1 FORMATION OF THE ATOMIC THEORY

Modern chemistry is based on the atomic theory. To understand the atomic theory, you must first learn the fundamental laws of chemistry. Such fundamental laws include the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. These are the basis of the atomic theory and at the same time represent conclusions drawn from the atomic theory.

However, the atomic theory by itself was incomplete. Chemistry could be a consistent system for the first time only when the atomic theory was combined with the concept of molecules. In the past, the existence of the atom was a mere hypothesis. At the beginning of the 20th century was the atomic theory finally proved. It also became clear that the atom itself consisted of several smaller particles.

Current atomic theory gradually developed along this line and constituted the framework of the material world.

1.1 The birth of chemistry

Modern chemistry was initiated by the French chemist Antoine Laurent Lavoisier (1743-1794). He established the law of conservation of mass for chemical reactions, and revealed the role of oxygen in combustion. Based on these principles, chemistry progressed in the right direction.

Actually oxygen was discovered independently by two chemists, the English chemist Joseph Priestley (1733-1804) and the Swedish chemist Carl Wilhelm Scheele (1742-1786), at the end of the 18th century. Thus, only about two hundred years have passed since modern chemistry was born. In this regard, chemistry is a relatively young science as compared with physics and mathematics, both of which have a history of several thousand years.

However, alchemy, metallurgy and pharmacy in ancient times might be regarded as the roots of chemistry. Many discoveries by those who engaged in these fields contributed to modern chemistry although alchemy was founded on a false theory. Moreover, prior to the 18th century, metallurgy and pharmacy were essentially based on experience alone and were devoid of theory. Thus, it seems impossible that these precursors were gradually developed into modern chemistry. Given these circumstances and the character of modern chemistry with its well-organized and systematic methodology, the true roots of modern chemistry may be found in ancient Greek philosophy.

The road from ancient Greek philosophy to modern atomic theory was not at all straightforward. In ancient Greece, atomic theory as well as denials of the existence of atoms were in severe conflict. In fact, atomic theory long remained unorthodox in the world of chemistry and science. Men of learning were not interested in atomic theory until the 18th century. At the beginning of the 19th century, the English chemist John Dalton (1766-1844) revived the atomic theory of ancient Greece. Even after this revival, not all scientists necessarily approved of the atomic theory. Not until the beginning of the 20th century the atomic theory was finally proved as a fact, not a mere hypothesis. It was accomplished by skilful experiments of the French physicist Jean Baptiste Perrin (1870-1942). Thus, it took a long time for the atomic theory to be established as the basis of modern chemistry.

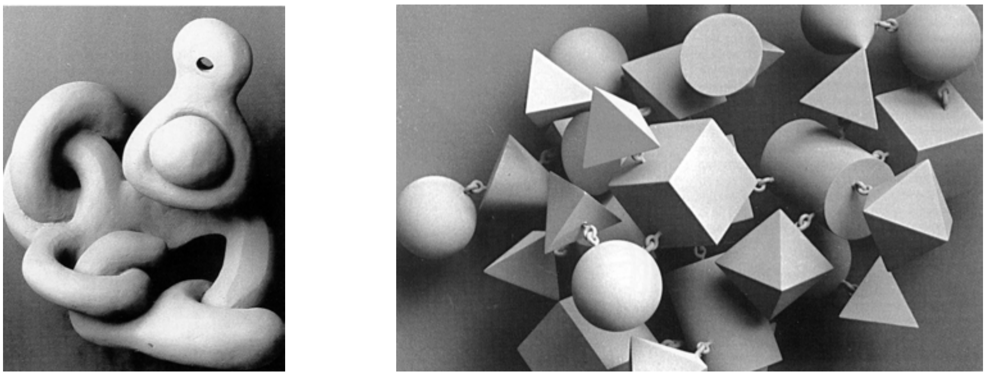
As noted earlier, chemistry is a relatively young science. Consequently, much remains to be done before chemistry can claim to fully understand matter, and through its understanding of matter, nature itself. So, it is very important upon beginning to study chemistry that we review briefly how chemistry has developed since its birth

(a) Ancient atomic theory

As stated above, the roots of modern chemistry lie in the atomic theory developed by ancient Greek philosophers. The philosophy of ancient Greek atomism is attributed to Democritos (*ca*. 460 BC-*ca*. 370 BC). However, no writings of Democritos are extant. Therefore, our resource must be the long poem "*De rerum natura*" written by the Roman poet Lucretius (*ca*. 96 BC-*ca*. 55 BC).

The atom described by Lucretius has some resemblance to the modern molecule. Wine and olive oil, for instance, had their own atoms. The atom was not an abstract entity. Rather, the atom had a definite shape with a function that appropriately corresponded to its shape. "The atom of wine is round and smoothly passes through your throat while the atom of wormwood is jagged and scratches your throat." Modern structural theory of molecules says that there is a very close relation between the structure of a molecule and its functions.

Though the philosophy as articulated by Lucretius was not supported by evidence gained by experiments, it was a precursor of modern chemistry.



(a) (b)

Figure 1. The world of Democritos.

We unfortunately cannot guess the image of atoms as conceived in Democritos' thought. A German chemist has suggested an image of atoms as they might have been conceived by Democritos. (a) An atom of sweet substance. (b) an atom of bitter substance. (reproduced from: F. Berr, W. Pricha, *Atommodelle,* Deutsches Museum, 1987.

In the long period ranging from ancient times to the end of the medieval age, the atomic theory remained heretical because the theory of four elements (water, earth, air and fire) by the ancient Greek philosopher Aristotole (384 BC-322 BC) reigned. When the authority of Aristotle declined at the beginning of the modern age, many philosophers and scientists began to develop theories influenced by ancient Greek atomic theory. The images of matter held by the French philosopher Rene Descartes (1596-1650), the German philosopher Gottfried Wilhelm Freiherr von Leibniz (1646-1716), and the English scientist Sir Issac Newton (1642-1727) were more or less influenced by that atomic theory.

(b) Dalton's atomic theory

At the beginning of the 19th century, atomic theory as a philosophy of matter was already well developed by Dalton who developed his atomic theory based on the role of atoms during chemical reactions. His atomic theory may be summarized as below.

Dalton's atomic theory

(i) The ultimate particle which composes the elements is the atom. All atoms of a particular element are identical.
(ii) The masses of atoms of the same type are equal but that of the other types are different.
(iii) The whole atom is involved in a chemical reaction. The whole atom forms a part of a compound. The type and number of atoms in a given compound are fixed.

The theoretical foundation of Dalton's theory was mainly based on the law of conservation of mass and the law of definite proportion[a], both of which had already been established, and the law of

[a] A given compound always contains exactly the same proportion of elements by weight.

multiple proportions[a] which was developed by Dalton himself.

The atom of Democritos may be said to be as a kind of miniature of matter. Hence the number of the type of atoms is equal to that of the type of matter. On the other hand, Dalton's atom is a constituent of matter, and many compounds are formed by the combination of a limited number of atoms. Hence, there should be a limitation in the number of types of atoms. Dalton's atomic theory requires the process in which two or more atoms combine to form matter. This is the reason why Dalton's atom is called the "chemical atom".

(c) Proof that atoms exist

When Dalton initially proposed his atomic theory, it attracted some attention. However, it failed to gain unanimous support. Some supporters of Dalton made considerable efforts to persuade those who were against the theory, but some opposition remained. The chemistry of his day was not sufficient to prove the existence of atoms by experiments. So the atomic theory remained a hypothesis. Moreover, science after the 18th century developed many experiments that made scientists were very skeptical of the hypothesis. For example, such noted chemists as Sir Humphry Davy (1778-1829) and Michael Faraday (1791-1867), both English, were skeptical towards the atomic theory.

While the atomic theory remained a hypothesis, great advances were made in the field of science. One was the rapid rise of thermodynamics in the 19th century. Structural chemistry of the day as represented by the atomic theory was purely academic with little possibility of practical application. But thermodynamics derived from such a practical issue as the efficiency of steam engines loomed more important. There was a very severe controversy between atomists and those who supported thermodynamics. The debate between the Austrian physicist Ludwig Boltzmann (1844-1906) and the German chemist Friedrich Wilhelm Ostwald (1853-1932) and with the Austrian physicist Ernst Mach (1838-1916) is noteworthy. It had a grave consequence, Boltzmann's suicide.

At the beginning of the 20th century, there was a great change in the popular interest in science. A series of important discoveries, including the discovery of radioactivity, aroused interest in the nature of atoms, and more generally, in structural science. That atoms exist was experimentally confirmed by the sedimentation equilibrium experiments carried out by Perrin

The English botanist Robert Brown (1773-1858) discovered an irregular motion of colloidal particles, and this movement was called **Brownian motion** after him. The Swiss physicist Albert Einstein (1879-1955) developed a theory of this motion based on the atomic theory. According to this theory, the Brownian motion could be expressed by an equation including Avogadro's number.

$$D = \frac{RT}{N}\frac{1}{6\pi\alpha\eta} \tag{1.1}$$

where D is the movement of the particles, R the gas constant, T the temperature, N Avogadro's number, α the radius of the particles and η the viscosity of the solution.

The essence of Perrin's idea is as follows. Colloidal particles move randomly by Brownian motion and simultaneously sink downward by gravitation. **Sedimentation equilibrium** results from the balance of these two motions, random motion and sedimentation by gravity. Perrin carefully observed the distribution of colloidal particles, and with the aid of Eq. 1.1 and this data, he obtained Avogadro's number. Surprisingly the value agreed well with Avogadro's number estimated by completely different methods. This agreement in turn proved the correctness of the atomic theory which was the basis of the theory of Brownian motion.

Needless to say, Perrin could not observe atoms directly. What scientists of those days, including Perrin, could do is to demonstrate that Avogadro's number obtained from various different methods based on the atomic theory is always identical. In other word, they proved the atomic theory rather indirectly by logical consistency.

In the frame of modern chemistry, this kind of methodology is still important. Even today it is

[a] When two elements A and B form a series of compounds, the ratios of the weight of B that combine with a given amount of A can always be reduced to small whole numbers.

impossible to observe directly such small particles as atoms with naked eyes or an optical microscope. To be observed directly by visible light, the size of the particles should be larger than the wavelength of visible light. The wavelength of visible light is in the range of 4.0×10^{-7}- 7.0×10^{-7} m, which is *ca.* 1000 times larger than the size of atoms. Hence it was beyond the range of optical instruments to observe atoms. With the aid of new devices such as the electron microscope (EM) or scanning tunneling microscope (STM), that impossibility was overcome. Though the principle of observing atoms by these devices is, however, different from what is involved in our observing the moon or flowers, we may be allowed to say that we can now observe the image of atoms almost directly.

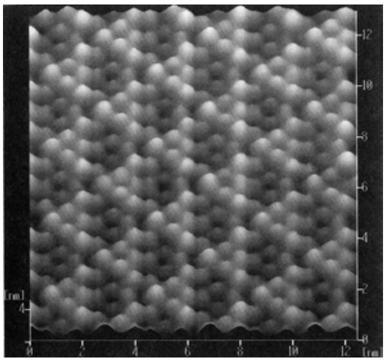


Figure 1.2 At last we observe atoms![a]

The photograph of the surface of silicon crystal observed by STM. Each cell-like block corresponds to a silicon atom. Scale; 2 nm. Reproduced with the permission of the Central Laboratory, Hitachi & Co.

1.2 The components of matter

(a) Atoms

In the world of chemistry based on the atomic theory, the ultimate smallest unit that constitutes matter is the **atom**. Matter is defined as a sum of atoms. The atom is the smallest component of an element that will not suffer from any change during a chemical reaction. All atoms are composed of common components, a **nucleus** and **electrons**. The diameter of a nucleus is *ca.* 10^{-15}-10^{-14} m, *i.e.*, *ca.* 1/10,000 that of an atom. More than 99 % of the mass of an atom is concentrated in the nucleus. The nucleus is composed of proton(s) and neutron(s), and their numbers determine the property of the element.

The mass of a proton is *ca.* 1.67×10^{-27} kg and bears a unit positive charge, 1.60×10^{-19} C (Coulomb). This amount is the smallest unit of electric charge and called the **elementary electric charge**. A nucleus has a positive electric charge whose amount depends on the number of protons it contains. The mass of a neutron is much the same as that of a proton, but a neutron does not possess any electric charge. An electron is a particle with unit negative charge, and a given atom contains the same number of electrons as that of protons in its nucleus. Hence an atom is electrically neutral. The properties of these particles that constitute atoms are summarized in Table 1.1

The number of protons in the nucleus is called the **atomic number** and the sum of the number of protons and the number of neutrons is called the **mass number**. Since a proton and a neutron have much the same mass, and the mass of the electron is negligible as compared with those of a proton and a neutron, the mass of an atom is almost proportional to its mass number.

When atomic number and mass number of a given atom are to be indicated, the atomic number is added to the lower left of the symbol of atoms as a subscript, and the mass number to the upper left of it as a superscript. For instance, the carbon atom is designated as $^{12}_{6}C$ since its atomic number is 6 and the mass number is 12. Sometimes only mass number is indicated, thus as ^{12}C.

[a] You can see more recent STM photos from the URL given below.
http://www.almaden.ibm.com/vis/stm/gallery.html

Table 1.1 Properties of particles that constitute atoms

	mass (kg)	relative mass	electric charge(C)
proton	1.672623×10^{-27}	1836	1.602189×10^{-19}
neutron	1.674929×10^{-27}	1839	0
electron	9.109390×10^{-31}	1	$-1.602189 \times 10^{-19}$

The number of protons and electrons possessed by an **element** determines the chemical properties of that element. The number of neutrons may vary. A given element will always have the same atomic number but may have different numbers of neutrons, and hence a different mass number. These variants are called **isotopes**. As an example, the isotopes of hydrogen are listed in Table 1.2.

Table 1.2 Isotopes of hydrogen

symbol and name	number of protons	number of neutrons
^{1}H hydrogen	1	0
^{2}H deuterium, D	1	1
^{3}H tritium, T	1	2

Many naturally occurring elements have isotopes. Some have more than two isotopes. The chemical properties of isotopes are much the same; only the mass number is different.

(b) Molecules

The smallest component of independently existing neutral matter is called **molecule**. Monatomic molecules are composed of one atom (*e.g.*, Ne). **Polyatomic molecules** are composed of more atoms (*e.g.*, CO_2). The kind of bond between atoms in polyatomic molecules is called a **covalent bond** (*cf.* Ch. 3.2(b)).

One reason why it took some time before atomic theory was fully accepted is as follows. In his theory Dalton admitted the existence of molecules (in modern terminology) which are formed by combinations of atoms of different types, but he did not accept the idea of diatomic molecules for elements such as oxygen, hydrogen or nitrogen which were being extensively investigated at that time. Dalton believed in the so-called "principle of the simplest"[a] and based on this, he automatically assumed that elements such as hydrogen and oxygen are monatomic.

The French chemist Joseph Louis Gay-Lussac (1778-1850) proposed a **law for gaseous reactions** which said that in gaseous reactions, the volume ratio was a whole number. Dalton's atomic theory could not give rationale for the law. In 1811, the Italian chemist Amedeo Avogadro (1776-1856) proposed that gaseous elements such as hydrogen and oxygen were not monatomic but diatomic. Furthermore, he also proposed that at constant temperature and pressure, all gases contained the same number of particles in a given volume. This hypothesis was initially termed **Avogadro's hypothesis**, but later was called **Avogadro's law**.

Avogadro's law could provide the basis for determining the relative masses of atoms, that is, the **atomic weights** (traditionally "weight" has been used instead of "mass"). Its significance was only tardily granted. The Italian chemist Stanislao Cannizzaro (1826-1910) realized the significance of Avogadro's hypothesis and strongly advocated its validity at the first International Chemical Congress held in Karlsruhe, Germany, in 1860, which was called to discuss international agreement for standard atomic weights. Thereafter, the validity of Avogadro's hypothesis was gradually accepted.

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[a] Nature likes the simplest when there is more than one possibility.

(c) Ions

An atom or a group of atoms bearing an electric charge is called an **ion**. A **cation** is an ion having a positive charge, an **anion** one having a negative charge. An electric attraction exists between a cation and an anion. In the crystal of sodium chloride (NaCl), sodium ion (Na^+) and chloride ion (Cl^-) are bonded with that electric attraction. This kind of bond is called an **ionic bond** (*cf*. Ch. 3.2 (a)).

1.3 Stoichiometry

(a) The early stage of stoichiometry

At the beginning of chemistry, the quantitative aspect of chemical change, *i.e*., the **stoichiometry** of a chemical reaction, was not given much attention. Even if some consideration were given, primitive experimental devices and techniques did not yield correct results.

One example involves the **phlogiston theory**. Phlogistonists attempted to explain the phenomenon of combustion in terms of “combustible earth”. According to them, combustion was the release of this combustible earth (from the combustible substance). It was later called “**phlogiston**”. Based on this theory, they defined combustion as the discharge of phlogiston from combustible matter. The change of mass of wood when it was burnt agreed nicely with this theory. However, the change of mass of metals when calcinated disagreed with the theory. In spite of that phlogistonists accepted that the nature of the two processes were essentially identical. The increase of mass of calcinated metals was a fact. The phlogistonists tried to explain this anomaly by assigning to phlogiston a negative mass

The Flanders philosopher Jan Baptista van Helmont (1579-1644) carried out a famous “willow” experiment. He grew a seedling of willow after measuring the combined weight of flowerpot and soil. Since there was no change in the weight of the flowerpot and the soil while the seedling grew up, he considered that the gain in weight was due only to the water he gave to the seedling. He concluded that “the root of all matter is water”. Judged from today’s viewpoint, his hypothesis and experiment were far from perfect, but this story is a good example of an emerging attitude for the quantitative aspects of chemistry. Helmont recognized the importance of stoichometry, and in that regard he was well ahead of his times.

At the end of the 18th century, the German chemist Jeremias Benjamin Richter (1762-1807) established the concept of **equivalent** (in modern terminology **chemical equivalent**) by careful examination of acid/base reactions, that is, the quantitative relation between the acid and the base in neutralization reactions. His equivalent, or now chemical equivalent, indicates a certain amount of matter in the reaction. An equivalent in neutralization corresponds to the relation between a given amount of acid and the amount of base to neutralize it. Knowledge of the correct equivalent was essential in order to produce good soap or gunpowder. Thus, such knowledge was of great practical importance.

At the same time Lavoisier established the law of conservation of mass, and gave a basis for the concept of equivalent by his ingenious and accurate experiments. Thus, stoichiometry which handles the quantitative aspects of chemical reactions became the fundamental methodology of chemistry. All fundamental laws of chemistry, from the laws of conservation of mass, of definite composition and of multiple proportions to the law for gaseous reactions, are all based on stoichiometry. These fundamental laws were the basis of atomic theory, and were consistently explained in terms of atomic theory. It is interesting to note, however, that the concept of equivalent was used before the atomic theory was introduced.

(b) Relative masses of atoms and atomic weights

Dalton recognized that it was necessary to determine the weight of each atom since it varied for each type of atom. Atoms were so small that it was impossible to determine the weight of one atom. So, he focused on the relative values of the weight and made a table of atomic weights (Fig. 1.3) for the first time in human history. In this table, the weight of the lightest element, hydrogen, was set to unity as the standard (H = 1). Atomic weight is a relative value, *i.e.,* a ratio without any dimension. Though some of his atomic weights were differed from modern values, most of his values were in the allowable range. This in turn shows that his ideas and experiments were correct.

Figure 1.3 Dalton's table which listed the symbols and atomic weight of elements
The table was prepared in 1807, and is one of the treasures in The Science Museum in London.

Later the Swedish chemist Jons Jakob Baron Berzelius (1779-1848) determined a series of atomic weights with oxygen as the standard (O = 100). Since Berzelius obtained these values chiefly by the analysis of oxides, he had a good reason to choose oxygen as the standard. However, the hydrogen standard is superior for simplicity. Today, after many discussions and modifications, the carbon standard is used. In it, the mass of carbon isotope ^{12}C with 6 protons and 6 neutrons is defined as 12.0000. The atomic weight of an atom is its mass relative to this standard. Though carbon has been chosen as the standard, in fact it may be regarded as a modified hydrogen standard.

Sample Exercise 1.1 The change of atomic weight caused by the change in the standard.

Calculate the atomic weight of hydrogen and carbon when the standard of Berzelius (O = 100) is adopted. The answer should be given to one decimal place.

Solution

atomic weight of hydrogen = 1 x (100/16) = 6.25 (6.3), atomic weight of carbon = 12 x (100/16) =75.0

The atomic weight of most elements is very close to an integer, that is, an integral multiple of the atomic weight of hydrogen. This is a natural consequence of the fact that the atomic weight of hydrogen is almost equal to the mass of a proton, which in turn is almost equal to the mass of a neutron, and the mass of an electron is negligibly small. However, most of the naturally occurring elements are mixtures of several isotopes, and the atomic weight depends on the distribution of the isotopes. For instance, the exact atomic weights of hydrogen and oxygen are 1.00704 and 15.9994, respectively. The atomic weight of oxygen is very close to 16 but slightly less.

Sample Exercise 1.2 Calculation of atomic weight.

Calculate the atomic weight of magnesium using the following isotope distribution
isotope distribution: ^{24}Mg; 78.70%; ^{25}Mg; 10.13%, ^{26}Mg; 11.17%.

Solution

0.7870 x 24 = 18.89
0.1013 x 25 = 2.533
0.1117 x 26 = 2.904

18.89 + 2.533 + 2.904 =24.327(amu; *cf* Ch. 1.3(e))
Atomic weight of Mg = 18.89 + 2.533 + 2.904 =24.327 (amu)

The small difference from the atomic weight found in the Periodic Table (24.305) resulted from a different way of rounding off the numbers.

(c) Chemical formula weight and molecular weight

Each compound is defined by a chemical formula indicating the type and the number of atoms constituting that compound. **Formula weight** (or chemical formula weight) is defined as the sum of atomic weight based on the type and the number of atoms defined by the chemical formula. The chemical formula of a molecule is called **molecular formula**, and its chemical formula weight is termed **molecular weight**.[a)]

For instance, the molecular formula of carbon dioxide is CO_2, and the molecular weight is 12 + (2 x 16) = 44. As with atomic weights, both formula weight and molecular weight are not necessarily integers. For instance, the molecular weight of hydrogen chloride HCl is 36.5. Even if the type and the number of atoms constituting the molecule are identical, the two molecules may have different molecular weights if different isotopes are involved.

Sample Exercise 1.3 Molecular weight of molecules containing isotopes.

Calculate the molecular weights of water H_2O and heavy water D_2O (^2H_2O) in round numbers.

Solution

Molecular weight of H_2O = 1 x 2 + 16 = 18, molecular weight of D_2O = (2 x 2) + 16 = 20

The difference in molecular weight of H_2O and D_2O is substantial, and the differences in physical and chemical properties between these two compounds are not negligible. H_2O is more readily electrolyzed than D_2O. Thus, the water remaining after electrolysis tends to contain more D_2O than natural water.

(d) Quantity of matter and mole

The most suitable method of quantitatively expressing the amount of matter is the number of particles such as atoms or molecules which constitute the matter in question. However, to count such small and invisible particles as atoms or molecules would be very difficult. Instead of counting the number of particles directly, we can use the mass of a definite number of particles.

Then, how was such a definite number chosen? To make a long story short, the number of particles contained in a gas of 22.4 dm^3 at standard pressure and temperature (0℃, 1atm) was chosen as this standard number. This number is called **Avogadro's number** for historical reasons. The name **Loschmidt number** was also proposed after the name of the Austrian chemist Joseph Loschmidt (1821-1895) who first determined this number by experiment (1865).

Definition of "mol"

(i) The amount of matter that contains the same number of particles as that contained in 12 g of ^{12}C.
(ii) One mol of matter contains the same number of particles as Avogadro's constant.
(iii) The amount of matter constituted in 6.02 x 10^{23} particles is one mol.

Since 1962, according to SI (Systeme Internationale) it was decided that, in the chemical world, **mole** (abbreviation mol) is to be used as the unit of the amount of matter. Avogadro's number was defined as the number of carbon atoms contained in 12 g of $^{12}_{6}C$ and renamed **Avogadro's constant**. There are several definitions of "mol".

[a)] It is impossible to define a molecule for compounds such as sodium chloride. The formula weight of NaCl is used as the alternative of the molecular weight.

(e) Atomic mass unit (amu)

As the standard of the mass of atoms in Dalton's system is the mass of hydrogen, the standard of the mass of atoms in SI is exactly 1/12 of the mass of ^{12}C. This value is called the **atomic mass unit** (amu) and is equal to $1.6605402 \times 10^{-27}$ kg and D (Dalton) was assigned as its symbol. The atomic weight of an atom is defined as the ratio of the average amu of the element with its natural isotope distribution *vs*. 1/12 of amu of ^{12}C.

Coffee Break: Who founded chemistry?

Priestley, the discoverer of oxygen, was a clergyman and Scheele ran a pharmacy. Lavoisier was a licensed tax collector and Dalton was a teacher at a private school. They had to provide the money for their research by themselves. Davy and Faraday were exceptional in that they were employed by the Royal Institution. It was the only place in England at that time where one could be paid for their services in the field of chemistry.

A systematic training of chemists was initiated at Giesen University, Germany. It was there that the German chemist Justus von Liebig (1803-1873) organized the first novel curriculum for chemists by combining lectures and experiments by students themselves, thus establishing the modern chemical education curriculum. He taught and trained a good number of chemists, and many of them entered chemical industry. Thus, chemistry became a profession.

Exercise

1.1 Isotopes

The naturally occurring carbon is a mixture of two isotopes, 98.90(3)% of ^{12}C and 1.10(3)% of ^{13}C. Calculate the atomic weight of carbon.

1.1 Answer

Atomic weight of carbon = 12 x 0.9890 + 13 x 0.0110 = 12.01(1)

1.2 Avogadro's constant

Diamond is a pure simple substance of carbon. Calculate the number of carbon atoms contained in 1 carat (0.2 g) of diamond.

1.2 Answer

The number of carbon atoms = $[0.2\ (g)/12.01\ (g\ mol^{-1})] \times 6.022 \times 10^{23} (mol^{-1}) = 1.00 \times 10^{22}$

1.3 Law of multiple proportions

The chemical compositions of three oxides of nitrogen A, B and C were examined. Show the results are consistent with the law of multiple proportions: the mass of nitrogen combined with 1 g of oxygen in each oxide. Oxide A; 1.750 (g), oxide B; 0.8750 (g), oxide C; 0.4375 (g).

1.3 Answer

If the law of multiple proportions is valid, the ratio of mass of nitrogen bonded to 1 g of oxygen should be a simple integer.

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1};\ \frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1};\ \frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

The result is in agreement with the law of multiple proportions.

1.4 The atomic weight

Naturally occurring copper was analyzed by a mass spectrometer. The results were: ^{63}Cu 69.09%, ^{65}Cu 30.91%. Calculate the atomic weight of Cu. The mass of ^{63}Cu and ^{65}Cu is 62.93 and 64.93 amu, respectively.

1.4 Answer

The atomic weight of Cu can be obtained as below.

62.93x (69.09/100) + 64.93x (30.91/100) = 63.55 (amu)

1.5 Mole

When a bee stings the victim, it discharges *ca*. 1 mg (1x 10^{-6} g) of isopentyl acetate $C_7H_{14}O_2$. This compound is the fragrant component of bananas, and serves as the information-transfer material to call other bees. How many molecules of isopentyl acetate are there in 1 mg of isopentyl acetate?

1.5 mole

The molecular weight of isopentyl acetate can be calculated as below.

$M = 7 \times 12.01 + 14 \times 1.008 + 2 \times 16.00 = 130.18$ (g mol^{-1})

mole: 1.0×10^{-6}(g)/130.18(g mol^{-1}) = 7.68×10^{-9}(mol)

the number of molecules of isopentyl acetate: 7.68×10^{-9}(mol) x 6.022×10^{23} (mol^{-1}) = 4.6×10^{15}

1.6 The mass of a hydrogen molecule

The atomic weight of hydrogen is 1.008. Calculate the mass of a hydrogen molecule.

1.6 Answer

The molar mass of hydrogen is 2.016×10^{-3} kg mol^{-1}. The mass of one molecule *m* is given below. m = [2.016×10^{-3} (kg mol^{-1})]/[6.022×10^{23}(mol^{-1}) = 3.35×10^{-27}(kg)