2 ATOMIC SYSTEMATICS AND NUCLEAR STRUCTURE

In this chapter the principles and systematics of atomic and nuclear physics are summarised briefly, in order to introduce the existence and characteristics of isotopes.

2.1 ATOMIC STRUCTURE AND THE PERIODIC TABLE OF THE ELEMENTS

Atoms consist of a nucleus surrounded by electrons. Compared to the diameter of an atom, which is of the order of 10^{-8} cm, the size of the nucleus is extremely small (~ 10^{-12} cm). The dense concentration of matter of the nucleus mainly consists of two kinds of particles, neutrons and protons, which have about the same mass. The neutron carries no electric charge, while the proton is positively charged. The number of protons (Z), the *atomic number*, is equal to the number of electrons surrounding the nucleus. Electrons have a mass that is about 1/1800 that of the proton mass and carry an equal but negative electrical charge, so that the atom as a whole is neutral. Atoms missing one or more electrons are referred to as positive ions, atoms with a number of electrons exceeding the atomic number are called negative ions.

Protons and neutrons, the building stones of the nucleus, are called *nucleons*. The sum of the number of protons and neutrons (N) in a nucleus is the nuclear *mass number*:

$$A = Z + N \qquad \textbf{(2.1)}$$

The notation describing a specific nucleus (= *nuclide*) of element X is:

$${}^{A}_{Z}X_{N}$$

Because the chemical properties of an element (X) are primarily determined by the number of electrons in the atom, the atomic number Z characterises the element. Therefore, writing ${}^{A}X$ alone defines the nuclide. The cloud of electrons circulating around the nucleus is well structured and consists of shells, each containing a maximum number of electrons. The chemical properties of an atom are now mainly determined by the number of electrons in the outer, incompletely filled electron shell. Because of this systematic, all atoms can be arranged in a *Periodic Table of the Elements* (part shown by Fig. 2.1).

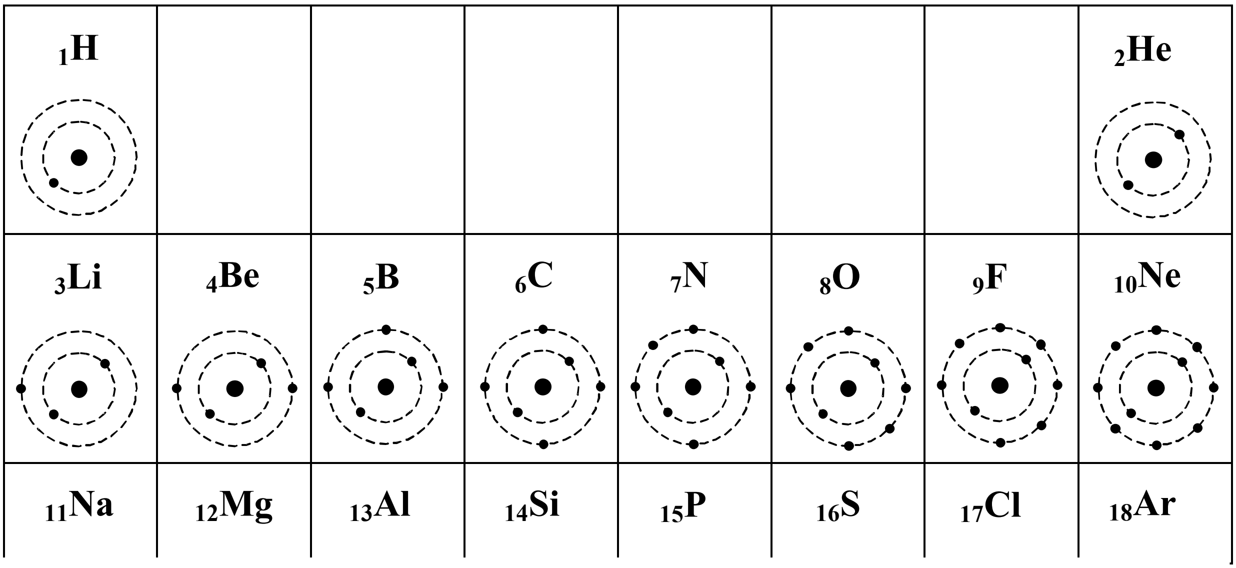

Fig. 2.1 Part of the periodic table of the elements, containing the light elements. Also shown are the electronic configurations of the respective atoms.

2.2 STRUCTURE OF THE ATOMIC NUCLEUS

The atomic nuclei are kept together by extremely strong forces between the nucleons (protons and neutrons) with a very small range. As repulsive electrical (Coulomb) forces exist between the protons, the presence of neutrons is required to stabilise the nucleus. In the most abundant nuclides of the *light elements*, the numbers of protons and neutrons are equal. Nuclei such as

$$^{2}_{1}H_{1} \qquad ^{4}_{2}He_{2} \qquad ^{12}_{6}C_{6} \qquad ^{14}_{7}N_{7} \qquad ^{16}_{8}O_{8}$$

are stable, as is the single proton (^{1}H = hydrogen). For the heavy elements the number of neutrons far exceeds the number of protons: ^{238}U contains only 92 protons, whereas the largest stable nuclide, the lead isotope ^{208}Pb has an atomic number of 82.

Instabilities are caused by an excess of protons or neutrons. Examples of such unstable or radioactive nuclei are

$$^{3}_{1}H_{2} \quad \text{and} \quad ^{14}_{6}C_{8}$$

For the light elements, a slight excess of neutrons does not necessarily result in unstable nuclei:

$$^{13}_{6}C_{7} \ , \ ^{15}_{7}N_{8} \ , \ ^{17}_{8}O_{9} \ , \ ^{18}_{8}O_{10}$$

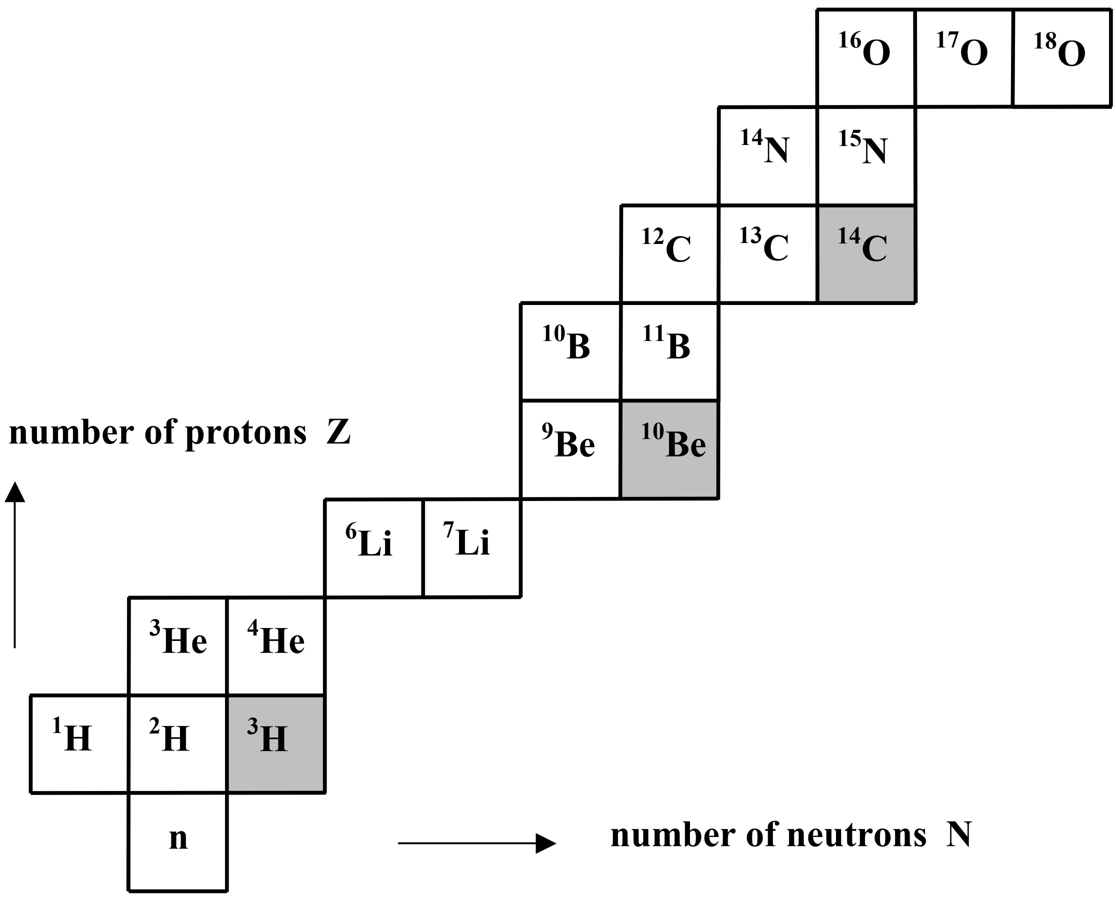

Fig. 2.2 Part of the chart of nuclides, containing the light elements. The isotopes of an element (equal Z) are found in a horizontal row, isobars (equal A) along diagonal lines, isotones (equal N) in vertical columns. The natural radioactive isotopes of H, Be, and C are marked grey.

are stable. For these "asymmetric" nuclei (Z≠N) however, the probability of formation during the "creation" of the elements, -nucleosynthesis-, was smaller, resulting in smaller natural concentrations for these nuclides.

The naturally occurring stable and radioactive isotopes of the light elements are shown in a part of the *Chart of Nuclides* (Fig. 2.2). Another part, showing some heavy nuclides is given by Fig.11.2. Note that in the higher A range N and Z are no longer about equal (^{238}U with Z=92 and N=146).

2.3 STABLE AND RADIOACTIVE ISOTOPES

The atomic nuclei of an element containing different numbers of neutrons are called *isotopes* (ισο τοπος = at the same place in the periodic table of the elements). Many elements have two or more stable, naturally occurring isotopes. In general, nuclei with even numbers of protons or/and neutrons are more stable. Nuclei of which the number of protons or/and neutrons corresponds to some specific even number, belonging to the series of so-called *magic numbers* 2, 8, 20, 28, 50, 82, and 126, have a relatively high stability and consequently large natural occurrence.

These magic numbers can be explained by a nuclear shell model with *closed nucleon shells* within the nucleus, similar to the closed electron shells in atoms, the basis of the periodicity in the periodic table of the elements. An example of the occurrence of the magic numbers is the large number of stable isotopes of lead: the largest stable nuclide, ^{208}Pb (with Z = 82 and N = 126) is double magic. Especially the uneven-uneven nuclei are unstable and have a small chance of natural occurrence. Most uneven-Z elements have only one or at most two stable isotopes.

The earlier statement that the chemical properties of an element only depend on the atomic number, implying that the chemical properties of isotopes are equal, needs revision if we look in detail. The fact is that in nature variable relative concentrations of isotopes are observed. There are two causes for this phenomenon:

1) the chemical and some physical properties of the isotopes of one element are not exactly equal, resulting in slightly different chemical and physical properties -and consequently different concentrations- of isotopic molecules, i.e. molecules that contain different isotopes of that element

2) if the isotopes concerned are radioactive, the process of radioactive decay causes the concentration of the isotopic molecules concerned to diminish in time; this may result in concentration differences that are much larger than as caused by the isotope processes as mentioned under 1).

These phenomena will be discussed separately in the next chapters.

2.4 MASS AND ENERGY

It is inconvenient to use the real mass of atoms and molecules. Instead we define the *atomic mass* as the mass expressed in atomic mass units (amu). Originally this was equivalent to the mass of a proton; later for practical reasons, the atomic mass unit has rather become defined as 1/12 times the mass of a ^{12}C atom:

$$1 \text{ amu} = 1.6605655 \times 10^{-27} \text{ kg} \qquad \textbf{(2.2)}$$

In chemistry it is now convenient to use the *mole* quantity, defined as the number of grams of the element equal to the atomic weight. The number of atoms in one mole of the element or the number of molecules in one mole of a chemical compound is then equal for any substance and given by Avogadro's number = 6.02252×10^{23}.

If we now add up the atomic masses of the building stones of a certain nucleus X, for instance of ^{12}C consisting of 6 protons and 6 neutrons, it appears that the atomic mass of X is smaller than the sum of the 12 constituent particles:

	6 protons = 6 x 1.007825 amu	= 6.04695 amu
	6 neutrons = 6 x 1.008665 amu	= 6.05199 amu
	total mass	= 12.09894 amu
compared to	A of ^{12}C	= 12.00000 amu (by definition)
	difference	= 0.09894 amu

This so-called *mass defect* has been converted into the *binding energy*, the potential energy "stored" in the nucleus to keep the particles together. This equivalence of mass and energy is defined by Einstein's special theory of relativity as

$$E_B = Mc^2 \qquad \textbf{(2.3)}$$

where E_B is the binding energy, M is the mass, and c is the velocity of light (2.997925×10^8 m/s).

According to this definition the equivalence between mass and energy is expressed as

$$1 \text{ amu} \equiv 931.5 \text{ MeV (million electronvolt)} \qquad \textbf{(2.4)}$$

where 1 eV = 1.602189×10^{-19} J (electrical charge unit = electron charge = 1.602189×10^{-19}C).

All energies in nuclear physics, such as the particle energies after nuclear decay, are presented in MeV or keV (kilo or 10^3 electronvolt). In our case of ^{12}C the binding energy turns out to be 0.09894 amu $\times$ 931.5 MeV/amu = 92.16 MeV or 7.68 MeV per nucleon.