INTRODUCTION TO CHEMISTRY

1.1 Approaching Chemistry: Experimentation

Chemistry is the study of the composition, properties, and transformations of matter. You encounter chemistry in most aspects of your everyday life, from the cooking of food in your kitchen and the medications that improve your life, to the battery in your telephone and the materials in your clothing.

Scientists explore the natural world by asking questions and performing experiments. The results of many experiments lead to a hypothesis that attempts to explain the results. The hypothesis, in turn, can be used to make more predictions and suggest more experiments until a consistent explanation or theory is proposed. Keep in mind that a theory represents the best explanation at some point in time. Theories are continually modified and often replaced altogether.

1.2 Chemistry and the Elements

There are 115 known elements as of the writing of these book. (Only 112 elements were known in 1998 when you were in primary class!) All matter is formed from these elements. An element is a fundamental substance that cannot be chemically changed or broken down into anything simpler. About 90 elements are naturally occurring; the rest have been produced using high-energy particle accelerators.
Elements are distributed in different amounts in nature. Hydrogen is quite common in the universe and the human body, while francium is extremely rare.

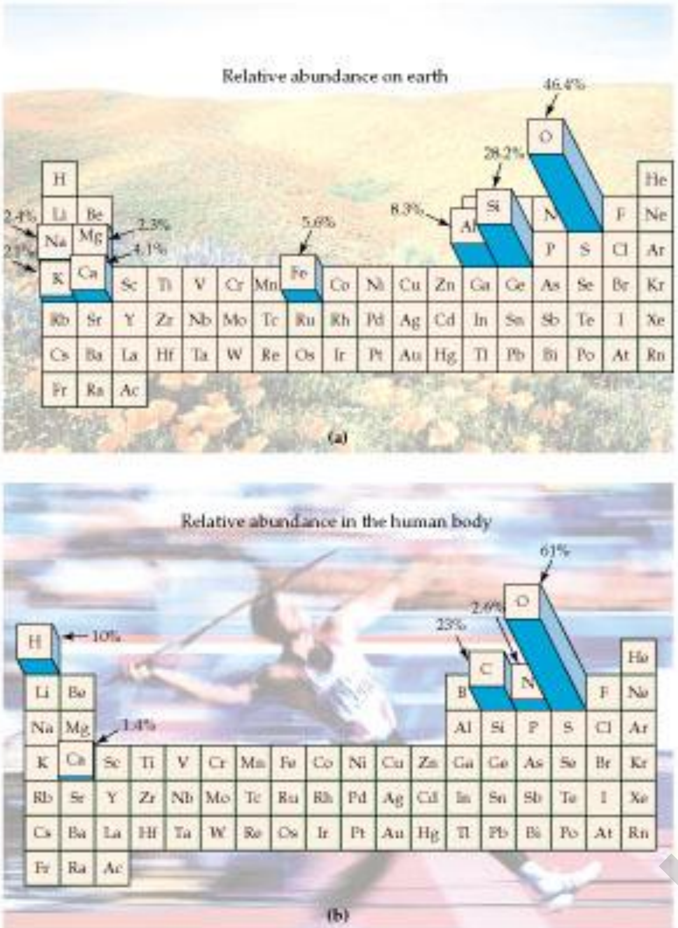
Elements have both a name and a symbol. The first letter of an element's symbol is capitalized, and the second letter, if present, is lowercased. The names and symbols of the following elements may be familiar:

TABLE 1.1 Names of Some Common Elements and Their Symbols

Aluminum	**Al**	Chlorine	**Cl**	Manganese	**Mn**	Copper (*cuprum*)	**Cu**
Argon	**Ar**	Fluorine	**F**	Nitrogen	**N**	Iron (*ferrum*)	**Fe**
Barium	**Ba**	Helium	**He**	Oxygen	**O**	Lead (*plumbum*)	**Pb**
Boron	**B**	Hydrogen	**H**	Phosphorus	**P**	Mercury (*hydrargyrum*)	**Hg**
Bromine	**Br**	Iodine	**I**	Silicon	**Si**	Potassium (*kalium*)	**K**
Calcium	**Ca**	Lithium	**Li**	Sulfur	**S**	Silver (*argentum*)	**Ag**
Carbon	**C**	Magnesium	**Mg**	Zinc	**Zn**	Sodium (*natrium*)	**Na**

1.3 Elements and the Periodic Table

Some elements (about 10) have been known since ancient times. The first tabulation of "pure substances" (now known to be elements) was published in 1789 by Antioine Lavoisier. Many elements were discovered in the late 1700s and early 1800s. Johann Döbereiner observed in 1829 that there were several triads, or groups of three elements with similar properties. The first periodic table was published by Dmitri Mendeleev in 1869
The periodic table has changed a lot since Mendeleev's time, but the changes have occurred much more slowly in the past two decades. The basic structure has remained constant. The horizontal rows are called periods and are numbered from 1 to 7 from top to bottom. The 18



vertical columns are called groups. Elements in a given group have similar chemical properties.

A summary of the main groups that have special names and properties is shown below.

Group 1A: Alkali metals Soft, shiny metals that react rapidly with water to form basic (alkali) solutions.

Group 2A: Alkaline earth metals Also lustrous metals, but less active than the Group 1A metals.
Group 7A: Halogens Colorful nonmetals that occur in gas, liquid, and solid phases.

Group 8A: Noble gases Very low-reactivity gases.

Elements with similar properties can also be grouped according to the categories metals, non-metals, and semimetals.

The large class of metals is found on the left and center of the periodic table. All metals except mercury are solids at room temperature. Metals have the following characteristics:
luster, or shiny appearance , malleable (can be pounded into thin sheets without breaking), ductile (can be pulled into the shape of thin wires), good conductors of heat and electricity

Nonmetals are found on the right side of the periodic table. Five are solids, one (bromine) is a liquid, and the remainder are gases. Several are brightly colored. In contrast to metals, solid nonmetals are brittle and are poor conductors of heat and electricity.

Seven semimetal elements are found adjacent to the boundary between metals and nonmetals. The properties of semimetals are intermediate between metals and nonmetals.

1.4 Some Physical & Chemical Properties of the Elements

Characteristics that can be used to describe or identify matter are termed properties. Chemists describe matter in terms of physical properties (such as boiling point, color, and amount) and chemical properties (such as reactivity with water). Properties can be intensive (such as density), which means the value of the property does not depend on the amount of a substance.

Properties can also be extensive (such as mass), which means they are dependent upon the amount of a substance present.

TABLE 1.2 Some Examples of Physical and Chemical Properties

Physical properties		Chemical properties
Temperature	Amount	Rusting (of iron)
Color	Odor	Combustion (of coal)
Melting point	Solubility	Tarnishing (of silver)
Electrical conductivity	Hardness	Hardening (of cement)

COMMON IONS

Positive Ions (Cations)

1+
Ammonium (NH_4^+)
Cesium (Cs^+)
Copper(I) or cuprous (Cu^+)
Hydrogen (H^+)
Lithium (Li^+)
Potassium (K^+)
Silver (Ag^+)
Sodium (Na^+)

2+
Barium (Ba^{2+})
Cadmium (Cd^{2+})
Calcium (Ca^{2+})
Chromium(II) or chromous (Cr^{2+})
Cobalt(II) or cobaltous (Co^{2+})
Copper(II) or cupric (Cu^{2+})
Iron(II) or ferrous (Fe^{2+})
Lead(II) or plumbous (Pb^{2+})
Magnesium (Mg^{2+})
Manganese(II) or manganous (Mn^{2+})
Mercury(I) or mercurous (Hg_2^{2+})
Mercury(II) or mercuric (Hg^{2+})
Strontium (Sr^{2+})
Nickel(II) (Ni^{2+})
Tin(II) or stannous (Sn^{2+})
Zinc (Zn^{2+})

3+
Aluminum (Al^{3+})
Chromium(III) or chromic (Cr^{3+})
Iron(III) or ferric (Fe^{3+})

Negative Ions (Anions)

1−
Acetate ($C_2H_3O_2^-$)
Bromide (Br^-)
Chlorate (ClO_3^-)
Chloride (Cl^-)
Cyanide (CN^-)
Dihydrogen phosphate ($H_2PO_4^-$)
Fluoride (F^-)
Hydride (H^-)
Hydrogen carbonate or bicarbonate (HCO_3^-)
Hydrogen sulfite or bisulfite (HSO_3^-)
Hydroxide (OH^-)
Iodide (I^-)
Nitrate (NO_3^-)
Nitrite (NO_2^-)
Perchlorate (ClO_4^-)
Permanganate (MnO_4^-)
Thiocyanate (SCN^-)

2−
Carbonate (CO_3^{2-})
Chromate (CrO_4^{2-})
Dichromate ($Cr_2O_7^{2-}$)
Hydrogen phosphate (HPO_4^{2-})
Oxide (O^{2-})
Peroxide (O_2^{2-})
Sulfate (SO_4^{2-})
Sulfide (S^{2-})
Sulfite (SO_3^{2-})

3−
Arsenate (AsO_4^{3-})
Phosphate (PO_4^{3-})

Balancing Chemical Equations

In your study of chemistry you will encounter numerous types and examples of chemical reactions. It simply is not practical to perform or demonstrate every chemical reaction of interest. Chemists write and interpret chemical equations that represent reactions. The following is the format used to represent a chemical reaction:

$$H_2 \quad + \quad Cl_2 \quad \longrightarrow \quad 2HCl$$

Both hydrogen and chlorine are diatomic molecules, so that's how we write them in the equation. Notice also that the same number of atoms are present on each side of the reaction arrow; the equation is balanced. According to the law of mass conservation, equations must be balanced. In other words, the number and kinds of atoms must be the same in the products as in the reactants.

Steps for Balancing Equations:

1. Write the unbalanced equation using the correct chemical formula for each reactant and product.

$$Al \quad + \quad O_2 \quad \longrightarrow \quad Al_2O_3$$

2. Find suitable coefficients. As you work to balance an equation, you can change the coefficients but not the formulas of the reactants or products. Adjust the coefficients until the same number of atoms of each element appear on both sides of the reaction arrow.

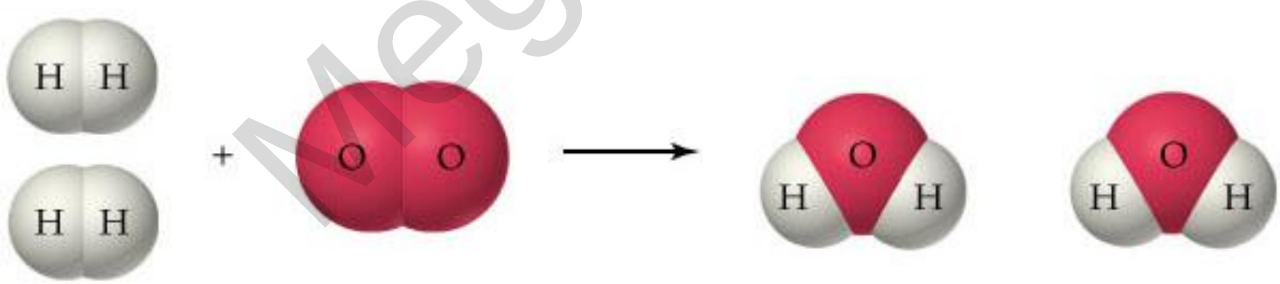
$$4Al \quad + \quad 3O_2 \quad \longrightarrow \quad 2Al_2O_3$$

3. If necessary, reduce the coefficients to their smallest whole-number values by dividing through by a common divisor.
4. Check your answer to see that the same number of atoms appear on both sides of the equation.

3.2 Chemical Symbols on Different Levels

A note about symbols and chemical reactions: In addition to learning to balance chemical equations, you must learn to visualize atoms at the microscopic level in order to fully understand chemistry. Let's consider the formation of water.

2H2 + O2 2H2O



You now know that this is a balanced chemical equation. You can also think about the equation in terms of the reaction of individual molecules: two molecules of hydrogen and one molecule of oxygen react to form two molecules of water.

Molecular Mass & Formula Mass:

The molecular mass is the average mass of the molecules of a particular substance. The term formula mass is used for ionic compounds. Both molecular mass and formula mass are calculated by adding up the atomic mass of all the atoms present in a species.

For ethylene, C_2H_4:
at. mass of 2 C = 2 x 12.0 = 24.0
at. mass of 4 H = 4 x 1.0 = 4.0
mol. mass of C_2H_4 = 28.
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For ethyl chloride, C_2H_5Cl:
at. mass of 2 C = 2 x 12.0 = 24.0
at. mass of 5 H = 5 x 1.0 = 5.0
at. mass of Cl = 35.5
mol. mass of C_2H_4 = 64.5

For hydrogen chloride, HCl:
at. mass of H = 1.0
at. mass of Cl = 35.5
mol. mass of HCl = 36.5

For sodium chloride, NaCl:
at. mass of Na = 23.0
at. mass of Cl = 35.5

mol. mass of NaCl = 58.5

Yields of Chemical Reactions

In the previous section, it was assumed that the reactions proceed "to completion,"—in other words, that all reactants are converted to products. However, this is not always the case. Side reactions can result in the formation of secondary products. Chemists calculate the percent yield of a reaction by comparing the amount of product formed to the theoretical yield predicted from stoichiometry.

$$\textbf{Percent yield} = \frac{\text{Actual yield of product}}{\text{Theoretica yield of product}} \times 100\%$$

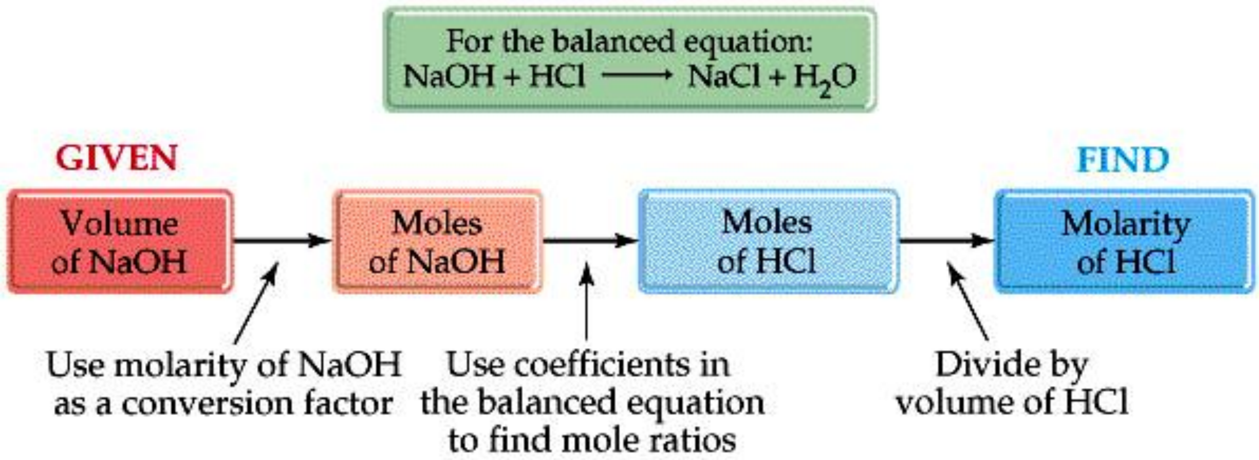
Example: What is the theoretical yield of Al_2S_3 when 10.0 g of aluminum is reacted with excess sulfur according to the equation below?

$$2Al(s) + 3\,S(s) \text{------------}\rightarrow Al_2S_3(s)$$

First, we need to convert grams of aluminum to moles:
Next, we relate moles of aluminum to moles of product using the stoichiometric coefficients as a mole ratio:
Finally, we calculate our theoretical yield of Al_2S_3 in grams.
Example:
A student performing the reaction above collected 18.7 g Al_2S_3. What is her percent yield?

Percent Composition and Empirical Formulas

The percentage of each element in a chemical formula is called percent composition. It is possible to determine chemical formulas from percent-composition data.



Let's consider a phosphorus oxygen compound with a mass that is 43.6% phosphorus and 54.6% oxygen. Because we are working with percent composition, let's assume we have 100 g of this compound.

Thus, we have 43.6 g phosphorus, and 54.6 g oxygen.

Now we can use the molar masses of each element to convert grams to moles.

43.6 g P (1 mol P / 31.0 g) = 1.41 mol P

56.4 g O (1 mol O / 16.0 g) = 3.52 mol O

By dividing the larger number of moles by the smaller number, we can determine the mole ratio of the elements.

Moles of O / Moles of P = 3.52 mol O / 1.41 mol P = 2.50 mol O / 1 mol P

We insert this ratio into the formula $PO_{2.5}$. Because formulas commonly use whole number values, we can multiply by 2 to get P_2O_5. This is the empirical formula, the simplest ratio of moles of P to moles of O determined from experimental data. To determine the actual molecular formula, we need to know the molecular mass of the substance—in this case, 284 The molecular formula is a multiple of the empirical formula, so dividing the molecular formula-weight by the empirical formula-weight will give us the multiplier.

Molecular formula mass / empirical formula mass = 284 / 142 = 2.0

Thus the molecular formula is twice the empirical formula, or P_4O_{10}.

Some Ways that Chemical Reactions Occur

A large number of important and interesting chemical reactions take place in aqueous solution. Three general categories of solution reactions can be distinguished according to the type of driving force behind them.

In precipitation reactions, soluble reactants yield an insoluble solid product. Many precipitation reactions involve the double displacement of a pair of ionic compounds. The driving force of a precipitation reaction is the production of the solid product.

In acid-base neutralization reactions, an acid reacts with a base to produce water and an ionic compound called a salt. The driving force of a neutralization reaction is the formation of water molecules.

In oxidation-reduction reactions (redox reactions), electrons are transferred between. The driving force of a redox reaction is a decrease in electrical potential.