Chapter 3
Chemical Reactions and Reaction Stoichiometry

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Stoichiometry

- The study of the numerical relationship between chemical quantities in a chemical reaction is called stoichiometry.

EX:

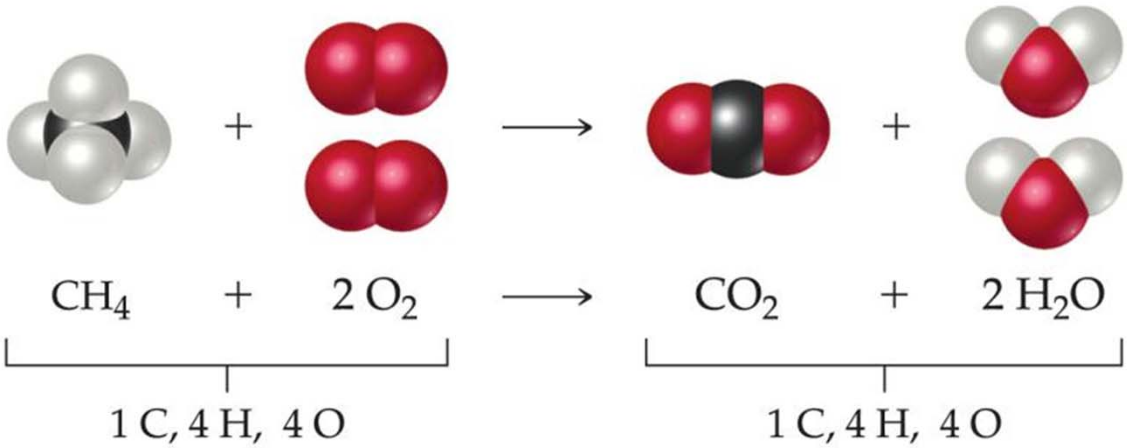
$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

2 molecules of C_8H_{18} react with 25 molecules of O_2 to form 16 molecules of CO_2 and 18 molecules of H_2O.

Chemical Equations

- Chemical equations are concise representations of **chemical reactions**

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$



Reactants appear on the left side of the equation.

Products appear on the right side of the equation.

The states of the reactants and products are written in parentheses to the right of each compound. (g) = gas; (l) = liquid; (s) = solid; (aq) = in aqueous solution

Chemical Equation

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

4 H atoms

2 H atoms

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

2 O atoms

2 O atoms + 1 O atom = 3 O atoms

Coefficients are inserted to balance the equation to follow the law of conservation of mass

How to Write Balanced Chemical Equations

- If an element occurs in only one compound on each side of the equation, try balancing this element first.
- When one of the reactants or products exists as the free element, balance this element last.
- In some reactions, certain groups of atoms (for example, polyatomic ions) remain unchanged. In such cases, balance these groups as a unit.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

2 O atoms

2 O atoms + 1 O atom = 3 O atoms

Question

- 1) When the following equation is balanced, the coefficients are ________.

$$C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$$

Quiz

- 1) When the following equation is balanced, the coefficients are ________.

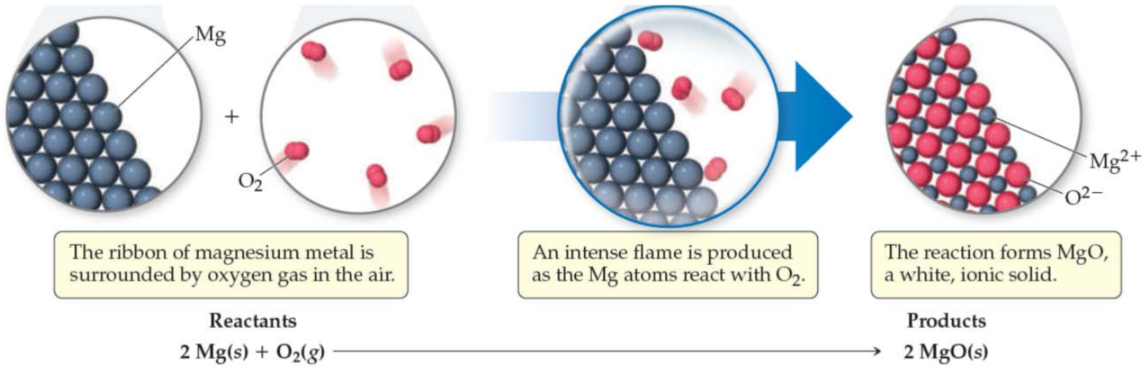
$$FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$$

Three Types of Reactions

- Combination reactions
- Decomposition reactions
- Combustion reactions

Combination reactions: two or more substances react to form one product.

- Examples:
  - $2\ Mg(s) + O_2(g) \longrightarrow 2\ MgO(s)$



Decomposition Reactions

Decomposition reaction: one substance breaks down into two or more substances

- Examples:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

Combustion Reactions

- Combustion reactions are generally rapid reactions that produce a flame.
- Combustion reactions most often involve oxygen in the air as a reactant.

- Examples:
  - $CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(g)$
  - $C_3H_8(g) + 5\ O_2(g) \longrightarrow 3\ CO_2(g) + 4\ H_2O(g)$

Formula Weight

- A formula weight is the sum of the atomic weights for the atoms in a chemical formula.
- The formula weight of calcium chloride, $CaCl_2$, would be

$$
\begin{array}{rl}
\text{Ca:} & 1\text{x}(40.08\ \text{amu}) \\
+\ \ \text{Cl:} & \underline{2\ \text{x}(35.453\ \text{amu})} \\
 & 110.99\ \text{amu}
\end{array}
$$

Molecular Weight

- A **molecular weight** is the sum of the atomic weights of the atoms in a molecule.
- For the molecule ethane, C_2H_6, the molecular weight would be
-

$$\begin{array}{rl} C: & 2 \times (12.011\ \text{amu}) \\ +\ H: & 6 \times (1.00794\ \text{amu}) \\ \hline & 30.070\ \text{amu} \end{array}$$

Quiz

- The formula weight of potassium dichromate ($K_2Cr_2O_7$) is _________ amu

- The formula weight of calcium nitrate $Ca(NO_3)_2$ is _________ amu.

Ionic Compounds and Formulas

- Ionic compounds exist with a three-dimensional order of ions.
- Ionic compounds use empirical formulas and formula weights (not molecular weights)

Percentage Composition from Chemical Formulas

- The mass percent composition or mass percent of an element is that element's percentage of the compound's total mass.

$$\%\ \text{composition of element} = \frac{\begin{pmatrix}\text{number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}\text{atomic weight}\\ \text{of element}\end{pmatrix}}{\text{formula weight of substance}} \times 100\%$$

Calculating Percentage Composition

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

Sol:

%C =(12x(12.0 amu)/ 342.0 amu)* 100% = 42.1%

%H =(22x(1.0 amu)/342.0 amu) * 100% = 6.4%

%O =(11x(116.0 amu)/342.0 amu) * 100% = 51.5%

Question

- What is the mass percent composition of halothane, $C_2HBrClF_3$?

The molar mass of $C_2HBrClF_3$ is 197.38 g/mol. The mass percents are

$$\% \text{ C} = \frac{\left(2 \text{ mol C} \times \dfrac{12.01 \text{ g C}}{1 \text{ mol C}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 12.17\% \text{ C}$$

$$\% \text{ H} = \frac{\left(1 \text{ mol H} \times \dfrac{1.01 \text{ g H}}{1 \text{ mol H}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 0.51\% \text{ H}$$

$$\% \text{ Br} = \frac{\left(1 \text{ mol Br} \times \dfrac{79.90 \text{ g Br}}{1 \text{ mol Br}}\right)}{197.38 \text{ } C_2HBrClF_3} \times 100\% = 40.48\% \text{ Br}$$

$$\% \text{ Cl} = \frac{\left(1 \text{ mol Cl} \times \dfrac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 17.96\% \text{ Cl}$$

$$\% \text{ F} = \left(\frac{3 \text{ mol F} \times \dfrac{19.00 \text{ g F}}{1 \text{ mol F}}}{197.38 \text{ g } C_2HBrClF_3}\right) \times 100\% = 28.88\% \text{ F}$$

Thus, the percent composition of halothane is 12.17% C, 0.51% H, 40.48% Br, 17.96% Cl, and 28.88% F.

Quiz

- The mass % of C in methane (CH_4) is _________?

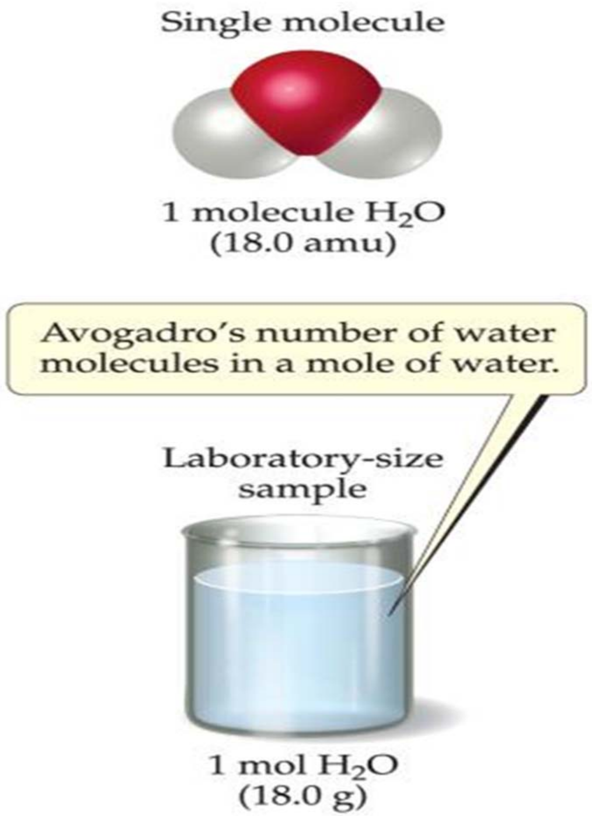
- The mass % of H in methane (C_2H_6) is _________?

Quiz

- Calculate the percentage by mass of Pb, N, O in $Pb(NO_3)_2$

Avogadro's Number and the Mole

- 6.02×10^{23} atoms or molecules is ONE Mole.
- Avogadro's number: 6.02×10^{23}
- One mole of 12C has a mass of 12.000 g.



Converting Moles to Number of Atoms

- Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

Sol:

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

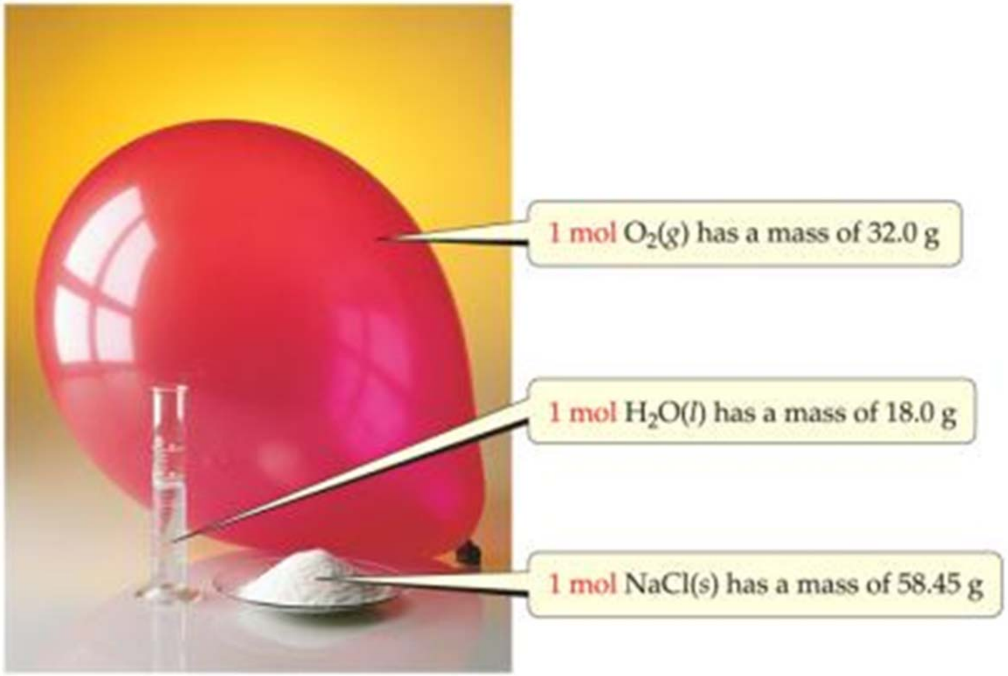
$$= 2.53 \times 10^{24}\ \text{H atoms}$$

Quiz

- Calculate the number of H atoms in 0.50 mol of $(NH_4)_2SO_4$.

Molar Mass

- A molar mass is the mass of 1 mol of a substance (i.e., g/mol).



Mole Relationships

Table 3.2 Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen	N_2	28.0	28.0	6.02×10^{23} N_2 molecules $2(6.02 \times 10^{23})$ N atoms
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	6.02×10^{23} $BaCl_2$ formula units 6.02×10^{23} Ba^{2+} ions $2(6.02 \times 10^{23})$ Cl^- ions

[a]Recall that the mass of an electron is more than 1800 times smaller than the masses of the proton and the neutron; thus, ions and atoms have essentially the same mass.

Calculating Molar Mass

- What is the molar mass of glucose, $C_6H_{12}O_6$?

Sol:

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

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

Converting Grams to Moles

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

Sol:

Using 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$ to write the appropriate conversion factor, we have

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

- Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

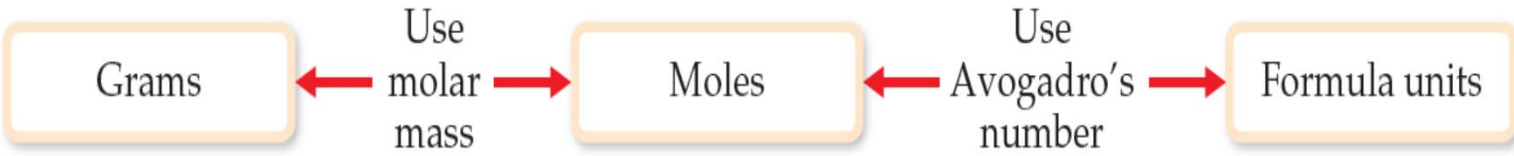
Sol:

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

Quiz

- How many moles of carbon dioxide (CO_2) are there in 52.06 g of carbon dioxide?

Interconverting Masses and Numbers of Particles



▲ Figure 3.12 **Procedure for interconverting mass and number of formula units.** The number of moles of the substance is central to the calculation. Thus, the mole concept can be thought of as the bridge between the mass of a sample in grams and the number of formula units contained in the sample.

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

$$\text{Molecules } C_6H_{12}O_6 = (5.23\ \cancel{g\ C_6H_{12}O_6})\left(\frac{1\ \cancel{mol\ C_6H_{12}O_6}}{180.0\ \cancel{g\ C_6H_{12}O_6}}\right)\left(\frac{6.02\times10^{23}\text{ molecules } C_6H_{12}O_6}{1\ \cancel{mol\ C_6H_{12}O_6}}\right)$$

$$= 1.75\times10^{22}\text{ molecules } C_6H_{12}O_6$$

$$\text{Atoms O} = (1.75\times10^{22}\ \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6\text{ atoms O}}{\cancel{\text{molecule } C_6H_{12}O_6}}\right)$$

$$= 1.05\times10^{23}\text{ atoms O}$$

Question

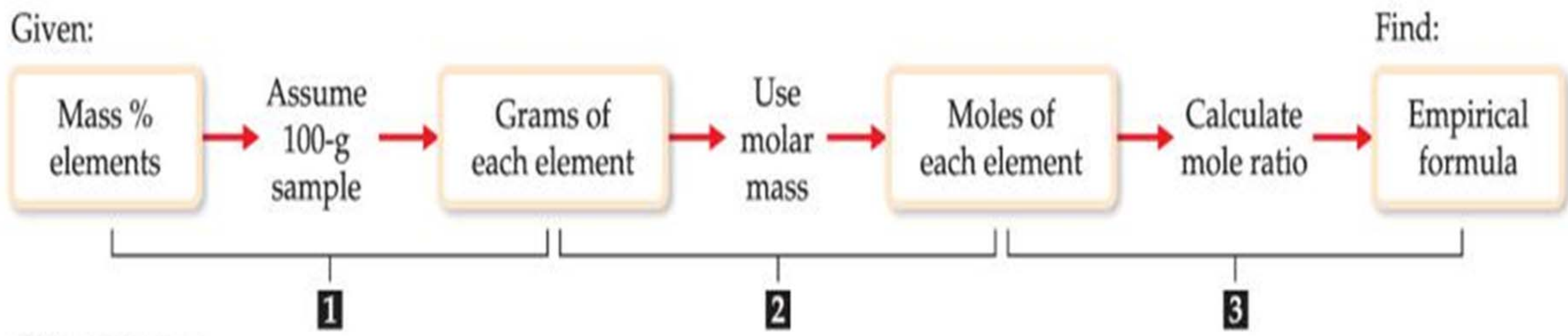
- An analytical balance can detect a mass of 0.1 mg. How many ions are present in this minimally detectable quantity of $MgCl_2$?

Solve:

$$? \text{ ions} = 0.1 \text{ mg } MgCl_2 \times \frac{1 \text{ g } MgCl_2}{1000 \text{ mg } MgCl_2} \times \frac{1 \text{ mol } MgCl_2}{95 \text{ g } MgCl_2} \times \frac{6.0 \times 10^{23} \text{ fu } MgCl_2}{1 \text{ mol } MgCl_2} \times \frac{3 \text{ ions}}{1 \text{ fu } MgCl_2} = 2 \times 10^{18} \text{ ions}$$

Empirical Formulas from Analyses

- One can determine the empirical formula from the percent composition by following these three steps.



Calculating an Empirical Formula

- 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?

Sol:

In 100.00 g of ascorbic acid we have 40.92 g C, 4.58 g H, and 54.50 g O.

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

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

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

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

$$C:H:O = (3 \times 1 : 3 \times 1.33 : 3 \times 1) = (3:4:3)$$

Thus, the empirical formula is $C_3H_4O_3$.

Quiz

- What is the empirical formula of a compound that contains 27.0% S, 13.4% O, and 59.6% Cl by mass?

Molecular Formulas from Empirical Formulas

- The number of atoms in a molecular formula is a multiple of the number of atoms in an empirical formula.

Whole-number multiple
$= molecular\ weight / empirical\ formula\ weight$

EX: The empirical formula of a compound was found to be CH. It has a molar mass of 78 g/mol. What is its molecular formula?

Solution:

- Whole-number multiple = 78/13 = 6

The molecular formula is C_6H_6.

Question

- Dibutyl succinate is an insect repellent used against household ants and roaches. Its composition is 62.58% C, 9.63% H, and 27.79% O. Its experimentally determined molecular mass is 230 u. What are the empirical and molecular formulas of dibutyl succinate?

Solve

1. Determine the mass of each element in a 100.00 g sample.	62.58 g C, 9.63 g H, 27.79 g O
2. Convert each of these masses to an amount in moles.	$? \text{ mol C} = 62.58 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 5.210 \text{ mol C}$
	$? \text{ mol H} = 9.63 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.55 \text{ mol H}$
	$? \text{ mol O} = 27.79 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.737 \text{ mol O}$
3. Write a tentative formula based on these numbers of moles.	$C_{5.21}H_{9.55}O_{1.74}$

4. Divide each of the subscripts of the tentative formula by the smallest subscript (1.74),

$$C_{\frac{5.21}{1.74}}\,H_{\frac{9.55}{1.74}}\,O_{\frac{1.74}{1.74}} = C_{2.99}\,H_{5.49}\,O$$

and round off any subscripts that differ only slightly from whole numbers; that is, round 2.99 to 3.

$$C_3H_{5.49}O$$

5. Multiply all subscripts by a small whole number to make them integral (here by the factor 2), and write the empirical formula.

$$C_{2\times 3}\,H_{2\times 5.49}\,O_{2\times 1} = C_6\,H_{10.98}\,O_2$$

$$2 \times 5.49 = 10.98 \approx 11$$

Empirical formula: $C_6H_{11}O_2$

To establish the molecular formula, first determine the empirical formula mass.

$$[(6 \times 12.0) + (11 \times 1.0) + (2 \times 16.0)]u = 115\ u$$

Since the experimentally determined formula mass (230 u) is twice the empirical formula mass, the molecular formula is twice the empirical formula.

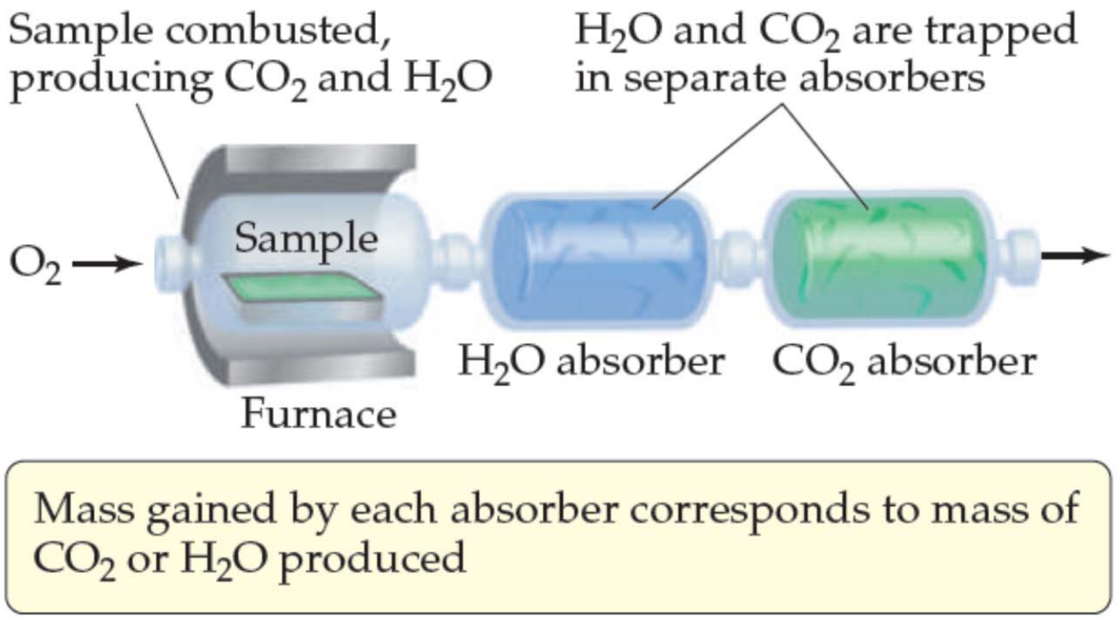
Molecular formula: $C_{12}H_{22}O_4$

Quiz

- A compound contains 40.0% C, 6.71% H, and 53.29% O by mass. The molecular weight of the compound is 60.05 amu. The molecular formula of this compound is ________.

Combustion Analysis

- One technique for determining empirical formulas in the laboratory is combustion analysis, commonly used for compounds containing principally carbon and hydrogen



▲ Figure 3.14 Apparatus for combustion analysis.

Determining an Empirical Formula by Combustion Analysis

- Isopropyl alcohol, sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

Sol:

- all of the carbon in the sample is converted to CO_2

Mass CO_2 produced $\xrightarrow{\text{Molar mass } CO_2 \text{ 44.0 g/mol}}$ Moles CO_2 produced $\xrightarrow{\text{1 C atom per } CO_2 \text{ molecule}}$ Moles of C original sample $\xrightarrow{\text{Molar mass C 12.0 g/mol}}$ Mass C in original sample

$$\text{Grams C} = (0.561\ \cancel{\text{g CO}_2})\left(\frac{1\ \cancel{\text{mol CO}_2}}{44.0\ \cancel{\text{g CO}_2}}\right)\left(\frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol CO}_2}}\right)\left(\frac{12.0\ \text{g C}}{1\ \cancel{\text{mol C}}}\right)$$

$$= 0.153\ \text{g C}$$

Determining an Empirical Formula by Combustion Analysis

- all of the hydrogen in the sample is converted to H_2O

Mass H_2O produced $\xrightarrow{\text{Molar mass } H_2O \text{ 18.0 g/mol}}$ Moles H_2O produced $\xrightarrow{\text{2 H atoms per } H_2O \text{ molecule}}$ Moles H in original sample $\xrightarrow{\text{Molar mass H 1.01 g/mol}}$ Mass H in original sample

$$\text{Grams H} = (0.306\ \cancel{\text{g } H_2O})\left(\frac{1\ \cancel{\text{mol } H_2O}}{18.0\ \cancel{\text{g } H_2O}}\right)\left(\frac{2\ \cancel{\text{mol}}\ \text{H}}{1\ \cancel{\text{mol } H_2O}}\right)\left(\frac{1.01\ \text{g H}}{1\ \cancel{\text{mol H}}}\right) = 0.0343\ \text{g H}$$

The mass of the sample, 0.255 g, is the sum of the masses of C, H, and O.

Mass of O = mass of sample – (mass of C + mass of H)

= 0.255 g – (0.153 g + 0.0343 g) = 0.068 g O

Moles C = (0.153 g)(1 mol C/12.0 g C) = 0.0128 mol C
Moles H = (0.0343 g)(1 mol H/1.01 g H)=0.0340 mol H
Moles O = (0.068 g)(1 mol O/16.0 g O)= 0.0043 mol O

$$\text{C:}\frac{0.0128}{0.0043} = 3.0 \quad \text{H:}\frac{0.0340}{0.0043} = 7.9 \quad \text{O:}\frac{0.0043}{0.0043} = 1.0$$

The first two numbers are very close to the whole numbers 3 and 8, giving the empirical Formula C_3H_8O.

Question

- Vitamin C is essential for the prevention of scurvy. Combustion of a 0.2000 g sample of this carbon– hydrogen–oxygen compound yields 0.2998 g CO_2 and 0.0819 g H_2O. What are the percent composition and the empirical formula of vitamin C?

Sol

Percent Composition

First, determine the mass of carbon in 0.2988 g CO_2, by converting to mol C,

$$? \text{ mol C} = 0.2998 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.010 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.006812 \text{ mol C}$$

and then to g C.

$$? \text{ g C} = 0.006812 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.08182 \text{ g C}$$

Proceed in a similar fashion for 0.0819 g H_2O to obtain

$$? \text{ mol H} = 0.0819 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.00909 \text{ mol H}$$

and

$$? \text{ g H} = 0.00909 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.00916 \text{ g H}$$

Obtain the mass of O in the 0.2000 g sample as the difference

$$? \text{ g O} = 0.2000 \text{ g sample} - 0.08182 \text{ g C} - 0.00916 \text{ g H} = 0.1090 \text{ g O}$$

Finally, multiply the mass fractions of the three elements by 100% to obtain mass percentage

$$\% \text{C} = \frac{0.08182 \text{ g C}}{0.2000 \text{ g sample}} \times 100\% = 40.91\% \text{ C}$$

$$\% \text{H} = \frac{0.00916 \text{ g H}}{0.2000 \text{ g sample}} \times 100\% = 4.58\% \text{ H}$$

$$\% \text{O} = \frac{0.1090 \text{ g O}}{0.2000 \text{ g sample}} \times 100\% = 54.50\% \text{ O}$$

$$? \text{ mol C} = 0.2998 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.010 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.006812 \text{ mol C}$$

$$? \text{ mol H} = 0.0819 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.00909 \text{ mol H}$$

$$? \text{ mol O} = 0.1090 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.006813 \text{ mol O}$$

From the numbers of moles of C, H, and O in the 0.2000 g sample, we obtain the tentative empirical formula

$$C_{0.006812}H_{0.00909}O_{0.006813}$$

Empirical formula of vitamin C: $C_3H_4O_3$

Quiz

- A compound is composed of only C, H, and O. The combustion of a 0.519-g sample of the compound yields 1.24 g of CO_2 and 0.255 g of H_2O. What is the empirical formula of the compound?

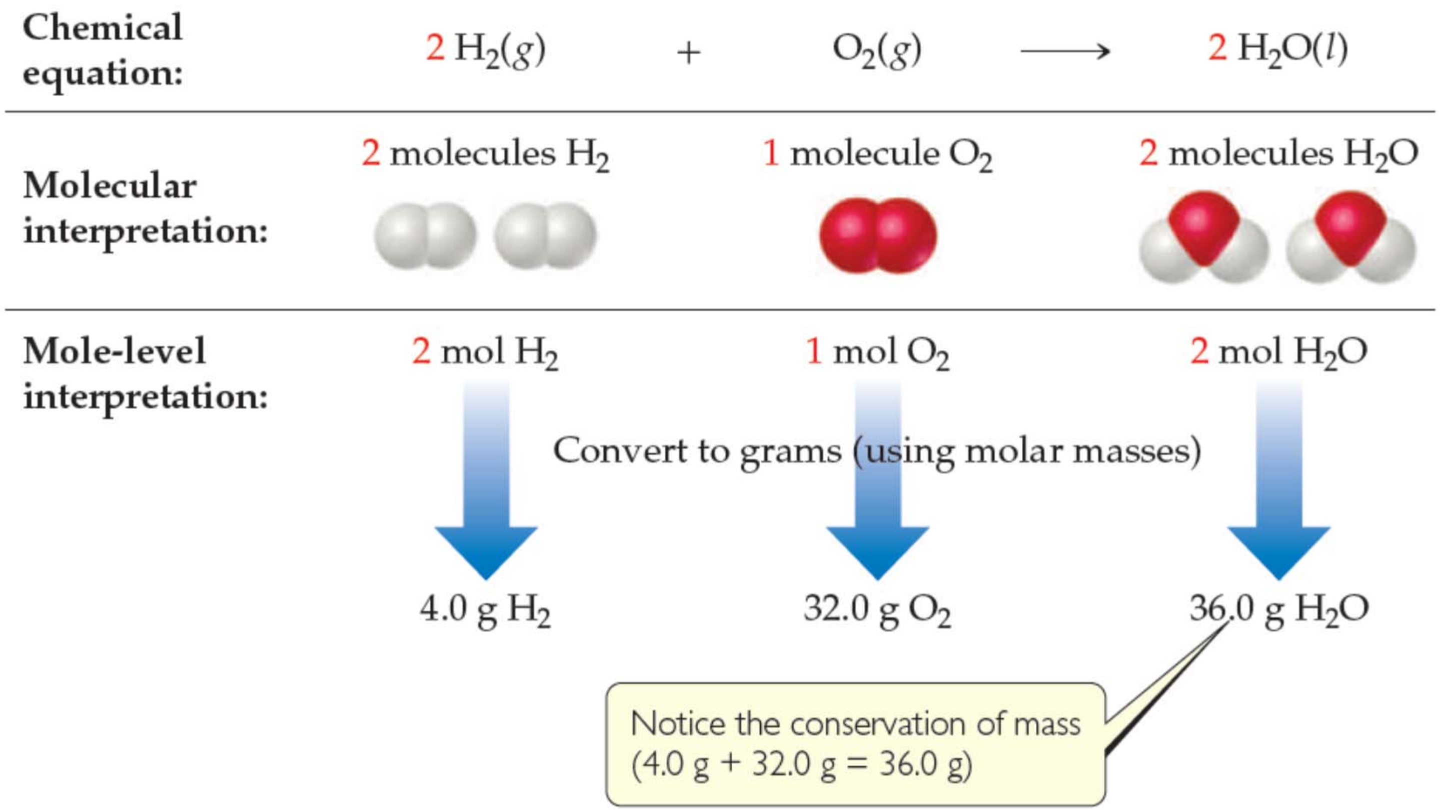
Quantitative Information from Balanced Equations

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

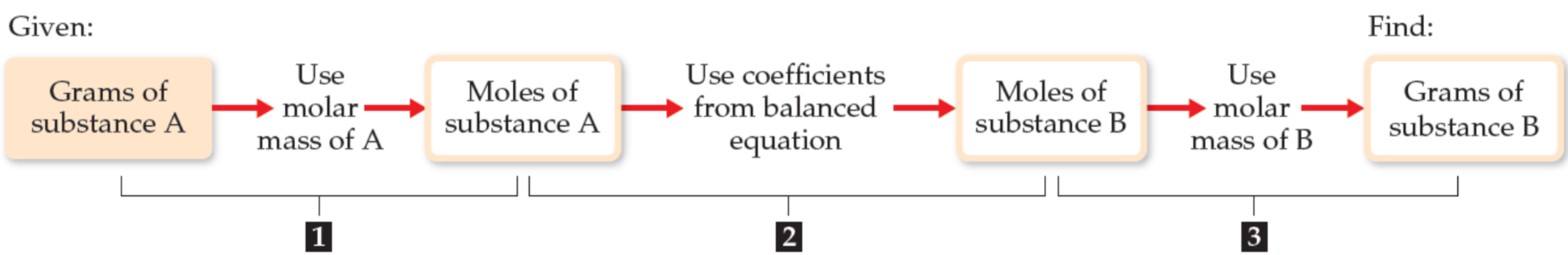
$2\,H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\,H_2O(l)$
2 molecules		1 molecule		2 molecules
$2(6.02 \times 10^{23}$ molecules)		$1(6.02 \times 10^{23}$ molecules)		$2(6.02 \times 10^{23}$ molecules)
2 mol		1 mol		2 mol

The coefficients in a balanced chemical equation indicate both the relative numbers of molecules (or formula units) in the reaction and the relative numbers of moles.
The coefficients are called stoichiometrically equivalent quantities.

Balanced chemical equation quantitatively.

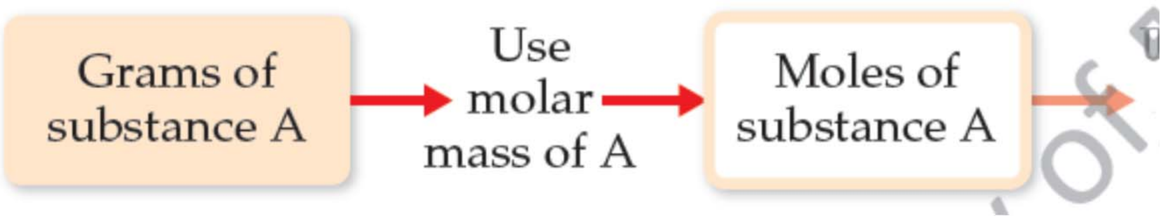


Balanced chemical equation quantitatively.



Ex: Determine how many grams of water are produced in the oxidation of 1.00 g of glucose, $C_6H_{12}O_6$

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(l)$$



Moles $C_6H_{12}O_6$ = (1.00 g $C_6H_{12}O_6$)(1 mol $C_6H_{12}O_6$/ 180.0 g $C_6H_{12}O_6$)

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(l)$$

Moles of substance A → Use coefficients from balanced equation → Moles of substance B

$$\text{Moles } H_2O = (1.00\text{ g } \cancel{C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6\text{ mol } H_2O}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

Moles of substance B → Use molar mass of B → Grams of substance B

$$\text{Grams } H_2O = (1.00\text{ g } \cancel{C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\text{ g } \cancel{C_6H_{12}O_6}}\right)\left(\frac{6\ \cancel{\text{mol } H_2O}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{18.0\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)$$

$$= 0.600\text{ g } H_2O$$

Quiz

The combustion of ammonia in the presence of excess oxygen yields NO_2 and H_2O:

$$4\,NH_3\,(g) + 7\,O_2\,(g) \rightarrow 4\,NO_2\,(g) + 6\,H_2O\,(g)$$

The combustion of 43.9 g of ammonia produces _________ g of NO_2.

Limiting Reactants

- The reactant that is completely consumed in a reaction is called the limiting reactant. The other reactants are sometimes called excess reactants

	$2\ H_2(g)$	$+$	$O_2(g)$	$\longrightarrow$	$2\ H_2O(g)$
Before reaction:	10 mol		7 mol		0 mol
Change (reaction):	−10 mol		−5 mol		+10 mol
After reaction:	0 mol		2 mol		10 mol

Sample Exercise 3.18

The most important commercial process for converting N_2 from the air into nitrogen-containing compounds is based on the reaction of N_2 and H_2 to form ammonia (NH_3):

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

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

Sol:

The number of moles of H_2 needed for complete consumption of 3.0 mol of N_2 is

$$\text{Moles } H_2 = (3.0\ \cancel{\text{mol } N_2})\left(\frac{3\ \text{mol } H_2}{1\ \cancel{\text{mol } N_2}}\right) = 9.0\ \text{mol } H_2$$

Because only 6.0 mol H_2 is available, we will run out of H_2 before the N_2 is gone, which tells us that H_2 is the limiting reactant. Therefore, we use the quantity of H_2 to calculate the quantity of NH_3 produced:

$$\text{Moles } NH_3 = (6.0\ \cancel{\text{mol } H_2})\left(\frac{2\ \text{mol } NH_3}{3\ \cancel{\text{mol } H_2}}\right) = 4.0\ \text{mol } NH_3$$

Theoretical and Percent Yields

- The quantity of product calculated to form when all of a limiting reactant is consumed is called the theoretical yield.
- The amount of product actually obtained, called the actual yield
- The percent yield of a reaction relates actual and theoretical yields

Percent yield = (actual yield/theoretical yield) * 100%

Calculating Theoretical Yield and Percent Yield

Adipic acid, $H_2C_6H_8O_4$, used to produce nylon, is made commercially by a reaction between cyclohexane (C_6H_{12}) and O_2:

$$2\,C_6H_{12}(l) + 5\,O_2(g) \longrightarrow 2\,H_2C_6H_8O_4(l) + 2\,H_2O(g)$$

(**a**) Assume that you carry out this reaction with 25.0 g of cyclohexane and that cyclohexane is the limiting reactant. What is the theoretical yield of adipic acid? (**b**) If you obtain 33.5 g of adipic acid, what is the percent yield for the reaction?

Sol:

(**a**) The theoretical yield is

$$\text{Grams } H_2C_6H_8O_4 = (25.0\ \cancel{\text{g } C_6H_{12}})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}}}{84.0\ \cancel{\text{g } C_6H_{12}}}\right)\left(\frac{2\ \cancel{\text{mol } H_2C_6H_8O_4}}{2\ \cancel{\text{mol } C_6H_{12}}}\right)\left(\frac{146.0\text{ g } H_2C_6H_8O_4}{1\ \cancel{\text{mol } H_2C_6H_8O_4}}\right)$$

$$= 43.5\text{ g } H_2C_6H_8O_4$$

(**b**) $$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{33.5\text{ g}}{43.5\text{ g}} \times 100\% = 77.0\%$$