

PHY140Y

33 Nuclear Properties

33.1 Overview

- Structure of the Nucleus
- Isotopes
- Atomic Mass and Number

33.2 Structure of the Nucleus

Lord Rutherford in 1909-1911 showed that the atom had to have a small, hard core that carried the bulk of the mass of the atom. Since atoms were neutral, the so-called nucleus of the atom had to have the same number of protons as the atom had electrons. This was the beginning of nuclear physics as we know it today.

It took about 20 years to sort out the basic constituents of the nucleus. Protons were thought to be one of the fundamental constituents, but it was only in 1932, when an American physicist by the name of Chadwick discovered the neutron, the two basic building blocks of the nucleus (or what we call “nucleons”) were identified. Together with the proton, these two particles are the constituents of all stable matter around us (excluding the much lighter electron for the moment). In order to explain the fact that the nucleus was stable despite the electrostatic repulsion of the protons, a new force known as the “strong force” had to be hypothesized that acted between protons and neutrons.

33.3 Isotopes

A nucleus is characterized by the number of protons and neutrons it has. The number of protons, the “atomic number” or Z , serves to define the type of nucleus it is and what element it is associated with, whereas the total number of protons and neutrons, the “atomic mass number” or A , is a measure of the total mass of the nucleus.

We denote a nucleus by the symbol ${}^A_Z\text{E}$, where E is the element corresponding to the specific value of Z . So, for example, the simplest nucleus is that of hydrogen, denoted ${}^1_1\text{H}$. The next lightest nucleus is that of helium, denoted ${}^4_2\text{He}$, with two protons and two neutrons. For a given type of nucleus (ie., a fixed Z), there are usually a number of different nuclei with slightly different values of A . These are known as “isotopes” of the given element, and reflect the fact that there are several possible stable configurations of neutrons and protons when we fix the number of protons. A large number of these isotopes are stable, in that they do not decay. Most of these are naturally recurring, so that one will find that a sample of a given element is usually a mixture of several different isotopes, each with the same chemical properties, but having a different nuclear composition.

However, what we find is that A increases faster than N for the stable isotopes. This implies that in order for a nucleus to be stable as the number of protons increases, the proportional number of neutrons must increase. For $N > 83$, there are no nuclei that are absolutely stable, and all configurations eventually decay. The stability of the nucleus can be traced to its binding energy: The greater the binding energy created by the strong force, the more stable the nucleus. The most stable nucleus turns out to be that of iron, ${}^{56}_{26}\text{Fe}$.

33.4 Atomic Mass and Number

The nucleus carries most of the mass of the atom. It is customary to use the “atomic mass unit” (amu or u) in this context. The reference for the amu is the most abundant form of carbon, ${}^{12}_6\text{C}$. We define this nucleus to have a mass equal to 12 amu. With this definition, we find that

$$1 \text{ amu} = 1.660540 \times 10^{-27} \text{ kg} = 931.494 \text{ MeV}/c^2. \quad (1)$$

In these units, the masses of the proton and neutron are

$$m_p = 938.27 \text{ MeV}/c^2 = 1.007276 \text{ amu} \quad (2)$$

$$m_n = 939.57 \text{ MeV}/c^2 = 1.008665 \text{ amu}, \quad (3)$$

respectively.

The mass of a nucleus is not just the mass of its constituents, but also reflects the binding energy of the strong force (which lowers the nucleus’ mass). The atomic mass of the most abundant helium isotope, ${}^4_2\text{He}$, is $m_{\text{He}} = 4.00188 \text{ amu}$. The binding energy of this nucleus is reflected in the mass difference

$$\Delta m = m_{\text{He}} - (2m_p + 2m_n) \quad (4)$$

$$= 4.00188 - [2(1.007276) + 2(1.008665)] = -0.030 \text{ amu}. \quad (5)$$

The binding energy per nucleon is a good figure of merit as it is a measure of how strongly bound the nucleus is. In this case, the binding energy would be $-\Delta mc^2 = 0.030 \text{ amu} c^2 = 27.95 \text{ MeV}$, so the binding energy per nucleon is

$$\frac{27.95}{4} = 7.00 \text{ MeV/nucleon}. \quad (6)$$

This corresponds to about 0.75% of the mass of the nucleus.

The binding energy per nucleon increases as N increases up to $N = 26$, or the element iron, which is the most stable of all nuclei. It then starts to fall slowly till at $N = 83$, it reaches a critical point where the number of neutrons and protons are barely stable. For larger values of N , all nuclei are unstable and decay in one of several ways. This means that the mass of the nucleus (in units of amu) is usually slightly below A for lighter nuclei with iron having the largest mass deficit. For heavier nuclei, the nuclear mass becomes very close to A .

The atomic mass quoted in the periodic table is in fact the atomic mass averaged over the naturally recurring isotopes, with the average weighted by the natural abundances of each isotope. So, for example, the atomic mass of uranium is quoted in the periodic table as 238.0289 amu, which reflects an isotopic composition of 99.26% ${}^{238}_{92}\text{U}$ and 0.74% ${}^{235}_{92}\text{U}$, each having masses of slightly more than 238 and 235 amu, respectively.