The atom and its nucleus

In ancient times, the Greek philosopher Demokritos asserted that all matter is made out of indivisible units. This chapter introduces the basic ideas and models that have given rise to our present understanding of the atom and its nucleus. We begin with Rutherford's experiment that provided the evidence for the existence of a small, massive and positively charged atomic nucleus and close with a discussion of the fundamental forces that operate within the nucleus.

Objectives

By the end of this chapter you should be able to:

- appreciate that *atomic spectra* provide evidence for an atom that can only take *discrete values in energy*;
- explain what *isotopes* are and how their existence implies that neutrons are present inside the nucleus;
- state the meaning of the terms *nuclide*, *nucleon*, *mass number* and *atomic number* (proton number);
- outline the properties of the *forces* that operate within the nucleus.

The discovery of the nuclear atom

In 1909, Geiger and Marsden, working under Rutherford's direction, performed a series of experiments in which they studied the scattering of alpha particles shot at a thin gold foil. Alpha particles have a mass approximately four times that of the hydrogen atom and a positive electric charge of two units ($2e$). Alpha particles are emitted when unstable elements decay; we will study them in more detail later.

Geiger and Marsden used radon as their source of alpha particles, which they directed in a fine beam toward the thin gold foil. The scattered alpha particles were detected (through a microscope) by the glow they caused on a fluorescent screen at the point of impact. As expected, most of the alpha particles were detected at very small scattering angles, such as at positions A, B and C in Figure 1.1.

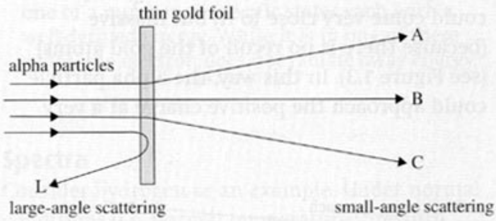

Figure 1.1 The majority of alpha particles are slightly deflected by the gold foil. Very occasionally, large-angle scatterings take place. The small deflections could be understood in terms of alpha particles approaching the nucleus at large distances. The large deflections were due to alpha particles approaching very close to the nucleus.

These small deflections could be understood in terms of the electrostatic force of repulsion between the positive charge of the gold atoms and the positive charge of the alpha particles. (Note that an alpha is about 8000 times more massive than the electron and so the effect of the electrons of the gold atoms on the path of the alphas is negligible.)

▶ To their great surprise, Rutherford, Geiger and Marsden found that, occasionally, alpha particles were detected at *very large scattering angles*, as can be seen in Figure 1.1. These large-angle scattering events could not be understood in terms of the prevailing model of the time – Thomson's model of the atom.

Consequences of the Rutherford (Geiger–Marsden) experiment

The very large deflection was indicative of an enormous force of repulsion between the alpha particle and the carrier of the positive charge of the atom. Such a large force could not be produced if the positive charge was distributed over the entire atomic volume, as Thomson had suggested earlier (see Figure 1.2). Rather, it suggested that the positive charge resided on an object that was tiny (thus the alpha particle could come very close to it) but massive (because there is no recoil of the gold atoms) (see Figure 1.3). In this way, the alpha particle could approach the positive charge at a very

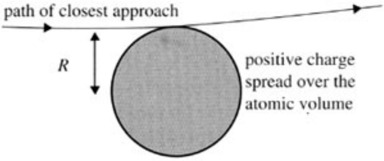

Figure 1.2 In Thomson's model, the closest an alpha particle can come to the atom's centre is a distance equal to the atomic radius.

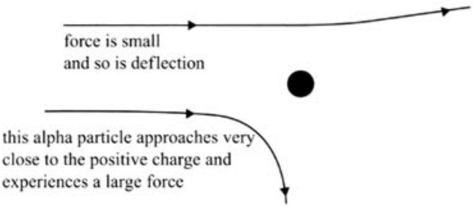

Figure 1.3 In Rutherford's model, the alpha particle can approach much closer if the nucleus is very small.

small distance, and the Coulomb force of repulsion, being proportional to the inverse of the square of the separation, would then be enormous. This force causes the large deflection in the alpha particle's path.

Rutherford calculated theoretically the number of alpha particles expected at particular scattering angles based on Coulomb's force law. He found agreement with his experiments if the positive atomic charge was confined to a region of linear size approximately equal to 10^{-15} m. This and subsequent experiments confirmed the existence of a nucleus inside the atom – a small, massive object carrying the positive charge of the atom.

Example question

Q1

Calculate the electric field at the surface of a nucleus of one unit of positive charge and radius 10^{-15} m. Compare this with the value of the electric field of the same charge that is now spread over a sphere of radius 10^{-10} m.

Answer

Applying the formula for the electric field $E = k\frac{Q}{r^2}$ we find

$$E = 9 \times 10^9 \times \frac{1.6 \times 10^{-19}}{(10^{-15})^2}$$
$$= 1.4 \times 10^{21}\ \mathrm{N\,C^{-1}}$$

Near the larger sphere, the electric field is $E = 1.4 \times 10^{11}\ \mathrm{N\,C^{-1}}$, which is a factor of 10^{10} smaller. This is why the deflecting forces in *Rutherford's model* are so large compared with what one might expect from *Thomson's model.*

The Rutherford model of the atom

These discoveries led to a new picture of the atom. A massive, positively charged nucleus occupied the centre of the atom and electrons orbited this nucleus in much the same way that planets orbit the sun: this was the Rutherford model (see Figure 1.4). The force keeping the electrons in orbit was the electrical force between the negative electron charge and the positive nuclear charge.

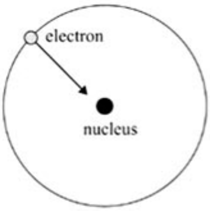

Figure 1.4 Rutherford's atomic model has the electrons orbiting the nucleus like planets orbiting the sun.

Immediately after these discoveries, difficulties arose with the Rutherford model.

▶ The main difficulty was that, according to the theory of electromagnetism, an accelerated charge should radiate electromagnetic waves and thus lose energy. The electrons in Rutherford's atom move in circular paths around the nucleus and so suffer centripetal acceleration. If they radiate and lose energy, it can be shown that this would lead to electron orbits that would spiral into the nucleus. The time required for the electron to fall into the nucleus is of the order of nanoseconds. Thus, the Rutherford model cannot explain why matter is stable, i.e. why atoms exist.

The Bohr model

The first attempt to solve this problem came from the Danish physicist Niels Bohr in 1911. These are the **Bohr postulates.**

▶ Bohr examined the simplest atom, that of hydrogen, and realized that the electron could exist in certain specific states of definite energy, without radiating away energy, if a certain condition was met by the orbit radius. The electron energy is thus *discrete* as opposed to continuous. The electron can only lose energy when it makes a transition from one state to another of lower energy. The emitted energy is then the difference in energy between the initial and final states. (See Figure 1.5.) The strongest piece of evidence in support of Bohr's idea is the existence of *emission* and *absorption spectra*.

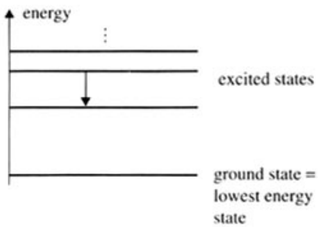

Figure 1.5 In the Bohr model the electron occupies one of a number of specific states each with a well-defined energy. While it is in one of these states, the electron does not radiate away energy.

Spectra

Consider hydrogen as an example. Under normal conditions (i.e. normal temperature, pressure, etc.) the electron in each hydrogen atom occupies the lowest energy state or **energy level** (the ground state). If the atoms are somehow excited (by increasing their temperature, for example) the electrons leave the ground state and occupy one of the higher energy, excited states. As soon as they do so, however, they make a transition back down to lower energy states, radiating energy in the process. The energy of the light emitted is very well defined since it corresponds to the difference in energy between the states involved in the transition. Knowing this energy difference allows us to calculate the wavelength of the emitted light, and so the wavelength, too, is well defined since the energy is. In this way it is found, for example, that hydrogen emits light of wavelengths 656 nm, 486 nm and 410 nm. Only hydrogen emits light of these wavelengths since only hydrogen has states whose energy differences lead to these wavelengths. Helium, for example, has energy states of different energy, so the wavelengths of light emitted when helium atoms are heated are different from those of hydrogen.

▶ The set of wavelengths of light emitted by the atoms of an element is called the *emission spectrum* of the element.

Conversely, consider atoms of hydrogen that are in their ground states and imagine sending light of a specific wavelength through a given quantity of hydrogen. If the wavelength of light does not correspond to any of the wavelengths in the emission **spectrum** of hydrogen, the light is transmitted through the atoms of hydrogen without any absorption. If, however, it matches one of the emission spectrum wavelengths, then this light is absorbed.

▶ The electrons simply take this energy and use it in order to make a transition to a higher energy state. The wavelengths that are so absorbed make up the *absorption spectrum* of the element and (as indicated above) they are the same wavelengths as those in the emission spectrum.

Thus, if white light (i.e. light containing all wavelengths) is sent through the gas and the transmitted light is analysed through a spectrometer, dark lines will be found at the position of the absorbed wavelengths.

Nuclear structure

Nuclei are made up of smaller particles, called protons and neutrons. The word *nucleon* is used to denote a proton or a neutron.

▶
- The number of protons in a nucleus is denoted by Z, and is called the atomic (or proton) number.
- The total number of nucleons (protons + neutrons) is called the mass (or nucleon) number, and is denoted by A.

Then the electric charge of the nucleus is $Z|e|$. The number of neutrons in the nucleus (the neutron number N) is thus $N = A - Z$. We will use the atomic and mass numbers to denote a nucleus in the following way: the symbol $^{A}_{Z}\mathrm{X}$ stands for the nucleus of element X, whose atomic number is Z and mass number is A. Thus $^{1}_{1}\mathrm{H}$, $^{4}_{2}\mathrm{He}$, $^{40}_{20}\mathrm{Ca}$, $^{210}_{82}\mathrm{Pb}$ and $^{238}_{92}\mathrm{U}$ are, respectively, the nuclei of hydrogen, helium, calcium, lead and uranium, with one, two, twenty, eighty-two and ninety-two protons. A nucleus with a specific number of protons and neutrons is also called a *nuclide*.

We can apply this notation to the nucleons themselves. For example, the proton (symbol p) can be written as $^{1}_{1}\mathrm{p}$ and the neutron (symbol n) as $^{1}_{0}\mathrm{n}$. We can even extend this notation to the electron, even though the electron has nothing to do with the nucleus and nucleons. We note that the atomic number is not only the number of protons in the nucleus but also its electric charge in units of $|e|$. In terms of this unit, the charge of the electron is -1 and so we represent the electron by $^{0}_{-1}\mathrm{e}$. The mass number of the electron is zero – it is so light with respect to the protons and neutrons that it is, effectively, massless. The photon (the particle of light) can also be represented in this way: the photon has the Greek letter gamma as its symbol, and since it has zero electric charge and (strictly) zero mass it is represented by $^{0}_{0}\gamma$. Table 1.1 gives a summary of these particles and their symbols.

Particle	Symbol
Proton	${}^{1}_{1}p$
Neutron	${}^{1}_{0}n$
Electron	${}^{0}_{-1}e$
Photon	${}^{0}_{0}\gamma$
Alpha particle	${}^{4}_{2}He$ or ${}^{4}_{2}\alpha$

Table 1.1

Isotopes

Nuclei that have the same number of protons (and therefore the same atomic number Z) but different number of neutrons (i.e. different N and mass number A) are called *isotopes* of each other. Since isotopes have the same number of protons, their atoms have the same number of electrons as well. This means that isotopes have identical chemical but different physical properties. The existence of isotopes can be demonstrated with an instrument called the mass spectrometer (this is discussed further in Chapter 6.6). The existence of isotopes is evidence for the existence of neutrons inside atomic nuclei.

The forces within the nucleus

The nucleons (i.e. protons and neutrons) are bound together in the nucleus by a force we have not yet met – *the strong nuclear force*. It is necessary to have a new force inside the nucleus, because otherwise the electrical repulsion between the positively charged protons would break the nucleus apart (see Figure 1.6).

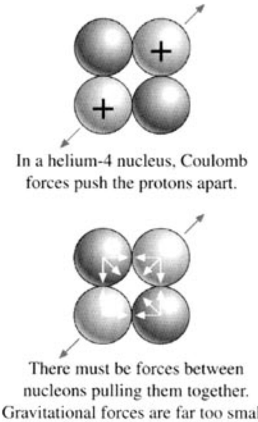

Figure 1.6 There is an attractive force between nucleons that keeps them bound inside the nucleus.

▶ The strong nuclear force is an attractive force and much stronger than the electrical force if the separation between two nucleons is very small (i.e. about $r = 10^{-15}$ m or less). For larger separations, the nuclear force becomes so small as to be negligible – we say that the nuclear force has a *short range*.

The experimental evidence for the properties of the nuclear force comes from scattering experiments in which electrons of energy equal to about 200 MeV (in later experiments neutrons were also used) are allowed to hit nuclei and their scattering is studied. If we make the assumption of short-range forces, we obtain agreement with the data. A result of these experiments is that the nuclear radius R is given by

$$R = 1.2 \times A^{1/3} \times 10^{-15}\ \text{m}$$

where A is the total number of protons and neutrons in the nucleus (the mass number). This implies that the nuclear density is the same for all nuclei (you will look at this further in the questions at the end of this chapter). The short range of the force implies that a given nucleon can only interact with a few of its immediate neighbours and not with all of the nucleons in the nucleus (see Figure 1.7).

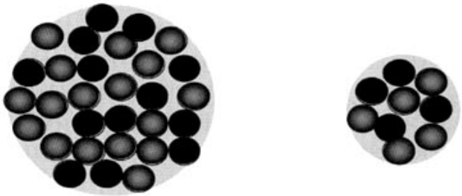

Figure 1.7 Irrespective of the size of the nucleus, any one nucleon is surrounded by the same number of neighbours, and only those act on it with the nuclear force.

Force	Electromagnetic	Strong nuclear	Weak nuclear
Acts on	Protons only	Protons and neutrons	Protons and neutrons
Nature	Repulsive	Attractive (mainly)	Attractive/repulsive
Range	Infinite	Short (10^{-15} m)	Short (10^{-17} m)
Relative strength	$\frac{1}{137}$	1	10^{-6}

Table 1.2 Forces operating in the nucleus.

There is one other force acting in the nucleus apart from the electrical and strong nuclear forces. This is the weak nuclear force, a force that is responsible for the decay of a neutron into a proton. The details of this decay (beta decay) will be examined in the next chapter. The forces acting in the nucleus are summarized in Table 1.2.

(Since the masses of subatomic particles are so small, the gravitational force is irrelevantly small compared with the other three forces.)

Questions

1 The radius of an atomic nucleus is given by the expression

$$R = 1.2 \times A^{1/3} \times 10^{-15}\ \text{m}$$

where A is the mass number of the nucleus.

(a) Use this expression to find the density of a nucleus of iron ($^{56}_{26}Fe$) in kg m^{-3}.

(b) How does this density compare with the normal density of iron?

(c) If a star with a mass equal to 1.4 times the mass of our sun (solar mass = 2.0×10^{30} kg) were to have this density, what should its radius be? (Such stars are formed in the end stage of the evolution of normal stars and are called neutron stars.)

2 Use the expression for the radius of a nucleus to show that all nuclei have the same density.

3 Describe carefully how the Geiger–Marsden–Rutherford experiment gave rise to the Rutherford model of the atom. Why is the experiment you just described inconsistent with Thomson's model of the atom?

4 Explain why the dark lines of an absorption spectrum have the same wavelengths as the bright lines of an emission spectrum for the same element.

5 What is an isotope? How do we know that isotopes exist?

6 Find the number of neutrons in these nuclei: $^{1}_{1}H$; $^{4}_{2}He$; $^{40}_{20}Ca$; $^{210}_{82}Pb$.

7 What is the electric charge of the nucleus $^{3}_{2}He$?

8 What is meant by the statement that the energy of atoms is discrete? What evidence is there for this discreteness?

9 What do you understand by the statement that the strong nuclear force has a short range?

10 What is the dominant force between two protons separated by a distance of:

(a) 1.0×10^{-15} m;

(b) 1.0×10^{-14} m?

11 Explain why a nucleon feels the strong force from roughly the *same number* of other nucleons, irrespective of the size of the nucleus.

HL only

12 Compare the gravitational force between two electrons a distance of 10^{-10} m apart with the electrical force between them when at the same separation.