GENERAL

chapter 7 review chemical formulas and chemical compounds

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Introduction

Welcome to the definitive review of Chapter 7, where we dive deep into the world of chemical formulas and the myriad compounds that make up the universe around us. Whether you're a high school chemistry student preparing for exams, a college sophomore polishing your foundational knowledge, or a lifelong learner curious about how substances are represented, this article will guide you through the core concepts, practical examples, and essential problem solving strategies.

In this review, we'll explore how chemical formulas are constructed, the distinctions between molecular, empirical, and ionic representations, and how to interpret these symbols to predict the behavior of real world substances. We'll also cover the major categories of compounds—ionic, covalent, acids, bases, salts, coordination, and organic—and examine their structural features and common applications. By the end, you'll have a solid grasp of the language that chemists use to describe matter and the tools to apply that knowledge in both academic and everyday contexts.

Understanding Chemical Formulas

Molecular vs. Empirical Formulas

When you look at a chemical formula, you're seeing a shorthand that encodes the elemental composition of a substance. Two of the most frequently encountered types are **molecular formulas** and **empirical formulas**.

- Molecular Formula – Represents the exact number of each type of atom in a single molecule. For example, the molecular formula of glucose is C6H12O6.
- Empirical Formula – Shows the simplest whole number ratio of atoms in a compound. The empirical formula for glucose is C1H2O1 because all elements are reduced to their lowest common ratio.

Understanding the difference is crucial when interpreting laboratory data or solving stoichiometric problems. Molecular formulas give you the precise composition, while empirical formulas are useful for quick comparisons and identifying common patterns among related substances.

Notation and Conventions

Accurate notation ensures clear communication. Here are some key conventions:

- Subscripts indicate the number of atoms of a particular element in a molecule.
- Superscripts are used for charges on ions (e.g., Na+, Cl-).
- Parentheses group atoms that repeat together (e.g., Ca3(PO4)2).
- Brackets denote complex ions or coordination compounds (e.g., [Fe(CN)6]4-).
- Prefixes such as mono-, di-, tri- are rarely used in modern notation but can help clarify ambiguous formulas.

By mastering these conventions, you'll be able to read and write formulas with confidence.

Types of Chemical Compounds

Ionic Compounds

Ionic compounds arise from the electrostatic attraction between cations and anions. These are typically formed between a metal and a non metal, and they crystallize into lattice structures. Key features include:

- High melting and boiling points due to strong lattice forces.
- Electrical conductivity when molten or dissolved in water.
- Common examples: NaCl (sodium chloride), CaCO3 (calcium carbonate).

Covalent Compounds

Covalent compounds involve sharing electrons between atoms, usually non metals. They can be further classified into:

- Nonpolar covalent – Equal sharing of electrons (e.g., H2).
- Polar covalent – Unequal sharing, creating partial charges (e.g., H2O).
- Network covalent – Extensive networks of shared electrons (e.g., SiO2).

These compounds generally have lower melting points compared to ionic compounds and are poor conductors of electricity.

Acids, Bases, and Salts

Acids and bases are defined by the **Arrhenius**, **Bronsted–Lowry**, and **Lewis**theories. Salts are ionic compounds formed from the neutralization of an acid and a base.

- Acids donate protons (H+) or accept electron pairs. Examples: HCl, H2SO4.
- Bases accept protons or donate hydroxide ions (OH-). Examples: NaOH, NH3.
- Salts result from neutralization, e.g., Na2SO4 from NaOH and H2SO4.

Coordination Compounds

Coordination compounds, also known as complex ions, involve a central metal atom or ion surrounded by ligands. The general formula is **[M(L)_n]^{m±}**, where *M* is the metal, *L* is a ligand, *n* is the coordination number, and *m* is the overall charge.

Examples:

- [Fe(CN)6]4- – Hexacyanoferrate(II).
- [Cu(NH3)4]2+ – Tetraamminecopper(II).

Organic Compounds

Organic chemistry focuses on carbon based molecules. Key structural features include:

- Single, double, and triple bonds between carbon atoms.
- Functional groups such as alcohols (-OH), ketones (-CO-), carboxylic acids (-COOH), and amines (-NH2).
- Isomerism—structural, stereochemical, and conformational variations.

Common organic molecules: **CH₄** (methane), **C₆H₁₂O₆** (glucose), **C₈H₁₀O₄** (aspirin).

Stoichiometry and Formula Calculations

Balancing Equations

Balancing chemical equations ensures conservation of mass. The general approach involves:

1. Write the unbalanced equation.
2. Count atoms of each element on both sides.

3. Adjust coefficients to equalize the counts.
4. Verify that all atoms and charge balance.

Example:

Unbalanced: *Fe + O₂ !’ Fe₂O₃*

Balanced: 4 Fe + 3 O₂ !’ 2 Fe₂O₃

Mole Concept

The mole is the SI unit for counting particles. One mole contains Avogadro’s number (**6.022 × 10²³**) of entities. Key relationships:

- n (moles) = mass (g) / molar mass (g/mol)
- mass (g) = n (moles) × molar mass (g/mol)
- molar mass = sum of atomic masses of all atoms in the formula unit

Using the mole concept allows conversion between mass, moles, and number of molecules.

Mass-to-Mass Calculations

When reacting substances, you often need to determine the limiting reagent and theoretical yield. Steps include:

1. Calculate moles of each reactant from given masses.
2. Use stoichiometric coefficients to find the mole ratio.
3. Identify the limiting reagent—the reactant that will be consumed first.
4. Use the limiting reagent’s moles to calculate the maximum product mass.

Example: If 10 g of Na and 20 g of Cl2 react to form NaCl, determine the theoretical yield of NaCl.

Common Chemical Formulas in Daily Life

Food Chemistry

Many everyday foods contain chemical compounds whose formulas reveal their nutritional and functional properties:

- Carbohydrates: C6H12O6 (glucose).

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