新世纪高等学校教材

化学系列教材

HUAXUE ZHUANYE

化学专业英语

汪辉亮 主编

北京师范大学出版集团
BEIJING NORMAL UNIVERSITY PUBLISHING GROUP
北京师范大学出版社

化学系列教材

化学专业英语

HUAXUE ZHUANYE YINGYU

汪辉亮　主　编

于海涛　舒　东　潘　睿　副主编

于海涛　王力元　李慧珍
汪辉亮　赵　扬　舒　东　潘　睿　编　者

图书在版编目(CIP)数据

化学专业英语／汪辉亮主编. —北京：北京师范大学出版社，2011.2
新世纪高等学校教材. 化学系列教材
ISBN 978-7-303-12100-7

Ⅰ. ①化… Ⅱ. ①汪… Ⅲ. ①化学－英语－高等学校－教材 Ⅳ. ①H31

中国版本图书馆CIP数据核字（2011）第016402号

营销中心电话　010-58802181 58808006
北师大出版社高等教育分社网　http://gaojiao.bnup.com.cn
电子信箱　beishida168@126.com

出版发行：北京师范大学出版社 www.bnup.com.cn
　　　　　北京新街口外大街19号
　　　　　邮政编码：100875
印　　刷：北京东方圣雅印刷有限公司
经　　销：全国新华书店
开　　本：170 mm×230 mm
印　　张：18
字　　数：312千字
版　　次：2011年2月第1版
印　　次：2011年2月第1次印刷
定　　价：30.00元

策划编辑：范　林　　　责任编辑：范　林　李洪波
美术编辑：毛　佳　　　装帧设计：毛　佳
责任校对：李　菡　　　责任印制：李　啸

前言

语言是思维和交流的工具。汉语无疑是世界上使用人口最多的语言，也是最古老、最优秀的语言之一。但另一个不得不承认的事实是，英语是当今世界上使用最广泛、影响力最大的语言。英语也是目前国际学术界进行交流的最重要的语言。对化学工作者而言，英语尤为重要，阅读文献、参加学术会议和发表论文等都离不开英语。

化学专业英语是各级各类高等院校化学专业学生的必修或选修课程。在教学过程中我们发现，有的学生即便修过一些双语化学课程，但对化学专业英语中一些基本的、简单的表达也不是很清楚。因此，我们认为，对学生进行系统的化学专业英语的训练是很有必要的。化学专业英语不同于普通英语之处在于其专业性，因此其重点为使学生掌握化学专业的基本词汇、术语和相关的表达方式，提高学生化学专业方面的英语能力。

本书共分9章，1～8章内容涉及了大学化学所涵盖的五个二级学科：无机化学、有机化学、分析化学、物理化学(含结构化学)、高分子化学，其中包括了元素、无机物、有机物和高分子的命名等，第9章为新材料。附录为实验室常用仪器和装置。本书在编写过程中参考了大量的相关各学科的英文原版教科书和期刊文章，同时还参考了很多网络资料。我们力求在本教材中将大学化学所涉及的化学专业词汇和表达都涵盖到。虽然化学专业英语的目的不是教给学生相关化学学科的专业知识，但我们在各章内容的选择上尽可能做到系统化、条理化。

本书由北京师范大学汪辉亮主编，河北师范大学于海涛、华南师范大学舒东、四川师范大学潘睿、北京师范大学王力元、河南师范大学赵扬和李慧珍等老师参与了部分章节的编写。

本书可用作综合性大学、师范大学及其他本科院校化学、化工专业的本科生及研究生的专业英语教材，也可用作与化学相关人员的参考资料。

由于编写时间有限，在本书中没有注释和练习题。同时，由于编者水平的限制，书中难免存在错误和不当之处，敬请读者指正。

最后对各位作者对本书作出的贡献表示衷心感谢。

编　者

2010 年 9 月

目 录

Chapter 1 Matter and Measurement

Chemistry is the science of matter and the changes it undergoes. Chemists study the composition, structure, and properties of matter. They observe the changes that matter undergoes and measure the energy that is produced or consumed during these changes. Chemistry provides an understanding of many natural events and has led to the synthesis of new forms of matter that have greatly affected the way we live.

Disciplines within chemistry are traditionally grouped by the type of matter being studied or the kind of study. These include **inorganic chemistry**, **organic chemistry**, **physical chemistry**, **analytical chemistry**, **polymer chemistry**, **biochemistry**, and many more specialized disciplines, e. g. **radiochemistry**, **theoretical chemistry**.

Chemistry is often called "**the central science**" because it connects the other natural sciences such as astronomy, physics, material science, biology and geology.

1.1 Classification of Matter

Matter is usually defined as anything that has mass and occupies space. **Mass** is the amount of matter in an object. The mass of an object does not change. The **volume** of an object is how much space the object takes up.

All the different forms of matter in our world fall into two principal categories: (1) pure substances; (2) mixtures. A **pure substance** can also be defined as a form of matter that has both definite composition and distinct properties. Pure substances are subdivided into two groups: elements and compounds. An **element** is the simplest kind of material with unique physical and chemical properties; it can not be broken down into anything simpler by either physical or chemical means. A **compound** is a pure substance that consists of two or more elements linked together in characteristic and definite proportions; it can be decomposed by a chemical change into simpler substances with a fixed mass ratio. **Mixtures** contain two or more chemical substances in variable proportions in which the pure substances retain their chemical identities. In principle, they can be separated into the component substances by physical means, involving physical changes. A sample is **homogeneous** if it always has the same composition, no matter what part of the sample is examined. Pure elements and pure chemical compounds are homogeneous. Mixtures can be homogeneous, too; in a **homogeneous mixture** the constituents are distributed uniformly and the composition and appearance of the mixture are uniform throughout. A **solutions** is a special type of homogeneous mixture. A **heterogeneous mixture** has physically distinct parts with different properties. The classification of matter is summarized in the diagram below:

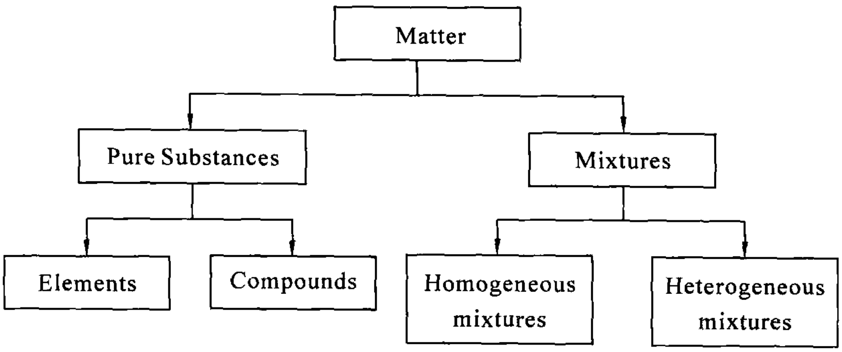

Figure 1.1 Classification of Matter

Matter can also be categorized into four distinct phases: solid, liquid, gas, and plasma. The **solid** phase of matter has the atoms packed closely together. An object that is solid has a definite shape and volume that cannot be changed easily. The **liquid** phase of matter has the atoms packed closely together, but they flow freely around each other. Matter that is liquid has a definite volume but changes shape quite easily. Solids and liquids are termed *condensed phases* because of their well-defined volumes. The **gas** phase of matter has the atoms loosely arranged so they can travel in and out easily. A gas has neither specific shape nor constant volume. The **plasma** phase of matter has the atoms existing in an **excited** state.

1.2 Properties of Matter

All substances have properties, the characteristics that give each substance its unique identity. We learn about matter by observing its properties. To identify a substance, chemists observe two distinct types of properties, physical and chemical, which are closely related to two types of change that matter undergoes.

Physical properties are those that a substance shows by itself, without changing into or interacting with another substance. Some physical properties are color, smell, temperature, boiling point, electrical conductivity, and density. A **physical change** is a change that does not alter the chemical identity of the matter. A physical change results in different physical properties. For example, when ice melts, several physical properties have changed, such as hardness, density, and ability to flow. But the sample has not changed its composition: it is still water.

Chemical properties are those that do change the chemical nature of matter. A **chemical change**, also called a chemical reaction, is a change that does alter the chemical identity of the substance. It occurs when a substance (or substances) is converted into a different substance (or substances). For example, when hydrogen burns in air, it undergoes a chemical change because it combines with oxygen to form water.

Separation of Mixtures

The separation of mixtures into its constituents in a pure state is an important process in chemistry. The constituents of any mixture can be separated on the basis of their differences in their physical and chemical properties e.g. particle size, solubility, effect of heat, acidity or basicity etc.

Some of the methods for separation of mixtures are:

(1) Sedimentation or decantation. To separate the mixture of coarse particles of a solid from a liquid e.g. muddy river water.

(2) Filtration. To separate the insoluble solid component of a mixture from the liquid completely i.e. separating the precipitate (solid phase) from any solution.

(3)Evaporation. To separate a non-volatile soluble salt from a liquid or recover the soluble solid solute from the solution.

(4)Crystallization. To separate a solid compound in pure and geometrical form.

(5)Sublimation. To separate volatile solids, from a non-volatile solid.

(6)Distillation. To separate the constituents of a liquid mixture, which differ in their boiling points.

(7) Solvent extraction method. Organic compounds, which are easily soluble in organic solvents but insoluble or immiscible with water forming two separate layers can be easily separated.

1.3 Atoms, Molecules and Compounds

The fundamental unit of a chemical substance is called an **atom**. The word is derived from the Greek *atomos*, meaning "undivisible" or "uncuttable". An atom is the smallest possible particle of a substance.

Molecule is the smallest particle of a substance that retains the chemical and physical properties of the substance and is composed of two or more atoms; a group of like or different atoms held together by chemical forces. A molecule may consist of atoms of a single chemical element, as with oxygen (O_2), or of different elements, as with water (H_2O).

A **chemical element** is a pure chemical substance consisting of one type of atom distinguished by its **atomic number**, which is the number of protons in its nucleus. The term is also used to refer to a pure chemical substance composed of atoms with the same number of protons. Until March 2010, 118 elements have been observed. 94 elements occur naturally on earth, either as the pure element or more commonly as a component in compounds. 80 elements have stable isotopes, namely all elements with atomic numbers 1 to 82, except elements 43 and 61 (*technetium* and *promethium*). Elements with atomic numbers 83 or higher (*bismuth* and above) are inherently unstable, and undergo *radioactive decay*. The elements from atomic number 83 to 94 have no stable nuclei, but are nevertheless found in nature, either surviving as remnants of the primordial stellar *nucleosynthesis* that produced the elements in the solar system, or else produced as short-lived daughter-isotopes through the natural decay of uranium and thorium. The remaining 24 elements are *artificial* or *synthetic* elements, which are products of man-induced processes. These synthetic elements are all characteristically unstable. Although they have not been found in nature, it is conceivable that in the early history of the earth, these and possibly other unknown elements may have been present. Their unstable nature could have resulted in their disappearance from the natural components of the earth, however.

The naturally occurring elements were not all discovered at the same time. Some, such as *gold*, *silver*, *iron*, *lead*, and *copper*, have been known since the days of earliest civilizations. Others, such as *helium*, *radium*, *aluminium*, and *bromine*, were discovered in the nineteenth century. The most abundant elements found in the earth's crust, in order of decreasing percentage, are *oxygen*, *silicon*, *aluminium*, and *iron*. Others present in amounts of 1% or more are *calcium*, *sodium*, *potassium*, and *magnesium*. Together, these represent about 98.5% of the earth's crust.

The nomenclature and their origins of all known elements will be described in Chapter 2.

A **chemical compound** is a pure chemical substance consisting of two or more different chemical elements that can be separated into simpler substances by chemical reactions. Chemical compounds have a unique and defined chemical structure; they consist of a fixed ratio of atoms that are held together in a defined spatial arrangement by *chemical bonds*. Compounds that exist as molecules are called **molecular compounds**. An **ionic compound** is a chemical compound in which ions are held together in a lattice structure by *ionic bonds*. Usually, the positively charged portion consists of metal **cations** and the negatively charged portion is an **anion** or *polyatomic ion*.

The relative amounts of the elements in a particular compound do not change: Every molecule of a particular chemical substance contains a characteristic number of atoms of its constituent elements. For example, every water molecule contains two hydrogen atoms and one oxygen atom. To describe this atomic composition, chemists write the **chemical formula** for water as H_2O.

The chemical formula for water shows how formulas are constructed.

The formula lists the symbols of all elements found in the compound, in this case H (hydrogen) and O (oxygen). A subscript number after an element's symbol denotes how many atoms of that element are present in the molecule. The subscript 2 in the formula for water indicates that each molecule contains two hydrogen atoms. No subscript is used when only one atom is present, as is the case for the oxygen atom in a water molecule. Atoms are indivisible, so molecules always contain whole numbers of atoms. Consequently, the subscripts in chemical formulas of molecular substances are always integers. We explore chemical formulas in greater detail in Chapter 2.

The simple formula that gives the simplest whole number ratio between the atoms of the various elements present in the compound is called its **empirical formula**. The simplest formula that gives the actual number of atoms of the various elements present in a molecule of any compound is called its **molecular formula**. **Elemental analysis** is an experiment that determines the amount (typically a weight percent) of an element in a compound. The elemental analysis permits determination of the empirical formula, and the molecular weight and elemental analysis permit determination of the molecular formula.

1.4 Numbers in Physical Quantities

1.4.1 Measurement

1. Physical Quantities

Physical properties such as *height*, *volume*, and *temperature* that can be measured are called **physical quantity**. A number and a unit of defined size are required to describe physical quantity, for example, 10 meters, 9 kilograms.

2. Exact Numbers

Exact Numbers are numbers known with certainty. They have unlimited number of significant figures. They arise by directly counting numbers, for example, the number of sides on a square, or by definition:

1 m = 100 cm, 1 kg = 1000 g

1 L = 1000 mL, 1 minute = 60 seconds

3. Uncertainty in Measurement

Numbers that result from measurements are *never* exact. Every experimental measurement, no matter how precise, has a degree of uncertainty to it because there is a limit to the number of digits that can be determined. There is always some degree of uncertainty due to experimental errors: limitations of the measuring instrument, variations in how each individual makes measurements, or other conditions of the experiment.

4. Precision and Accuracy

In the fields of engineering, industry and statistics, the **accuracy** of a measurement system is the degree of *closeness* of measurements results to its *actual* (*true*) value. The **precision** of a measurement system, also called *reproducibility* or *repeatability*, is the degree to which repeated measurements under unchanged conditions show the same results. Although the two words can be synonymous in colloquial use, they are deliberately contrasted in the context of the scientific method.

A measurement system can be accurate but not precise, precise but not accurate, neither, or both. A measurement system is called *valid* if it is both accurate and precise. Related terms are **bias** (non-random or directed effects caused by a factor or factors unrelated by the independent variable) and **error** (random variability), respectively. **Random errors** result from uncontrolled variables in an experiment and affect precision; **systematic errors** can be assigned to definite causes and affect accuracy. For example, if an experiment contains a systematic error, then increasing the sample size generally increases precision but does not improve accuracy. Eliminating the systematic error improves accuracy but does not change precision.

1.4.2 Significant Figures

The number of digits reported in a measurement reflects the accuracy of the measurement and the precision of the measuring device. **Significant figures** in a number include all of the digits that are known with certainty, plus the first digit to the right that has an uncertain value. For example, the uncertainty in the mass of a powder sample, e.g. 3.1267g as read from an "analytical balance" is ±0.0001g.

In any calculation, the results are reported to the fewest significant figures (for multiplication and division) or fewest decimal places (addition and subtraction).

1. Rules for deciding the number of significant figures in a measured quantity

The number of significant figures is found by counting from left to right, beginning with the first nonzero digit and ending with the digit that has the uncertain value, e.g.

459 (3) 0.206 (3) 2.17(3) 0.00693 (3) 25.6 (3) 7390 (3) 7390. (4)

(1) All nonzero digits are significant, e.g. 1.234 g has 4 significant figures, 1.2 g has 2 significant figures.

(2) Zeroes between nonzero digits are significant: e.g. 1,002 kg has 4 significant figures, 3.07 mL has 3 significant figures.

(3) Leading zeroes to the left of the first nonzero digits are not significant; such zeroes merely indicate the position of the decimal point: e.g. 0.001 m has only 1 significant figure, 0.012 g has 2 significant figures.

(4) Trailing zeroes that are also to the right of a decimal point in a number are significant: e.g. 0.0230 mL has 3 significant figures, 0.20 g has 2 significant figures.

(5) When a number ends in zeroes that are not to the right of a decimal point, the zeroes are not necessarily significant: e.g. 190 miles may be 2 or 3 significant figures, 50,600 calories may be 3, 4, or 5 significant figures.

The potential ambiguity in the last rule can be avoided by the use of *standard exponential*, or *"scientific" notation*. For example, depending on whether the number of significant figures is 3, 4, or 5, we would write 50,600 calories as:

5.06×10^4 calories (3 significant figures),

5.060×10^4 calories (4 significant figures), or

5.0600×10^4 calories (5 significant figures).

2. Rules for rounding off numbers

(1) If the digit to be dropped is greater than 5, the last retained digit is increased by one. For example, 12.6 is rounded to 13.