

13.Nuclei

Module 1 : Are Electrons Present Inside the Nucleus?

This module is designed to help you to:

- explain the difference between atomic mass and mass number;
- present arguments to rule out the existence of electrons inside the nucleus;
- explain the origin of beta rays due to the transformation of neutrons to protons and vice-versa;
- explain why neutrinos are emitted along with beta particles.

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13.1.1 Introduction

We know that nuclei of atoms consist of neutrons and protons, together known as nucleons. The number of protons, denoted by Z , is equal to the number of orbiting electrons. The number of protons together with the number of neutrons (N) equals the mass number of the nucleus. We shall learn the difference between atomic mass and mass number and also between nuclides and nuclei.

Certain nuclei, especially the nuclei with large mass numbers, are observed to emit three kinds of radiation. These nuclei are called radioactive nuclei and the radiations emitted are called alpha, beta, and gamma rays. Of these alpha rays are positively charged particles, beta rays are negatively charged particles whereas gamma rays are high energy photons.

The detailed study of the properties of beta rays shows that they consist of negatively charged electrons or positively charged positrons.

As electrons (as beta rays) are emitted by nuclei, a question was raised regarding the presence of electrons in the nuclei. In this module, we will present arguments to show that electrons could not be present inside the nucleus.

13.1.2 Atomic mass and mass number

The terms, atomic mass and mass number, for a given element are sometimes assumed to be synonymous. However, this is not so. The differences between the two terms are given below.

- (i) The atomic mass of an element is expressed in terms of the unit u , the atomic mass unit ($1 u = 1.66 \times 10^{-27} \text{ kg}$). This unit is derived on the basis that the mass of the carbon nucleus ${}^{12}_6\text{C}$ is exactly 12 u . On this scale, the mass of a neutron and a proton respectively are 1.008665 u and 1.007276 u . Atomic mass deals with many decimal values. The mass number is, however, just a number representing the total number of nucleons in a given nucleus.
- (ii) For a given element, the atomic number has only one unique value. Its mass number, however, can have more than one value, as the nuclei of the same element can have different numbers of nucleons. Such atoms of an element having nuclei with the same atomic number (Z) but different mass numbers (A) are called isotopes. ${}^{12}_6\text{C}$ and ${}^{14}_6\text{C}$ are the more well-known of the isotopes of carbon.
- (iii) The numerical value of the atomic mass of a given element is generally a fractional value. The different mass number values for a given element, however, are only integral numbers.

We can understand these differences by noticing that a given element can exist in more than one stable form. All these different forms have similar chemical properties. Nuclei have the same atomic number (and hence the same chemical properties) but different mass numbers are said to be isotopes of the given nucleus. The atomic mass of an element determined by chemical methods is the weighted average of its different isotopic forms.

This generally has a fractional value depending on the number of isotopes and their relative abundances.

For example, chlorine exists in two isotopic forms ^{35}Cl (mass = 34.9689 u) and ^{37}Cl (mass = 36.9659 u), whose relative abundances are nearly 75% and 25%, respectively. The atomic mass of chlorine would, therefore, be

$$\frac{34.9689 \times 75 + 36.9659 \times 25}{100} = 35.4682 \text{ u}$$

The mass numbers of a given element are stated in terms of the total number of nucleons in the nuclei of its different isotopes. These are, therefore, always integral values that are different for different isotopes. For example, the mass number of ^{35}Cl is 35 while that of ^{37}Cl is 37.

At present, it is possible to determine very precisely the masses of all the isotopes of a given element. This is done by using mass spectrometers. Because of the high level of precision associated with these measurements, the atomic mass seldom has an integral value. The mass number for each isotope, however, has only an integral value.

13.1.3 Nuclides and nuclei

We need to explain and understand the difference between the two terms: nuclides and nuclei. We try to understand this difference through the specific example of the element sodium.

The element sodium, represented by the symbol Na , has an atomic number 11, i.e. a total of 11 protons in each of the nuclei of its atoms. However, this element has two stable forms, one with a mass number 23 and the other with a mass number 22. The nucleus of the former, therefore, has 12 neutrons and that of the latter has 11 neutrons. Each of these two different types of nuclei, of the same element (sodium), is referred to as nuclide of sodium. Each of these isotopes of sodium is a nuclide that has a specific number of neutrons and protons.

All the different nuclides of an element have the same atomic number but different mass number. Symbolically a nuclide X is written as A_ZX . Thus, the two nuclides of sodium would be represented simply as ${}^{23}_{11}\text{Na}$ and ${}^{22}_{11}\text{Na}$ and the three nuclides of carbon would be represented as ${}^{11}_6\text{C}$, ${}^{13}_6\text{C}$, and ${}^{14}_6\text{C}$.

Remember that this way of representation is only a convention. In many books, you would find uranium of mass number 235 and atomic number 92 written as ${}^{235}_{92}\text{U}$ or ${}^{235}_{92}\text{U}$.

Sometimes in symbolic representation of a nuclide only the mass number is indicated. For example the nuclide of carbon with mass number 14 is represented by ${}^{14}\text{C}$.



Let Us Reflect

Q.13.1.1 Why does the mass number always have an integral value?

Q.13.1.2 Why is atomic mass often a fractional number?

Q.13.1.3 Two stable isotopes of lithium ${}^6_3\text{Li}$ and ${}^7_3\text{Li}$ have respective abundances of 7.5% and 92.5%. These isotopes have masses 6.01512 u and 7.01600 u, respectively. Find the average atomic mass of lithium.

Note that the atomic mass in this case is a fraction.

Q.13.1.4 Boron has two stable isotopes, ${}^{10}_5\text{B}$ and ${}^{11}_5\text{B}$. Their respective masses are 10.01294 u and 11.00931 u. If the atomic mass of boron is found to be 10.811 u, find the percentage abundances of ${}^{10}_5\text{B}$ and ${}^{11}_5\text{B}$.

Q.13.1.5 Table 13.1.1 and 13.1.2 below contain data for stable isotopes of elements neon (atomic number = 10) and zinc (atomic number = 30). Calculate the atomic mass of these elements.

Table 13.1.1: Stable isotopes of ${}_{10}\text{Ne}$

Mass Number of the (stable) isotope	Mass, (u)	Natural Percentage abundance (or availability)
20	19.99244	90.48%
21	20.99384	0.27%
22	21.99138	9.25%

Table 13.1.2: Stable isotopes of ${}_{30}\text{Zn}$

Mass Number of the (stable) isotope	Mass, (u)	Natural Percentage abundance (or availability)
64	63.92914	48.63%
66	65.92605	27.90%
67	66.92715	04.10%
68	67.92486	18.75%

70	69.92535	00.62%
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A.13.1.1 The mass number of an isotope of a given element is stated in terms of the total number of nucleons in its nucleus. These are, therefore, always integral values that are different for different isotopes. For example, the mass number of ^{35}Cl is 35 while that of ^{37}Cl is 37.

A.13.1.2 The atomic mass of an element determined by chemical methods is the weighted average of its different isotopic forms. This, therefore, generally has a fractional value depending on the number of isotopes and their relative abundances.

For example, chlorine exists in two isotopic forms ^{35}Cl (mass = 34.9689 u) and ^{37}Cl (mass = 36.9659 u), whose relative abundances are nearly 75% and 25%, respectively. The atomic mass of chlorine would, therefore, be

$$\frac{34.9689 \times 75 + 36.9659 \times 25}{100} = 35.4682 \text{ u.}$$

A.13.1.3 Given

Relative abundance of $^6_3\text{Li} = 7.5\%$

Relative abundance of $^7_3\text{Li} = 92.5\%$

Mass of $^6_3\text{Li} = 6.01512 \text{ u}$

Mass of $^7_3\text{Li} = 7.01600 \text{ u}$

$\therefore$ Average atomic mass of lithium

$$= \frac{6.01512 \times 7.5 + 7.01600 \times 92.5}{100} = 6.94093 \text{ u}$$

A.13.1.4 Using relative abundances as $x\%$ and $(100-x)\%$, calculate x as in Q.13.1.3 above.

A.13.1.5

(i) 20.18005 u

(ii) 65.39557 u

Calculate the atomic mass of the remaining elements as done in Q.13.1.3 above.

The results of these exercises show that the atomic masses of different elements have fractional values.

13.1.4 Are electrons present in atomic nuclei?

Studies on nuclear stability show that some nuclei, especially those with high mass numbers, are unstable. These unstable nuclei decay into product nuclei which are more stable than the original nuclei. During their decay, the unstable nuclei emit α rays (nuclei of helium), β rays (electrons) and γ rays (electromagnetic radiation). The emission of β rays was particularly intriguing because it indicated the possible existence of electrons inside the nuclei. However, there are powerful arguments that indicate that electrons cannot exist as electrons in nuclei.

We have a lot of experimental and theoretical evidence which indicates that electrons cannot be a constituent of atomic nuclei. We give below the basic ideas about some of this evidence.

13.1.4.1 Evidence based on the uncertainty principle

Heisenberg's principle of uncertainty implies that the position and momentum of a subatomic particle like an electron cannot be precisely known. If Δx is the uncertainty in the position and Δp is the uncertainty in the momentum, then according to the uncertainty principle, $\Delta x \cdot \Delta p \approx \frac{1}{2} \frac{h}{2\pi}$

This equation implies that if we try to locate the particle more precisely, that is, reduce Δx , the uncertainty in the momentum, Δp , flares up. Similarly, attempts to reduce Δp result in an increase in Δx .

Now if the electron were a constituent of the nucleus, the uncertainty in its position (Δx) would be of the order of the size of the nucleus, which has been determined to be of the order of 10^{-14} m. It follows from the uncertainty principle that the uncertainty in its momentum (Δp) would be of the order of

$$\Delta p = \frac{h}{4\pi\Delta x} = \frac{6.63 \times 10^{-34}}{4\pi \times 10^{-14}} \text{ kg ms}^{-1},$$

,

which is of the order of $5 \times 10^{-21} \text{ kg ms}^{-1}$

When there is $5 \times 10^{-21} \text{ kg ms}^{-1}$ uncertainty in the momentum of the electron, its momentum itself would have a magnitude of this order i.e. $5 \times 10^{-21} \text{ kg ms}^{-1}$.

When the velocity of an electron approaches that of light, the electron is said to be relativistic. The energy of a relativistic electron is given by

$$E = \sqrt{p^2 c^2 + m_0^2 c^4}$$

where $m_0 c^2$ is the rest mass energy of the electron due to its rest mass m_0 .

However, as $v \rightarrow c$, the rest mass energy is quite small compared to its kinetic energy.

If we neglect this term in the above expression, then $E = pc$, which is mostly kinetic energy.

So, if the momentum of the electron, derived on the assumption of its existence inside the nucleus, is $5 \times 10^{-21} \text{ kg ms}^{-1}$, then its kinetic energy must be pc or, $5 \times 10^{-21} \text{ kg ms}^{-1} \times 3 \times 10^8 \text{ ms}^{-1} \approx 10 \text{ MeV}$

This order of magnitude of the electron's kinetic energy is not only inconsistent with the observed values of the energy of β -rays but is also much higher than the average value of the binding energy/nucleon in different nuclei.

Heisenberg's uncertainty principle, therefore, rules out the possibility of the existence of free electrons inside atomic nuclei.

13.1.4.2 Evidence based on the values of spin

The electron, the proton, and the neutron are all considered to have an intrinsic angular momentum or spin associated with them. All these particles have the same spin value $= \frac{1}{2} \frac{h}{2\pi}$

It turns out that if we think of nuclei as consisting of protons and electrons, the observed values of the spin of a number of nuclei are inconsistent with the total spin of their constituents. Such discrepancies disappear if we regard the electron as not being a constituent of the nucleus.

The observed spin values of nuclei, therefore, also appear to rule out the possibility of the existence of electrons inside the nucleus.

13.1.4.3 Evidence based on the values of magnetic moments

We know that we can associate a magnetic moment with an electron. This magnetic moment for the electron in an atom is contributed by its orbital motion as well as by its intrinsic angular momentum or spin. The spin part of the magnetic moment of the electron is an intrinsic characteristic of the electron.

Like the electron, the nucleons (protons as well as neutrons) also have intrinsic magnetic moments. It is found that the magnetic moment of a nucleon is only of the order of 10^{-3} times that of an electron.

The observed values of the magnetic moments of different nuclei are of the order of the magnetic moment of the nucleons only. This fact is again inconsistent with the possibility of the existence of electrons inside atomic nuclei.

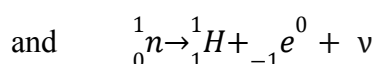
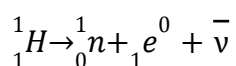
We thus realize that we have a number of reasons to rule out the possibility of electrons inside atomic nuclei. Nuclei, as we now say, are made-up of protons and neutrons.

13.1.4.4 Evidence based on the decay of radioactive nuclei

We first note that radioactivity is a nuclear property. The emission of β rays, therefore, appears to suggest that atomic nuclei may have electrons as their constituents. We now have evidence that rules out the possibility of electrons residing within the nucleus. Therefore, we need to explain how atomic nuclei are able to emit electrons, which they do not contain.

We can understand β emission or β decay of a nucleus by postulating that proton and neutron are the two states of the same species known as the nucleon. Whenever transformation from one state to the other comes about, a very light charged particle also needs to be emitted. This is necessary to ensure charge and mass conservation. Since the neutron and the proton have mass values that are quite close to each other, the emitted particle must be very light.

We can understand β^- decay in terms of the following nuclear reactions:



Note that while a free neutron decays to a proton, the decay of a proton to a neutron (the first reaction above) is possible only inside the nucleus since protons have a smaller mass than neutrons.

The additional particles, the $\bar{\nu}$ and ν , associated with these reactions, are known as the antineutrino and the neutrino, respectively. The existence of neutrinos was predicted on the basis of the conservation of energy and angular momentum in β -decay. The existence of neutrinos has now been confirmed experimentally.

Nuclei are thus able to emit β^- particles (electrons/positrons), in spite of not having them as their constituents.

13.1.5 Summary

- The nucleus of an atom carries almost all its mass.
- The positive charge is provided by protons. The number of orbital electrons is equal to the number of protons in the nucleus. This makes an atom electrically neutral.
- The number of protons is denoted by Z , called the atomic number of the nucleus.
- The remaining mass of the nucleus is provided by N neutrons.
- The sum $N+Z$ is the mass number of the nucleus. It is always a whole number.
- The actual mass of the nucleus is the weighted average mass of all its stable isotopes. It is generally a fractional number.
- Nuclei of different isotopes of an element are referred to as nuclides.

- Electrons are not present in the atomic nuclei. This can be explained by the evidences based on the
 - (i) uncertainty principle
 - (ii) values of spin
 - (iii) values of magnetic moments
 - (iv) decay of radioactive nuclei.

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