Stoichiometry

Units of Measurement
Standard International Unites (SIU)

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

[a]The abbreviation sec is frequently used.

Metric System

Prefixes convert the base units into units that are appropriate for the item being measured.

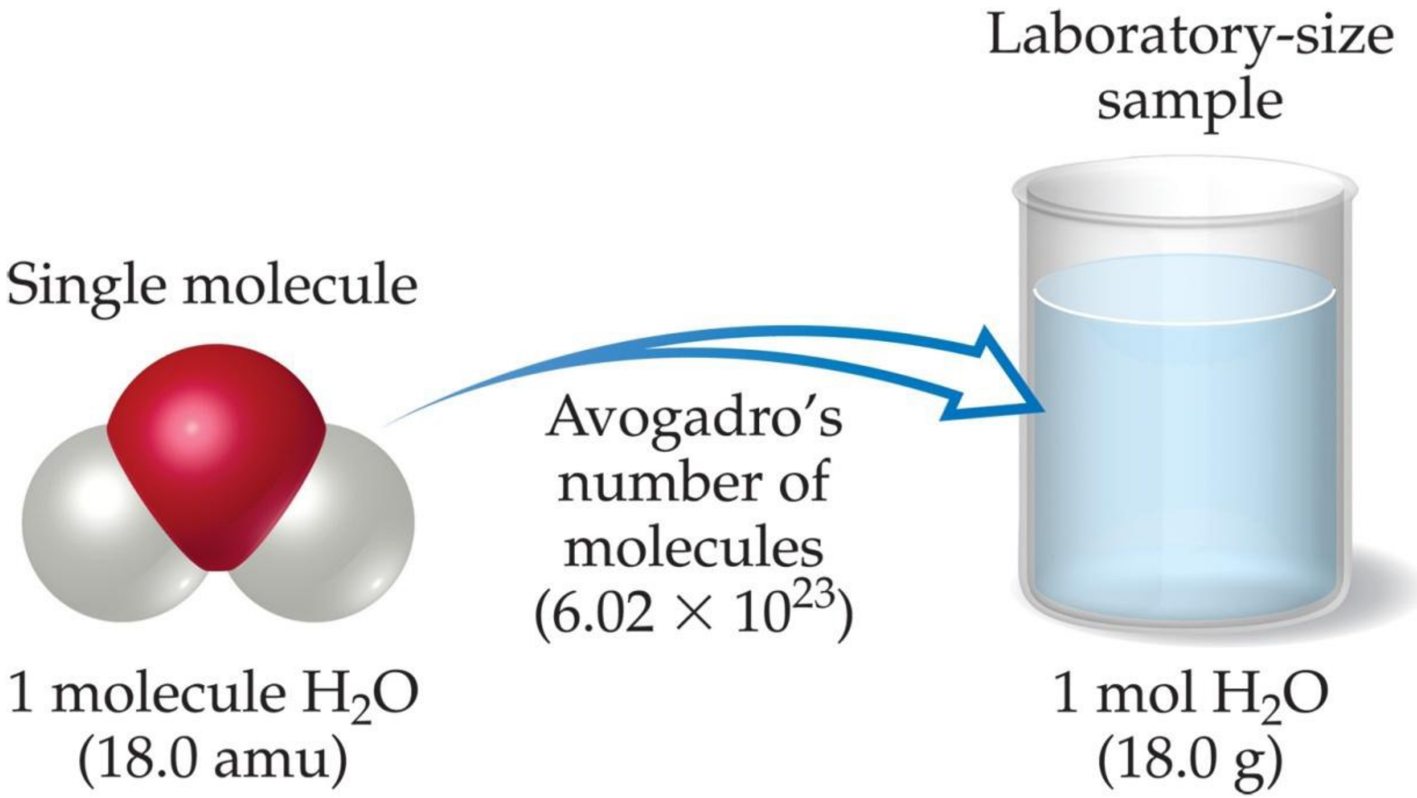
Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ m
Deci	d	10^{-1}	1 decimeter (dm) $= 0.1$ m
Centi	c	10^{-2}	1 centimeter (cm) $= 0.01$ m
Milli	m	10^{-3}	1 millimeter (mm) $= 0.001$ m
Micro	μ[a]	10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

[a]This is the Greek letter mu (pronounced "mew").

More Derived Units

Quantity	Definition	Units
Area	Length × width	m^2
Volume	Length × width × height	m^3
Density	Mass / volume	kg/m^3, g/cm^3, g/mL
Speed	Distance / time	$m\ s^{-1}$
Acceleration	Change in speed / time	$m\ s^{-2}$
Frequency	Event / time	s^{-1}
Force	Mass × acceleration	$kg\ m\ s^{-2}$ (newton, N)
Pressure	Force / area	$kg\ m^{-1}\ s^{-2}$ (pascal, Pa)
Energy	Force × distance	$kg\ m^2\ s^{-2}$ (joule, J)

Avogadro's Number



- $N_A = 6.02 \times 10^{23}$

1 mole of ^{12}C has a mass of 12 g.

The Mole

The amount of substance that contains as many atoms as there are in exactly 12 grams of pure 12C. That is 6.022×10^{23} atoms

1 mole of particles contains 6.022×10^{23} particles (Avogadro's number)

For example

One mole of hydrogen atoms = 6.023×10^{23} atoms of hydrogen

One mole of hydrogen molecules = 6.023×10^{23} molecule of hydrogen

One mole of electrons = 6.023×10^{23} electrons

One mole of sodium ions (Na^+) = 6.023×10^{23} Na^+ ions

Molar Mass

A substance's molar mass (molecular weight) is the mass in grams of one mole of the compound.

C=12 , O=16 amu

CO_2 = 44.01 grams / mole

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

1 atom of ^{12}C = 12.00 amu

1 mole ^{12}C atoms = 12.00 g ^{12}C

For any element
atomic mass (amu) = molar mass (grams)

Ex: How many H atoms are in 72.5 g of C_3H_8O ?

1 mol C_3H_8O molecules = 8 mol H atoms

1 mol H = 6.023×10^{23} atoms H

1 mol C_3H_8O = (3 x 12) + (8 x 1) + 16 = 60 g C_3H_8O

$$\frac{72.5 \text{ g } C_3H_8O}{60 \text{ g } C_3H_8O} \times 8 \text{ mol H atoms} \times 6.023 \times 10^{23} \text{ H atoms} =$$

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

Types of Formulas

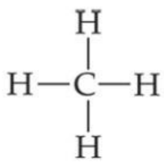
- **Empirical Formula: (E.F.)**
- Simplest whole-number ratio of atoms in a compound where “empirical” means derived from experiment

- **Molecular Formula: (M.F.)**
- Chemical formula of a compound that expresses the actual number of atoms present in one molecule.

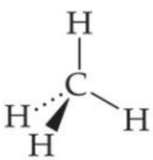
-The molecular formula will either be exactly the same or some multiple of the empirical formula!

H_2O E.F same as M.F

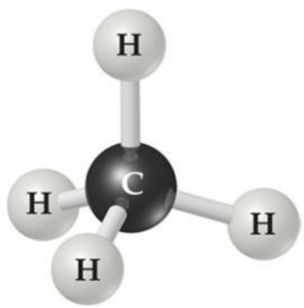
Types of Formulas



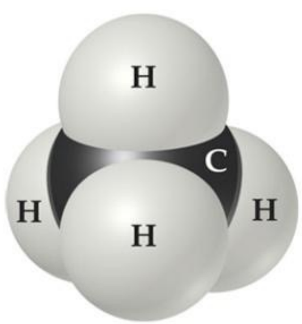
Structural formula



Perspective drawing



Ball-and-stick model



Space-filling model

- **Structural formulas**

  show the order in which atoms are bonded.

- Perspective drawings also show the three-dimensional array of atoms in a compound.

Percent Composition

Percent composition by mass:

The mass of one element in a compound divided by the mass of the entire compound

Steps to determine percentage composition:

1. Calculate the mass of each individual element in the compound
2. Add up all the masses of each element to get the total mass of compound
3. Divide the mass of each individual element with the total mass of compound

Percent Composition

$$\% \text{ element} = \frac{(\text{number of atoms})(\text{atomic weight})}{(\text{FW of the compound})} \times 100$$

The Percent of C in C_2H_6

$$\%C = \frac{(2)(12.0 \text{ amu})}{(30.0 \text{ amu})}$$

$$= \frac{24.0 \text{ amu}}{30.0 \text{ amu}} \times 100$$

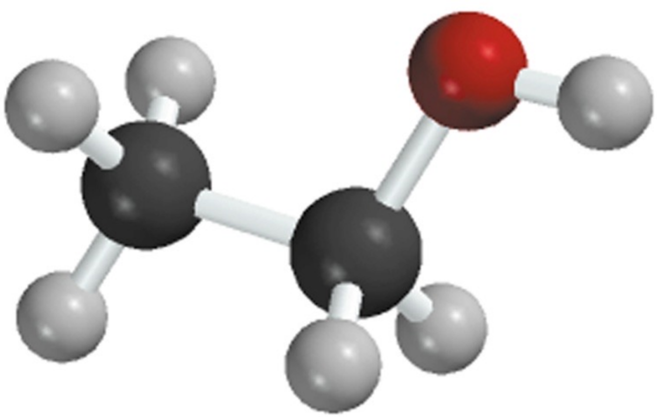
$$= 80.0\ \%$$

Percentage composition

Percent composition of an element in a compound

$$= \frac{n \times \text{molar mass of element}}{\text{molar mass of compound}} \times 100\%$$

n is the number of moles of the element in 1 mole of the compound



C_2H_6O

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

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

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

$52.14\% + 13.13\% + 34.73\% = 100.00\%$

Calculating Empirical Formulas

The compound *para*-aminobenzoic acid is composed of carbon (61.31%), hydrogen (5.14%), nitrogen (10.21%), and oxygen (23.33%). Find the empirical formula of PABA.

Assuming 100.00 g of *para*-aminobenzoic acid,

C: $$61.31 \text{ g} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 5.105 \text{ mol C}$$

H: $$5.14 \text{ g} \times \frac{1 \text{ mol}}{1.01 \text{ g}} = 5.09 \text{ mol H}$$

N: $$10.21 \text{ g} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 0.7288 \text{ mol N}$$

O: $$23.33 \text{ g} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 1.456 \text{ mol O}$$

Calculate the mole ratio by dividing by the smallest number of moles:

$$\text{C:} \quad \frac{5.105\ \cancel{\text{mol}}}{0.7288\ \text{mol}} = 7.005 \approx 7$$

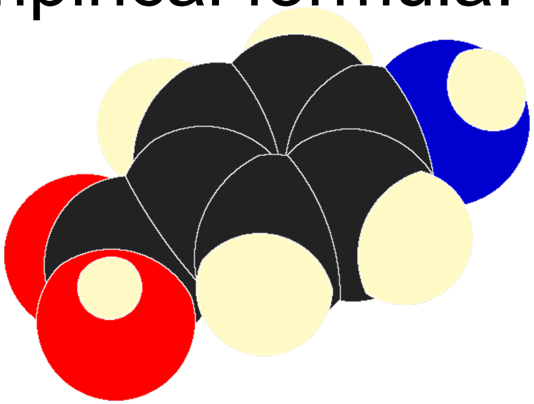
$$\text{H:} \quad \frac{5.09\ \cancel{\text{mol}}}{0.7288\ \cancel{\text{mol}}} = 6.984 \approx 7$$

$$\text{N:} \quad \frac{0.7288\ \cancel{\text{mol}}}{0.7288\ \cancel{\text{mol}}} = 1.000$$

$$\text{O:} \quad \frac{1.458\ \cancel{\text{mol}}}{0.7288\ \cancel{\text{mol}}} = 2.001 \approx 2$$

These are the subscripts for the empirical formula:

$$\mathbf{C_7H_7NO_2}$$



Molecular Formula

Molecular formula of a compound that expresses the actual number of atoms present in one molecule.

The molecular formula will either be exactly the same or some multiple of the empirical formula

Determining Molecular Formula

1. Follow the same steps for determining empirical formula.
2. For the Molecular Formula, you need to be given the molar mass of the compound. Find the molar mass of the empirical formula.
3. Divide the molar mass of the compound by the molar mass of the empirical formula to get the factor with which to multiply each subscript in the empirical formula

A compound contains 75% carbon and 25% hydrogen. Determine its empirical formula. The molecular mass of this compound is 16 amu. Determine its molecular formula also. The atomic masses are: C = 12 amu, H = 1 amu.

Element	% Composition	Atomic mass	Atomic Ratio	Simplest atomic Ratio
C H	75 25	12 1	75/12 = 6.25 25/1=25	6.25 / 6.25 = 1 25/6.25 = 4

So, The empirical formulae of the compound = C_1H_4 or CH_4

The empirical formula mass = (1 x 12) + (4 x 1) = 12 + 4 = 16 amu

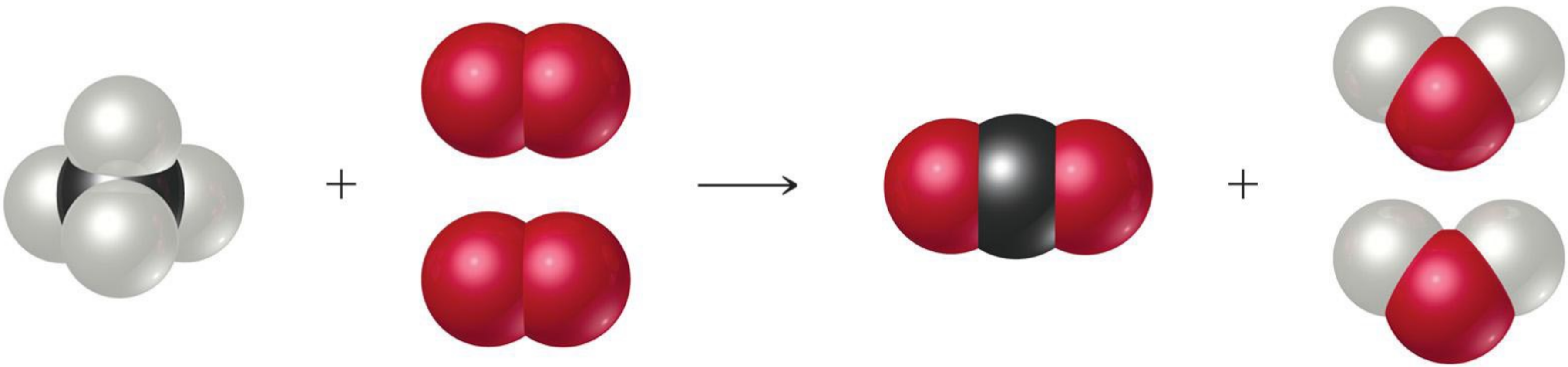
Molecular mass (given) = 16 amu

$$\text{So, } N = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{16\text{ amu}}{16\text{ amu}}$$

Therefore, Molecular formula = 1 x Empirical formula = 1 x CH_4 = CH_4

Chemical Equation

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



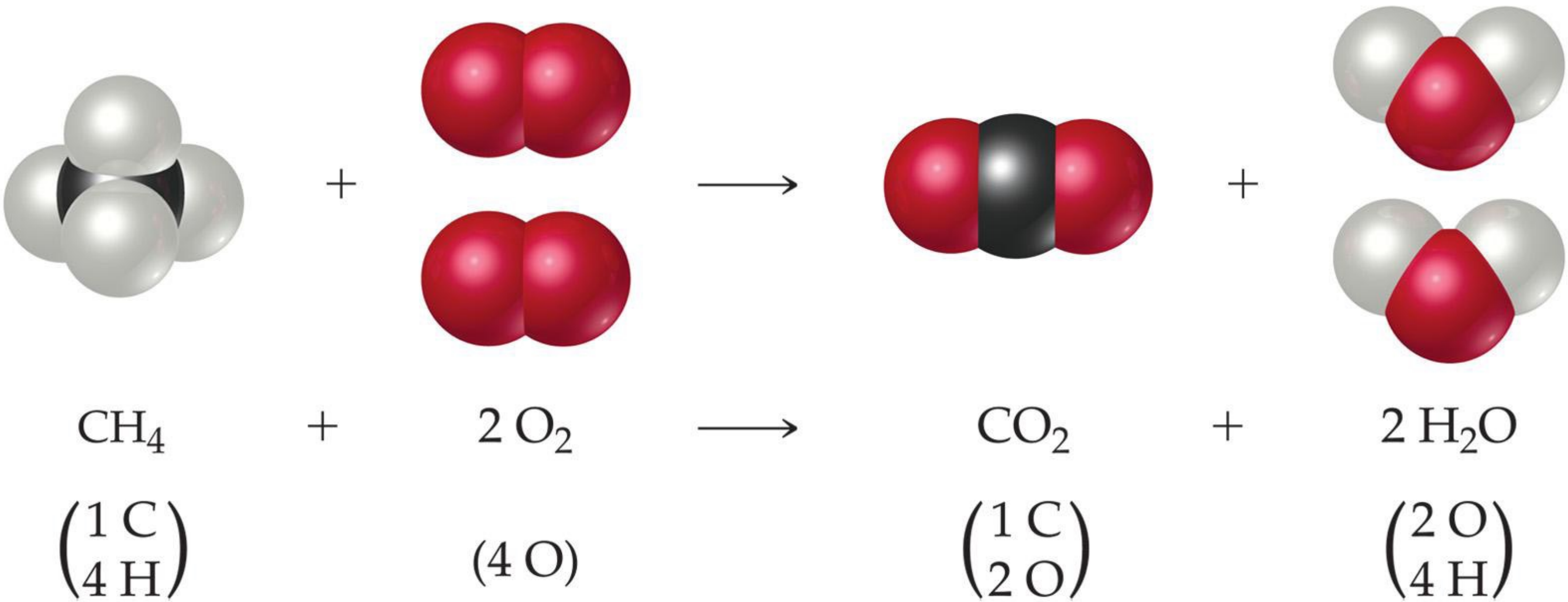
$$\underset{\begin{pmatrix}1\,C\\4\,H\end{pmatrix}}{CH_4} \quad + \quad \underset{(4\,O)}{2\,O_2} \quad \longrightarrow \quad \underset{\begin{pmatrix}1\,C\\2\,O\end{pmatrix}}{CO_2} \quad + \quad \underset{\begin{pmatrix}2\,O\\4\,H\end{pmatrix}}{2\,H_2O}$$

Reactants appear on the left side of the equation

Products appear on the right side of the equation

Chemical Equation

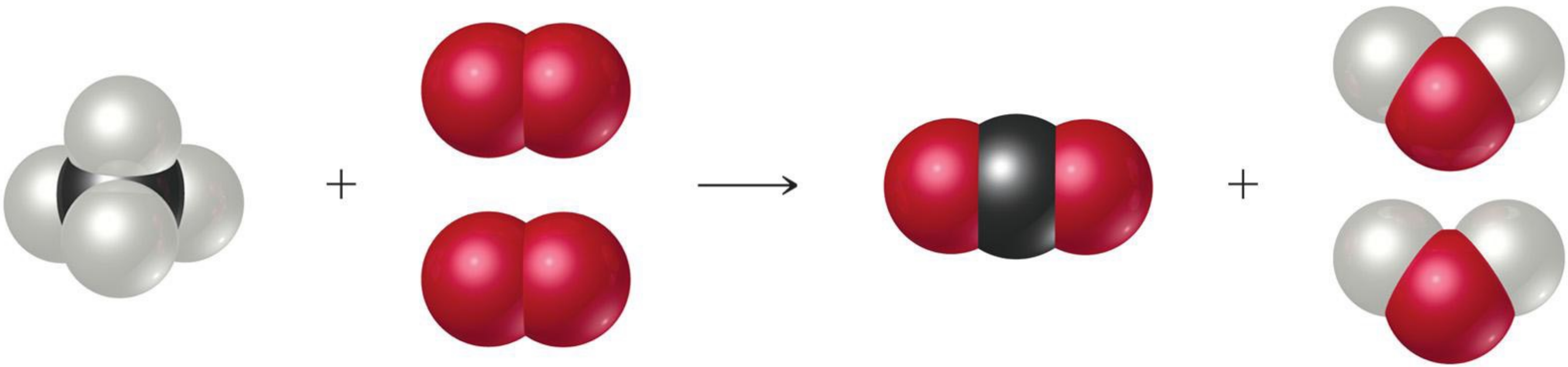
$$CH_{4\ (g)} + 2\ O_{2\ (g)} \longrightarrow CO_{2\ (g)} + 2\ H_2O_{\ (g)}$$



The states of the reactants and products are written in parentheses to the right of each compound

Chemical Equation

$$CH_{4\ (g)} + 2\ O_{2\ (g)} \longrightarrow CO_{2\ (g)} + 2\ H_2O_{\ (g)}$$

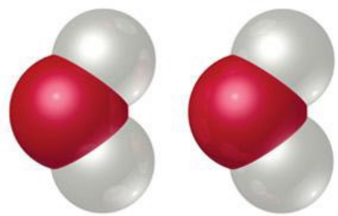


$$CH_4 \quad + \quad 2\,O_2 \quad \longrightarrow \quad CO_2 \quad + \quad 2\,H_2O$$

$$\begin{pmatrix}1\,C\\4\,H\end{pmatrix} \qquad (4\,O) \qquad \begin{pmatrix}1\,C\\2\,O\end{pmatrix} \qquad \begin{pmatrix}2\,O\\4\,H\end{pmatrix}$$

Coefficients are inserted to balance the equation

Subscripts and Coefficients

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
- Coefficients tell the number of molecules.

BALANCING EQUATION

Balancing by Inspection GUIDELINES

1. Count the number of elements on both sides of the equation
2. Change the coefficients (NEVER the subscripts) to get the same number of elements on both sides of the equation

BALANCING EQUATION

A representation of a chemical reaction:

$C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

reactants products

Unbalanced !

$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

The equation is balanced.

1 mole of ethanol reacts with 3 moles of oxygen

to produce

2 moles of carbon dioxide and 3 moles of water

Stoichiometric Calculations

Ex.: 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?

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

$$2\ \text{mol} \longrightarrow 4\ \text{mol}$$

$$2x(12+4+16)\ \text{g} \longrightarrow 4x(2+16)\ \text{g}$$

$$209\ \text{g} \longrightarrow m$$

$$m = \frac{209 \times 4(2+16)}{2(12+4+16)} = 235\ \text{g}\ H_2O$$

Theoretical Yield

•The theoretical yield is the maximum amount of product that can be made.

– In other words it's the amount of product possible as calculated through the stoichiometry problem.

•This is different from the actual yield, which is the amount one actually produces and measures.

•Actual yield is less than theoretical yield

Percent Yield

- One finds the percent yield by comparing the amount actually obtained (actual yield) to the amount it was possible to make (theoretical yield).

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

Ex: Without using a calculator, arrange the following samples in order of increasing numbers of carbon atoms:

12 g ^{12}C, 1 mol C_2H_2, 9×10^{23} molecules of CO_2

12 g ^{12}C (6×10^{23} C atoms) < 9×10^{23} CO_2 molecules (9×10^{23} C atoms) < 1 mol C_2H_2 (12×10^{23} C atoms).

Ex: What is the mass in grams of 1.000 mol of glucose, $C_6H_{12}O_6$?

$$\begin{aligned} 6\text{ C atoms} &= 6(12.0\text{ amu}) &= 72.0\text{ amu} \\ 12\text{ H atoms} &= 12(1.0\text{ amu}) &= 12.0\text{ amu} \\ 6\text{ O atoms} &= 6(16.0\text{ amu}) &= 96.0\text{ amu} \\ & & \overline{180.0\text{ amu}} \end{aligned}$$

$C_6H_{12}O_6$ has a molar mass of 180.0 g/mol

Ex: Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$

$$\text{Moles } C_6H_{12}O_6 = (5.380\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\text{ mol } C_6H_{12}O_6}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right) = 0.02989\text{ mol } C_6H_{12}O_6$$

Ex: Calculate the mass, in grams, of 0.433 mol of calcium nitrate

$$\text{Grams } Ca(NO_3)_2 = (0.433 \cancel{\text{mol } Ca(NO_3)_2})\left(\frac{164.1 \text{ g } Ca(NO_3)_2}{1 \cancel{\text{mol } Ca(NO_3)_2}}\right) = 71.1 \text{ g } Ca(NO_3)_2$$

Ex: How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$? (b) How many oxygen atoms are in this sample?

$$= (5.23 \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \cancel{\text{mol } C_6H_{12}O_6}}{180.0 \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6.02 \times 10^{23} \text{ molecules } C_6H_{12}O_6}{1 \cancel{\text{mol } C_6H_{12}O_6}}\right)$$
$$= 1.75 \times 10^{22} \text{ molecules } C_6H_{12}O_6$$

$$\text{Atoms O} = (1.75 \times 10^{22} \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6 \text{ atoms O}}{1 \cancel{\text{molecule } C_6H_{12}O_6}}\right)$$
$$= 1.05 \times 10^{23} \text{ atoms O}$$

Ex: Ascorbic acid (vitamin C) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?

40.92 g C, 4.58 g H, and 54.50 g O.

$$\text{Moles C} = (40.92\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}}\right) = 3.407\ \text{mol C}$$

$$\text{Moles H} = (4.58\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}}\right) = 4.54\ \text{mol H}$$

$$\text{Moles O} = (54.50\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}}\right) = 3.406\ \text{mol O}$$

$$\text{C: } \frac{3.407}{3.406} = 1.000 \qquad \text{H: } \frac{4.54}{3.406} = 1.33 \qquad \text{O: } \frac{3.406}{3.406} = 1.000$$

$$\text{C:H:O} = 3(1:1.33:1) = 3:4:3$$

$C_3H_4O_3$

Ex: Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in $C_{12}H_{22}O_{11}$

$$\%C = \frac{(12)(12.0\ \text{amu})}{342.0\ \text{amu}} \times 100\% = 42.1\%$$

$$\%H = \frac{(22)(1.0\ \text{amu})}{342.0\ \text{amu}} \times 100\% = 6.4\%$$

$$\%O = \frac{(11)(16.0\ \text{amu})}{342.0\ \text{amu}} \times 100\% = 51.5\%$$

Ex: Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$

Solve

$$\text{H atoms} = (0.350\ \cancel{\text{mol } C_6H_{12}O_6})\left(\frac{6.02 \times 10^{23}\ \cancel{\text{molecules } C_6H_{12}O_6}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{12\ \text{H atoms}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)$$

$$= 2.53 \times 10^{24}\ \text{H atoms}$$

Ex: Mesitylene, a hydrocarbon that occurs in small amounts in crude oil, has an empirical formula of C_3H_4. The experimentally determined molecular weight of this substance is 121 amu. What is the molecular formula of mesitylene?

$$\frac{\text{Molecular weight}}{\text{Empirical formula weight}} = \frac{121}{40.0} = 3.02$$

therefore multiply each subscript in the empirical formula by 3 to give the molecular formula:

$$C_9H_{12}$$