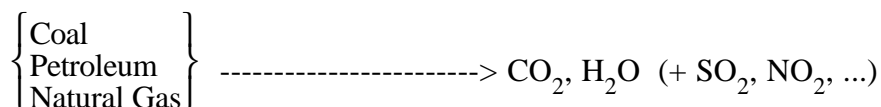# Chapter 12

NUCLEAR ENERGY: INTRODUCTION

$$E = mc^2$$

(Einstein)

In the preceding chapters we discussed the conversion of chemical energy of fossil fuels into more useful energy forms such as heat, work and electricity. This conversion involves a rearrangement of atoms of C, H, O, N and S to form new molecules:

$$\left.\begin{array}{l}\text{Coal} \\ \text{Petroleum} \\ \text{Natural Gas}\end{array}\right\} \ \text{------------------------>} \ CO_2, H_2O \ (+ SO_2, NO_2, ...)$$

The atoms maintain their integrity in this transformation. The energy released is called the heating value of the fuel.

In addition to fossil fuels, "nuclear fuels" can also be found in the Earth's crust. Uranium is the best known among them. The energy of these fuels can also be used by society: their potential energy can be converted to more useful forms, primarily into heat and subsequently to electricity. This will be discussed in Chapters 13 and 14.

Conversion of nuclear energy to more useful forms involves the nuclear reactions of fission and fusion. In these reactions the atoms lose their integrity. New atoms, either smaller or larger, are formed by the rearrangement of their nuclei. This process is accompanied by a much greater release of energy than that corresponding to the heat of combustion of fossil fuels. We shall see, however, that potentially harmful forms of energy ('radioactivity') are also obtained in this process (Chapter 15). Here we introduce the minimum of atomic physics necessary to understand the main issues concerning this very controversial energy form. We show that not much is required to grasp the essence of what happened at Three Mile Island or Chernobyl, or even to be able to take an *informed* stand on the future of nuclear power in our society.

A Brief Historical Prelude: Isotopes

In 1895, the German physicist Wilhelm Conrad Röntgen (1845-1923) accidentally discovered x-rays, which were to revolutionize medicine. One year later the French physicist Antoine Henri Becquerel (1852-1908) discovered another form of invisible radiation which was to have even more dramatic consequences for humanity. He first thought that compounds containing uranium emit this radiation because of their phosphorescence. But then he realized that a photographic plate had been revealed by these highly penetrating rays even after it was wrapped in black paper and left in a drawer to exclude sunlight. Several years later, the Polish-born French physicist Marie Sklodowska-Curie (1867-1934) gave the name *radioactivity* to this spontaneous emission of radiation by uranium. She and her husband Pierre Curie (1859-1906) were the pioneers in discovering and dealing with radioactive materials. She died of leukemia as a consequence of frequent exposure to them; Pierre Curie also suffered from radiation sickness, until his untimely death in a traffic accident.

In 1903 the New Zealand-born British physicist Ernest Rutherford (1871-1937) found that radioactivity consists of at least two types of radiation: positively charged alpha (α) particles and negatively charged beta (β) particles, After Marie and Pierre Curie had isolated a new and intensely radioactive element (radium) from pitchblende, the Frenchman Paul Villard (1860-1934) observed that radium emits very powerful rays, similar to x-rays but ones that readily pass through metal sheets; in keeping with Rutherford's terminology, they were called gamma (γ) rays. The penetrating power of the various forms of radiation is illustrated in Figure 12-1: the power of gamma rays is about 50,000 times greater than that of alpha particles. The consequences of this fact are discussed in more detail in Chapter 15.

FIGURE 12-1. Penetrating power of the various forms of radiation.

The dramatic discovery of radioactivity at the turn of the century, together with the simultaneous discovery of the electron in 1897 by the Englishman J. J. Thomson (1856-1940), led scientists to re-evaluate the structure of the atom. In 1911 Rutherford formulated the 'nuclear' model of the atom. Until then it was thought that electrons are embedded in a spherical volume of positive charge much like the seeds in a watermellon. Rutherford and his collaborators discovered, however, that all the positive charge of an atom is

concentrated in a nucleus whose volume is a very tiny fraction of the overall size of the atom. Two years later, the Danish physicist Niels Bohr (1885-1962) formulated the 'planetary' model of the atom, according to which electrons revolve around the nucleus much like the planets orbit around the Sun. This atomic structure, illustrated in Figure 12-2, marked the beginning of the fascinating journey into modern atomic physics, but its most rudimentary form will be the end of our journey into physics. Although now known to be simplistic, we shall adopt this model here because it is sufficiently accurate for the purposes of our discussion. The only missing ingredients for our story are the discoveries of the neutron in 1932 by the Englishman James Chadwick (1891-1974) and of neutron-induced (artificial) radioactivity in 1934 by the Italian Enrico Fermi (1901-1954).

FIGURE 12-2
The simplified Rutherford-Bohr model of the structure of an atom (not drawn to scale).

An atom of any element (carbon, hydrogen, oxygen, uranium, etc.) consists therefore of a *nucleus*, which is positively charged, and of *electrons*, which revolve around the nucleus. Electrons are negatively charged and their mass is much smaller than that of the nucleus. The nucleus consists of *protons* and *neutrons*, which are collectively and appropriately called *nucleons*. Protons are positively charged; neutrons do justice to their name and are neutral, that is, uncharged. The number of protons in an atom is equal to the number of electrons; atoms are therefore also neutral under normal circumstances.

Because neutrons are neutral, their number in the nucleus of an atom is not necessarily equal to the number of protons in the nucleus. It is not uncommon, therefore, to find in nature different forms of an element with the same number of protons, but different numbers of neutrons. Such different forms of an atom, which are chemically identical but

can be very different physically, are called *isotopes*. Table 12-1 lists the isotopes that are most important for our purposes.

The isotopes of hydrogen are important in nuclear fusion because they are light and can combine to form larger (and heavier) atoms. This combination requires that two positively charged nuclei come together. We shall see in Chapter 14 that this is not easy to accomplish on Earth; it does happen all the time inside the Sun. The isotopes of uranium are important in nuclear fission because they are heavy and can break up to form smaller (and lighter) atoms. In particular, they can break up quite easily when they are bombarded by the uncharged neutron particles.

TABLE 12-1
Properties of selected isotopes

Isotope [a]	Mass Number (p + n)	Atomic Number (p)	Neutron Number (n)
$_1H^1$	1	1	0
$_1H^2$	2	1	1
$_1H^3$	3	1	2
$_{92}U^{235}$	235	92	143
$_{92}U^{238}$	238	92	146

[a] The subscript to the left of the atomic symbol corresponds to the atomic number (number of protons or electrons); the superscript to the right of the symbol corresponds to the mass number (number of protons plus neutrons). For simplicity, we often write H-1, H-2, H-3, U-235, U-238, etc. and remember that there are 92 protons (and electrons) in uranium, one proton in hydrogen, etc.

Nuclear Fission and Fusion: Introduction

It is a fact of life that some isotopes are 'stable' and others are unstable or 'radioactive'. For example, approximately one in every 7000 hydrogen atoms in water (whose molecules consist of two hydrogen atoms and one oxygen atom) is H-2, a stable form of hydrogen called deuterium. The rest is H-1, or 'normal' hydrogen. The radioactive tritium, H-3, is not found in nature, but it can be produced in nuclear reactions (see Chapter 14).

What do we mean by stability? To answer this question, let us remind ourselves of the universal law of nature that we discussed in Chapter 3: we saw that there exists in nature a spontaneous tendency to *maximize the entropy* of a 'system'. Another law of nature is the spontaneous tendency to *minimize the energy* of a system. We understood the entropy maximization principle from everyday experience and by recalling some elements of a poker game. We can understand the energy minimization principle from daily life as well. For example, we would rather sit than stand; this is because by sitting we minimize our (gravitational) potential energy. *Stable isotopes* are simply those that have attained this

energy minimum. *Radioactive isotopes* have excess energy which they seek to release in order to achieve such stability. This is a natural process that occurs spontaneously; indeed, archeological dating of objects relies on this process of 'decay' of a radioactive isotope of carbon, C-14. This process can also be induced in a nuclear reaction or in an atomic bomb.
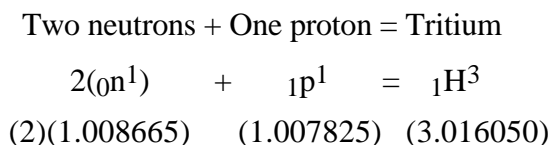
The energy released by radioactive isotopes takes different forms. Society is primarily interested in the heat released, as we shall see in Chapters 13 and 14. However, some alpha, beta and gamma radiation inevitably accompanies this release of heat. We can view this radiation as "nuclear pollution," by analogy with the use of fossil fuels. In the combustion of fossil fuels we perform a chemical reaction to obtain heat. Inevitably, pollution is produced as well, in the form of sulfur oxides, nitrogen oxides, etc. In nuclear fuel utilization we perform a nuclear reaction also to obtain heat. However, some of the nuclear energy is converted to radiation. Dealing with this source of nuclear pollution is a complex problem. We shall summarize it in Chapter 15.

We now need to discuss briefly how the stability of atoms is measured, because this allows us to calculate how much energy is released in nuclear reactions. Following the analogy between chemical reactions and nuclear reactions, the concept of *binding energy* is taken to be analogous to the heating value of fossil fuels.

The binding energy of an atom is the minimum amount of energy required to break up its nucleus into individual components.

Figure 12-3 shows the binding energies of all known stable and radioactive isotopes. The larger the binding energy per nucleon, the more stable the isotope is. This can be understood by realizing that these isotopes, which are seen to have an intermediate mass (50-75 atomic mass units), have released the greatest amount of energy in the process of their formation.

An elementary calculation will best illustrate this key concept and explain the energy units shown in Figure 12-3. The nucleus of tritium, the least abundant of the three isotopes of hydrogen and the only radioactive one, consists of two neutrons (n) and one proton (p), as shown in Table 12-1. Its formation from the individual nucleons can be represented as follows:

$$\text{Two neutrons} + \text{One proton} = \text{Tritium}$$

$$2(_0\text{n}^1) \quad + \quad _1\text{p}^1 \quad = \quad _1\text{H}^3$$

$$(2)(1.008665) \quad (1.007825) \quad (3.016050)$$

The numbers in parentheses correspond to the masses of the neutron, proton and tritium particles. The units are appropriately called atomic mass units (a.m.u.). One a.m.u. is, understandably, a very small mass, equivalent to 1.66056×10^{-27} kilograms. (Note that, in contrast to what was said on p. 16 – which is generally true – here the exact numbers *are* important.)

FIGURE 12-3. Binding energies of all known stable and radioactive isotopes.

In Chapter 6 we saw that mass is conserved in chemical reactions. In other words, the sum of the masses of the reactants (for example, methane and oxygen) is equal to the sum of the masses of the products (carbon dioxide and water). This law of mass conservation is not valid for nuclear reactions. We see in the nuclear reaction above that the total mass of the reactants (two neutrons and one proton) is larger than the mass of the product (tritium). There is another law, a very familiar one too, that governs this transformation. In nuclear reactions, mass is converted into energy according to the Einstein equation:

$$\text{Change in energy} = [\text{Change in mass}] [\text{Speed of light}]^2$$

or

$$\Delta E = \Delta m \, c^2$$

The speed of light is a universal constant, equivalent to some 300 million meters per second.
     So the total energy of the nucleus, or of the bound system, is lower than the total energy of the separated nucleons, and one must expend energy to separate them. In the example of the reaction of a proton and two neutrons to form a tritium nucleus, we have:

$$\Delta m = [(2)(1.008665 \text{ a.m.u.}) + 1.007825 \text{ a.m.u.}] - 3.016050 \text{ a.m.u.} = 0.009105 \text{ a.m.u.}$$

The energy expended in this process is the binding energy of tritium. It is calculated using the Einstein equation, as follows:

$$\Delta E = \Delta m\ c^2 = (0.009105\ \text{a.m.u.})\ (\frac{1.66056 \times 10^{-27}\ \text{kg}}{1\ \text{a.m.u.}})\ (3 \times 10^8\ \frac{\text{m}}{\text{s}})^2 =$$

$$= 1.360745 \times 10^{-12}\ \text{J} =$$

$$= (1.360745 \times 10^{-12}\ \text{J})\ (\frac{6.242 \times 10^{12}\ \text{MeV}}{1\ \text{J}}) = 8.494\ \text{MeV}.$$

This is the total binding energy of the nucleus. Because tritium has three nucleons, its binding energy per nucleon is thus:

$$\text{Binding energy per nucleon} = \frac{\text{Total binding energy}}{\text{Number of nucleons}} = \frac{8.494\ \text{MeV}}{3\ \text{nucleons}} = 2.831\ \text{MeV}$$

This number is shown in Figure 12-4, along with the binding energies of other isotopes that are of interest to us. These binding energies are obtained in an analogous fashion.

Illustration 12-1. Calculate the binding energy of uranium-235 ($_{92}U^{235}$), if its atomic mass is 235.043943 a.m.u.

Solution.
The sum of the masses of the 92 protons and 143 neutrons in uranium-235 is

$$(92)\ (1.007825\ \text{a.m.u.}) + (143)\ (1.008665\ \text{a.m.u.}) = 236.958995\ \text{a.m.u.}$$

This is in excess of the atomic mass of uranium-235. The total nuclear binding energy is thus

$$236.958995 - 235.043943 = 1.915052\ \text{a.m.u.} =$$

$$= (1.915052\ \text{a.m.u.})\ (\frac{931\ \text{MeV}}{1\ \text{a.m.u.}}) = 1782.9\ \text{MeV}$$

The binding energy per nucleon is therefore $1782.9/235 = 7.59$ MeV/nucleon. This number is also shown in Figure 12-4.

FIGURE 12-4. Binding energies of isotopes used in fission and fusion reactions.

Now we are in a position to define the processes of nuclear fission and fusion. Figure 12-4 shows that the binding energy per nucleon is about 7.6 MeV for heavy nuclei (those having a mass number approximately equal to 240, e.g., uranium) and about 8.5 MeV for nuclei of intermediate mass.

Nuclear fission is a nuclear reaction in which a heavy nucleus is fragmented into intermediate-mass nuclei.

In the process of fission, the stability of the intermediate-mass nuclei is approached from the right of the diagram in Figure 12-4. An excess binding energy of about 0.9 MeV per nucleon (8.5 minus 7.6) is released. If one approaches the stability of the intermediate-mass nuclei from the low-mass-number end of Figure 12-4, much more energy per nucleon is released, because of the greater difference in binding energies between the light nuclei (for example, 2.83 MeV per nucleon for $_1H^3$) and the intermediate-mass nuclei (8.5 MeV per nucleon).

Nuclear fusion is a nuclear reaction in which two light nuclei are combined to form a heavier nucleus.

It took scientists and engineers about half a century, from Becquerel's discovery of radioactivity in 1896, to come up with devices that can take advantage of this tremendous amount of energy. Their activity was particularly intense during World War II, giving rise

to operations such as the Manhattan Project in the U.S. The first practically significant fission reaction was achieved in 1942 at the University of Chicago by a team of scientists led by Enrico Fermi. The first device was an atomic bomb because each one of the warring parties was afraid that the enemy would be the first one to develop this new weapon. The development of better controlled and more useful devices followed a decade later. In 1954 President Eisenhower broke ground for the world's first nuclear electric power plant in Shippingport, Pennsylvania. In 1957 the plant began delivering electricity and Lewis Strauss, chairman of the U. S. Atomic Energy Commission, announced to the world that nuclear power would soon be "too cheap to meter."

The principles and applications of nuclear fission and fusion – the two basic ways of using nuclear energy – are discussed in more detail in the following two chapters. Why Lewis Strauss was wrong is explored in Chapters 13 and 15.

REVIEW QUESTIONS

12-1. The nucleus of a helium atom (4.0028 a.m.u.) weighs less than the sum of its parts. In the fusion of 2 protons and 2 neutrons, the helium nucleus is formed and its binding energy is released. Calculate how much energy is released in MeV/nucleon (see Figure 12-4) and in BTU per pound of helium. A proton weighs 1.0078 a.m.u. and a neutron weighs 1.0090 a.m.u. Four kilograms of helium are equivalent to 6×10^{26} a.m.u.

12-2. Indicate whether the following statements are true or false:
(a) Deuterium and tritium are isotopes of uranium.
(b) Different forms of an element having the same number of neutrons but with different numbers of protons are called isotopes.
(c) Society's goal in performing a nuclear reaction is to release the radiation that accompanies it.
(d) The heat released per pound of a nuclear fuel is much greater than the heat released per pound of a fossil fuel.

INVESTIGATIONS

12-1. Today nuclear scientists take great care when they deal with radiation. The pioneers in the investigations of nuclear energy were probably not aware of the harmful effect of radiation on their health. We mentioned the notorious cases of the Curies. Find out about the causes of death of W. Röntgen, H. Becquerel, E. Rutherford, J. Chadwick and E. Fermi.

12-2. The equation on p. 219 that Einstein formulated in 1905 was part of a revolutionary theory of space and time. What is the name of this theory? Summarize in a few understandable sentences what this theory is all about.