Chapter 3

Chemical Reactions and Reaction Stoichiometry

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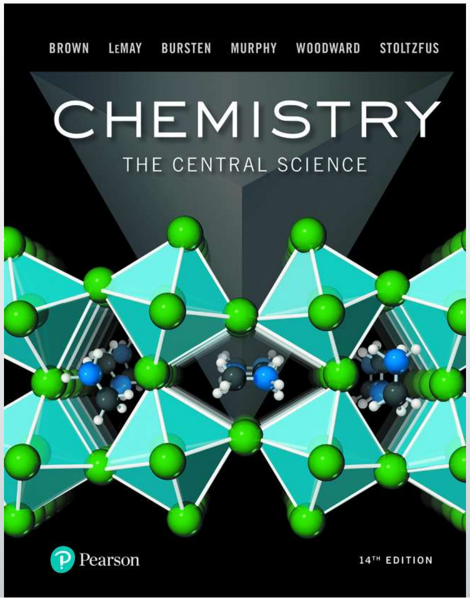


3.1 Chemical Equations

- We represent chemical reaction by **chemical equations**.

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

- We read the + sign as **reacts with** and the arrow as **produces**.
- The chemical formulas to the left of the arrow represent the starting substances, called reactants.
- The chemical formulas to the right of the arrow represent substances produced in the reaction, called products.
- The numbers in front of the formulas, called coefficients, indicate the relative numbers of molecules of each kind involved in the reaction.

3.1 Chemical Equations

Balancing Equations

- Because atoms are neither created nor destroyed in any reaction, a balanced chemical equation must have an equal number of atoms of each element on each side of the arrow.
- We balance the equation by determining the coefficients that provide equal numbers of each type of atom on both sides of the equation.
- For most purposes, a balanced equation should contain the smallest possible whole-number coefficients.

3.1 Chemical Equations

A Step-by-Step Example of Balancing a Chemical Equation

- Consider the reaction that occurs when methane (CH_4) burns in air to produce carbon dioxide gas (CO_2) and water vapor (H_2O).
- It is usually best to balance first those elements that occur in the fewest chemical formulas in the equation.

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O \quad \text{(unbalanced)}$$

$$CH_4 + O_2 \longrightarrow CO_2 + 2\,H_2O \quad \text{(unbalanced)}$$

$$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O \quad \text{(balanced)}$$

3.1 Chemical Equations

Indicating the States of Reactants and Products

- We use the symbols (g), (l), (s), and (aq) for substances that are gases, liquids, solids, and dissolved in aqueous solution, respectively.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

- A delta (Δ) above the reaction arrow indicates the addition of heat.

Sample Exercise 3.2

Balance the equation

$$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

$$Na(s) + 2\,H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

$$Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

$$2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

3.2 Simple Patterns of Chemical Reactivity

Combination and decomposition Reactions

- In combination reactions, two or more substances react to form one product.
- In a decomposition reaction one substance undergoes a reaction to produce two or more other substances.

Combination Reactions
$A + B \longrightarrow C$
$C(s) + O_2(g) \longrightarrow CO_2(g)$
$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)$
$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$
Decomposition Reactions
$C \longrightarrow A + B$
$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$
$PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$
$Cu(OH)_2(s) \longrightarrow CuO(s) + H_2O(g)$

Sample Exercise 3.3

Write a balanced equation for

a. the combination reaction between lithium metal and fluorine gas

$$2\,Li(s) + F_2(g) \longrightarrow 2\,LiF(s)$$

b. the decomposition reaction that occurs when solid barium carbonate is heated (two products form, a solid and a gas).

$$BaCO_3(s) \longrightarrow BaO(s) + CO_2(g)$$

Combustion Reactions

- Combustion reactions are rapid reactions that produce a flame.
- Hydrocarbons combusted in air react with O_2 to form CO_2 and H_2O.

$$C_3H_8(g) + 5\,O_2(g) \longrightarrow 3\,CO_2(g) + 4\,H_2O(g)$$

Sample Exercise 3.4

Write the balanced equation for the reaction that occurs when methanol, $CH_3OH(l)$, is burned in air.

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

$$CH_3OH(l) + \tfrac{3}{2}\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

$$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

3.3 Formula Weights

Formula and Molecular Weights

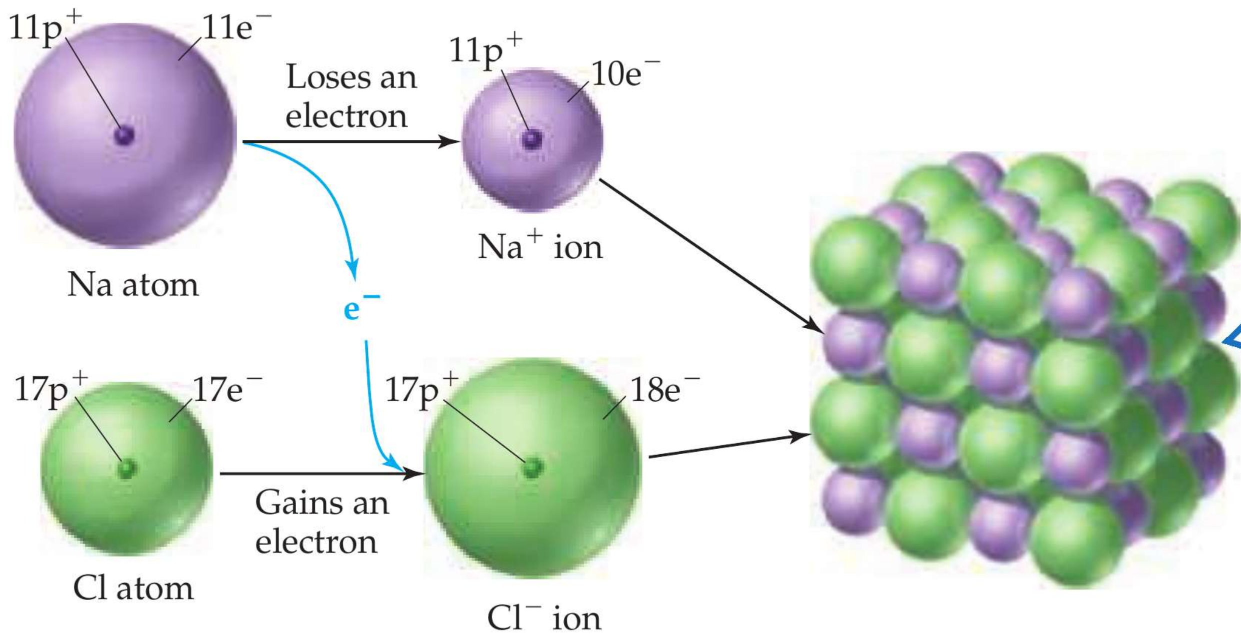
- The formula weight (*FW*) of a substance is the sum of the **atomic weights** (***AW***) of the atoms in the chemical formula of the substance.

$$\begin{aligned} FW \text{ of } \mathrm{H_2SO_4} &= 2(AW \text{ of H}) + AW \text{ of S} + 4(AW \text{ of O}) \\ &= 2(1.01) + 32.06 + 4(16.00) \\ &= 98.01 \text{ amu} \end{aligned}$$

- If the chemical formula is that of a molecule, the formula weight is also called the molecular weight (*MW*).

3.3 Formula Weights

- Because ionic substances exist as three-dimensional arrays of ions, it is inappropriate to speak of molecules of these substances.
- Instead we use the empirical formula as the formula unit.



Sample Exercise 3.5

Calculate the formula weight of

a. sucrose, $C_{12}H_{22}O_{11}$ (table sugar)

$$FW = 12(AW \text{ of C}) + 22(AW \text{ of H}) + 11(AW \text{ of O})$$

$$= 12(12.01) + 22(1.01) + 11(16.00) = 342.34 \text{ amu}$$

b. calcium nitrate, $Ca(NO_3)_2$.

$$FW = AW \text{ of Ca} + 2[AW \text{ of N} + 3(AW \text{ of O})]$$

$$= 40.08 + 2[14.01 + 3(16.00)] = 164.10 \text{ amu}$$

3.3 Formula Weights

Percentage Composition from Chemical Formulas

- The percentage composition (sometimes called the **elemental composition** of a substance) of a compound is the percentage by mass contributed by each element in the substance.

$$\begin{matrix}\text{\% mass composition}\\ \text{of element}\end{matrix} = \frac{\begin{pmatrix}\text{Number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}AW\\ \text{of element}\end{pmatrix}}{FW\ \text{of substance}} \times 100\%$$

Sample Exercise 3.6

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

$$\begin{matrix}\%\text{ mass composition}\\ \text{of element}\end{matrix} = \frac{\begin{pmatrix}\text{Number of atoms}\\ \text{of element}\end{pmatrix}\begin{pmatrix}AW\\ \text{of element}\end{pmatrix}}{FW \text{ of substance}} \times 100\%$$

$$\%\ \mathrm{C} = \frac{(12)(12.01)}{342.34} \times 100\% = 42.10\%$$

$$\%\ \mathrm{H} = \frac{(22)(1.01)}{342.34} \times 100\% = 6.49\%$$

$$\%\ \mathrm{O} = \frac{(11)(16.00)}{342.34} \times 100\% = 51.41\%$$

3.4 Avogadro's Number and the Mole

- One mole (n) is the amount of matter that contains as many objects (atoms, molecules, or whatever other objects we are considering) as the number of atoms in exactly 12 g of isotopically pure ^{12}C.
- This number is determined to be 6.0221415×10^{23}, which is called Avogadro's number, N_A.

Sample Exercise 3.8

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

$$\text{No. of atoms} = 0.350\ \text{mol}\ C_6H_{12}O_6 \left(\frac{6.022 \times 10^{23}\ \text{molecules}\ C_6H_{12}O_6}{1\ \text{mol}\ C_6H_{12}O_6}\right) \times \left(\frac{12\ \text{atoms H}}{1\ \text{molecule}\ C_6H_{12}O_6}\right)$$

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

3.4 Avogadro's Number and the Mole

Molar Mass

- The atomic weight of an element in atomic mass units is numerically equal to the mass in grams of 1 mol of that element.
- The mass in grams of one mole of a substance is called the molar mass (*MM*) of the substance.
- The molar mass in grams per mole of any substance is numerically equal to its formula weight in atomic mass units.

Interconverting Masses and Moles

Interconverting Masses and Numbers of Particles

Sample Exercise 3.9

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

$$MM = 6(AW \text{ of C}) + 12(AW \text{ of H}) + 6(AW \text{ of O})$$

$$= 6(12.01) + 12(1.01) + 6(16.00) = 180.18 \text{ amu}$$

Sample Exercise 3.10

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

$$n = 5.380\ \text{g}\ C_6H_{12}O_6 \left(\frac{1\ \text{mol}\ C_6H_{12}O_6}{180.18\ \text{g}\ C_6H_{12}O_6}\right)$$

$$= 0.02986\ \text{mol}\ C_6H_{12}O_6$$

Sample Exercise 3.11

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

$$m = 0.433\ \mathrm{mol\ Ca(NO_3)_2}\left(\frac{164.10\ \mathrm{g\ Ca(NO_3)_2}}{1\ \mathrm{mol\ Ca(NO_3)_2}}\right)$$

$$= 71.1\ \mathrm{g\ Ca(NO_3)_2}$$

Sample Exercise 3.12

a. How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$?

$$\text{No. of molecules} = 5.23 \text{ g } C_6H_{12}O_6 \left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.18 \text{ g } C_6H_{12}O_6}\right)\left(\frac{6.022 \times 10^{23} \text{ molecules } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6}\right)$$

$$= 1.75 \times 10^{22} \text{ molecule } C_6H_{12}O_6$$

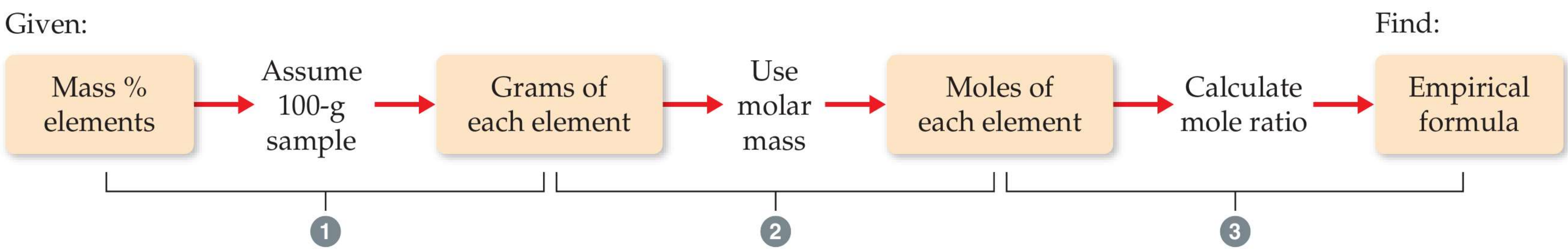
b. How many oxygen atoms are in this sample?

$$\text{No. of atoms} = 1.75 \times 10^{22} \text{ molecule } C_6H_{12}O_6 \left(\frac{6 \text{ atoms O}}{1 \text{ molecule } C_6H_{12}O_6}\right)$$

$$= 1.05 \times 10^{23} \text{ atoms O}$$

3.5 Empirical Formulas from Analyses

- The ratio of the numbers of moles of all elements in a compound gives the subscripts in the compound's empirical formula.



Sample Exercise 3.13

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?

- For simplicity we assume we have exactly 100 g of material; thus, we have 40.92 g C, 4.58 g H, and 54.50 g O.

$$n\ \mathrm{C} = 40.92\ \mathrm{g\ C}\left(\frac{1\ \mathrm{mol\ C}}{12.01\ \mathrm{g\ C}}\right) = 3.407\ \mathrm{mol\ C}$$

$$n\ \mathrm{H} = 4.58\ \mathrm{g\ H}\left(\frac{1\ \mathrm{mol\ H}}{1.01\ \mathrm{g\ H}}\right) = 4.535\ \mathrm{mol\ H}$$

Sample Exercise 3.13

$$n\ \mathrm{O} = 54.50\ \mathrm{g\ O}\left(\frac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g\ O}}\right) = 3.406\ \mathrm{mol\ O}$$

$$\mathrm{C}_{\frac{3.407}{3.406}}\mathrm{H}_{\frac{4.535}{3.406}}\mathrm{O}_{\frac{3.406}{3.406}}$$

$$\mathrm{C}_{1.000}\mathrm{H}_{1.331}\mathrm{O}_{1.000}$$

$$\mathrm{C_3H_4O_3}$$

3.5 Empirical Formulas from Analyses

Molecular Formulas from Empirical Formulas

- The subscripts in the molecular formula of a substance are always whole-number multiples (N) of the subscripts in its empirical formula.

$$N = \frac{\text{molecular weight } (MW)}{\text{empirical formula weight}}$$

Sample Exercise 3.14

Mesitylene, a hydrocarbon found in crude oil, has an empirical formula of C_3H_4 and an experimentally determined molecular weight of 121 amu. What is its molecular formula?

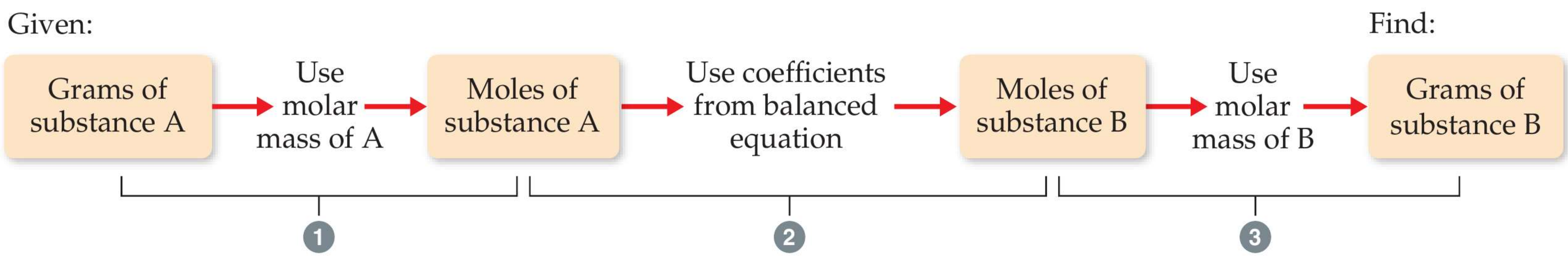
$$N = \frac{\text{molecular weight } (MW)}{\text{empirical formula weight}}$$

$$= \frac{121}{3(12.01) + 4(1.01)} = 3.02$$

C_9H_{12}

3.6 Quantitative Information from Balanced Equations

- Stoichiometry is the area of study that examines the quantities of substances consumed and produced in chemical reactions.
- 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.



Sample Exercise 3.16

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

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

$$m\ H_2O = 1.00\ g\ C_6H_{12}O_6\left(\frac{1\ mol\ C_6H_{12}O_6}{180.18\ g\ C_6H_{12}O_6}\right)\left(\frac{6\ mol\ H_2O}{1\ mol\ C_6H_{12}O_6}\right)\left(\frac{18.02\ g\ H_2O}{1\ mol\ H_2O}\right)$$

$$= 0.600\ g\ H_2O$$

Sample Exercise 3.17

Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide gas exhaled by astronauts. The hydroxide reacts with the carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

$$2\,\mathrm{LiOH}(s) + \mathrm{CO_2}(g) \longrightarrow \mathrm{Li_2CO_3}(s) + \mathrm{H_2O}(l)$$

$$m\ \mathrm{CO_2} = 1.00\ \mathrm{g\ LiOH}\left(\frac{1\ \mathrm{mol\ LiOH}}{23.95\ \mathrm{g\ LiOH}}\right)\left(\frac{1\ \mathrm{mol\ CO_2}}{2\ \mathrm{mol\ LiOH}}\right)\left(\frac{44.01\ \mathrm{g\ CO_2}}{1\ \mathrm{mol\ CO_2}}\right)$$

$$= 0.919\ \mathrm{g\ CO_2}$$