Abdelhafid Boussouf University Centre of Mila

Institute of Natural and Life Sciences,

Common Core Department

Abdelhafid Boussouf University Centre of Mila

General Chemistry

Course Support

CHAPTER I : ATOMISTICS

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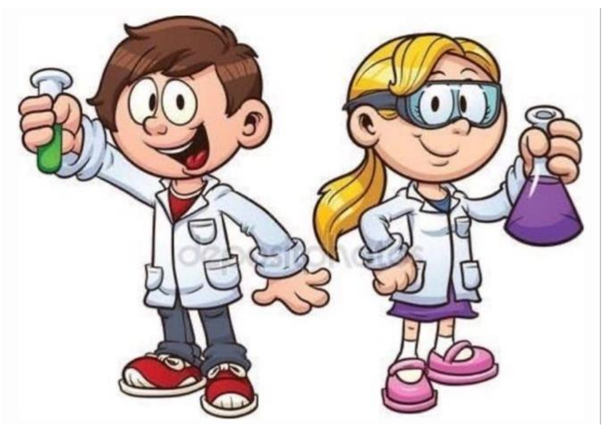

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I. The atoms:

This chapter will describe some of the fundamental chemical principles related to the composition of matter, including those central to the concept of molecular identity.

I.1 The atom:

The atom is the smallest unit of matter that is composed of three sub-atomic particles: the proton, the neutron, and the electron. Protons and neutrons make up the nucleus of the atom, a dense and positively charged core, whereas the negatively charged electrons can be found around the nucleus in an electron cloud.

I.2 The Mole:

The mole is the unit of measurement in the International System of Units (SI) for amount of substance. It is defined as the amount of a chemical substance that contains as many elementary entities (e.g., atoms, molecules, ions, electrons, or photons). This number is expressed by the Avogadro constant, which has a value of $6.022140857 \times 10^{23} mol^{-1}$. The mole is one of the base units of the SI, and has the unit symbol mol.

I.3 Avogadro's Number

Avogadro's number, is a fundamental constant in chemistry and physics that represents the number of atoms, molecules, or entities in one mole of a substance. It is named after the Italian scientist Amedeo Avogadro, who first proposed the concept that equal volumes of gases, at the same temperature and pressure, contain the same number of molecules.

Avogadro's number is approximately 6.022×10^{23} entities per mole. This means that in 12 grams of carbon-12 (a specific isotope of carbon), there are precisely Avogadro's number of carbon atoms.

I.4 Atomic Mass Unit

An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12. The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of 4.0026 amu. An atom of sulfur-32 has a mass of 31.972 amu.

II. Fundamental Chemical Laws:

II.1 Law of Conservation of Mass (Antoine Lavoisier (1743-1794)

"**Nothing comes from nothing**" is an important idea in ancient Greek philosophy that argues that what exists now has always existed, since no new matter can come into existence where there was none before.

In the 1790s, a greater emphasis began to be placed on the quantitative analysis of chemical reactions. Accurate and reproducible measurements of the masses of reacting elements and the compounds they form led to the formulation of several basic laws. One of these is called the law of conservation of mass, which states that during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants. In other words, **mass cannot be created or destroyed during a chemical reaction, but is always conserved.**

If we react 12 grams of Carbon with 32 grams of Oxygen, we find that we have formed 44 grams of carbon dioxide CO_2. So, the total mass of reactants equals the total mass of products, a proof of the law of conservation of mass.

II.2 Definite Proportions (Joseph Proust (1754-1826)

Joseph Proust (1754-1826) formulated the law of definite proportions (also called the Law of Constant Composition or Proust's Law). This law states that if a compound is broken down into its constituent elements, the masses of the constituents will always have the same proportions, regardless of the quantity or source of the original substance

The Law of Definite Proportions applies when elements are reacted together to form the same product. As an example, any sample of pure water contains 11.19% hydrogen and 88.81% oxygen by mass. It does not matter where the sample of water came from or how it was prepared. Its composition, like that of every other compound, is fixed.

II.3 Multiple Proportions (John Dalton about 1803)

The law of multiple proportions states that if two elements form more than one compound between them, the masses of one element combined with a fixed mass of the second element form in ratios of small integers.

<u>Example: Oxides of Carbon</u>

Consider two separate compounds are formed by only carbon and oxygen. The first compound contains 42.9% carbon and 57.1% oxygen (by mass) and the second compound contains 27.3% carbon and 72.7% oxygen (again by mass). Is this consistent with the law of multiple proportions?

Solution

The Law of Multiple Proportions states that the masses of one element which combine with a fixed mass of the second element are in a ratio of whole numbers. Hence, the masses of oxygen in the two compounds that combine with a fixed mass of carbon should be in a whole-number ratio.

Thus, for every 1 g of the first compound there are 0.57 g of oxygen and 0.429 g of carbon. The mass of oxygen per gram carbon is:

$$\frac{0.571\text{g oxygen}}{0.429\text{g carbon}}=1.33\,\frac{\text{g oxygen}}{\text{g carbon}}$$

Similarly, for 1 g of the second compound, there are 0.727 g oxygen and 0.273 g of carbon. The ration of mass of oxygen per gram of carbon is

$$\frac{0.727\text{g oxygen}}{0.273\text{g carbon}}=2.66\,\frac{\text{g oxygen}}{\text{g carbon}}$$

Dividing the mass of oxygen per g of carbon of the second compound: $\frac{2.66}{1.33}=2$

Hence the masses of oxygen combine with carbon in a 2:1 ratio which consistent with the Law of Multiple Proportions since they are whole numbers.

II.4 Conservation of Elements

Both the initial and final substances are composed of atoms because all matter is composed of atoms. According to the law of conservation of matter, matter is neither created nor destroyed, so we must have the same number and kind of atoms after the chemical change as were present before the chemical change.

For example, hydrogen (H_2) and oxygen (O_2) can react to produce water (H_2O): $2H_2(g)+O_2(g)=>2H_2O(g)$

We can see from the reaction above that atoms of the elements hydrogen and oxygen are present in both the starting materials and the product.

II.5 The Law of Combining Volumes (Gay-Lussac 1808)

The law of combining volumes is also known as Gay-Lussac's law. When gases combine at constant temperature and pressure, the volumes involved are always in the ratio of simple whole numbers.

Example:

In the reaction $2H_2(g)+O_2(g)=>2H_2O(g)$

means that for every 2 mol $H_2(g)$ consumed there will be 1 mol $O_2(g)$ consumed and 2 mol $H_2O(g)$ produced.2 volumes of H_2 react with 1 volume of O_2 to produce 2 volumes of H_2O.

III. The Structure of Atoms

An atom consists of a positively charged nucleus surrounded by one or more negatively charged particles called electrons. The number of protons found in the nucleus equals the number of electrons that surround it, giving the atom a neutral charge (neutrons have zero charge). Most of an atom's mass is in its nucleus; the mass of an electron is only 1/1836 the mass of the lightest nucleus, that of hydrogen. Although the nucleus is heavy, it is small compared with the overall size of an atom.

The radius of a typical atom is around 1 to 2.5 angstroms (Å), whereas the radius of a nucleus is about 10^{-5} Å. If an atom were enlarged to the size of the earth, its nucleus would be only 200 feet in diameter and could easily fit inside a small football stadium.

The nucleus of an atom contains protons and neutrons. Protons and neutrons have nearly equal masses, but they differ in charge. A neutron has no charge, whereas a proton has a positive charge that exactly balances the negative charge on an electron. Table 1 lists the charges of these three sub atomic particles, and gives their masses expressed in atomic mass units. The atomic mass unit (amu) is defined as exactly one-twelfth the mass of a carbon atom that has six protons and six neutrons in its nucleus. On this scale, protons and neutrons have masses that are close to, but not precisely, 1 u each. In fact, there are 6.022×10^{23} u in 1 gram. This number is known as Avogadro's number, $\mathcal{N}$. The number of protons in the nucleus of an atom is known as the atomic number, Z. It is the same as the number of electrons around the nucleus of the electrically-neutral atom. The mass number of an atom is equal to the total number of protons and neutrons.

Table 1: Charge and mass of three sub atomic particles

Particle	Charge	Mass (grams)
Electrons	-1	9.1094×10^{-28}
Protons	+1	1.6726×10^{-24}
Neutrons	0	1.6749×10^{-24}

III.1 Nucleus

The **nucleus** of an atom is comprised of protons and neutrons; it is therefore positively charged. The number of protons within the nucleus of a given atom is equal to the **atomic number** of the corresponding element, which can be found on the periodic table. For example, the atomic number of helium is two. Therefore, the number of protons is also two. The number of neutrons within the nucleus of a given atom can be found by subtracting the atomic number from the atomic mass. The **mass number** is the sum of protons and neutrons.

Atomic Mass Number = Number of Protons + Number of Neutrons

To find the number of neutrons, subtract the atomic number, the number of protons, from the mass number.

Notation of a specific element follows this format: $^{A}_{Z}\mathbf{X}$

where E is a specific element, A is mass number, Z is the atomic number, and C is the charge.

For helium, the notation is as follows: $^{4}_{2}\mathbf{He}$

Helium has 2 protons, 2 neutrons and a charge of zero.

III.2 Isotopes

Atoms of the same element that have a different number of neutrons are known as **isotopes**. Most elements have several naturally occurring isotopes. The atomic mass of a particular element is equal to the average of the relative abundance of all its isotopes found in nature.

$$\textbf{average mass} = \frac{\sum \mathbf{MiXi}}{\mathbf{100}}$$

For example, there are three naturally occurring isotopes of carbon: carbon-12, carbon-13, and carbon-14. Carbon-12 is the most common of these three, making up about 98.89% of all carbon, whereas carbon-13 has 1.11% natural abundance. Carbon-14 occurs rarely in nature 10^{-4} %. Atomic masses for other elements use the carbon-12 scale as a reference. Early physicists assigned the atomic mass of 12 to the carbon-12 isotope (which is the most common carbon isotope) so that it would be easier to determine the atomic masses of other atoms. Using this information, we can determine the average atomic mass of carbon. (Use 13 for the approximate mass of carbon-13.)

average atomic mass = mass of carbon-12 x (% natural abundance/100) + mass of carbon-13 x (% natural abundance/100)

$$= 12 \times 0.9889 + 13 \times 0.0111 + 14 \times 10^{-4} = 12.0111$$

An example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure III-1. Most hydrogen atoms have just one proton and one electron and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

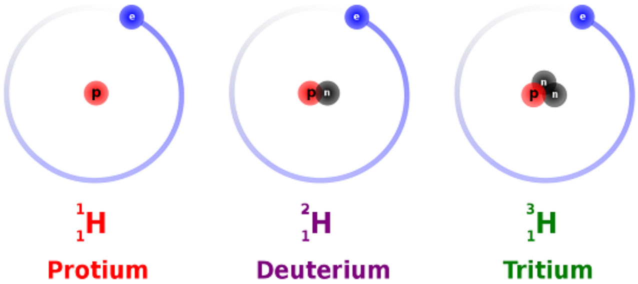


Figure III-1The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3) Table 2 gives examples of common isotopes and their percent abundances in nature:

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below:

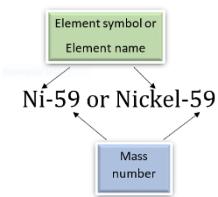
Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}Ni\ ,\ ^{238}_{92}U$$

In the nickel nucleus represented above, the atomic number 28 indicates the nucleus contains 28 protons, and therefore, it must contain 31 neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons as do all uranium nuclei and this particular uranium nucleus has 146 neutrons.

Another way of representing isotopes is by adding a hyphen and the mass number to the chemical name or symbol. Thus, the two nuclei would be Nickel-59 or Ni-59 and Uranium-238 or U-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) are given to the side of the name.



One way to represent isotopes: The element symbol or name comes first, then a hyphen, then the mass number

Example

How many protons, electrons, and neutrons are in an atom of ${}^{40}_{19}K$?

Solution

atomic number= (number of protons) =19

For all atoms with no charge, the number of electrons is equal to the number of protons.

number of electrons=19

The mass number, 40 is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

number of neutrons=40−19=21.

III.3 Binding Energy

As a simple example of the energy associated with the strong nuclear force, consider the helium atom composed of two protons, two neutrons, and two electrons. The total mass of these six subatomic particles may be calculated as:(Example 1)

$$\underbrace{(2 \times 1.0073 \text{ amu})}_{\text{protons}} + \underbrace{(2 \times 1.0087 \text{ amu})}_{\text{neutrons}} + \underbrace{(2 \times 0.00055 \text{ amu})}_{\text{electrons}} = 4.0331 \text{ amu}$$

However, mass spectrometric measurements reveal that the mass of an atom is 4.0026 amu, less than the combined masses of its six constituent subatomic particles. This difference between the calculated and experimentally measured masses is known as the mass defect of the atom. In the case of helium, the mass defect indicates a “loss” in mass of 4.0331 amu – 4.0026 amu = 0.0305 amu. The loss in mass accompanying the formation of an atom from protons, neutrons, and electrons is due to the conversion of that mass into energy that is evolved as the atom forms. The nuclear binding energy is the energy produced when the atom’s nucleons are bound together; this is also the energy needed to break a nucleus into its constituent protons and neutrons.

The conversion between mass and energy is most identifiably represented by the mass-energy equivalence equation as stated by Albert Einstein: $E=mC^2$

where E is energy, m is mass of the matter being converted, and c is the speed of light in a vacuum. This equation can be used to find the amount of energy that results when matter is converted into energy. Using this mass-energy equivalence equation, the binding energy of a nucleus may be calculated from its mass defect, as demonstrated in Example 1. A variety of units are commonly used for nuclear binding energies, including electron volts (eV), with 1 eV equaling the amount of energy necessary to the move the charge of an electron across an electric potential difference of 1 volt, making 1eV= 1,602x 10^{-19}j.

III.4 Stability

A nucleus is stable if it cannot be transformed into another configuration without adding energy from the outside. Of the thousands of nuclides that exist, about 250 are stable. A plot of the number of neutrons versus the number of protons for stable nuclei reveals that the stable isotopes fall into a narrow band. This region is known as the band of stability (also called the belt, zone, or valley of stability). Note that the lighter stable nuclei, in general, have equal numbers of protons and neutrons. For example, nitrogen-14 has seven protons and seven neutrons. Heavier stable nuclei, however, have increasingly more neutrons than protons. For example: iron-56 has 30 neutrons and 26 protons, an n:p ratio of 1.15, whereas the stable nuclide lead-207 has 125 neutrons and 82 protons, an n:p ratio equal to 1.52. This is because larger nuclei have more proton-proton repulsions, and require larger numbers of neutrons to provide compensating strong forces to overcome these electrostatic repulsions and hold the nucleus together.

The relative stability of a nucleus is correlated with its binding energy per nucleon, the total binding energy for the nucleus divided by the number or nucleons in the nucleus. For a $^{4}_{2}He$ instance, the binding energy for a nucleus is therefore:

$$\frac{28.4\ \text{MeV}}{4\ \text{nucleons}} = 7.10\ \text{MeV/nucleon}$$

Example

The iron nuclide $^{56}_{26}Fe$ lies near the top of the binding energy curve and is one of the most stable nuclides. What is the binding energy per nucleon (in MeV) for the nuclide $^{56}_{26}Fe$ (atomic mass of 55.9349 amu)?

Solution

As in Example, we first determine the mass defect of the nuclide, which is the difference between the mass of 26 protons, 30 neutrons, and 26 electrons, and the observed mass of an $^{56}_{26}Fe$ atom:

Mass defect = [(26×1.0073 amu) +(30×1.0087 amu) +(26×0.00055 amu)] −55.9349 amu

= 56.4651 amu−55.9349 amu

= 0.5302 amu

We next calculate the binding energy for one nucleus from the mass defect using the mass-energy equivalence equation:

$$\begin{aligned} E &= mc^2 = 0.5302\ \text{amu} \times \frac{1.6605 \times 10^{-27}\ \text{kg}}{1\ \text{amu}} \times (2.998 \times 10^8\ \text{m/s})^2 \\ &= 7.913 \times 10^{-11}\ \text{kg}\cdot\text{m/s}^2 \\ &= 7.913 \times 10^{-11}\ \text{J} \end{aligned}$$

We then convert the binding energy in joules per nucleus into units of MeV per nuclide:

$$7.913 \times 10^{-11}\ \text{J} \times \frac{1\ \text{MeV}}{1.602 \times 10^{-13}\ \text{J}} = 493.9\ \text{MeV}$$

Finally, we determine the binding energy per nucleon by dividing the total nuclear binding energy by the number of nucleons in the atom:

$$\text{Binding energy per nucleon} = \frac{493.9\ \text{MeV}}{56} = 8.820\ \text{MeV/nucleon}$$