CHEMISTRY

an atoms-focused approach

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Chapter 7

Stoichiometry

Mass Relationships and Chemical Reations

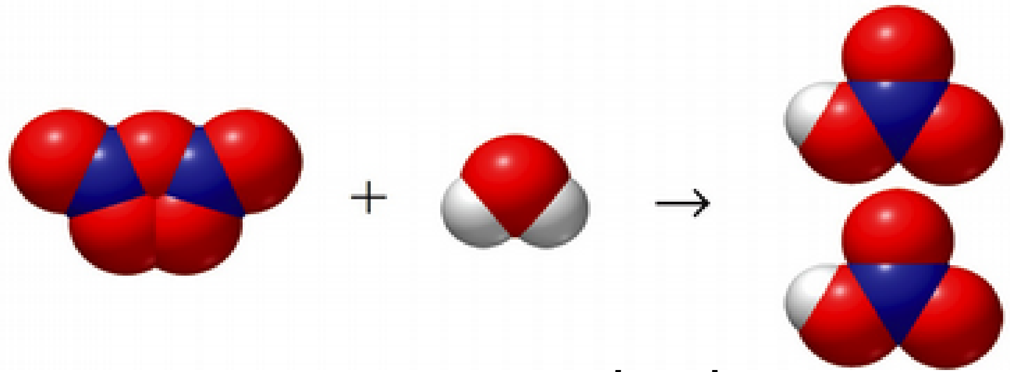
Chapter Outline

- 7.1 Chemical Reactions and the Conservation of Mass
- 7.2 Balancing Chemical Equations
- 7.3 Combustion Reactions
- 7.4 Stoichiometric Calculations and the Carbon Cycle
- 7.5 Percent Composition and Empirical Formulas
- 7.6 Empirical and Molecular Formulas
- 7.7 Combustion Analysis
- 7.8 Limiting Reagents and Percent Yield

Chemical Reactions

Combination Reaction: two or more substances combine to form one product.

$$N_2O_5(g) + H_2O(\ell) \rightarrow 2\,HNO_3(aq)$$



Reactants → Products

Chemical Equation

- Chemical equation:
  - Describes proportions of **reactants (**the substances that are consumed) and **products** (the substances that are formed) during a chemical reaction.
  - Describes the changes on the atomic level.
    - $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$
    - $Fe_2O_3(s) + 3H_2SO_4(aq) \rightarrow 3H_2O(l) + Fe_2(SO_4)_3(aq)$
  - Physical state of reactants and products:
    - (*s*) = solid; (*l*) = liquid; (*g*) = gas; (*aq*) = aqueous soln.

States of Substances

- States are shown by abbreviations in parenthesis after each chemical

  H_2O (s), H_2O (l), H_2O (g)

- Standard phases are:
  - (s) – solid
  - (l) – liquid
  - (g) – gas
  - (aq) – aqueous – dissolved in water
  - (↑) – gas produced from aqueous phase
  - (↓) – solid produced from aqueous phase

Types of Reactions

- Synthesis – compound formed from its base elements:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

- Decomposition – compound decomposes into its base elements:

$$2NH_3 \rightarrow N_2 + 3H_2$$

Types of Reactions

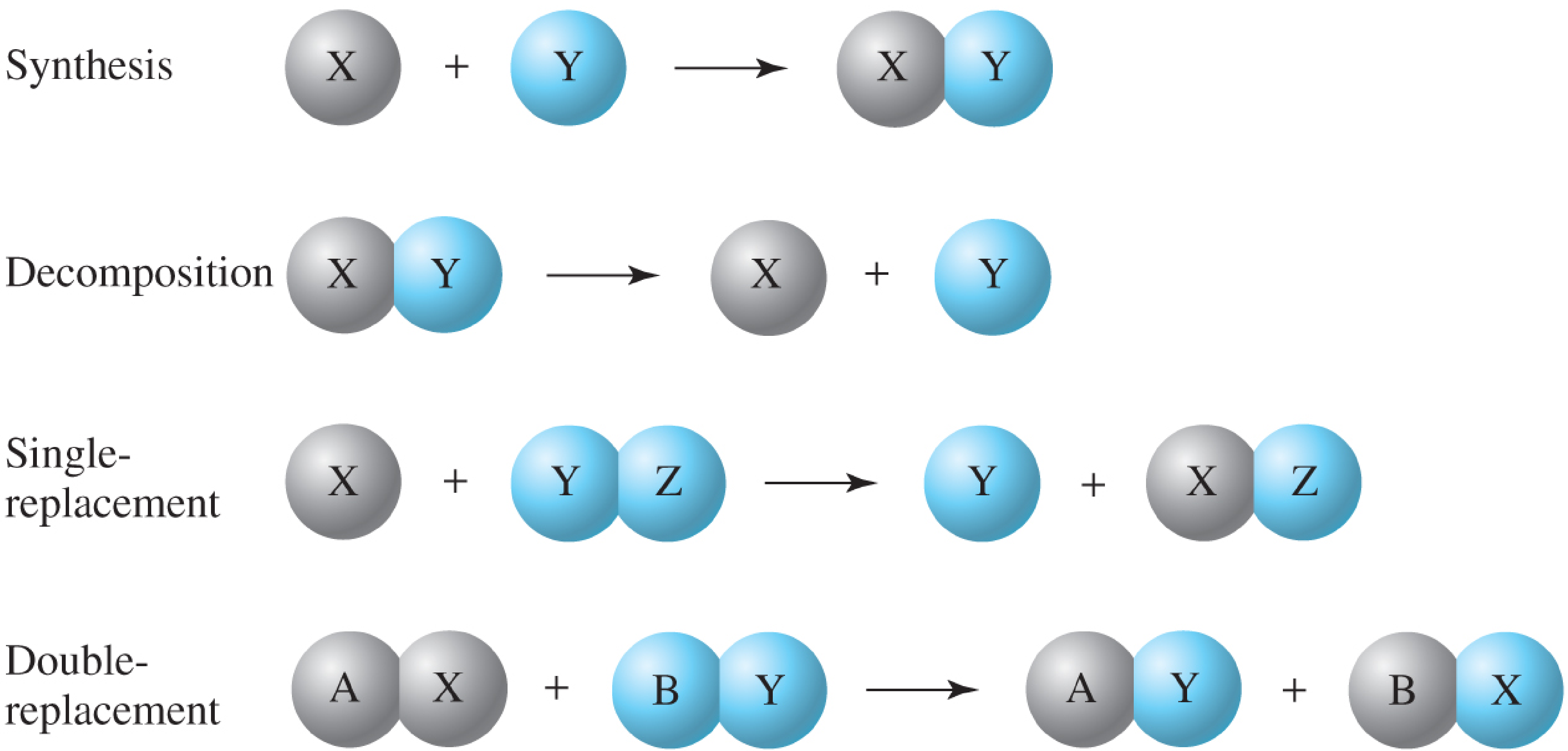
- Single replacement – an element replaces another in a compound:

$$2NaBr + I_2 \rightarrow 2NaI + Br_2$$

- Double replacement – two elements or polyatomic ions in two separate compounds switch places:

$$KNO_3 + Ca(OH)_2 \rightarrow KOH + Ca(NO_3)_2$$

Types of Reactions



Types of Reactions

- Complete combustion – fuel and oxygen produce water and carbon dioxide:

$$CH_4 + O_2 \rightarrow H_2O + CO_2$$

- Incomplete combustion – fuel and oxygen produce water and carbon Monoxide:

$$CH_4 + O_2 \rightarrow H_2O + CO$$

Law of Conservation of Mass

- Law of conservation of mass
  - The sum of the masses of the reactants in a chemical equation is equal to the sum of the masses of the products.
- Stoichiometry
  - Quantitative relation between reactants and products in a chemical equation
  - Indicated in chemical equation by coefficients

Moles and Chemical Equations

$$Fe_2O_3(s) + 3H_2SO_4(aq) \rightarrow 3H_2O(l) + Fe_2(SO_4)_3(aq)$$

- Chemical Equation
  - Indicates substances involved (reactants, products)
  - Coefficients
    - Indicate proportions of reactants and/or products
    - On macroscale, indicate number of moles of each substance.

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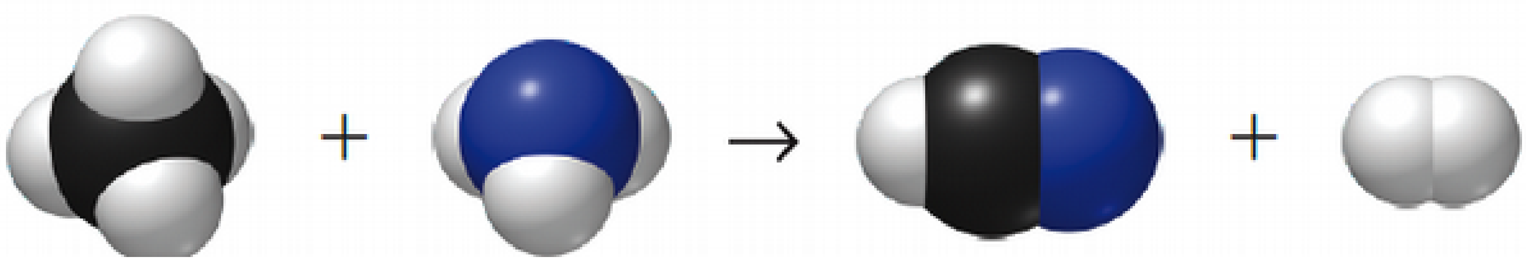
Balancing Chemical Equations

- Balanced chemical equations follow the law of conservation of mass.
  - (not balanced)

$$CH_4(g) + NH_3(g) \rightarrow HCN(g) + H_2(g)$$

$$\text{reactants} \rightarrow \text{products}$$



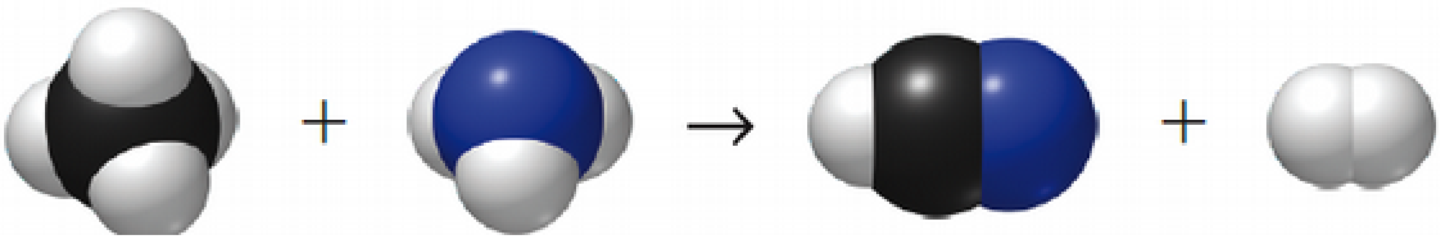
Balancing Chemical Equations

- Three Step Approach:
  - Write correct formulas for reactants and products, including physical states.
  - Balance an element that appears in <u>only one</u> reactant and product first.
  - Choose coefficients to balance other elements as needed.
  - Reduce coefficients to lowest whole numbers.

Balancing Chemical Equations

$$CH_4(g) + NH_3(g) \rightarrow HCN(g) + H_2(g)$$

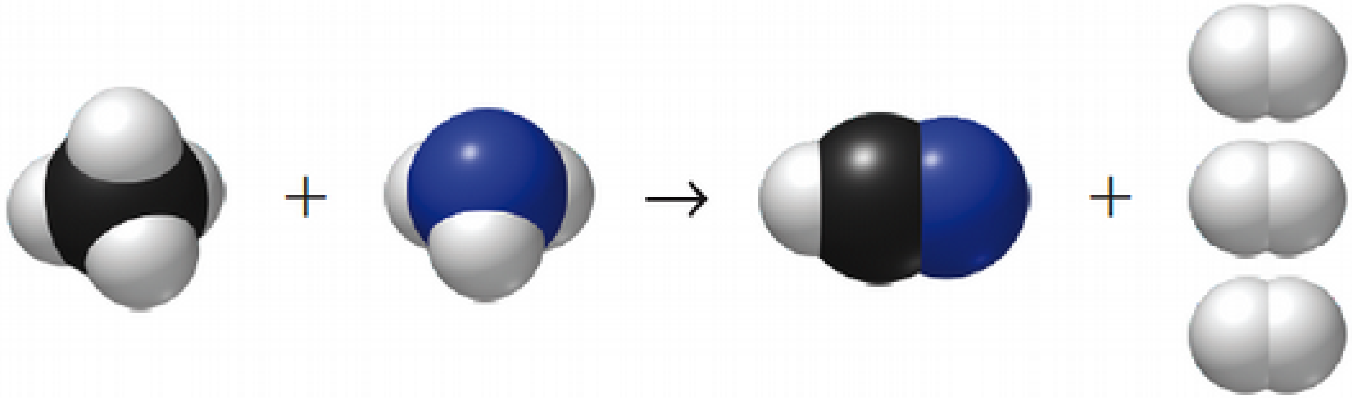
reactants → products



Element	Reactant Side	Product Side	Balanced?
C	1	1	✔
N	1	1	✔
H	4 + 3 = 7	1 + 2 = 3	✗

Balancing Chemical Equations

$$CH_4(g) + NH_3(g) \rightarrow HCN(g) + \underline{3}\ H_2(g)$$



Element	Reactant Side	Product Side	Balanced?
C	1	1	✔
N	1	1	✔
H	$4 + 3 = 7$	$1 + (3 \times 2) = 7$	✔

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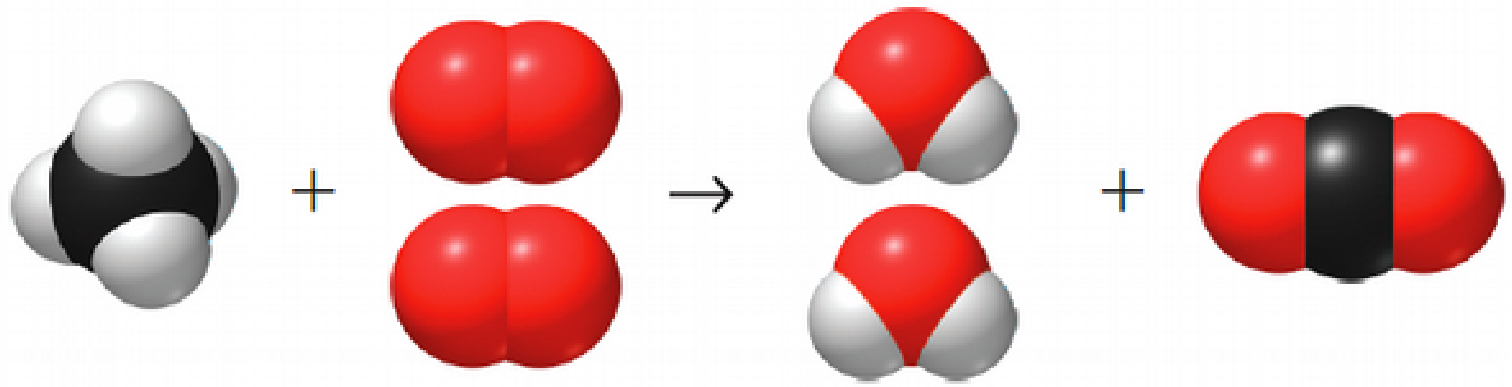
Combustion Reactions

- Hydrocarbons:
  - molecular compounds composed of only hydrogen and carbon.
  - "organic" compounds.
  - Combustion products are CO_2 and H_2O.

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

Combustion Reactions

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



Element	Reactant Side	Product Side	Balanced?
C	1	1	✔
H	4	$2 \times 2 = 4$	✔
O	$2 \times 2 = 4$	$(2 \times 1) + 2 = 4$	✔

Practice: Combustion Reactions

Balance the following combustion reaction.

a) $C_5H_{12} + O_2 \rightarrow CO_2 + H_2O$

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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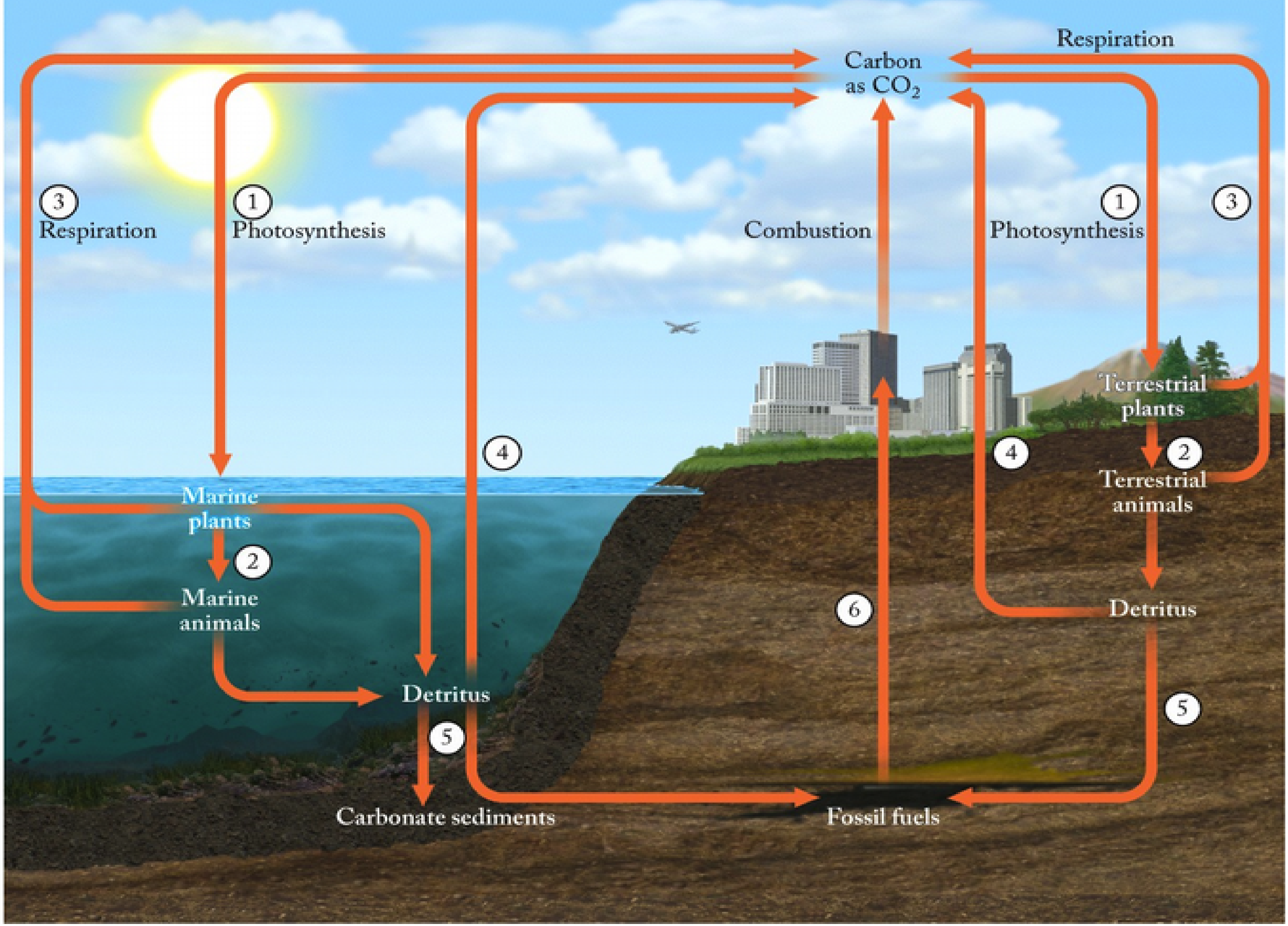
Carbon as CO_2
Respiration
3 Respiration
1 Photosynthesis
Combustion
1 Photosynthesis
3
Terrestrial plants
2
Terrestrial animals
4
4
Marine plants
2
Marine animals
Detritus
6
Detritus
5
5
Carbonate sediments
Fossil fuels

Photosynthesis and Respiration

- Photosynthesis:
  - Plants convert CO_2 and H_2O into glucose:
    - $CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
- Respiration (reverse of photosynthesis):
  - Living organisms use glucose as a source of energy:
    - $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow CO_2(g) + 6H_2O(l)$
- Combustion of Hydrocarbons
  - Returns 6.8×10^{12} kg/yr of C to atmosphere.

Stoichiometric Calculations

- Calculating the mass of a product from the mass of a reactant requires:
  - The **mole ratio** from the balanced chemical equation.
  - Molar mass of the reactant.
  - Molar mass of the product.

Stoichiometry Example

- How much CO_2 enters the atmosphere annually from the combustion of 6.8 x 10^{12} kg of Carbon?
- Balanced Eq'n: C(s) + O_2(g) → CO_2(g)

$$\left(6.8\times10^{12}\ \cancel{\text{kg C}}\right) \longrightarrow \ = ?$$

$$2.5\times10^{13}\ \text{kg } CO_2$$

$$\left(\frac{1000\ \cancel{\text{g}}}{\cancel{\text{kg}}}\right)\left(\frac{1\ \cancel{\text{mole C}}}{12\ \cancel{\text{g}}}\right)\left(\frac{1\ \cancel{\text{mole } CO_2}}{1\ \cancel{\text{mole C}}}\right)\left(\frac{44.0\ \cancel{\text{g } CO_2}}{1\ \cancel{\text{mole } CO_2}}\right)\left(\frac{1\ \text{kg}}{1000\ \cancel{\text{g}}}\right)$$

Practice: Stoichiometry

How much carbon dioxide would be formed if 10.0 grams of C_3H_8 were completely burned in oxygen?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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Percent Composition

- **Percent Composition**:
  - the composition of a compound expressed in terms of the percentage by mass of each element
  - $$\frac{\text{mass of element in compound}}{\text{mass of compound}} \times 100\%$$

Percent Composition

- Example: percent C in CH_4 and $C_{10}H_{16}$:

$$\left(\frac{\text{mass C}}{\text{mass CH}_4}\right)\left(\frac{(12.01\ \text{g/mol C})}{16.04\ \text{g/mol CH}_4}\right)\times 100 = 74.88\%$$

$$\left(\frac{\text{mass C}}{\text{mass C}_{10}\text{H}_{16}}\right)\left(\frac{(12.01\ \text{g/mol C})\times 10}{136.23\ \text{g/mol C}_{10}\text{H}_{16}}\right)\times 100 = 88.16\%$$

Empirical Formula

- **Empirical Formula**:
  - Formula based on the lowest whole number ratio of its component elements
  - C_4H_8 reduces to CH_2
  - Na_2O_2 reduces to NaO

Mass % to Empirical Formula

- Approach:
  1. Assume 100 g of substance.
  2. Convert mass of each element to moles.
  3. Compute mole ratios.
  4. If necessary, convert to smallest whole number ratios by dividing by smallest number moles.

Mass % to Empirical Formula

Example:

Compound is 74.88% C and 25.12 % H

1. In 100 g sample, 74.88 g C, 25.12 g H
2. 6.23 moles C, 24.92 moles H
3. Ratio of 24.92 mol H to 6.23 mol C
4. Reduces to 4 moles H:1 mole C
5. Empirical formula of CH_4

Practice: Empirical Formulas

- For thousands of years the mineral chalcocite has been a highly prized source of copper. Its chemical composition is 79.85% Cu and 20.15% S. What is its empirical formula?
  - Collect and Organize:
  - Analyze:
  - Solve:
  - Think about It:

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Empirical vs. Molecular Formulas

- Empirical Formula:
  - Simplest whole-number molar ratio of elements in a compound
- Molecular Formula:
  - Actual molar ratio of elements in a compound
  - Equal to whole # multiple of empirical formula
  - Need empirical formula and molecular mass.
  - Both C_2H_2 and C_6H_6 have the same empirical formula, CH
    - Glucose: empirical formula = CH_2O

      molecular formula = $(CH_2O) \times 6 = C_6H_{12}O_6$

Empirical vs. Molecular Formulas

- Glycoaldehyde (60.05 g/mol):
  - Elemental analysis
    - = 40.00% C, 6.71% H, 53.28% O
  - Mole ratios
    - C:H:O = 3.33 : 6.66 : 3.33, which simplifies to 1:2:1
  - Empirical formula = CH_2O (30.02 g?)
  - Molecular formula = CH_2O x 2 = $C_2H_4O_2$

Practice: Empirical to Molecular Formula

Asbestos was used for years as an insulating material in buildings until prolonged exposure to asbestos was demonstrated to cause lung cancer. Asbestos is a mineral containing magnesium, silicon, oxygen, and hydrogen. One form of asbestos, chrysotile (520.27 g/mol), has the composition 28.03% magnesium, 21.60% silicon, 1.16% hydrogen. Determine the molecular formula of chrysotile.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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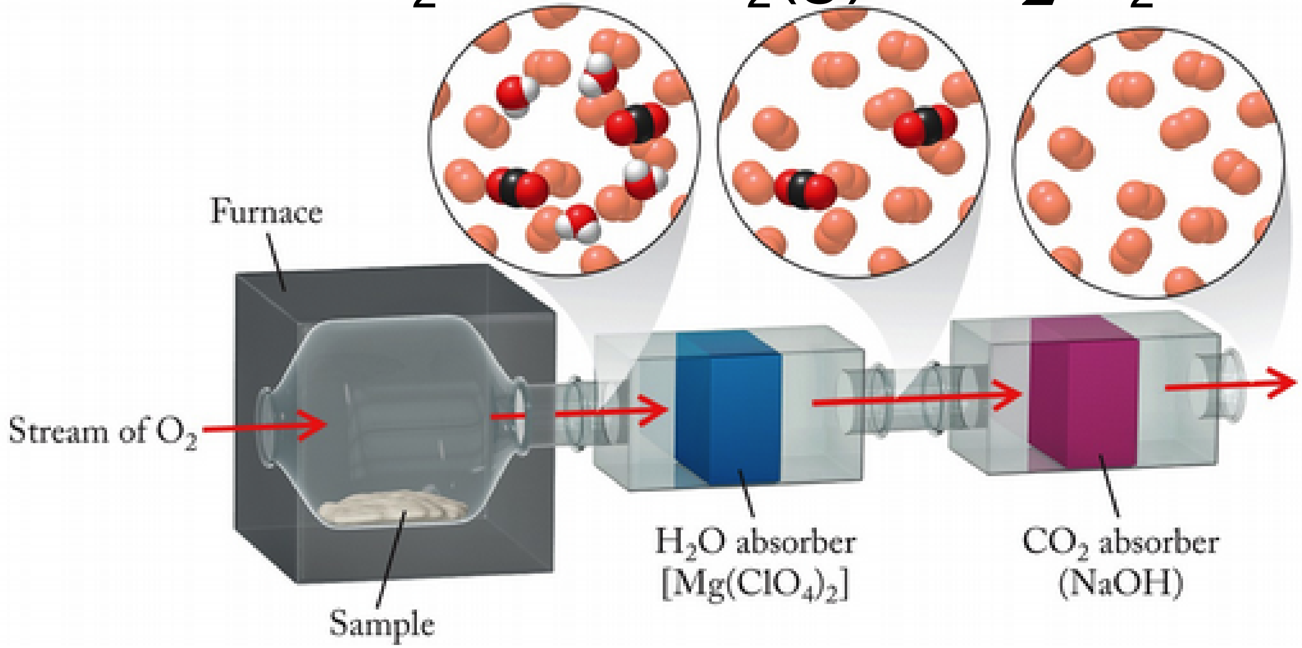
Combustion Analysis

- The % of C and H in C_aH_b determined from the mass of H_2O and CO_2 produced by combustion:
  - C_aH_b + excess $O_2 \rightarrow a\,CO_2(g) + \frac{b}{2}H_2O$



Practice: Combustion Analysis

Combustion analysis of an unknown compound indicated that it is 92.23% C and 7.82% H. The mass spectrum indicated the molar mass is 78 g/mol. What is the molecular formula of this unknown compound?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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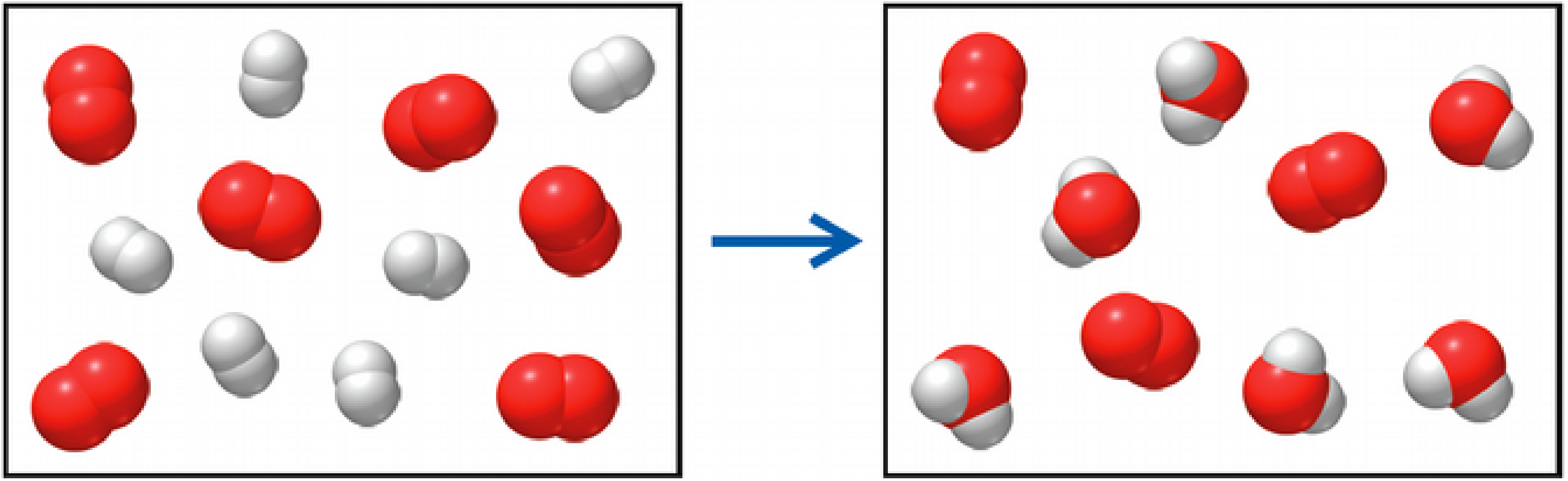
Limiting Reactant

Hydrogen and Oxygen react to form water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$



H_2 (g) = white; O_2(g) = red. Which runs out first?

Limiting Reactant

- Limiting Reactant:
  - Substance that is completely consumed in the chemical reaction
  - Determines the amount of product that can be formed during the reaction
  - Identified by:
    - # of moles of reactants
    - Stoichiometry of balanced chemical equation

Limiting Reactants

- $SO_3(g) + H_2O(g) \rightarrow H_2SO_4(l)$
- How much product is obtained from reaction of 20.00 g SO_3 and 10.00 g H_2O?
- Find limiting reactant:
  - SO_3 : 20.00 g/(80.06 g/mol) = 0.2498 moles
  - H_2O: 10.00 g/(18.02 g/mol) = 0.5549 moles
- Stoichiometry: need 1 mol SO_3: 1 mol H_2O
- Limiting reactant = SO_3

Practice: Limiting Reactant

If 10.0 g of calcium hydroxide are reacted with 10.0 g of carbon dioxide to produce calcium bicarbonate:

a. What is the limiting reactant?

b. How many grams of calcium bicarbonate will be produced?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Percent Yield

- Theoretical Yield:

  The maximum amount of product possible in a chemical reaction for given quantities of reactants

- Actual Yield:

  The measured amount of product formed

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

Practice: Percent Yield

Aluminum burns in bromine liquid, producing aluminum bromide. In one experiment, 6.0 g of aluminum reacted with an excess of bromine to yield 50.3 g aluminum bromide. Calculate the theoretical and percent yields.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

This concludes the Lecture PowerPoint presentation for Chapter 7

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