the following samples in order of increasing number of atoms:

0.50 mol H_2O

23 g Na

6.0×10^{23} N_2 molecules.

Problems from Text

3.94

An organic compound was found to contain only C, H, and Cl. When a 1.50-g sample of the compound was completely combusted in air, 3.52 g of CO_2 was formed. In a separate experiment the chlorine in a 1.00-g sample of the compound was converted to 1.27 g of AgCl. Determine the empirical formula of the compound.

Problems from Text

3.96

An element X forms an iodide (XI_3) and a chloride (XCl_3). The iodide is quantitatively converted to the chloride when it is heated in a stream of chlorine:

$$2\,XI_3 + 3\,Cl_2 \longrightarrow 2\,XCl_3 + 3\,I_2$$

If 0.5000 g of XI_3 is treated, 0.2360 g of XCl_3 is obtained.

a) Calculate the molar mass of the element X.

b) Identify the element X.

How to Set-up a Limiting Reactant Problem Using a Reaction Table

Given 15.2 g of H_2 and 63.0 g of N_2, how many grams of NH_3 can be produced in the reaction: $N_2 + 3H_2 \longrightarrow 2NH_3$?

Let's see a problem done using a reaction table. Learning this method can prove to be very useful! (Equation must first be balanced!)

Reaction	3 H_2	+ N_2	—>	2 NH_3
Initial (g)	15.2	63.0	given amounts	0
Initial (mol)	$7.5\underline{4}0$	$2.2\underline{4}8$	use molar mass to convert given amounts	0
Moles of Reaction (moles rxn)	$2.5\underline{1}3$	**$2.2\underline{4}8$** (limiting)	= mole/coefficient	0
Change in Mole	$-6.7\underline{4}4$	$-2.2\underline{4}8$	= (limiting moles rxn) *coefficient	$+4.4\underline{9}6$
Final (mol)	$0.7\underline{9}6$	**0**	= initial + change	$4.4\underline{9}6$
Final (g)	1.6	**0**	use molar mass	76.6