Basic chemistry

冯清

华中科技大学出版社

图书在版编目(CIP)数据

Basic Chemistry/冯清　刘绍乾　主编. 一武汉：华中科技大学出版社，2008 年 9 月
ISBN 978-7-5609-4594-1

Ⅰ. B…　Ⅱ. ①冯…　②刘…　Ⅲ. 化学-高等学校-教材-英文　Ⅳ. O6

中国版本图书馆 CIP 数据核字(2008)第 078343 号

Basic Chemistry　　冯清　刘绍乾　主编

策划编辑：胡章成
责任编辑：胡　芬　　封面设计：范翠璇
责任校对：张　琳　　责任监印：周治超

出版发行：华中科技大学出版社(中国·武汉)
武昌喻家山　邮编：430074　电话：(027)87557437

录　排：华中科技大学惠友文印中心
印　刷：华中科技大学印刷厂

开本：787mm×1092mm　1/16　印张：26　字数：650 000
版次：2008 年 9 月第 1 版　印次：2008 年 9 月第 1 次印刷　定价：40.00 元
ISBN 978-7-5609-4594-1/O·448

Basic Chemistry

for Students of Medicine and Biology

Chief editor Feng Qing, Liu Shaoqian

Vice-chief editor Sun Yaliang, Ma Ruhai, Wang Xing po

Editor (Alphabetical List)

Feng Qing (Huazhong University of Science and Technology)
Hu Guozhi (Huazhong University of Science and Technology)
Hu Yongxiang (Huazhong University of Science and Technology)
Liu Min (Huazhong University of Science and Technology)
Liu Shaoqian (Central South University)
Ma Ruhai (China Medical University)
Qian Pin (Central South University)
Sun Yaliang (Huazhong University of Science and Technology)
Wang Xingpo (Shandong University)
Yang Xiaolan (Chongqing Medical University)
Zhang Wenhua (Huazhong Normal University)

China • Wuhan
Huazhong University of Science and Technology Press

Preface

In order to meet the development for all-English or bilingual teaching of basic chemistry in medical colleges of our country, we are now grouped to edit this book. The sequence of chapters and related contents in this book are consistent to the sixth edition of Basic Chemistry in Chinese which is published by People's Medical Publishing House as a NHO's programming book material. We have applied "three basic contents—basic theories, basic knowledge, basic skills" and "five features—thoughtful, scientific, advanced, heuristic, adaptable" to this book. That means precise demonstrations, clear arrangements, visual illuminations and fluent language. The whole book is arranged in a proper and gradual sequence to avoid destroying the coherence or connections of chapters and furthermore unnecessary repetition.

In the present edition the first five chapters introduce concepts of solutions to the students, including colligative properties of solutions (Chapter 2), electrolyte solutions (Chapter 3), buffer solutions (Chapter 4), and colloids (Chapter 5). The next three chapters examine the factors that determine the speed and extent of chemical reactions: thermodynamics (Chapter 6), kinetics (Chapter 7), and electrochemistry (Chapter 8). The next two chapters deal with atomic structure (Chapter 9) and molecular structure (Chapter 10). After a discussion of coordination compounds (Chapter 11), the final chapters survey analytic chemistry of titrimetric methods (Chapter 12), ultraviolet visible spectrophotometry (Chapter 13), and modern instrumental analysis (Chapter 14).

The essay at the end of chapters present is about interesting applications of basic chemistry relevant to the main chapter subject. Including topics from science, biological science, medicine, and day-to-day life, these applications enliven and reinforce the material presented in each chapter. We have made our efforts to make this book as effective, clear and readable as possible, to show the beauty and logic of basic chemistry; and to make the subject interesting to learn.

The authors thank the editors of Huazhong University of Science and Technology Press, who have an excellent job of putting the final manuscript into book form; and most of all, the editor Zhangcheng Hu. Her encouragement, hard work, and advices made this book possible.

Finally, we express our indebtedness to our students who gave the reason for beginning, the courage to continue, and the fortitude to complete the project.

We are sincerely grateful to them all.

Authors

Contents

Chapter 1　Introduction

In this chapter we address the question "What exactly is chemistry". In addition, we consider common terminology associated with the field of chemistry. Like all other sciences, chemistry has its own specific language. It is necessary to restrict the meanings of some words so that all chemists (and those who study chemistry) can understand a given description of a chemical phenomenon in the same way.

Chemists have "tools of the trade". The tool they use most is called measurement. Measurement is indispensable in the study of chemistry. This chapter will help you to learn what you need to know to deal properly with measurement.

Solutions are common in nature, and they represent an abundant form of matter. Solutions carry nutrients to the cells of our bodies and carry away waste products. The general term concentration refers to the quantity of solute in a standard quantity of solution. A large percentage of all chemical reactions take place in solution, including most of those discussed in later chapters in this text.

1.1　Introduction to Chemistry

1.1.1　Chemistry：The Central Science

The dictionary defines **chemistry** as "the science of the composition, structure, properties and reactions of matter, especially of atomic and molecular systems". Chemistry has traditionally been subdivided into four areas：organic, inorganic, analytical and physical chemistry.

Chemistry is at the core of our scientific knowledge. All of the natural sciences (physical science, life science, and earth science) explore different relationships between materials, their interactions with each other, and their interactions with energy. The relationships between chemistry and other natural sciences are illustrated in Figure 1-1.

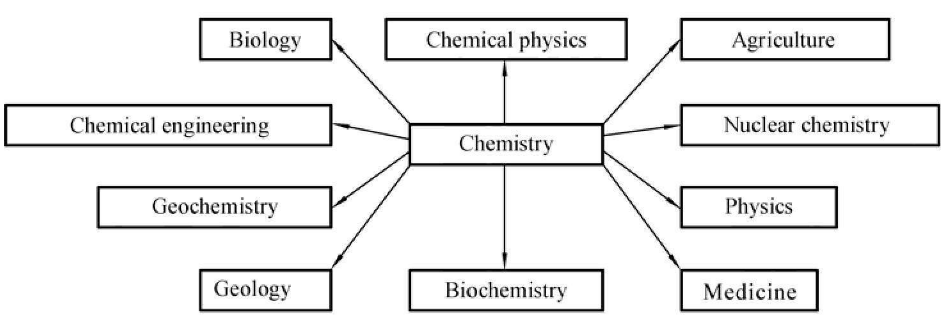
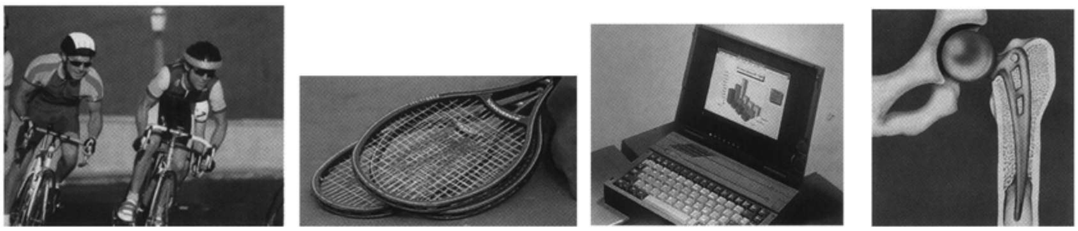


Figure 1-1　Chemistry and other natural sciences

With the development of the society, it is especially exciting time to be studying chemistry. Chemistry helps us understand the nature of materials, so that we can make better use of them and create new ones： specialized steels, advanced composites, synthetic polymers, and countless other new materials (Figure 1-2).



(a)　(b)　(c)　(d)

Figure 1-2　Photos

(a) Specialized steels in bicycles, synthetic polymers in clothing and helmets; (b) Advanced composites in sports equipment; (c) Liquid crystals in computer screens; (d) High-tension polymers in synthetic hip joints

The influence of chemistry extends to the natural environment as well. The air, water, land, and organisms that thrive there form a remarkably complex system of chemical interactions. There is no question that modern chemical products have enhanced the quality of our lives, but the manufacture and use of them also pose increasing dangers, such as toxic waste, acid rain, global warming and ozone depletion. If our heedless application of chemical principles has led to some of these problems, our thoughtful application of the same principles will play a major role in solving them.

Perhaps the significance of chemistry is most amazing when you ponder the underlying chemical nature of biology. The most vital biological questions—How did life arise and evolve? How does an organism reproduce, grow, and age? What is the essence of health and disease?—Ultimately have chemical answers.

1.1.2　Advice for Learning and Studying Chemistry

A. Time Management

A master schedule should be prepared. One study period should follow as soon as possible after the lecture. Select shorter but frequent study periods and review the contents.

B. Attendance

The importance of regular class attendance cannot be overemphasized. This text is a great ally, but what material and in what depth your instructor covers can be discovered only in class. The key to what will be on the exams is found in the lecture. You need to be there.

C. Asking Questions

Few students feel confident enough to ask questions in class. This is natural enough—we all tend to think our questions might be “dumb”. Nevertheless, you still need to ask the questions. If you hesitate to ask questions in class, take advantage of your instructor’s office hours. You may find that the instructor is really human and helpful in the office.

D. Study Skills

(1) Memorization. Chemistry has a vocabulary of its own, and in this respect, it must be approached as a foreign language. Sometimes definitions must be memorized before understanding.

(2) Reading ahead. Even in the high school, football and basketball opponents are scouted before the game. Every team likes to know what they are up against and what lies ahead. In studying science, the equivalent to scouting is reading ahead. Even if you do not grasp the concepts, reading ahead will give you a feeling for some of the material that you will discuss in the next class. If you know something that seems confusing is coming, you will be more alert when the concept is discussed. Reading ahead is also a time saver. When you know that certain definitions or tables are in the book, you can save notetaking time.

(3) Notetaking. A useful suggestion is to leave about one-third of the page as a blank margin when you take notes in class. This can be used for adding any thoughts, notes, or questions later and for summarizing in the review.

We are now ready to continue the study of chemistry. It is a fascinating subject. I hope you enjoy it. Please let the text and the instructor help you to get as much out of it as you possibly can.

1.2 Measurement and SI Units

Scientific progress is based on gathering and interpreting careful observations. These observations can be expressed either in qualitative way, like “that car is going fast”, or in a quantitative manner: “That car is going 95 miles per hour.” Qualitative descriptions involve terms such as hot, cold, fast, slow, heavy and light. Quantitative descriptions that use a number to describe a property transfer much more information. Units are standards that are used for quantitative comparison between measurements for the same type of quantity. The first measurements were probably based on the human body. In 17th and 18th centuries, scientists found that the lack of standard units was a problem. In 1960, the international committee met in France to establish the International System of Units, a revised metric system is now accepted by scientists throughout the world. The units of this system are called SI units, from the French Système International d’unitès.

1.2.1 General Features of SI Units

As Table 1-1 shows, the **SI system** is based on a set of seven fundamental units, or base units, each of which is identified with a physical quantity. All other units, called **derived units**, are combinations of these base units. For example, the derived unit for speed, meter per second ($m \cdot s^{-1}$), is the base unit for length (m) divided by the base unit for time (s).

Table 1-1　The SI base units

Quantity	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	Kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

For quantities that are much smaller or much larger than the base unit, we use decimal prefixes and exponential (**scientific) notation**. Because these prefixes are based on powers of 10, SI units are easier to use in calculations than English units, such as pounds and inches. Table 1-2 shows the most important prefixes. Many of the calculations you'll perform involve numbers written in exponential notation:

$$A\times 10^{n}$$

where A is greater than or equal to 1 and less than 10 ($1\leqslant A\leqslant 10$) and n is an integer.

Table 1-2　Common decimal prefixes

Prefix	Prefix symbol	Meaning		Multiple
		Number	Word	
tera	T	1 000 000 000 000	trillion	10^{12}
giga	G	1 000 000 000	billion	10^{9}
mega	M	1 000 000	million	10^{6}
kilo	k	1 000	thousand	10^{3}
hecto	h	100	hundred	10^{2}
deka	da	10	ten	10
deci	d	0.1	tenth	10^{-1}
centi	c	0.01	hundredth	10^{-2}
milli	m	0.001	thousandth	10^{-3}
micro	μ	0.000 001	millionth	10^{-6}
nano	n	0.000 000 001	billionth	10^{-9}
pico	p	0.000 000 000 001	trillionth	10^{-12}
femto	f	0.000 000 000 000 001	quadrillionth	10^{-15}

1.2.2　Some Important SI Units in Chemistry

A. Volume

Volume is the space that a given quantity of matter occupies. The SI unit for volume is the cubic meter (m^3). Since this is a rather large volume for typical laboratory situations, the metric unit, known as the liter, is used. One liter is defined as the exact volume of one cubic decimeter (i.e., 1 L = 1 dm^3). On a smaller scale, one milliliter is the exact volume of one cubic centimeter (1 mL = 1 cm^3 = 1 cc). Thus the units milliliter and cubic centimeter can be used interchangeably when expressing volume.

$$1\ m^3 = 10^3\ L = 10^6\ cm^3 \qquad 1\ L = 1000\ mL = 1000\ cm^3$$

B. Temperature

Temperature is familiar quantity to us. Temperature is generally measured in the unit of degree Celsius, abbreviated ℃. The Celsius scale is defined by assigning the freezing point of water a value of 0 ℃ and the boiling point of water as 100 ℃.The SI unit of temperature is Kelvin. The relationship between Celsius and Kelvin temperature is：

$$T = t + 273.15$$

1.3 Ways of Expressing Concentration

1.3.1 The Mole

A. Definition of Mole

The **mole** (abbreviated mol) is the SI unit for amount of substance. It is defined as the amount of substance in a system that contains as many elementary particles (atoms, molecules, or formula units) as the quantity of $^{12}_{6}C$ atom in exactly 12 grams of carbon-12. This number is called Avogadro's number, in honor of Italian physicist Amedeo Avogadro.

One mole (1 mol) contains 6.022×10^{23} entities (to four significant figures). Thus,

1 mol carbon-12	contains	6.022×10^{23} atoms
1 mol H_2O	contains	6.022×10^{23} molecules
1 mol NaCl	contains	6.022×10^{23} formula units

When using the term mole, it is important to specify the formula of unit to avoid any misunderstanding.

For example,

a mole of hydrogen atoms (formula H) contains 6.022×10^{23} H atoms,

a mole of hydrogen molecules (formula H_2) contains 6.022×10^{23} H_2 molecules,

a mole of sodium carbonate, Na_2CO_3, contains 6.022×10^{23} Na_2CO_3 units.

However, the mole is not just a counting unit, like the dozen, which specifies only the number of objects. The definition of the mole specifies the number of objects in a fixed mass of substance. Therefore, the molecular mass (or formula mass) of a compound expressed in amu (atomic mass unit, 1 amu equals 1.66054×10^{-24} g) is numerically the same as the mass of 1 mol of the compound expressed in gram. Thus, for example,

1 molecule of H_2O has a molecular mass of 18.0 amu,

1 mol H_2O (6.022×10^{23} molecules) has a mass of 18.0 g.

Similarly, 1 mol H_2SO_4 (6.022×10^{23} formula units) has a mass of 98.0 g, we can write the following relationship:

$$98.0 \text{ g } H_2SO_4 = 1 \text{ mol } H_2SO_4$$

B. Molar Mass

The **molar mass** (M) of a substance is the mass per mole of its entities (atoms, molecules, or formula units). Thus, molar mass has unit of gram per mole ($g \cdot mol^{-1}$). For all substances, the molar mass in grams per mole is equal to the formula weight in atomic mass grams (see Table 1-3).

Table 1-3 Molar mass

Unit	Number	Mass
1.00 mol	6.022×10^{23} atoms	atomic mass of element in grams
	6.022×10^{23} molecules	formula or molecular weight in grams
	6.022×10^{23} ions	formula weight of ionic compound in grams

C. Interconvertion of Mole and Mass

The mole is such a convenient unit for laboratory work, one of the reasons is that it allows to calculate the mass (or the number of chemical entities) of a substance in a sample if we know the amount (number of moles) of the substance. Conversely, if we know the mass (or number of entities) of a substance, we can calculate the number of moles.

The molar mass, which expresses the equivalent relationship between 1 mol of a substance and its mass in grams, can be used as a conversion factor. We multiply a given amount (in moles) by the molar mass of an element or compound (M, in $g \cdot mol^{-1}$) to convert it to mass (in grams).

$$\text{Mass(g)} = \text{Number of moles} \times \text{molar mass}(g \cdot mol^{-1}) \tag{1-1}$$

Or, we divide a given mass (in grams) by the molar mass (multiply by $1/M$) to convert the given mass to amount (in moles).

$$\text{Number of moles} = \frac{\text{mass(g)}}{\text{molar mass}(g \cdot mol^{-1})} \tag{1-2}$$

[Example 1-1] How many moles of NaOH are present in 80 g ?

Solution

The molar mass of NaOH is $40\ g \cdot mol^{-1}$. Since 80 g are about twice the mass of one mole, we can estimate that our final answer will be about two moles of NaOH. The molar mass is used for the exact conversion of grams into moles.

$$\text{Moles of NaOH} = \frac{80\ g}{40\ g \cdot mol^{-1}} = 2\ mol$$

1.3.2 Expression of Concentration in Terms of Molarity

When we dissolve a substance in a liquid, we call the substance the solute and the liquid the solvent. When a solution forms, the solute's individual chemical entities become evenly dispersed throughout the available volume and surrounded by solvent molecules. The general term concentration refers to the quantity of solute in a standard quantity of solution. The quantity of solvent or solution can be expressed in terms of volume or in terms of mass or amount of substance. Thus, there are several ways of expressing the concentration of a solution, such as molarity, molality and mole fraction. Qualitatively, we say that a solution is diluted when the solute concentration is low and concentrated when the solute concentration is high.

A. Molarity

Molarity (molar concentration), denoted M, expresses the concentration in terms of moles of solute per liter of solution.

$$\text{Molarity} = \frac{\text{amount of substance of solute(mol)}}{\text{volume of solution(L)}} \tag{1-3}$$

The molarity of a solution is based on the total volume of solution but not on the volume of solvent. In order to find the molarity of a solution, we need to know the solution volume in liters and the number of moles of solute present.

[Example 1-2] What is the molar concentration (molarity) of a solution prepared by dissolving 4.5 g of NaCl in enough water to form 500 mL of solution?

Solution

To solve this problem, the mass of solute must be converted to moles, and the volume of solution must be expressed in liters.

$$\text{Molarity of NaCl} = \frac{\text{moles of solute}}{\text{volume of solution}} = \frac{4.5\ \text{g}/(58.5\ \text{g} \cdot \text{mol}^{-1})}{0.5\ \text{L}} = 0.154\ \text{mol} \cdot \text{L}^{-1}$$

B. Conversion of Percent Concentration to Molarity

There are three different ways of representing percent concentration:

(1) Percent by mass (or mass-mass percent);

(2) Percent by volume (or volume-volume percent);

(3) Mass-volume percent.

Percent by mass (or mass-mass percent) is the percentage unit most often used in chemical laboratories. It is equal to the mass of solute divided by the total mass of solution multiplied by 100 (to put the value in terms of percentage).

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{1-4}$$

The mass of the solution is equal to the mass of the solute plus the mass of the solvent.

A solution whose mass percent concentration is 5.0% would contain 5.0 g of solute per 100.0 g of solution (5.0 g of solute and 95.0 g of solvent). Thus percent by mass directly gives the number of grams of solute in 100 g of solution. The percent by mass concentration unit is often abbreviated as % (m/m).

Percent by volume (or volume-volume percent), which is abbreviated % (v/v), is used as a concentration unit in situations where the solute and solvent are both liquids or both gases. In these cases, it is more convenient to measure volumes than masses. Percent by volume is equal to the volume of solute divided by the total volume of solution, multiplied by 100%.

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{1-5}$$

The third type of percentage unit is mass-volume percent. This unit, which is often encountered in clinical and hospital settings, is particularly convenient to use when you work with a solid solute, which is easily weighed, and a liquid solvent. Solutions of drugs for internal and external use, intravenous and intramuscular injectables, and reagent solutions for testing are usually labeled in mass-volume percent.

Mass-volume percent, which is abbreviated % (m/v), is equal to the mass of solute (in grams) divided by the total volume of solution (in milliliters), multiplied by 100%.

$$\text{Mass} - \text{volume percent} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\% \tag{1-6}$$

Note that in the definition of mass-volume percent, specific mass and volume units are given. This is necessary because the units do not cancel, as was the case with mass percent and volume percent.

Mass-volume percent indicates the number of grams of solute dissolved in each 100 mL of solution. Thus a 2.3% (m/v) solution of any solute contains 2.3 g of solute in each 100 mL of solution, and a 5.4%(m/v) solution contains 5.4 g of solute in each 100 mL of solution.

[Example 1-3] Concentrated laboratory acid is 35.0% HCl by mass and has a density of 1.18 $g \cdot mL^{-1}$. What is its molarity?

Solution

If we know the number of moles of HCl in a given volume of solution, we can calculate the molarity by using Equation (1-3). Since a volume was not given, we can start with any volume we wish. The molarity will be the same for 1 mL as for 25 L. To make the problem as simple as possible, assume that we have exactly 1 L of solution (V=1 L). The number of moles of HCl (n) in 1 L can be obtained as follows:

(1) Find the mass of 1 L of solution from the density;

(2) Find the mass of HCl in 1 L of solution using the percent by mass;

(3) Convert the mass of HCl to moles of HCl.

The mass of 1 L (10^3 mL) of solution is

$$10^3\ \text{mL} \times 1.18\ \text{g} \cdot \text{mL}^{-1} = 1\ 180\ \text{g}$$

There are 35 g HCl per 100 g of solution since the solution is 35% HCl by mass. The mass of HCl in 1 L of solution is

$$1\ 180\ \text{g} \times 35\% = 413\ \text{g}$$

The number of moles of HCl is

$$\frac{413\ \text{g}}{36.6\ \text{g} \cdot \text{mol}^{-1}} = 11.3\ \text{mol}$$

$$\text{Molarity of HCl} = 11.3\ \text{mol} \cdot \text{L}^{-1}$$

The calculation can all be carried out in one step as follows:

$$\frac{1.18\ \text{g solution}}{1\ \text{mL solution}} \times 1\ 000\ \text{mL solution} \times \frac{35\ \text{g HCl}}{100\ \text{g solution}} \times \frac{1\ \text{mol HCl}}{36.5\ \text{g HCl}} = 11.3\ \text{mol} \cdot \text{L}^{-1}$$

$$\text{Molarity} = \frac{\text{density} \times 1000 \times \%(\text{m/m})}{\text{molar mass}}$$

[Example 1-4] What is the molarity of physiological saline (0.9% (m/v) NaCl)?

Solution

If we know the number of moles of NaCl in a given volume of solution, we can calculate the molarity by using Equation (1-3). Since physiological saline contains 0.9 g of NaCl in each 100 mL of solution, we can start with 100 mL of solution.

$$\text{Molarity of NaCl} = \frac{\text{moles of solute}}{\text{volume of solution}} = \frac{0.9\ \text{g}/(58.5\ \text{g} \cdot \text{mol}^{-1})}{0.1\ \text{L}} = 0.154\ \text{mol} \cdot \text{L}^{-1}$$

$$\text{Molarity} = \frac{10 \times \%(\text{m/v})}{\text{molar mass}}$$

1.3.3 Expression of Concentration in Terms of Molality

The **molality (molal concentration)** of a solution, denoted m, is the moles of solute per kilogram of solvent.

$$\text{Molality(m)} = \frac{\text{moles of solute}}{\text{kilogram of solvent}} \tag{1-7}$$

Thus, if we form a solution by mixing 0.200 mol of NaOH and 0.500 kg of water, the concentration of the solution is 0.200 mol/(0.500 kg) = 0.400 m (that is, 0.400 molal) in NaOH.

The definitions of molarity and molality are similar enough that they may be easily confused. Molarity depends on the volume of solution, whereas molality depends on the mass of solvent. When water is the solvent, the molality and molarity of dilute solutions are numerically about the same because 1 kg of solvent is nearly the same as 1 kg of solution, and 1 kg of the solution has a volume of about 1 L.

The molality of a given solution does not vary with temperature because mass does not vary with temperature. Molarity, however, changes with temperature because the expansion or contraction of the solution changes its volume. Thus molality is often the concentration unit of choice when a solution is to be used over a range of temperatures.

[Example 1-5] Glucose, $C_6H_{12}O_6$, is a sugar that occurs in fruits. It is also known as "blood sugar" because it is found in blood and is the body's main source of energy. What is the molality of a solution containing 5.67 g of glucose dissolved in 25.2 g of water?

Solution

We first convert mass of glucose to moles, because the molality equals moles of solute (glucose) divided by mass of solvent (water) in kilograms.

The moles of glucose (M = 180.2 g • mol^{-1}) in 5.67 g are found as follows:

$$\frac{5.67\ \text{g}}{180.2\ \text{g} \cdot \text{mol}^{-1}} = 0.031\ 5\ \text{mol}$$

The mass of water is 25.2 g(25.2×10^{-3} kg).

$$\text{Molality} = \frac{0.031\ 5\ \text{mol}}{25.2 \times 10^{-3}\ \text{kg}} = 1.25\ \text{mol} \cdot \text{kg}^{-1} = 1.25\ \text{m}$$