



Chapter Thirteen

NUCLEI



13.1 INTRODUCTION

In the previous chapter, we have learnt that in every atom, the positive charge and mass are densely concentrated at the centre of the atom forming its nucleus. The overall dimensions of a nucleus are much smaller than those of an atom. Experiments on scattering of α -particles demonstrated that the radius of a nucleus was smaller than the radius of an atom by a factor of about 10^4 . This means the volume of a nucleus is about 10^{-12} times the volume of the atom. In other words, an atom is almost empty. If an atom is enlarged to the size of a classroom, the nucleus would be of the size of pinhead. Nevertheless, the nucleus contains most (more than 99.9%) of the mass of an atom.

Does the nucleus have a structure, just as the atom does? If so, what are the constituents of the nucleus? How are these held together? In this chapter, we shall look for answers to such questions. We shall discuss various properties of nuclei such as their size, mass and stability, and also associated nuclear phenomena such as radioactivity, fission and fusion.

13.2 ATOMIC MASSES AND COMPOSITION OF NUCLEUS

The mass of an atom is very small, compared to a kilogram; for example, the mass of a carbon atom, ^{12}C , is 1.992647×10^{-26} kg. Kilogram is not a very convenient unit to measure such small quantities. Therefore, a

different mass unit is used for expressing atomic masses. This unit is the atomic mass unit (u), defined as $1/12^{\text{th}}$ of the mass of the carbon (^{12}C) atom. According to this definition

$$\begin{aligned} 1\text{u} &= \frac{\text{mass of one } ^{12}\text{C atom}}{12} \\ &= \frac{1.992647 \times 10^{-26} \text{ kg}}{12} \\ &= 1.660539 \times 10^{-27} \text{ kg} \end{aligned} \quad (13.1)$$

The atomic masses of various elements expressed in atomic mass unit (u) are close to being integral multiples of the mass of a hydrogen atom. There are, however, many striking exceptions to this rule. For example, the atomic mass of chlorine atom is 35.46 u.

Accurate measurement of atomic masses is carried out with a mass spectrometer. The measurement of atomic masses reveals the existence of different types of atoms of the same element, which exhibit the same chemical properties, but differ in mass. Such atomic species of the same element differing in mass are called *isotopes*. (In Greek, isotope means the same place, i.e. they occur in the same place in the periodic table of elements.) It was found that practically every element consists of a mixture of several isotopes. The relative abundance of different isotopes differs from element to element. Chlorine, for example, has two isotopes having masses 34.98 u and 36.98 u, which are nearly integral multiples of the mass of a hydrogen atom. The relative abundances of these isotopes are 75.4 and 24.6 per cent, respectively. Thus, the average mass of a chlorine atom is obtained by the weighted average of the masses of the two isotopes, which works out to be

$$\begin{aligned} &= \frac{75.4 \times 34.98 + 24.6 \times 36.98}{100} \\ &= 35.47 \text{ u} \end{aligned}$$

which agrees with the atomic mass of chlorine.

Even the lightest element, hydrogen has three isotopes having masses 1.0078 u, 2.0141 u, and 3.0160 u. The nucleus of the lightest atom of hydrogen, which has a relative abundance of 99.985%, is called the proton. The mass of a proton is

$$m_p = 1.00727 \text{ u} = 1.67262 \times 10^{-27} \text{ kg} \quad (13.2)$$

This is equal to the mass of the hydrogen atom (= 1.00783u), minus the mass of a single electron ($m_e = 0.00055 \text{ u}$). The other two isotopes of hydrogen are called deuterium and tritium. Tritium nuclei, being unstable, do not occur naturally and are produced artificially in laboratories.

The positive charge in the nucleus is that of the protons. A proton carries one unit of fundamental charge and is stable. It was earlier thought that the nucleus may contain electrons, but this was ruled out later using arguments based on quantum theory. All the electrons of an atom are outside the nucleus. We know that the number of these electrons outside the nucleus of the atom is Z, the atomic number. The total charge of the

atomic electrons is thus $(-Ze)$, and since the atom is neutral, the charge of the nucleus is $(+Ze)$. The number of protons in the nucleus of the atom is, therefore, exactly Z , the atomic number.

Discovery of Neutron

Since the nuclei of deuterium and tritium are isotopes of hydrogen, they must contain only one proton each. But the masses of the nuclei of hydrogen, deuterium and tritium are in the ratio of 1:2:3. Therefore, the nuclei of deuterium and tritium must contain, in addition to a proton, some neutral matter. The amount of neutral matter present in the nuclei of these isotopes, expressed in units of mass of a proton, is approximately equal to one and two, respectively. This fact indicates that the nuclei of atoms contain, in addition to protons, neutral matter in multiples of a basic unit. This hypothesis was verified in 1932 by James Chadwick who observed emission of neutral radiation when beryllium nuclei were bombarded with alpha-particles (α -particles are helium nuclei, to be discussed in a later section). It was found that this neutral radiation could knock out protons from light nuclei such as those of helium, carbon and nitrogen. The only neutral radiation known at that time was photons (electromagnetic radiation). Application of the principles of conservation of energy and momentum showed that if the neutral radiation consisted of photons, the energy of photons would have to be much higher than is available from the bombardment of beryllium nuclei with α -particles. The clue to this puzzle, which Chadwick satisfactorily solved, was to assume that the neutral radiation consists of a new type of neutral particles called *neutrons*. From conservation of energy and momentum, he was able to determine the mass of new particle 'as very nearly the same as mass of proton'.

The mass of a neutron is now known to a high degree of accuracy. It is

$$m_n = 1.00866 \text{ u} = 1.6749 \times 10^{-27} \text{ kg} \quad [13.3]$$

Chadwick was awarded the 1935 Nobel Prize in Physics for his discovery of the neutron.

A free neutron, unlike a free proton, is unstable. It decays into a proton, an electron and a antineutrino (another elementary particle), and has a mean life of about 1000s. It is, however, stable inside the nucleus.

The composition of a nucleus can now be described using the following terms and symbols:

Z - *atomic number* = number of protons [13.4(a)]

N - *neutron number* = number of neutrons [13.4(b)]

A - *mass number* = $Z + N$
= total number of protons and neutrons [13.4(c)]

One also uses the term nucleon for a proton or a neutron. Thus the number of nucleons in an atom is its mass number A .

Nuclear species or nuclides are shown by the notation ${}^A_Z X$ where X is the chemical symbol of the species. For example, the nucleus of gold is denoted by ${}^{197}_{79} \text{Au}$. It contains 197 nucleons, of which 79 are protons and the rest 118 are neutrons.

The composition of isotopes of an element can now be readily explained. The nuclei of isotopes of a given element contain the same number of protons, but differ from each other in their number of neutrons. Deuterium, ${}^2_1\text{H}$, which is an isotope of hydrogen, contains one proton and one neutron. Its other isotope tritium, ${}^3_1\text{H}$, contains one proton and two neutrons. The element gold has 32 isotopes, ranging from $A = 173$ to $A = 204$. We have already mentioned that chemical properties of elements depend on their electronic structure. As the atoms of isotopes have identical electronic structure they have identical chemical behaviour and are placed in the same location in the periodic table.

All nuclides with same mass number A are called *isobars*. For example, the nuclides ${}^3_1\text{H}$ and ${}^3_2\text{He}$ are isobars. Nuclides with same neutron number N but different atomic number Z , for example ${}^{198}_{80}\text{Hg}$ and ${}^{197}_{79}\text{Au}$, are called *isotones*.

13.3 SIZE OF THE NUCLEUS

As we have seen in Chapter 12, Rutherford was the pioneer who postulated and established the existence of the atomic nucleus. At Rutherford's suggestion, Geiger and Marsden performed their classic experiment: on the scattering of α -particles from thin gold foils. Their experiments revealed that the distance of closest approach to a gold nucleus of an α -particle of kinetic energy 5.5 MeV is about 4.0×10^{-14} m. The scattering of α -particle by the gold sheet could be understood by Rutherford by assuming that the coulomb repulsive force was solely responsible for scattering. Since the positive charge is confined to the nucleus, the actual size of the nucleus has to be less than 4.0×10^{-14} m.

If we use α -particles of higher energies than 5.5 MeV, the distance of closest approach to the gold nucleus will be smaller and at some point the scattering will begin to be affected by the short range nuclear forces, and differ from Rutherford's calculations. Rutherford's calculations are based on pure coulomb repulsion between the positive charges of the α -particle and the gold nucleus. From the distance at which deviations set in, nuclear sizes can be inferred.

By performing scattering experiments in which fast electrons, instead of α -particles, are projectiles that bombard targets made up of various elements, the sizes of nuclei of various elements have been accurately measured.

It has been found that a nucleus of mass number A has a radius

$$R = R_0 A^{1/3} \quad (13.5)$$

where $R_0 = 1.2 \times 10^{-15}$ m (=1.2 fm; 1 fm = 10^{-15} m). This means the volume of the nucleus, which is proportional to R^3 is proportional to A . Thus the density of nucleus is a constant, independent of A , for all nuclei. Different nuclei are like a drop of liquid of constant density. The density of nuclear matter is approximately 2.3×10^{17} kg m $^{-3}$. This density is very large compared to ordinary matter, say water, which is 10^3 kg m $^{-3}$. This is understandable, as we have already seen that most of the atom is empty. Ordinary matter consisting of atoms has a large amount of empty space.

Example 13.1 Given the mass of iron nucleus as 55.85u and A=56, find the nuclear density?

Solution

$$m_{\text{Fe}} = 55.85, \quad u = 9.27 \times 10^{-26} \text{ kg}$$

$$\text{Nuclear density} = \frac{\text{mass}}{\text{volume}} = \frac{9.27 \times 10^{-26}}{(4\pi/3)(1.2 \times 10^{-15})^3} \times \frac{1}{56} \\ = 2.29 \times 10^{17} \text{ kg m}^{-3}$$

The density of matter in neutron stars (an astrophysical object) is comparable to this density. This shows that matter in these objects has been compressed to such an extent that they resemble a *big nucleus*.

13.4 MASS-ENERGY AND NUCLEAR BINDING ENERGY

13.4.1 Mass – Energy

Einstein showed from his theory of special relativity that it is necessary to treat mass as another form of energy. Before the advent of this theory of special relativity it was presumed that mass and energy were conserved separately in a reaction. However, Einstein showed that mass is another form of energy and one can convert mass-energy into other forms of energy, say kinetic energy and vice-versa.

Einstein gave the famous mass-energy equivalence relation

$$E = mc^2 \tag{13.6}$$

Here the energy equivalent of mass m is related by the above equation and c is the velocity of light in vacuum and is approximately equal to $3 \times 10^8 \text{ m s}^{-1}$.

Example 13.2 Calculate the energy equivalent of 1 g of substance.

Solution

$$\text{Energy, } E = 10^{-3} \times (3 \times 10^8)^2 \text{ J}$$

$$E = 10^{-3} \times 9 \times 10^{16} = 9 \times 10^{13} \text{ J}$$

Thus, if one gram of matter is converted to energy, there is a release of enormous amount of energy.

Experimental verification of the Einstein's mass-energy relation has been achieved in the study of nuclear reactions amongst nucleons, nuclei, electrons and other more recently discovered particles. In a reaction the conservation law of energy states that the initial energy and the final energy are equal provided the energy associated with mass is also included. This concept is important in understanding nuclear masses and the interaction of nuclei with one another. They form the subject matter of the next few sections.

13.4.2 Nuclear binding energy

In Section 13.2 we have seen that the nucleus is made up of neutrons and protons. Therefore it may be expected that the mass of the nucleus is equal to the total mass of its individual protons and neutrons. However,