Chapter 3
Stoichiometry
Ratios of combination

Topics

- Molecular and formula masses
- Percent composition of compounds
- Chemical equations
- Mole and molar mass
- Combustion analysis (Determining the formula of a compound)
- Calculations with balanced chemical equations
- Calculations involving limiting reactant

Chemical Stoichiometry

- **Stoichimetry** from Greek "**measuring elements**".

  That is "Calculation of quantities in chemical reactions"
- **Stoichiometry** - The study of **quantities** of materials **consumed** and **produced** in chemical reactions.

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 one formula unit of an ionic compound

Calculating the molecular mass

- Calculate the molecular mass 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

Calculating the formula mass

- Calculate the formula mass for sodium nitrate, $NaNO_3$
- Write down each element; multiply by atomic mass
  - Na = 1 x 23.0 = 23.0 amu
  - N = 1 x 14.01 = 14.01 amu
  - O= 3 x 16.0 = 48.0 amu

  - Total: 23.0 + 14.01 + 48.0 = 85.01 amu

3.2 Percent Composition of Compounds

- % composition is calculated by dividing the total mass of each element in a compound by the molecular mass of the compound and multiplying by 100
- % composition allows verification of purity of a sample

Percent Composition of Compounds

$$\frac{n \text{ x atomic mass of element}}{\text{Molecular or formula mass of compound}} \text{ x } 100\%$$

% 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

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

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

% Composition

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

$$\text{What is the \% oxygen in this sample? (hint: 100\%)}$$

Practice

Find the percent composition of each element in C_2H_6O

$$\%C = \frac{2 \times (12.01\ g)}{46.07\ g} \times 100\% = 52.14\%$$

$$\%H = \frac{6 \times (1.008\ g)}{46.07\ g} \times 100\% = 13.13\%$$

$$\%O = \frac{1 \times (16.00\ g)}{46.07\ g} \times 100\% = 34.73\%$$

52.14% + 13.13% + 34.73% = 100.0%

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

Reading Chemical Equations

- The plus sign (+) means "react" and the arrow points towards the substance produced in the reaction.
- The chemical formulas on the left side of the equation are called reactants (consumed) and after the arrow (right side) are called products (formed).
- The numbers in front of the formulas are called stoichiometric coefficients.

$$2Na + \underbrace{2H_2O}_{\text{Reactants}} \rightarrow 2\underbrace{NaOH}_{\text{Products}} + H_2$$

Coefficient

- Stoichiometric coefficients: numbers in front of the chemical formulas; give ratio of reactants and products.

Chemical Equations

- A representation of a chemical reaction:
- State symbols:
  - (*s*) solid (*l*) liquid (*g*) gas
  - (*aq*) water solution
- $C_2H_5OH\ (l) + 3O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3H_2O\ (l)$

  reactants products

$HCl(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + NaCl(aq)$

Catalysts, photons (hν), or heat (Δ) may stand above the reaction arrow ($\rightarrow$).

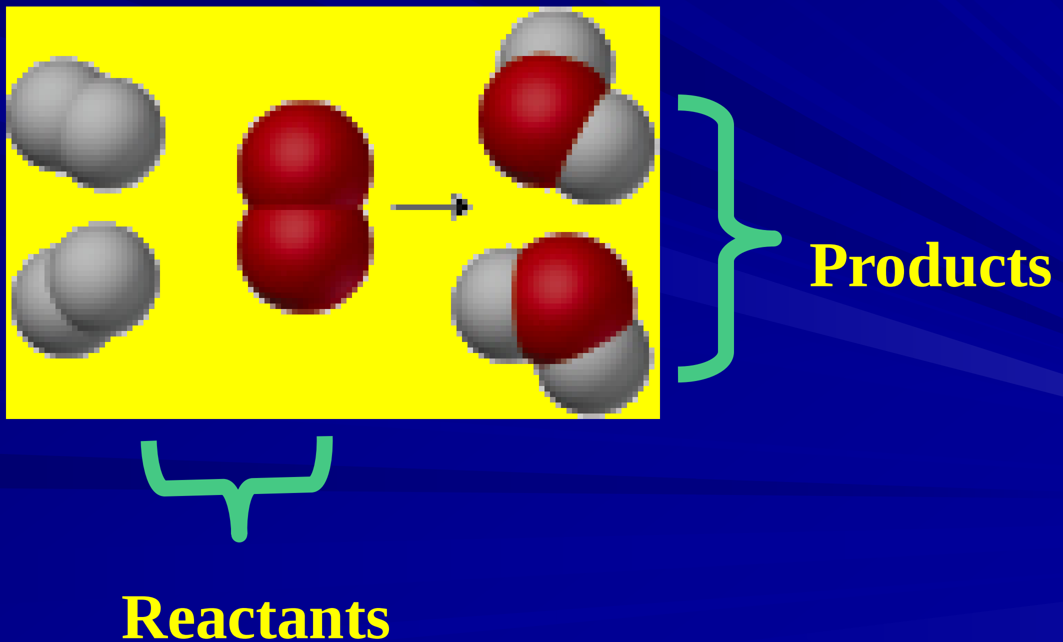
Balanced chemical equation

- $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
- The equation is balanced.
- 1 molecule of ethanol reacts with 3 molecules of oxygen to produce

2 molecules of carbon dioxide and 3 molecules of water

- **The chemical equation for the formation of water can be visualized as two hydrogen molecules reacting with one oxygen molecule to form two water molecules:**

$$2H_2 + \mathbf{O_2} \rightarrow 2H_2\mathbf{O}$$



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

  $2\ KClO_{3(s)} \xrightarrow{\Delta} 2\ KCl_{(s)} + 3\ O_{2(g)}$

Balancing chemical equations

- When balancing a chemical reaction coefficients are added in front of the compounds to balance the reaction, but subscripts should not be changed

- Changing the subscripts changes the compound

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.

Balancing Chemical Equations
By inspection (Trial and error)

- ❖ **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

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

- ❖ **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.**

$$2C_2H_6 \quad \textit{NOT} \quad C_4H_{12}$$

Balancing Chemical Equations

- Start by balancing those elements that appear in only one reactant and one product.

$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ start with C or H but not O

2 carbon on left

1 carbon on right

multiply CO_2 by 2

$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

6 hydrogen on left

2 hydrogen on right

multiply H_2O by 3

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

Balancing Chemical Equations

- Balance those elements that appear in two or more reactants or products.

$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$ 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
multiply both sides by 2

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

Balancing Chemical Equations

- Check to make sure that you have the same number of each type of atom on both sides of the equation.

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

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)

Balancing Chemical Equations

- Check to make sure that you have the same number of each type of atom on both sides of the equation.

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

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

3.4 The Mole and Molar Masses

- Balanced equations tell us what is reacting and in what relative proportions on the **molecular level**.
- However, **chemists must work with the chemical reactions on a macroscopic level.**

The Mole

- The *mole (mol)* is a number equal to the number of carbon atoms in exactly 12.00 grams of ^{12}C (Counting by weight??)
- Techniques such as mass spectrometry were used to count this number
- The number was found as $6.02214X10^{23}$
- This number was known as "Avogadro's number"
- Thus, one mole of a substance contains of $6.022X10^{23}$ units of that substance
- So, as a dozen of eggs is 12; a mole of eggs is Avogadro's number of eggs.

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 $H_2O_{(l)}$

This relationship can be made because of Avogadro's number (N_A)

How the mole is used in chemical calculations?

- 12 grams of ^{12}C contain Avogadro's number of atoms

  = $6.022 X 10^{23}$ C atoms
- 12.01 g of natural C (^{12}C, ^{13}C, ^{14}C) contains $6.022 X 10^{23}$ C atoms
- Atomic mass of C atom =12.01 amu

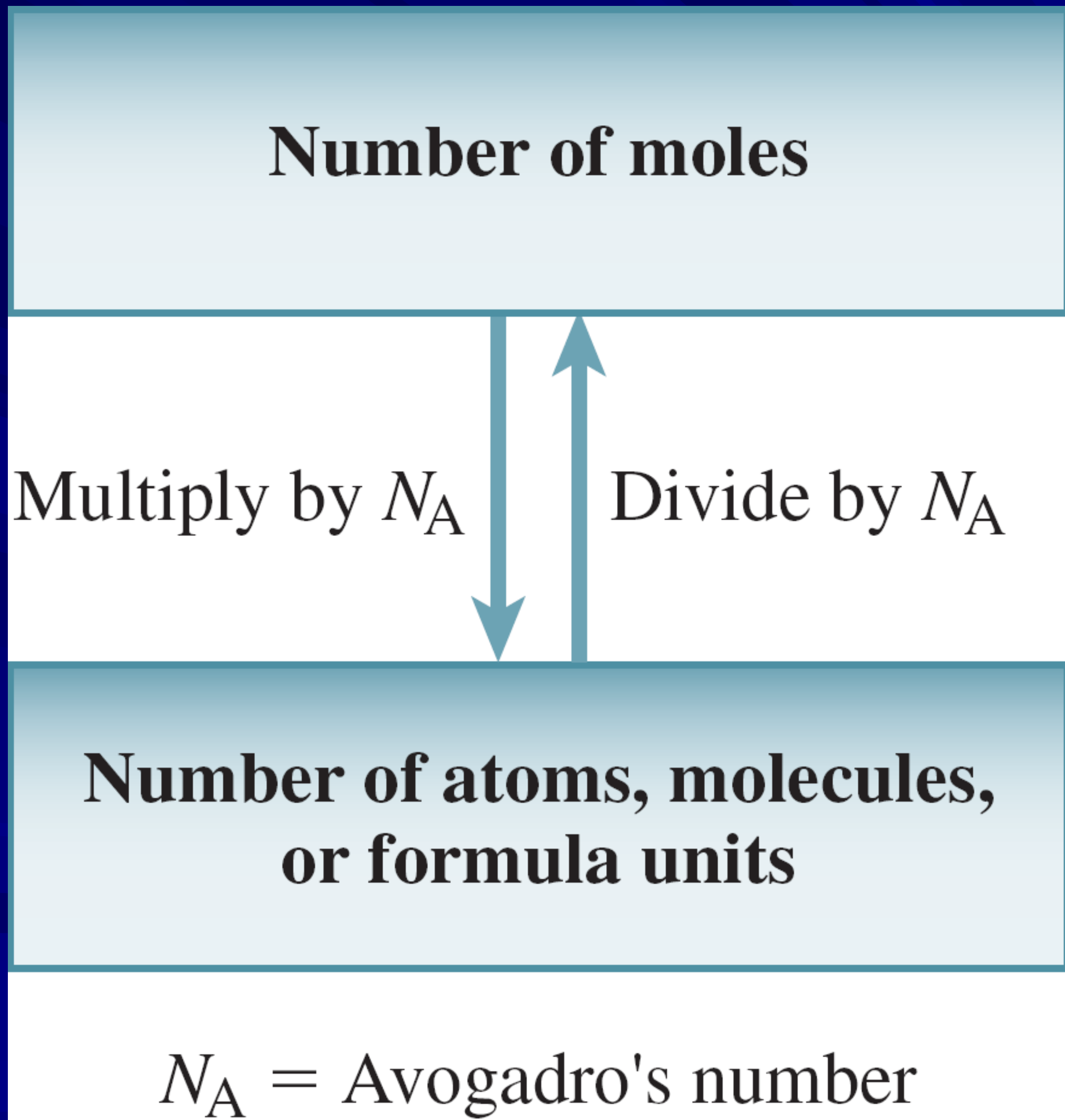
Express amu in grams

$6.022\ \mathrm{X}\ 10^{23}$ atoms of C (each has a mass of 12 amu) have a mass of 12 g

$$? \text{ g in} = 6.022 \times 10^{23} \text{ amu } X \frac{1 \text{ C atom}}{12 \text{amu}} X \frac{12\text{g}}{6.022 \times 10^{23} \text{ C atoms}} = 1\text{g}$$

$$Avogadro's \text{ number of amu} = 1\text{g C}$$

Moles and Atoms



Molar mass

- Molar mass is the mass of 1 mole of a substance expressed in grams.
- Often called molecular weight.
- To determine the molar mass of a compound: add up the molar masses of the elements (taking # moles of each element into consideration) that makes it up.

Molar Mass

- Molar masses of elements:

Carbon = 12.0 grams/mole

Sodium = 22.9 grams/mole

- What is the relationship between molar mass and atomic mass?

Molar Mass for Compounds

Calculate the molar mass for each of

H_2O

H 2 x 1.01 g/mol =   2.02

O 1 x 16.00 g/mol = <u>16.00</u>

Molar mass = 18.02 g/mol

Conversions between grams, moles and atoms

Grams → Divide by molar mass → **Moles**

$$\frac{\text{g}}{\text{g/mol}} = \text{mol}$$

Moles → Multiply by molar mass → **Grams**

$$\text{mol} \times (\text{g/mol}) = \text{g}$$

Particles (Atoms, Molecules, Formula units) → Divide by N_A → **Moles**

$$\text{particles} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ particles}} = \text{mol}$$

Moles → Multiply by N_A → **Particles** (Atoms, Molecules, Formula units)

$$\text{mol} \times \left(\frac{6.022 \times 10^{23} \text{ particles}}{1 \text{ mol}}\right) = \text{particles}$$

Interconverting mass, moles and number of particles

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

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

Example

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

(1 mole = 6.022×10^{23})

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

Empirical and Molecular Formulas

Empirical formula: **the lowest whole number ratio of atoms in a compound.**

Molecular formula: **the true number of atoms of each element in the formula of a compound.**

- **molecular formula = (empirical formula)$_n$**
- **[n = integer]**
- **molecular formula = C_6H_6 = $(CH)_6$**
- **empirical formula = CH**

Formulas for ionic compounds are ALWAYS empirical (lowest whole number ratio).

Examples:

$NaCl$ $MgCl_2$ $Al_2(SO_4)_3$ K_2CO_3

Formulas for molecular compounds MIGHT be empirical (lowest whole number ratio).

Molecular:	H_2O	$C_6H_{12}O_6$	$C_{12}H_{22}O_{11}$
	↓	↓	↓
Empirical:	H_2O	CH_2O	$C_{12}H_{22}O_{11}$

Calculating Empirical Formula from percent composition

- Empirical formulas can be determined from % composition data; a good analysis tool.
- Assume you have a 100 g. Then the percentages become grams.
- Convert grams to moles
- Find the 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.*

Example

- Calculate the empirical formula of a compound composed of 38.67 % C, 16.22 % H, and 45.11 %N.
- Assume 100 g so
- $38.67\text{ g C} \times \dfrac{1\text{mol C}}{12.01\text{ gC}} = 3.220\text{ mole C}$
- $16.22\text{ g H} \times \dfrac{1\text{mol H}}{1.01\text{ gH}} = 16.09\text{ mole H}$
- $45.11\text{ g N} \times \dfrac{1\text{mol N}}{14.01\text{ gN}} = 3.219\text{ mole N}$

Now divide each value by the smallest value

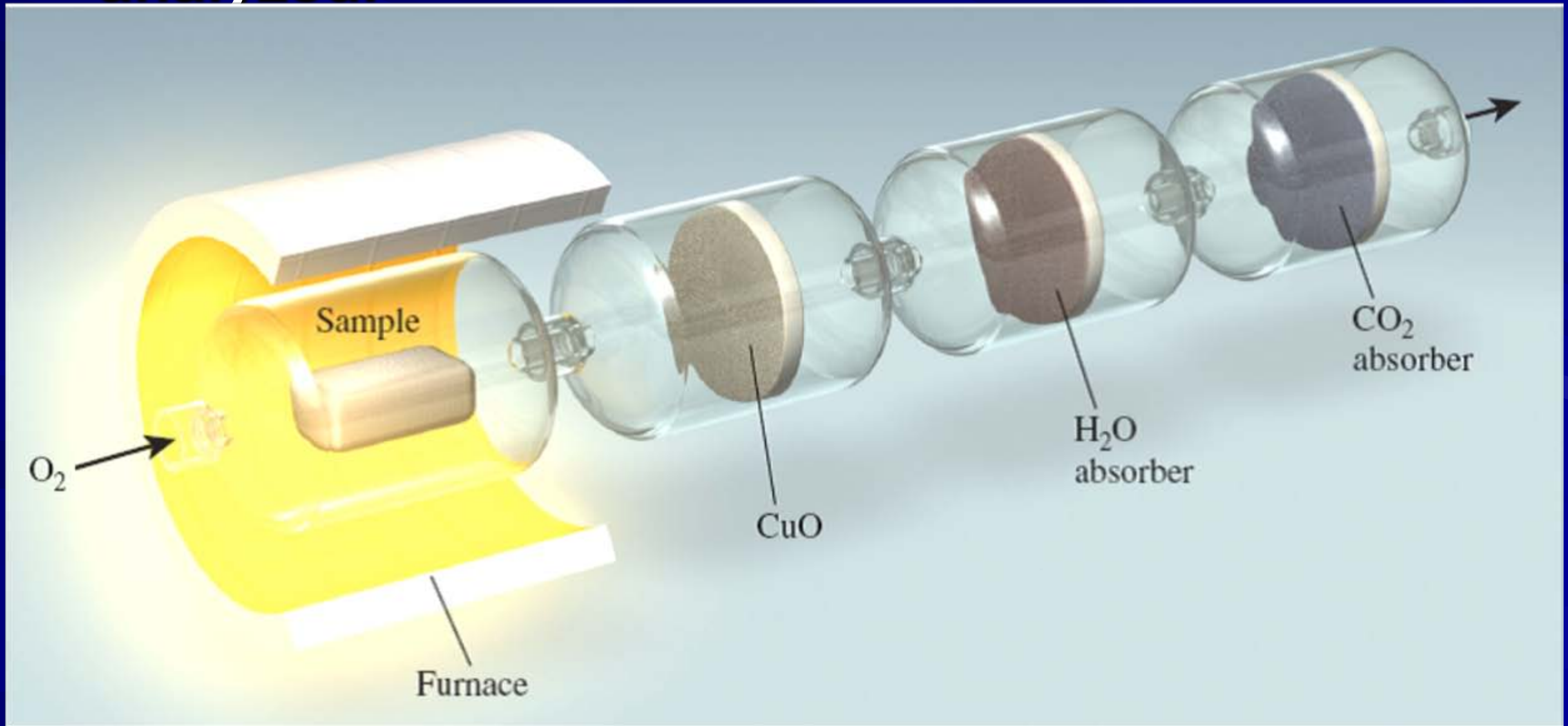
- The ratio is $\frac{3.220 \text{ mol C}}{3.219 \text{ mol N}} = \frac{1 \text{ mol C}}{1 \text{ mol N}}$

- The ratio is $\frac{16.09 \text{ mol H}}{3.219 \text{ mol N}} = \frac{5 \text{ mol H}}{1 \text{ mol}}$

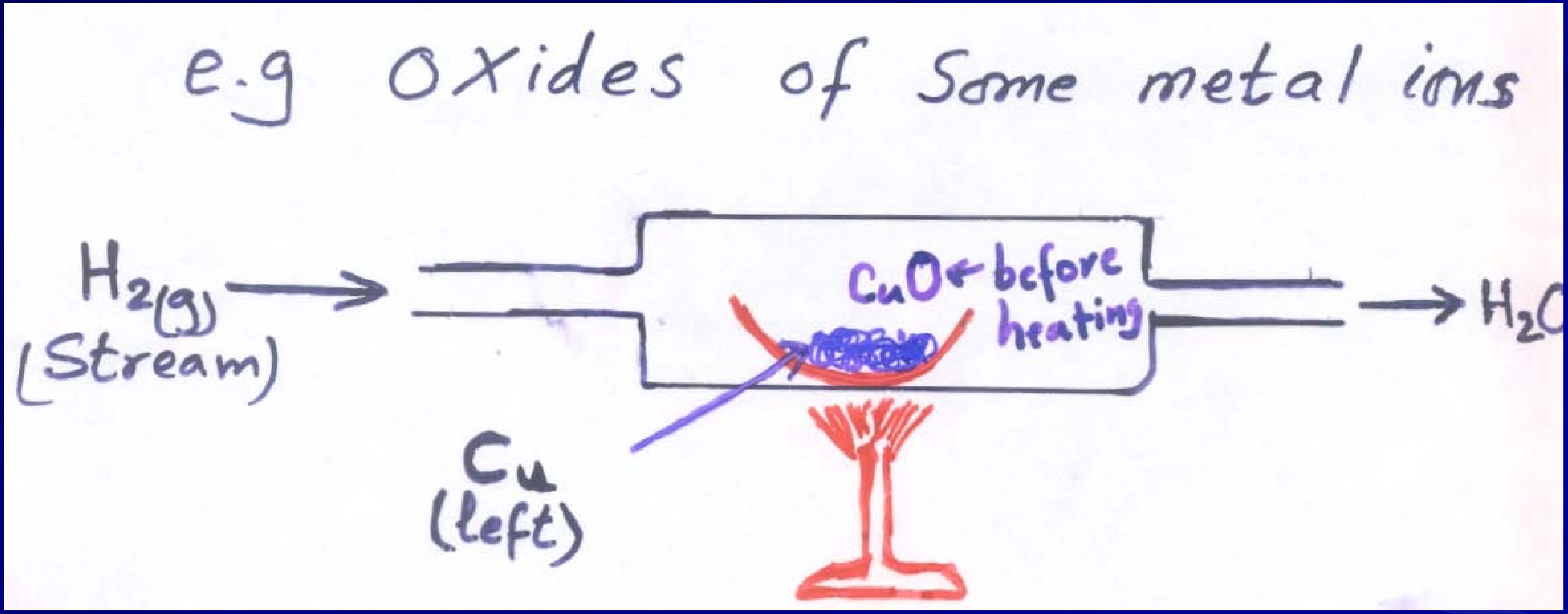
- = $C_1H_5N_1$

3.5 Combustion Analysis

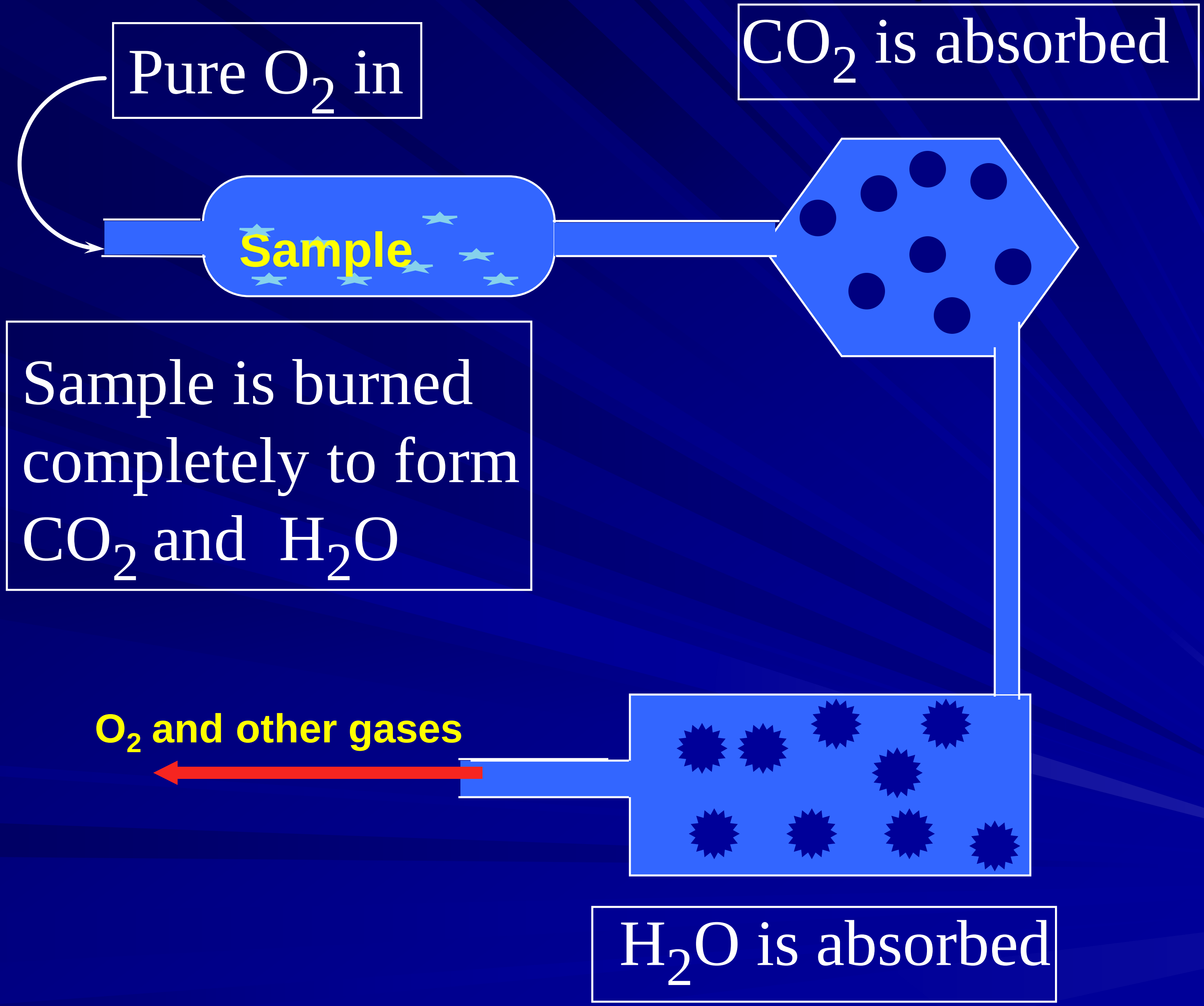
- Analysis of organic compounds (C,H and sometimes O) are carried using an apparatus like the one below
- A compound of unknown composition is decomposed by heat. The elements are carefully trapped and the number of moles of each are analyzed.



Determination of percent composition and simplest formula from experiment



Experimental Determination of the formula of a compound



Experimental Determination of the formula of a compound by elemental analysis

- A sample of a compound composed of carbon oxygen and hydrogen are combusted in a stream of O_2 to produce CO_2 and H_2O. The H_2O and CO_2 are trapped and the masses of each are measured.

Calculating empirical formula

A sample of certain material contains only C, H and O has a mass of 0.255g. When the reaction is complete, 0.561 g of CO_2 and 0.306g of H_2O are produced. What is the empirical formula of the compound?

1. Determine the mass of C in the sample.

- For each mole of CO_2, there is one mole of C. Convert moles of C to grams of C.
- 0.561 g CO_2 x (1 mol CO_2 / 44.01 g CO_2) x (1 mol C / 1 mol CO_2) (12.01 g C/ mol C)

= 0.153 g C

2. **Determine the mass of H in the sample.**

- There are 2 moles of hydrogen per mole of H_2O.
- 0.306 g H_2O x (1 mol H_2O / 18.0 g H_2O) x

(2 mole H / mole H_2O) x (1.01 g H / mol H)

= 0.0343 g H

3. Mass O = mass sample - mass H - mass C

- Mass sample = 0.255 g
- Mass O = 0.255 - 0.153 - 0.0343 = **0.068 g O**

- To get empirical formula, convert g back to moles
- 0.153 g C x (1 mol C / 12.01 g C) = 0.0128 mol C
- 0.0343 g H x (1 mol H / 1.01 g H) = 0.0340 mol H
- 0.068 g O x (1 mol O / 16.0 g O) = 0.0043 mol O
- Divide each by 0.0043 to get ratio of each element to O
- C: 0.0128 mol C / 0.0043 mol O = 2.98 ~ 3
- There are 3 moles of carbon for each mole of oxygen
- H: 0.0340 mol H / 0.0043 mol O = 7.91 ~ 8
- There are 8 moles of hydrogen per mole of oxygen
- Empirical Formula C_3H_8O

Determining molecular formula from empirical formula

- Since the empirical formula is the lowest ratio, the actual molecule would weigh more.
  - By a whole number multiple.
- <u>Divide</u> the actual molar mass by the empirical formula mass

- Benzene has the emperical formula CH and a molar mass of 78.0 g. what is its molecular formula?

Emperical formula mass = 12.0 + 1.00 = 13.00 g/mol

- n= **actual molar mass** / empirical **formula mass**
- **n = 78.0/13.0 = 6**
- molecular formula = (empirical formula)$_n$

  n = integer]

- molecular formula = C_6H_6

Example

A 2.103 g Copper oxide when heated in a stream of $H_{2(g)}$ yields 0.476 g H_2O. What is the formula of Copper oxide

Mass of O in H_2O formed = mass of O in the oxide

$$\#g\ O = 0.476\ g\ H_2O \times \frac{16\ g\ O}{18.0\ g\ H_2O} = 0.423\ g\ O$$

Mass of Cu = 2.103 g − 0.423 g = 1.680 g

$\# \text{ mol } O = 0.423 g\, O \times \frac{1 \text{ mol } O}{16.00 g\, O} = 0.0264 \text{ mo}$

$\# \text{ mol } Cu = 1.680 g\, Cu \times \frac{1 \text{ mol } Cu}{63.55 g\, Cu}$

$= 0.0264 \text{ mol}$

The Simplest ratio of O : Cu is

$0.0264 \text{ mol } O : 0.0264 \text{ mol } Cu$

$\frac{0.0264}{0.0264} \text{ mol } O : \frac{0.0264}{0.0264} \text{ mol } Cu$

$1\, O : 1\, Cu$

∴ The Simplest formula of copper oxide is CuO

If the Simplest mole ratio involves one or more fractional numberes

e.g : 3.50 : 2.33

Multiply through by the Smallest integer that will give Whole numbe ratio

$$\frac{3.50 \times 2/2}{2.33 \times 3/3} = \frac{7/2}{\underset{7}{\cancel{6.99}}/3} \approx \boxed{\frac{3}{2}}$$

3.6 Calculations with balanced chemical equations

- Chemical reactions involve changes in matter, making of new materials with new properties, or energy changes.
- Atoms cannot be created or destroyed
- Chemical reactions are described using the chemical equations
- Given an amount of either starting material or product, other quantities can be determined.
- use conversion factors from
  - molar mass (g - mole)
  - balanced equation (mole - mole)

Calculations of moles of reactants or products from Balanced Equations

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

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

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

Calculations of mass of reactants or products from Balanced Equations

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

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

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

Example

$$C_2H_4(g) + HCl(g) \longrightarrow C_2H_5Cl(g)$$

If we have 15.0 g of C_2H_4 **how many moles of HCl** are needed to carry out the reaction to completion?

The balanced equation tells us it takes 1 mole of HCl for each mole of C_2H_4 reacted.

We must begin by converting the number of grams of C_2H_4 which has a molar mass of 28.08 g/mol, to moles

$$15.0\ \text{g}\ \cancel{C_2H_4} \times \frac{1\ \text{mol}\ C_2H_4}{28.08\ \cancel{\text{g}\ C_2H_4}} = 0.534\ \text{mol}\ C_2H_4$$

What moles of HCl is needed to carry the reaction through to completion?

$$C_2H_4\,(g) + HCl(g) \longrightarrow C_2H_5Cl(g)$$

$$0.534\ \cancel{\text{mol } C_2H_4} \times \frac{1\text{mol HCl}}{\cancel{1\text{mol } C_2H_4}} = 0.534 \text{ mol HCl}$$

What mass of HCl is needed to carry the reaction through to completion?

$$0.534 \text{ mol HCl} \times \frac{36.5 \text{ g HCl}}{1\text{mol HCl}} = 19.6 \text{ g HCl}$$

Example

$$C_2H_4(g) + \cancel{HCl}(g) \longrightarrow C_2H_5Cl(g)$$

How many moles of product are made when 15.0 g of C_2H_4 is reacted with an excess of HCl?

$C_2H_4(g) + HCl(g) \longrightarrow C_2H_5Cl(g)$

15.0 g ?

Molar mass of C_2H_4 (ethylene) = 28.08 g/mol

$$15.0 \text{ g } C_2H_4 \times \frac{1\text{mole } C_2H_4}{28.06 \text{ g } C_2H_4} = 0.534 \text{ mol } C_2H_4$$

Stoichiometry of the balanced equation indicates a mole ratio of 1:1 reactants

$$C_2H_4\,(g) + HCl(g) \longrightarrow C_2H_5Cl(g)$$

0.534 mol 0.534 mol

(molar mass C_2H_5Cl = 64.51 g/mol)

$$0.543 \text{ mol } C_2H_5Cl \times \frac{64.51 \text{ g } C_2H_5Cl}{1 \text{ mol } C_2H_5Cl} = 35.0 \text{ g } C_2H_5Cl$$

Example

How many grams of NH_3 will be produced from 4.8 mol H_2?

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

Excess 4.8 mol ?

Begin with what you are given. You have 4.8 mol H_2

$$4.8\ \cancel{\text{mol}}\ H_2 \times \frac{2\ \text{mol}\ NH_3}{3\ \cancel{\text{mol}}\ H_2} = 3.2\ \text{mol}\ NH_3$$

$$3.2\ \text{mol}\ NH_3 \times \frac{17.0\ \text{g}\ NH_3}{1\ \text{mol}\ NH_3} = 54.4\ \text{g}\ NH_3$$

3.7 Limiting reactants

- **If the reactants are not present in stoichiometric amounts, at end of reaction some reactants are still present (in excess).**
- **Limiting Reactant: one reactant that is consumed**

Limiting Reactant: Reactant that limits the amount of product formed in a chemical reaction

Excess reactant - the one that is left overr

Limiting Reactants/ Example

$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

when 3.5g of Cu is added to a solution containing 6.0g of $AgNO_3$ what is the limiting reactant?

What is the mass Ag produced
and what is the mass of the excess reagent?

Calculations involving a limiting reactant

$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

1. $3.5\text{g Cu} \times \dfrac{1\text{ mol Cu}}{63.5\text{g Cu}} \times \dfrac{2\text{ mol Ag}}{1\text{ mol Cu}} =$ **0.11 mol of Ag**

2. $6.0\text{g } AgNO_3 \times \dfrac{1\text{mol } AgNO_3}{170\text{g AgNO3}} \times \dfrac{2\text{mol Ag}}{2\text{mol } AgNO_3} =$ **0.035mol Ag**

The Limiting Reactant is $AgNO_3$.

Limiting reactant is:
The reactant that produces the least amount of product

Determining the limiting reactant by comparing the mole ratio of Cu and $AgNO_3$ required by the balanced equation with the mole ratio actually present

$$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$$

$$3.5g\ Cu \times \frac{1\ mol\ Cu}{63.5g\ Cu} \times = \ 0.055\ mol\ Cu$$

$$6.0g\ AgNO_3 \times \frac{1 mol\ AgNO_3}{170g\ AgNO_3} = 0.0353 mol\ AgNO_3$$

$$\text{Mole ratio required by the balanced equation} = \frac{\text{mol AgNO}_3}{\text{mol Cu}} = \frac{2}{1} = 2.0$$

$$\text{Actual mole ration} = \frac{\text{mol AgNO}_3}{\text{mol Cu}} = 0.64$$

Thus $AgNO_3$ is the limiting reactant

What is the mass of Ag produced?

Take the limiting reactant:

$$6.0g\ AgNO_3 \times \frac{1mol\ AgNO_3}{170g\ AgNO3} \times \frac{2mol\ Ag}{2mol\ AgNO_3} \times \frac{108g\ Ag}{1molAg}$$

= 3.8g Ag

Excess reagent

- $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

Excess reagent is ????

Calculate amount of Cu needed to react with the limiting reactant

$$6.0g\ AgNO_3 \times \frac{1mol\ AgNO_3}{170g\ AgNO3} \times \frac{1mol\ Cu}{2mol\ AgNO_3} \times \frac{63.5\ Cu}{1molCu}$$

$$= 1.12\ g\ Cu$$

Amount of Cu left (Excess reagent) = 3.5 – 1.12 = 2.38 g Cu

The Reaction Yield

- The amount of stuff you make in the experiment is the yield.
- The **theoretical yield** is the amount you would get if everything went perfect.
- The **actual yield** is what you make in the lab.

The percent yield

- **The percent yield relates the actual yield (amount of material recovered in the laboratory) to the theoretical yield:**

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

Learning outcomes

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

Learning outcomes

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