Chapter 3

Stoichiometry: Ratios of Combination

3.1 Molecular and Formula Masses

3.2 Percent Composition of Compounds

3.3 Chemical Equations

3.4 The Mole and Molar Mass

3.5 Combustion Analysis

3.6 Calculations with Balanced Chemical Equations

3.7 Limiting Reactants

1

3.1 Molecular and Formula Masses

- Molecular mass - (molecular weight)
  - The mass in amu's of the individual molecule
  - Multiply the atomic mass for each element in a molecule by the number of atoms of that element and then total the masses
- Formula mass (formula weight)-
  - The mass in amu's of an ionic compound

Calculating Molecular Mass

- Calculate the molecular mass (Mwt) for carbon dioxide, **CO_2**
- Write down each element; multiply by atomic mass
  - C = 1 x 12.01 = 12.01 amu
  - O = 2 x 16.00 = 32.00 amu

  - **Total**: 12.01 + 32.00 = 44.01 amu

Your Turn!

- Calculate the molecular mass or formula mass for each of the following:
  – Sulfur trioxide
  – Barium phosphate
  – Sodium nitrite
  – Acetic acid

3.2 Percent Composition of Compounds

- Calculated by dividing the total mass of each element in a compound by the molecular mass of the compound and multiplying by 100

$$\text{percent by mass of an element} = \frac{n \times \text{atomic mass of element}}{\text{molecular or formula mass of compound}} \times 100\%$$

% Composition

- Calculate the percent composition of iron in a sample of iron (III) oxide
- Formula: Fe_2O_3
- Calculate formula mass
  - Fe = 2 x 55.85 = **111.70 amu**
  - O = 3 x 16.00 = **48.00 amu**
  - **Total mass**: 111.70 + 48.00 = **159.70 amu**

% Composition

$$\%\text{ by mass} = \frac{111.70}{159.70} \times 100 = 69.9\%\text{ Fe}$$

What is the % oxygen in this sample? (hint :100%)

Example

Glucose ($C_6H_{12}O_6$)

What is the mass percent of each element in glucose?

- For each mole of glucose, we have 6 moles of C, 12 moles H, and 6 moles O

6 C	**6 x 12.01 = 72.06 amu**	
12 H	**12 x 1.008 = 12.096 amu**	**+**
6 O	**6 x 16.00 = 96.00 amu**	**+**
	180.16 amu	

$$\text{mass percent of C} = \frac{72.06\ \text{C}}{\mathbf{180.16}\ \text{glucose}} = 0.3999 \times 100 = 39.99 \text{ mass \%C}$$

$$\text{mass percent of H} = \frac{12.096\ \text{H}}{\mathbf{180.16}\ \text{glucose}} = 0.06714 \times 100 = 6.714 \text{ mass \%H}$$

$$\text{mass percent of O} = \frac{96.00\ \text{O}}{\mathbf{180.16}\ \text{glucose}} = 0.5329 \times 100 = 53.29 \text{ mass \%O}$$

Your Turn!

- Calculate the percent oxygen in a sample of potassium chlorate

- $KClO_3$

3.3 Chemical Equations

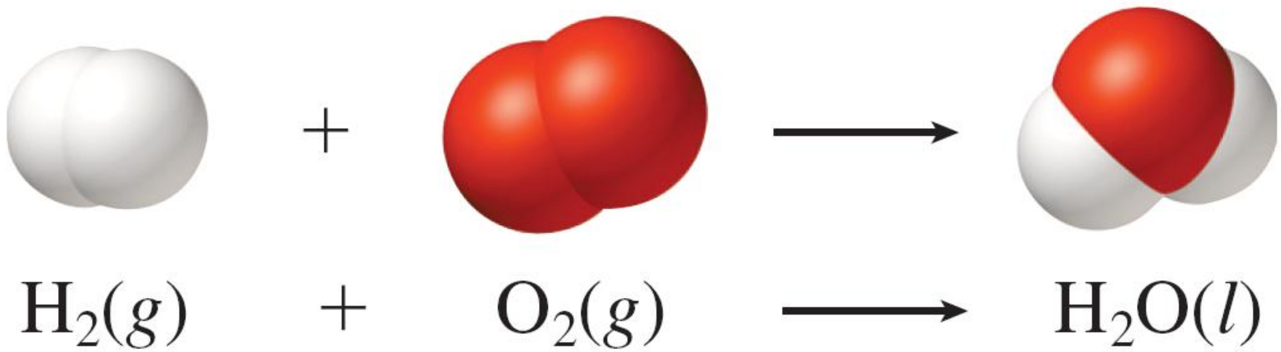
- Chemical equations represent chemical “sentences”
- Read the following equation as a sentence
  - $NH_3 + HCl \rightarrow NH_4Cl$
  - “ammonia reacts with hydrochloric acid to produce ammonium chloride”

Chemical Equations

- **Reactant:** any species to the left of the arrow (consumed)
- **Product:** any species to the right of the arrow (formed)
- **State symbols:**
  - (s) solid (l) liquid (g) gas
  - (aq) water solution

Balancing Equations

- **Balanced:** same number and kind of atoms on each side of the equation



$$2\,H_{2\,(g)} + O_{2\,(g)} \longrightarrow 2\,H_2O_{(l)}$$

Balancing Equations

- **Steps for successful balancing**

1. Change coefficients for compounds before changing coefficients for elements.(never change subscripts!)
2. Treat polyatomic ions as units rather than individual elements.
3. Count carefully, being sure to recount after each coefficient change.

- In chemical reaction, ***all atoms*** present in reactants must be accounted for among the products. this is the called ***balanced chemical equation***

$2\ H_2 + O_2 \longrightarrow 2\ H_2O$		**Balanced equation**
$2\ Mg + O_2 \longrightarrow 2\ MgO$		**Balanced equation**
$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$		**Unbalanced equation**

Balancing Chemical Equations

1. Write the **correct** formula(s) for the reactants on the left side and the **correct** formula(s) for the product(s) on the right side of the equation.

Ethane reacts with oxygen to form carbon dioxide and water

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

2. Change the numbers in front of the formulas (***coefficients***) to make the number of atoms of each element the same on both sides of the equation. Do not change the subscripts.

$2C_2H_6$ **NOT** C_4H_{12}

$2H_2O$ **NOT** H_4O_2

Balancing Chemical Equations

$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$

start with C or H but not O

2 carbon on left

1 carbon on right

multiply CO_2 by 2

$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

multiply H_2O by 3

6 hydrogen on left

2 hydrogen on right

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

Balancing Chemical Equations

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

multiply O_2 by $\frac{7}{2}$

2 oxygen on left

4 oxygen (2x2) + 3 oxygen (3x1) = **7** oxygen on right

$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

remove fraction

$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

multiply both sides by 2

Double check!

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

4 C (2 x 2)	4 C
12 H (2 x 6)	12 H (6 x 2)
14 O (7 x 2)	14 O (4 x 2 + 6)

Reactants	Products
4 C	4 C
12 H	12 H
14 O	14 O

Balancing Equations

- Balance the equation representing the combustion of hexane

$$\_\_C_6H_{14(l)} + \_\_O_{2(g)} \rightarrow \_\_CO_{2(g)} + \_\_H_2O_{(l)}$$

(Hint: Make a list of all elements and count to keep track)

Balancing Equations

- Balance the equation representing the combustion of hexane

$$C_6H_{14(l)} + 7/2O_{2(g)} \rightarrow 6CO_{2(g)} + 7H_2O_{(l)}$$

Or…multiply through the entire equation to eliminate fractions

$$2C_6H_{14(l)} + 7O_{2(g)} \rightarrow 12CO_{2(g)} + 14H_2O_{(l)}$$

Chemical Equations

- Equations can represent physical changes

$KClO_{3(s)} \rightarrow KClO_{3(l)}$

Or chemical changes

- Note the symbol for heat above the arrow
- $KClO_{3(s)} \xrightarrow{\Delta} KClO_{3(l)}$

3.4 The Mole and Molar Masses

- Balanced equations tell us what is reacting and in what relative proportions on the molecular level.

The Mole

- The unit of measurement used by chemists in the laboratory
- **The mole** is "the number equal to number of carbon atoms in exactly 12 grams of pure ^{12}C"
- **Avogadro** has determined this number to be

  $$N_A = 6.02214 \times 10^{23} \quad \text{Particle/mole}$$

- 1 mole ^{12}C atoms = 6.022×10^{23} atoms = 12.00g

The Mole

$$2\ H_{2(g)}\ +\ O_{2(g)}\ \rightarrow 2\ H_2O_{(l)}$$

2 molecules $H_{2(g)}$ + 1 molecule $O_{2(g)} \rightarrow$ 2 molecules $H_2O_{(l)}$

2 moles $H_{2(g)}$ + 1 mole $O_{2(g)}$ $\rightarrow$ 2 moles $H_2O_{(l)}$

Molar Mass

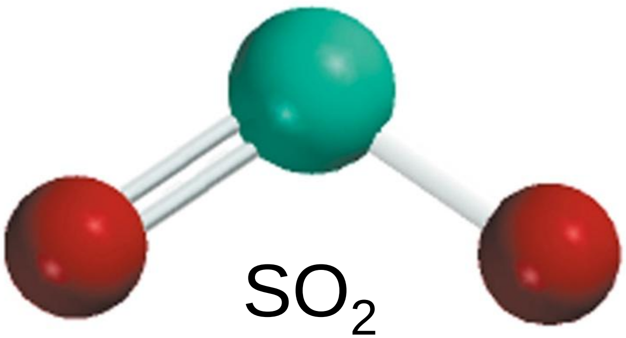
Molar mass is the mass of 1 mole of anything in grams

- Carbon = 12.0 grams/mole
- Sodium = 22.9 grams/mole
- 1 mole lithium atoms = 6.941 g of Li

For any element

atomic mass (amu) = molar mass (grams)

Molecular mass (or molecular weight) is the sum of the atomic masses (in amu) in a molecule.



SO_2

1S	32.07 amu
2O	+ 2 x 16.00 amu
SO_2	64.07 amu

> For any molecule
>
> molecular mass (amu) = molar mass (grams)

1 molecule SO_2 = 64.07 amu

1 mole SO_2 = 64.07 g SO_2

Molar Mass for Compounds

Calculate the molar mass for each of the following:

H_2O

H 2 x 1.01 g/mol = 2.02

O 1 x 16.00 g/mol = $\underline{16.00}$

Molar mass = 18.02 g/mol

Example

How many atoms are in 0.551 g of potassium (K) ?

$$1 \text{ mol K} = 39.10 \text{ g K}$$

$$1 \text{ mol K} = 6.022 \times 10^{23} \text{ atoms K}$$

$$0.551 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} \times \frac{6.022 \times 10^{23} \text{ atoms K}}{1 \text{ mol K}} =$$

$$8.49 \times 10^{21} \text{ atoms K}$$

Example

How many H atoms are in 72.5 g of C_3H_8O ?

$$1 \text{ mol } C_3H_8O = (3 \times 12) + (8 \times 1) + 16 = 60 \text{ g } C_3H_8O$$

$$1 \text{ mol } C_3H_8O \text{ molecules} = 8 \text{ mol H atoms}$$

$$1 \text{ mol H} = 6.022 \times 10^{23} \text{ atoms H}$$

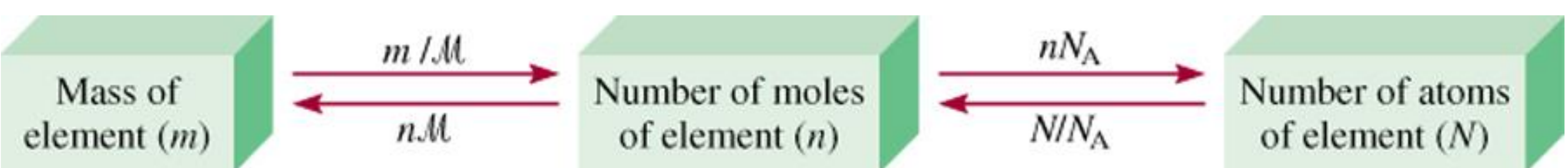
$$72.5 \text{ g } C_3H_8O \times \frac{1 \text{ mol } C_3H_8O}{60 \text{ g } C_3H_8O} \times \frac{8 \text{ mol H atoms}}{1 \text{ mol } C_3H_8O} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} =$$

$$5.82 \times 10^{24} \text{ atoms H}$$

Conversions between grams, moles and atoms

$$\textbf{No of Moles} = \frac{\textbf{Mass (g)}}{\textbf{Molar Mass (g/mole)}}$$

$$\textbf{No of Atoms} = \textbf{No of Moles} \times \frac{\mathbf{6.022 \times 10^{23}}\ \textbf{atom}}{\textbf{1 mole}}$$



Interconverting mass, moles and number of particles

Determine the number of moles in 85.00 grams of sodium chlorate, $NaClO_3$

$$85.00\ \mathbf{g\ NaClO_3} \times \frac{1\ \text{mole}\ NaClO_3}{106.44\ \text{g}\ NaClO_3} = 0.7986\ \text{mol}\ NaClO_3$$

Another

Determine the number of molecules in 4.6 moles of ethanol, C_2H_5OH.

(1 mole = 6.022×10^{23})

$$4.6 \text{ mol } C_2H_5OH \times \frac{6.02 \times 10^{23} \text{ molecules } C_2H_5OH}{1 \text{mol } C_2H_5OH} = 2.8 \times 10^{24} \text{ molecules}$$

Another

- Determine how many H atoms are in 4.6 moles of ethanol.

$$2.8 \times 10^{24} \text{ molecules } C_2H_5OH \times \frac{6 \text{ H atoms}}{1 \text{ molecule } C_2H_5OH} = 1.7 \times 10^{25} \text{ atoms H}$$

Another

- How many grams of oxygen are present in 5.75 moles of aluminum oxide, Al_2O_3?

Strategy:

Example

Determine the number of fluorine atoms in 24.24 grams of sulfur hexafluoride.

Empirical and Molecular Formulas

- **Empirical-** simplest whole-number ratio of atoms in a formula
- **Molecular** - the "true" ratio of atoms in a formula; often a whole-number multiple of the empirical formula
- We can determine **empirical formulas from % composition data.**

Empirical and Molecular Formulas

molecular formula = (empirical formula)$_n$ [n = integer]

molecular formula = C_6H_6 = $(CH)_6$

empirical formula = CH

If the Molar mass for any compound is known one can get the chemical or molecular formula by the following:

1- calculate empirical mass = $\sum$ atomic mass

1- find the multiplication factor = MM/EM

1- molecular formula = multiplication factor x EF

Note: if MM = EM

molecular formula = empirical formula

Example

Caffeine contain **49.48 % C, 5.159% H, 28.87 % N, and 16.49 % O** by mass has a molar mass **194.2 g/mol**. Determine the molecular formula of caffeine.

Solution

Assume you have 100g of the caffeine, Then you have

49.48 g C 5.159 g H 28.87 g N 16.49 g O

- Divide by the molar mass of each element

$$\frac{49.48 \text{ g C}}{12.01 \text{ g/mole C}} = \mathbf{4.12} \text{ mol C}$$

5.12 mol H **2.06** mol N **1.03** mol O

Find the ratio of the moles (divide by the smallest number)

$$\frac{\mathbf{4.12}}{\mathbf{1.03}} \qquad \frac{\mathbf{5.12}}{\mathbf{1.03}} \qquad \frac{\mathbf{2.06}}{\mathbf{1.03}} \qquad \frac{\mathbf{1.03}}{\mathbf{1.03}}$$

4 C **5 H** **2 N** **1 O**

Continue

4 C **5 H** **2 N** **1 O**

-The empirical formula is $C_4H_5N_2O$

- The empirical Atomic mass = 12.01 x 4 + 1.008 x 5 + 14.01 x 2 + 16 x1 = 97.08 g/mol

- The empirical Atomic mass = 97.08 g/mol

- The multiplication factor = $\dfrac{\mathbf{194.2\ g/mol}}{97.08\ \text{g/mol}} = 2$

The molecular formula = $(C_4H_5N_2O)_2$ = $\mathbf{C_8H_{10}N_4O_2}$

Example

- lactic acid (***M***=90.08 g/mol) contains 40.0 mass% C, 6.71 mass% H, and 53.3 mass% O.
- Determine the molecular formula

Solution

Assuming there are 100. g of lactic acid

$$\frac{40 \text{ g C}}{12.01\text{g C}} \qquad \frac{6.71 \text{ g H}}{1.008 \text{ g H}} \qquad \frac{53.3 \text{ g O}}{16.00 \text{ g O}}$$

3.33 mol C   6.66 mol H   3.33 mol O

$$\frac{\mathbf{3.33}}{3.33} \quad \frac{\mathbf{6.66}}{3.33} \quad \frac{\mathbf{3.33}}{3.33} \longrightarrow$$

CH_2O empirical formula EM = 30.03 g/mol

$$\frac{90.08 \text{ g/mol}}{30.03 \text{ g/mol}} \longrightarrow 3$$

$C_3H_6O_3$ is the molecular formula

Empirical Formulas

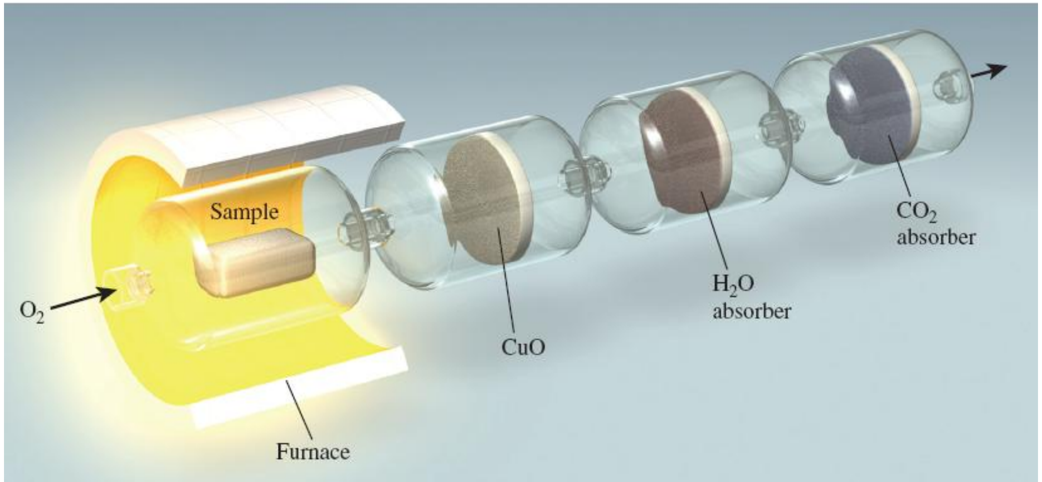
- Steps for success
  - Convert given amounts to moles
  - Mole ratio (divide all moles by the smallest number of moles)
  - The numbers represent subscripts.
    - If the numbers are not whole numbers, multiply by some factor to make them whole.

Empirical Formula

- Determine the empirical formula for a substance that is determined to be 85.63% carbon and 14.37% hydrogen by mass.

3.5 Combustion Analysis

- Analysis of organic compounds (C,H and sometimes O) are carried using an apparatus like the one below



$$\text{Organic Compound} \xrightarrow{O_2} CO_2,\ H_2O$$

Combustion Analysis

- The data given allows an empirical formula determination with just a few more steps.
- The mass of products (carbon dioxide and water) will be known so we work our way back.

Example

- Vitamin C (**M=176.12g/mol**) is a compound of C,H, and O. When a **1.000-g** sample of vitamin C is placed in a combustion chamber and burned, the following data are obtained:

Mass of CO_2 = 1.50 g

Mass of H_2O = 0.41 g

What is the molecular formula of vitamin C?

Solution

There are 1 mole of C per 1 mole of CO_2

$$\frac{1.50 \text{ g } CO_2}{44.01 \text{ g/mol } CO_2} = 0.034 \text{ mole of } CO_2$$

Moles of C = moles of CO_2 = 0.034 mole

Mass of C = Moles of C x Molar mass of C

0.034 mole of C x 12.01 g/mol = **0.409 g C**

Continue

There are 2 mole of H per 1 mole of H_2O

$$\text{Moles of } H_2O = \frac{0.41\text{ g}}{18.02\text{ g } H_2O} = 0.023\text{ mole}$$

Moles of H = 2 x Moles of H_2O
= 2 x 0.023
= 0.046 mole H

Mass of H = Mole of H x Molar mass of H
0.046 mole H x 1.008 g/mole = **0.046 g of H**

Mass of O = 1 – (0.409 g C + 0.046 g of H) = **0.545 g of O**

$$\frac{0.409\text{ g C}}{12.01\text{ g C}} = 0.0341\text{ mol C} \qquad \frac{0.046\text{ g H}}{1.008\text{ g H}} = 0.0456\text{ mol H} \qquad \frac{0.545\text{ g O}}{16.00\text{ g O}} = 0.0341\text{ mol O}$$

$$C_1H_{1.3}O_1 \longrightarrow C_3H_4O_3 \longrightarrow \frac{176.12\text{ g/mol}}{88.06\text{ g}} = 2.000$$

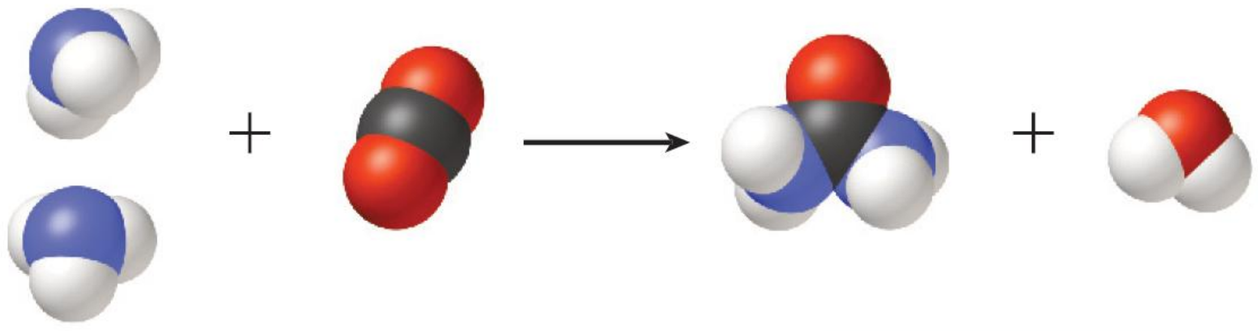
$C_6H_8O_6$

3.6 Calculations with Balanced Chemical Equations

- Balanced equations allow chemists and chemistry students to calculate various amounts of reactants and products.
- The coefficients in the equation are used as mole ratios.

Stoichiometry

- Stoichiometry- using balanced equations to find amounts
- How do the amounts compare in the reaction below?

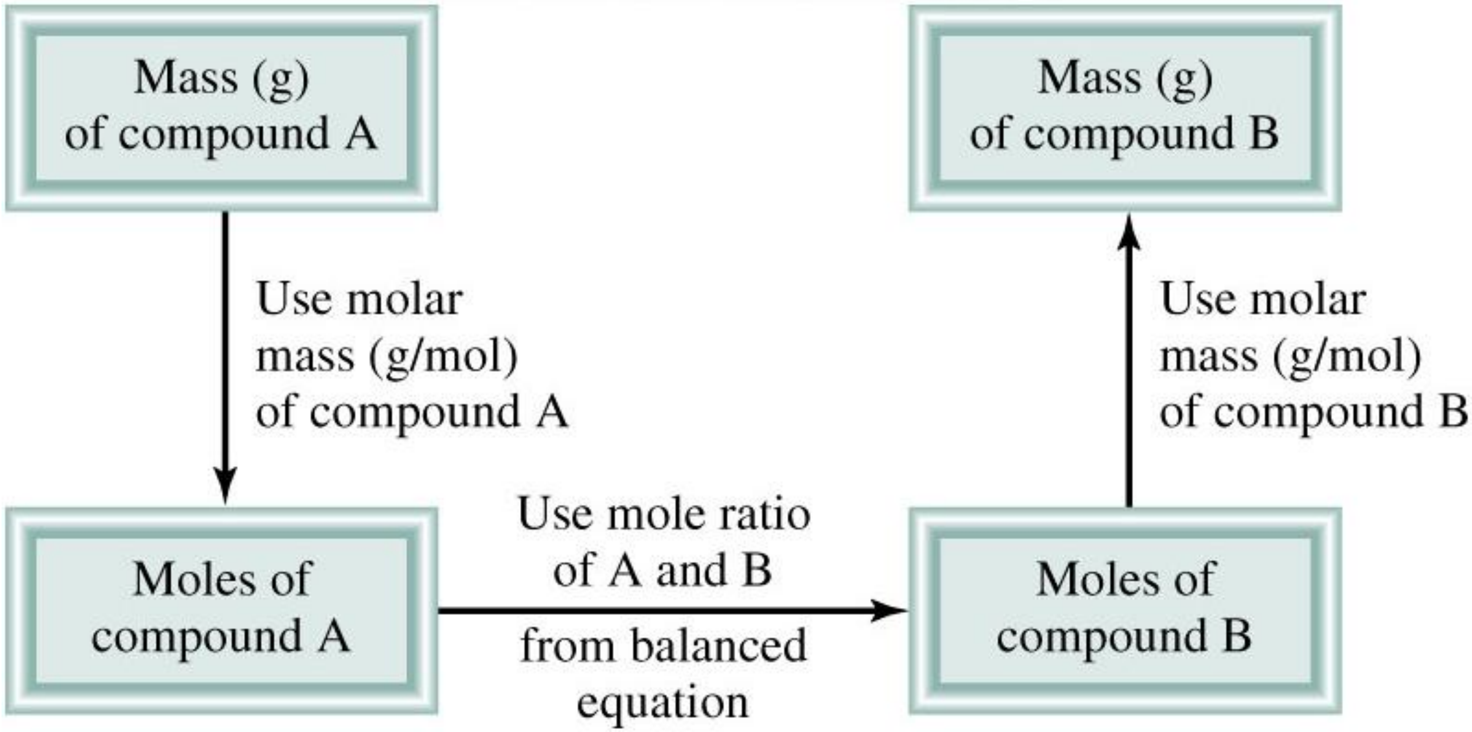


$$2NH_3(g) + CO_2(g) \longrightarrow (NH_2)_2CO(aq) + H_2O(l)$$

Mole Ratios

- Many mole ratios can be written from the equation for the synthesis of urea
- Mole ratios are used as conversion factors

$$\frac{2\ \text{mol}\ NH_3}{1\ \text{mol}\ CO_2} \quad \text{or} \quad \frac{1\ \text{mol}\ CO_2}{2\ \text{mol}\ NH_3}$$



1. Write balanced chemical equation
2. Convert quantities of known substances into moles
3. Use coefficients in balanced equation to calculate the number of moles of the required quantity
4. Convert moles of required quantity into desired units

Calculations with Balanced Equations

- How many moles of urea could be formed from 3.5 moles of ammonia?

$$2NH_{3(g)} + CO_{2(g)} \rightarrow (NH_2)_2CO_{(aq)} + H_2O_{(l)}$$

$$3.5 \text{ mol } NH_3 \times \frac{1 \text{ mol } (NH_2)_2CO}{2 \text{ mol } NH_3} = 1.8 \text{ mol } (NH_2)_2CO$$

Mass to Mass

A chemist needs 58.75 grams of urea, how many grams of ammonia are needed to produce this amount?

Strategy:

Grams → moles → mole ratio → grams

$$2NH_{3(g)} + CO_{2(g)} \rightarrow (NH_2)_2CO_{(aq)} + H_2O_{(l)}$$

$$58.75 \text{ g } (NH_2)_2CO \times \frac{1 \text{ mol} (NH_2)_2CO}{58.06 \text{ g} (NH_2)_2CO} \times \frac{2 \text{ mol } NH_3}{1 \text{ mol } (NH_2)_2CO} \times \frac{17.04 \text{ g } NH_3}{1 \text{ mol } NH_3} = 34.49 \text{ g } NH_3$$

You Try!

How many grams of carbon monoxide are needed to produce 125 grams of urea?

Example

Methanol burns in air according to the equation

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

If 209 g of methanol are used up in the combustion, what mass of water is produced?

$$\text{grams } CH_3OH \xrightarrow[\text{molar mass } CH_3OH]{} \text{moles } CH_3OH \xrightarrow[\text{coefficients chemical equation}]{} \text{moles } H_2O \xrightarrow[\text{molar mass } H_2O]{} \text{grams } H_2O$$

molar mass CH_3OH

coefficients chemical equation

molar mass H_2O

$$209\ \cancel{\text{g } CH_3OH} \times \frac{1\ \cancel{\text{mol } CH_3OH}}{32.0\ \cancel{\text{g } CH_3OH}} \times \frac{4\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol } CH_3OH}} \times \frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}} =$$

$\boxed{\textbf{235 g } H_2O}$

3.7 Limiting Reactants

- **Limiting reactant -** the reactant that is used up first in a reaction (limits the amount of product produced)

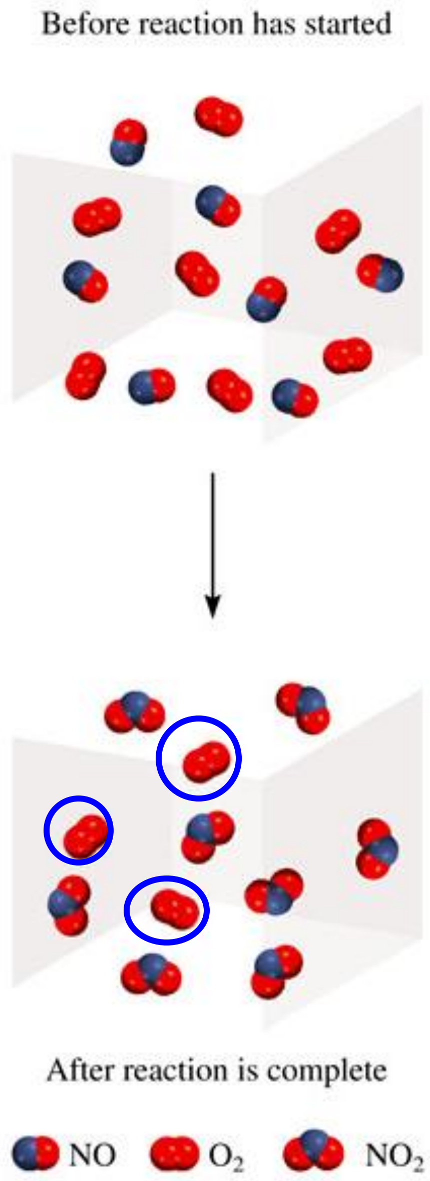
- **Excess reactant -** the one that is left over

Limiting Reagents

$2NO + 2O_2 \longrightarrow 2NO_2$

NO is the limiting reagent

O_2 is the excess reagent



Calculation involve limiting reactant

1. Balance the equation.
2. Convert masses to moles.
3. Determine which reactant is limiting.
4. Use moles of limiting reactant and mole ratios to find moles of desired product.
5. Convert from moles to grams.

How to calculate Limiting Reactant

In one process, **124 g** of Al are reacted with **601 g** of Fe_2O_3

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

Calculate the mass of Al_2O_3 formed?

g Al ⟶ mol Al ⟶ mol Fe_2O_3 needed ⟶ g Fe_2O_3 needed

OR

g Fe_2O_3 ⟶ mol Fe_2O_3 ⟶ mol Al needed ⟶ g Al needed

$$\cancel{124\ \text{g Al}} \times \frac{1\ \cancel{\text{mol Al}}}{\cancel{27.0\ \text{g Al}}} \times \frac{1\ \cancel{\text{mol } Fe_2O_3}}{2\ \cancel{\text{mol Al}}} \times \frac{160.\ \text{g } Fe_2O_3}{1\ \cancel{\text{mol } Fe_2O_3}} = \mathbf{367}\ \text{g } Fe_2O_3$$

Start with 124 g Al ⟶ need 367 g Fe_2O_3

Have more Fe_2O_3 (601 g) so Al is limiting reagent

Continue

Use limiting reagent (Al) to calculate amount of product that can be formed.

$$\text{g Al} \longrightarrow \text{mol Al} \longrightarrow \text{mol } Al_2O_3 \longrightarrow \text{g } Al_2O_3$$

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

$$124 \cancel{\text{ g Al}} \times \frac{1 \cancel{\text{ mol Al}}}{27.0 \cancel{\text{ g Al}}} \times \frac{1 \cancel{\text{ mol } Al_2O_3}}{2 \cancel{\text{ mol Al}}} \times \frac{102.\text{ g } Al_2O_3}{1 \cancel{\text{ mol } Al_2O_3}} = 234 \text{ g } Al_2O_3$$

Example

- When 35.50 grams of nitrogen react with 25.75 grams of hydrogen, how many grams of ammonia are produced?
- How many grams of excess reagent remain in the reaction vessel?

$$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$$

$$\text{Moles of } N_2 = \frac{35.5\text{ g}}{28.01\text{g/mol}} = 1.27 \text{ mole of } N_2$$

$$\text{Moles of } H_2 = \frac{25.75\text{ g}}{2.02\text{g/mol}} = 12.75 \text{ mole of } H_2$$

$$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$$

Continue

1.27 mole of N_2 12.75 mole of H_2

1- Assume the N_2 is the limiting reactant

2- Calculate how many moles of H_2 needed

$$\text{Moles of } H_2 \text{ needed} = \text{moles of } N_2 \times \frac{3 \text{ Moles of } H_2}{1 \text{ Moles of } N_2}$$

$$1.27 \text{ mole } N_2 \times \frac{3 \text{ Moles of } H_2}{1 \text{ Moles of } N_2} = \mathbf{3.81 \text{ mole } H_2}$$

N_2 is the limiting reactant

Continue

$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$

$g\ N_2 \rightarrow moles\ N_2 \rightarrow moles\ NH_3 \rightarrow g\ NH_3$

$$\text{Moles of } NH_3 = 1.27 \text{ mole } N_2 \times \frac{2 \text{ Moles of } NH_3}{1 \text{ Moles of } N_2} = 2.54 \text{ mole } NH_3$$

Mass of NH_3 = 2.54 mole × 17.03 g/mole = 43.25 g

Moles of H_2 unreacted = 12.75 – 3.81 = 8.94 moles

Mass of H_2 remains = 8.94 mole × 2.02 g/mole = **18.06 g H_2**

Reaction Yield

- **Theoretical yield**: the maximum amount of product predicted by stoichiometry
- **Actual yield:** the amount produced in a laboratory setting
- **Percent yield:** a ratio of actual to theoretical (tells efficiency of reaction)

$$\textbf{\% Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Example

3.75 g of zinc (Zn) reacted with **excess** hydrochloric acid (HCl), **5.58** g of zinc chloride ($ZnCl_2$) were collected. What is the percent yield for this reaction?

$$Zn_{(s)} + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

g of Zn $\rightarrow$ moles of Zn $\rightarrow$ moles of $ZnCl_2$ $\rightarrow$ g of $ZnCl_2$

$$3.75 \text{ g of Zn} \times \frac{\text{mole}}{65.38 \text{ g}} \times \frac{1 \text{ mole of } ZnCl_2}{1 \text{ mole of Zn}} \times \frac{136.3 \text{ g } ZnCl_2}{\text{mole}}$$

= 7.82 g of $ZnCl_2$ (Theoretical Yield)

Continue

Theoretical yield = 7.82 g $ZnCl_2$

Actual yield = 5.58 g $ZnCl_2$

$$\% \text{ yield} = \frac{5.58 \text{ g of } ZnCl_2}{7.82 \text{ g of } Zncl_2} \text{ X } 100 = 71.4\ \%$$

A Few Reaction Types

- **Combination:** one product is formed
- **Decomposition:** one reactant produces more than one product
- **Combustion:** a hydrocarbon reacts with oxygen to produce carbon dioxide and water

Combination Reaction

General formula: $A + B \rightarrow AB$

Sodium + chlorine $\rightarrow$ sodium chloride

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

Sulfur dioxide + water $\rightarrow$ sulfurous acid

$$SO_2 + H_2O \rightarrow H_2SO_3$$

Decomposition Reaction

General formula: $AB \rightarrow A + B$

Copper (II) carbonate decomposes with heat into copper (II) oxide and carbon dioxide

$CuCO_3 \rightarrow CuO + CO_2$

Potassium bromide decomposes into its elements

$2KBr \rightarrow 2K + Br_2$

Combustion (hydrocarbons)

General formula: $C_xH_y + O_2 \rightarrow CO_2 + H_2O$

Methane gas burns completely

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

Butane liquid in a lighter ignites

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

Review

- Molecular mass
- Percent composition
- Chemical equations
  - Reactants
  - Products
  - State symbols
  - Balancing

Review continued

- Mole concept and conversions
- Empirical and molecular formulas
  - Combustion analysis
- Stoichiometry
- Limiting reactant
- % yield
- Types of reactions