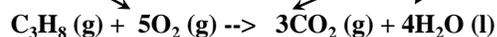


Stoichiometry

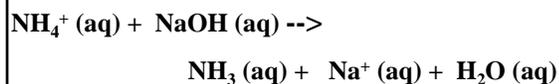
Relationship between the moles of the reactants and the moles of the products in a chemical equation is indicated by Coefficients in the Chemical Equation



Balanced Equations

Mass Balance

Charge Balance



Avogadro's Number

$$6.022137 \times 10^{23} = 1 \text{ Mole}$$

$$1 \text{ g} = 6.02 \times 10^{23} \text{ amu}$$

Amedeo Avogadro (1776-1856) Turin, Italy - Avogadro's hypothesis
1811- equal volumes of gas contain equal numbers of particles

**Atomic Weight
and
Molecular Weight**

**Grams/Mole
(Daltons)**

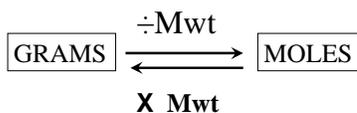
weight is inaccurate these are
actually masses

**Conversion of Mass to Moles
(Moles are the Currency of
Exchange for chemists)**

Example:

20 g of Glucose ($C_6H_{12}O_6$ - Mwt = 180 g/mol)

$$20.0 \text{ g} \div 180 \text{ g/mol} = 0.111 \text{ mol}$$



Moles to Grams (& volume)

Example:

0.500 moles of Ethanol (C_2H_6O - Mwt = 46 g/mol)

$$0.500 \text{ g} \times 46 \text{ g/mol} = 23.0 \text{ g}$$

$$D_{\text{EtOH}} = 0.90 \text{ g/mL}$$

$$23.0 \text{ g} \div 0.900 \text{ g/mL} = 25.6 \text{ mL EtOH}$$

Problem

How many atoms of Oxygen are present in 4.9 g of H_2SO_4 ?

$$4.9 \text{ g} \div 98 \text{ g/mol} = 0.05 \text{ mol } \text{H}_2\text{SO}_4$$

$$4 \text{ mol of O atoms per } \text{H}_2\text{SO}_4 \times 0.05 = 0.20 \text{ moles O}$$

$$6.02 \times 10^{23} \text{ atoms/mole} \times 0.20 \text{ moles} = \\ 1.2 \times 10^{23} \text{ atoms}$$

Percent Composition

The composition of chemical compounds is expressed in percentage by mass



27.3% C

76.7% O

Calculating Percent Composition

Since percent composition is determined on a mass basis we must calculate the mass of each element in a chemical formula and the mass of the total chemical formula

Example: What is the % composition of aniline ($\text{C}_6\text{H}_7\text{N}$)?

	C	H	N
Mol wt of aniline =	$(6 \times 12.01) +$	$(7 \times 1.008) +$	14.007
	72.06 g	7.056g	14.007g
÷ Total molecular weight (93.12 g/mol) x 100%			
percentage =	77.38%	7.58%	15.04%

Calculating Empirical Formulas

The Empirical formula of a compound is the simplest whole number ratio which conforms to the percentage composition.

The Molecular Formula may be equal to the Empirical formula or a whole number multiple of the Empirical formula.

Example: Glucose $C_6H_{12}O_6$ Molecular Weight = 180 g/mol

The simplest whole number ratio of the elements is $C_1H_2O_1$ and this is the result that will be obtained from a calculation from the percent composition.

The Glucose Empirical Formula

The percentage composition must be converted to number of moles for each element.

	C	H	O
Percentage	40.03%	6.667%	53.33%

To do this assume you have 100g of the substance having the composition given. Then you will have masses corresponding to the percentage of each element.

Mass	40.03g	6.667g	53.33g
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convert to moles by dividing each by the *correct* atomic weight.

Moles	3.333 mole	6.614 mole	3.333 mole
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The Glucose Empirical Formula

	C	H	O
Moles	3.333 mole	6.614 mole	3.333 mole

We always express the composition in terms of whole numbers. To convert to whole numbers divide by the smallest of the values obtained (3.333) often (but not always) this is the *largest common factor* of the numbers

Whole Number Ratio	1	1.98	1
After Rounding	1	2	1

The Empirical Formula is $C_1H_2O_1$

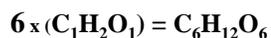
Converting Empirical Formulas to Molecular Formulas

Glucose Empirical Formula = $C_1H_2O_1$

the Empirical Formula weight = 30.025 g/formula weight

the Molecular Weight (determined by a separate experiment, see Chapter 11) = 180 g/mole

an Integer factor of $n = (180 \div 30.025) = 6$ relates the Empirical formula to the Molecular formula



A Tricky Example

A 20.882 g sample is found to have 6.072 g Na, 8.474 g S, & 6.336 g of O.

element (at wt)	Na (22.99)	S (32.07)	O (15.999)
Mass (g)	6.072	8.474	6.336
moles	0.2641	0.2642	0.3960
ratio ($\div 0.2641$)	1	1	1.5

for O we do not have an integer \rightarrow (0.2641 is not the largest common factor)

note the decimal $0.5 \times 2 =$ the integer 1

multiply all values by 2 $\implies Na_2S_2O_3$

Combustion analysis

A 2.00 g sample of a compound produced 4.86 g of CO_2 and 2.03 g of H_2O upon combustion. What is the Empirical Formula?

$4.86 \div 44.01 \text{ g/mol} = 0.1104 \text{ mol } CO_2 \implies 0.1104 \text{ mol C} \implies 1.326 \text{ g C}$

$2.03 \div 18.02 \text{ g/mol} = 0.1127 \text{ mol } H_2O \implies 0.2254 \text{ mol H} \implies 0.2272 \text{ g H}$

determine O by difference

mass of H + C = 1.553 g therefore $2.0 - 1.553 = 0.4468 \text{ g O}$

0.02792 mol O

Combustion cont'd

	C	H	O
Moles	0.1104	0.2254	0.02792
Ratio	3.95	8.07	1

Empirical Formula = C₄H₈O

Empirical formula weight = 72.143 g/mol
