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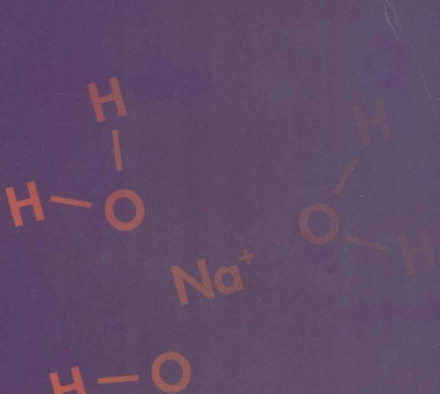
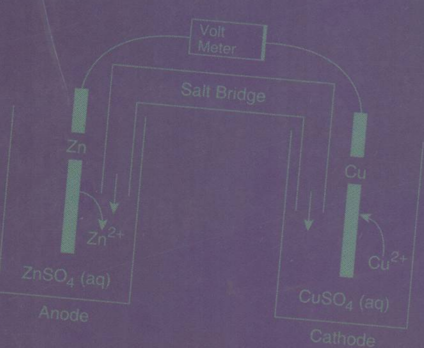
HIGH-YIELD™

General Chemistry

by STEPHEN BRESNICK, M.D.





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COLUMBIA REVIEW

HIGH-YIELD GENERAL CHEMISTRY

Stephen D. Bresnick, M.D.
President and Director
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Rose Tree Corporate Center
1400 North Providence Road
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Columbia Review

HIGH-YIELD GENERAL CHEMISTRY

Preface

High-Yield General Chemistry is an easy-to-read, efficient, and high-quality review book for first-year, college-level general chemistry. The book focuses on a conceptual review of core general chemistry topics and covers an amazing amount of material for its size. In addition to a content review, many examples and sample problems are given. For mastery of review material, over 200 practice questions with solutions are provided. The book is designed for all college students or others wishing to review and understand the major concepts in general chemistry. Students who are pre-health majors, chemistry majors, or even nonscience majors will all benefit from this book.

This is one of four books in the *High-Yield College Science Review Series* by Williams & Wilkins. The series includes *High-Yield General Chemistry, High-Yield Organic Chemistry, High-Yield Physics,* and *High-Yield Biology.* This series has been designed to make these four important college sciences easier to understand and master. All the High-Yield books contain a great science review, many examples and sample problems, and several hundred practice questions with answers and explanations.

The author of this series, Dr. Stephen Bresnick, is an expert in helping students understand, review, and retain basic college science material. Many students work their way through college courses without really comprehending the material they are supposed to be learning, and Dr. Bresnick has designed this series to help students **understand science better** and **improve their course grades.** In addition, the series has been developed to help students prepare for **post-graduate and pre-professional tests,** such as the GRE, MCAT, DAT, PCAT, VET, OAT, AHPAT, and other tests. Dr. Bresnick is an academic physician who teaches and writes science review material for college students. He is currently Director of **Columbia Review,** a national test preparation company specializing in science and English skills review for students interested in entering medical school.

Organization

There are three sections in this book. Each section corresponds to the quarter- or semester-specific topics that most college students study in general chemistry courses. Each of the three sections presents a review of important general chemistry topics. The topic review emphasizes conceptual learning, complete with numerous sample problems and examples. Each section of the book is followed by several sets of review questions, complete with solutions.

Acknowledgments

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SECTION

Chemical Properties of the Elements

1 Introductory Concepts

I. Basic Definitions and Relationships Used in Chemistry

Before conceptual and quantitative problems in chemistry can be solved, some basic definitions and concepts must be reviewed. This first chapter provides a review of the basics of stoichiometry and oxidation–reduction, and the balancing of chemical equations. An understanding of these topics is important both for success in chemistry courses in college and for conceptual understanding of the chapters that follow.

An **atom** is the smallest neutral component of an element that has all the chemical properties of the element. As will be reviewed in Chapter 2, atoms are composed of subatomic particles: **protons, neutrons,** and **electrons.** For all practical purposes, an atom can be thought of as having a very small core, the **nucleus,** which contains both protons and neutrons. Electrons form a cloud around the nucleus. The space occupied by electrons (electron cloud) is massive compared to the nucleus. The **atomic number** of an atom is the number of protons that the atom contains.

In many substances, groups of atoms are joined together by chemical bonds to form **molecules**. The composition of a molecule can be expressed by its **molecular formula**—by writing the symbols of the atoms it contains, with numeric subscripts showing how many of that kind of atom are present in the molecule.

A **mole** is the amount of a substance that contains **Avogadro's number** of particles. Avogadro's number is **approximately 6.02×10^{23}**. Thus, a mole of atoms contains 6.02×10^{23} atoms.

The **atomic weight** of an element is an average of the weights of all the isotopes of the element. **Isotopes** are atoms of the same element that differ in mass but not in atomic number (atoms with the same number of protons but a different number of neutrons). **Stoichiometry** is the study of the molar relationships between atoms and compounds.

A. ATOMIC AND MOLECULAR WEIGHT

In chemical calculations, you often need to calculate the mass of an element or a molecule. This calculation is accomplished by finding the atomic weight of the element on the periodic table. The atomic weight is the larger of the two numbers given for each element. The smaller number is usually the atomic number. Molecular weight is calculated by adding the individual atomic weights of the elements composing the compound. Example 1-1 illustrates the calculation of molecular weight.

One common mistake in the calculation of molecular weight is not multiplying through by the subscripts in a molecular formula. For example, a molecule with a formula of $Al_2(SO_4)_3$ would actually be thought of as $Al_2S_3O_{12}$ for calculating molecular weight.

Example 1-1: What is the molecular weight of C_6H_6?

Solution: Look up the atomic weights of both carbon and hydrogen (C = 12 and H = 1). Multiply each by the number of that type of atom, then add the result. Molecular weight = $(12 \times 6) + (1 \times 6) = 78$ g/mol.

B. EMPIRICAL FORMULA VERSUS MOLECULAR FORMULA

Chemical formulas are derived empirically, which means that experiments must be done to determine the actual mass of the elements in a compound or their percentage by weight. The **empirical formula** (EF) of a compound is the smallest possible integer ratio of the different kinds of atoms present in a compound.

The **molecular formula** (MF) is an integral multiple of the empirical formula. It expresses the actual number of atoms joined by chemical bonds to form a molecule. For example, the molecular formula for benzene is C_6H_6. Notice the 1 : 1 ratio of carbon and hydrogen, making the empirical formula CH.

The molecular formula can be determined if the molecular weight and the empirical formula of a compound are known. The ratio of the molecular weight of the compound to the molecular weight of the empirical formula provides the integral multiple from which the molecular formula can be determined. Example 1-2 shows how a molecular formula is determined from the molecular weight and empirical formula.

How do you experimentally determine the molecular weight of a compound? Usually from studies of the compound in its gaseous state. If a compound is weighed and then vaporized at constant temperature and volume, the gaseous molecular weight can be determined.

C. DESCRIPTION OF COMPOSITION BY PERCENT MASS

Many analytic methods do not give an empirical formula directly. Instead, a **percent mass** is provided for each of the elements in the sample. The empirical formula can then be determined from the composition by percent mass (Example 1-3). To determine the empirical formula from percent composition, follow these steps:

1. Assume 100 g of the compound. (This amount makes the calculation easier.)
2. Find the number of moles of each element in the compound.
3. Divide each of these numbers of moles by the smallest value arrived at in step 2.
4. Finally, multiply all by the smallest factor that provides whole numbers. These numbers are the subscripts in the empirical formula.

Example 1-2: Glucose has an empiric formula CH_2O and a molecular weight of 180 g/mol. What is its molecular formula?

Solution: The empiric weight of glucose is 12 + 1 + 1 + 16 = 30. 180/30 = 6 of the empiric unit in the molecule. The molecular formula is $C_6H_{12}O_6$.

Example 1-3: A hydrocarbon was determined to contain 20% hydrogen by mass. Determine its empiric formula.

Solution: First, realize that a hydrocarbon contains only carbon and hydrogen, so the actual percent composition is 20% hydrogen and 80% carbon. Then, follow these steps:

1. **Assume** 100 g of the hydrocarbon: Because 20% of 100 g is 20 g, and 80% of 100 g is 80 g, there are 20 g of hydrogen and 80 g of carbon.
2. **Convert** each of these masses to moles:
   (80 g) (1mol/12 g C) = 6.67 mol
   (20 g) (1mol/1 g H) = 20 mol
3. **Divide** each of these by the smallest number:
   6.67 / 6.67 = 1
   20 / 6.67 = 3
4. **Multiply** by the smallest factor that makes these whole numbers (1 in this case). Thus, the empiric formula is CH_3.

D. MOLE CONCEPT AND AVOGADRO'S NUMBER

Although some of these terms have already been defined, a more detailed discussion is warranted. An understanding of the mole concept is critical in understanding chemistry.

Elements react in certain ratios by weight. This observation led to the modern understanding of moles and stoichiometry. Because molecules and atoms are extremely small, it is not convenient to use quantities such as dozens or scores to define a quantity of atoms or molecules. The number of dozens of atoms needed to add up to 1 gram is an extremely large number. The mole is a quantity that makes dealing with huge numbers of molecules much easier. **A mole of atoms is the number of atoms of C-12 that weigh exactly 12.0 g.** Also, as defined previously, a mole of atoms is equivalent to 6.02×10^{23} atoms.

Conveniently, a mole of any type of atom or compound is the number of atoms that has a weight equal to the atomic or molecular weight. Therefore, an appropriate unit for the atomic weights on the periodic table is grams per mole (g/mol). Examples 1-4 and 1-5 demonstrate how to calculate the number of moles in a gram quantity of a substance.

Because a mole of a substance contains Avogadro's number of particles of that substance, it is easy to calculate the number of atoms that are in a mass quantity of a compound. A good example of this calculation is given in Example 1-6.

Example 1-4: How many moles of O_2 are in 48 g of O_2?

Solution: Moles O_2 = (48 g)(1mol/32 g) = 1.5 mol

Example 1-5: How many moles is 50.0 g CH_4?

Solution: Moles CH_4 = (50.0 g)(1 mol/16.04 g) = 3.12 moles

Example 1-6: How many carbon and hydrogen atoms are in 25 g of CH_4?

Solution: $[25\text{ g } CH_4 / (16\text{ g/mol } CH_4)](1\text{ mol C}/1\text{ mol } CH_4)(6.02 \times 10^{23})$
$= 9.41 \times 10^{23}$ C atoms

$25\text{ g } CH_4 / (16\text{ g/mol } CH_4)(4\text{ mol H}/1\text{ mol } CH_4)(6.02 \times 10^{23})$
$= 3.76 \times 10^{24}$ H atoms

E. DENSITY

Matter has mass and occupies space or volume. Density refers to the way in which mass is related to volume, and is represented by the formula:

density = m/v **m = mass (unit: usually grams)**
v = volume (unit: liters, milliliters, or centimeters cubed)

Density is an **intrinsic property** of a substance, meaning that it is not dependent on the amount of matter. The volume of a substance does change with temperature, however, so density is dependent on temperature and therefore is usually reported at a given temperature.

You are probably aware of density differences in things around you. For example, you know that a block of iron weighs more than a block of wood of equal volume and that oil floats above water. You have also seen examples of density differences in the organic laboratory, when a liquid like chloroform ($CHCl_3$) tends to form a layer beneath a less dense aqueous layer.

F. OXIDATION NUMBER

Before beginning a review of oxidation numbers, you should review the meaning of oxidation and reduction.

Oxidation: A loss of electrons or an increase in oxidation number.
Reduction: A gain of electrons and a decrease in oxidation number.

The **oxidation number or oxidation state of an atom in a compound is an assigned numeric representation of the positive or negative character of the atom.** In other words, it is the number of electrons that an atom appears to have gained over or lost from its normal complement when it is combined with other atoms.

Some basic rules for **assigning oxidation numbers** are as follows:

—The sum of the oxidation numbers of atoms in a molecule or ion must equal the overall charge on the species (e.g., for O_2, the sum of the oxidation states of the two oxygen atoms must be zero; for SO_4^{2-}, the sum of the oxidation states of the sulfur and four oxygen atoms must be -2).
—The oxidation number of a free, uncharged element is zero (e.g., O_2, H_2, Na, Cl_2).
—Alkali metals, found in the first column of the periodic table (group I), have an oxidation number of $+1$ in compounds (e.g., Na^+, K^+).
—Alkaline earth metals, found in the second column of the periodic table (group II), have an oxidation number of $+2$ in compounds (e.g., Ca^{+2}, Mg^{+2}).