Chapter 7

Nuclear Energy

The various types of energy discussed so far—mechanical, thermal, chemical, electrical, radiant—are sufficient to account for nearly all of the energy transformations that we observe in the universe. There are a few important exceptions, however. Nuclear reactors extract a million times more energy from each uranium atom than we would expect from a typical chemical reaction. Nuclear bombs do the same only more quickly, while natural radioactive decay processes similarly release huge amounts of energy on a per-atom basis. And the sun itself, in order to maintain such an enormous energy output over billions of years, must be drawing on some highly concentrated energy source in its interior.

The explanation of all these phenomena lies in the atomic **nucleus**, the tiny dot of dense matter that forms the core of every atom. Whereas a *chemical* reaction is a rearrangment of the atoms that make up one or more molecules, a **nuclear reaction** is a rearrangment of the much smaller particles that make up an atomic nucleus.

Most of the space in every atom is taken up by what I visualize as a fuzzy cloud of electrons, typically about 10^{-10} meters in diameter. The negatively charged electrons are held in place by the positive charge of the nucleus at the center of the atom. Ordinarily, there are just enough electrons to precisely balance the positive charge of the nucleus. The outermost electrons are rather easily removed, however. This allows them to move around in metallic crystals, jump from atom to atom during frictional interactions, and overlap with other atoms to form molecular bonds.

Although atoms are small, atomic nuclei are much smaller still: about 100,000 times smaller in diameter, or roughly 10^{-15} meters. Within this tiny space are the positively charged protons that pull on the electrons electrostatically. But the protons also *repel* each other, and this repulsion is extremely strong because they're so close together. Why doesn't the nucleus fly apart due to this enormous electrostatic force? There must be some other, attractive force that binds the protons to each other. This force is called the **strong nuclear force**, because it must be very strong.

The mass of an individual proton (which can be measured by firing a beam of protons through a magnetic field and seeing how strongly the beam is deflected) turns out to be almost the same as that of a hydrogen atom; the mass of an individual electron is nearly 2000 times less, enough to make up the rest of the atom's mass. A simple hydrogen atom, then, consists of just a single electron in a cloud surrounding a single-proton nucleus.

The next simplest atom, helium, contains two electrons and therefore two protons, but its mass is *four* times as great as that of a hydrogen atom. The extra mass comes from particles called **neutrons**, which are about as heavy as protons but carry no electric charge. Like protons, neutrons live inside the atomic nucleus and are subject to the strong nuclear force. An ordinary helium nucleus contains two protons and two neutrons.

Most of the other elements contain about as many neutrons as protons, though the balance usually isn't exact, and the very heavy elements (like uranium) tend to contain a significant excess of neutrons. Furthermore, a nucleus with a given number of protons need not always contain the same number of neutrons. For instance, although ordinary hydrogen contains no neutrons, there is an uncommon variety of hydrogen, called **deuterium** or **heavy hydrogen**, whose nucleus contains one neutron (in addition to the proton). The extra neutron doesn't affect the atom's chemical behavior (which is governed entirely by the number of electrons present), but does increase the atom's mass, hence the name heavy hydrogen. A still heavier variety of hydrogen contains *two* neutrons per atom; it is called **tritium**, because it is about three times as heavy as ordinary hydrogen.

The different varieties of an element, differing only in the number of neutrons present in the nucleus, are called **isotopes**. Some elements have more than a dozen known isotopes. When we want to specify a particular isotope, we normally write a number to the upper-left of the symbol for the element, as in $^{2}\mathrm{H}$, or $^{14}\mathrm{C}$ or $^{238}\mathrm{U}$. The number indicates the *total* number of protons and neutrons in the nucleus, and is thus a rough measure of the isotope's mass (in units of the mass of a single proton or neutron). Thus, $^{2}\mathrm{H}$ is deuterium, with one proton and one neutron; $^{14}\mathrm{C}$ is a rare isotope of carbon, with six protons and eight neutrons; $^{238}\mathrm{U}$ is the most common isotope of uranium, with a total of 238 protons and neutrons (92 protons and 146 neutrons). The number of protons is determined by the name of the element, but sometimes we write it to the bottom-left of the symbol, as in $^{2}_{1}\mathrm{H}$, $^{14}_{6}\mathrm{C}$, or $^{238}_{92}\mathrm{U}$.

Although different isotopes of the same element behave the same chemically, their nuclei can behave quite differently. The most important example of such behavior is the spontaneous **decay** of many isotopes. For example, a $^{238}\mathrm{U}$ nucleus can spontaneously split into a $^{234}_{90}\mathrm{Th}$ (thorium-234) nucleus and a $^{4}\mathrm{He}$ (helium-4) nucleus. A $^{14}\mathrm{C}$ nucleus, on the other hand, can spontaneously convert to $^{14}_{7}\mathrm{N}$ (nitrogen-14), while emitting an electron. The reaction equations for these decays can be written

$$^{238}_{92}\mathrm{U} \longrightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He} \tag{7.1}$$

and

$$^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e}. \tag{7.2}$$

Notice that I've given the electron (e) a superscript of 0, since it contains no protons or neutrons, and a subscript of -1, since it contains one unit of negative electric charge. (For the various nuclei in the reactions, the number of units of electric charge is equal to the number of protons.)

These two examples illustrate the most common types of **radioactive decay**. Many other isotopes decay in similar ways, emitting either a helium-4 nucleus (sometimes called an **alpha particle**) or an electron (sometimes called a **beta particle**). Notice that in both examples, the *total* number of protons plus neutrons is the same after the reaction as before. This is true of all nuclear reactions, so we can suppose that there is a law of nature requiring that the number of protons plus neutrons is *conserved* in any reaction. Similarly, the net amount of electric charge is conserved in any reaction: when a neutron converts to a proton in reaction 7.2, it must emit an electron so that the net electric charge doesn't change. Note that the second reaction does *not* conserve the number of protons or the number of neutrons separately. The conserved quantities are the *total* number of protons plus neutrons (collectively called **nucleons**), and the net electric charge.

Hundreds of isotopes are known to undergo spontaneous radioactive decay. You can never predict *when* a particular radioactive nucleus will decay, but each species of isotope has a certain intrinsic *probability* of decaying in any given time period. Usually we characterize this probability in terms of the time that must pass before a nucleus has a 50/50 chance of decaying. This time is called the isotope's **half-life**. The half-life of uranium-238 is one of the longest known, about 4.5 billion years; the half-life of carbon-14 is about six thousand years. Some isotopes are so unstable that their half-lives are measured in seconds or even milliseconds or microseconds. On the other hand, many isotopes are completely stable, never decaying at all (as far as we can tell).

Spontaneous radioactive decay converts a considerable amount of energy from nuclear energy (stored in the original nucleus) into kinetic energy (mostly carried away by the emitted ^{4}He nucleus or electron). This energy is the main source of the earth's internal thermal energy, as certain long-lived isotopes, left over from the earth's formation billions of years ago, slowly decay away. Radioactive decay is also extremely useful in medical diagnosis, although there the isotopes used are chosen for their half-lives measured in hours or days.

Although alpha and beta emission are the most common types of radioactive decay, other types of decays are possible. The most important of these is when a heavy nucleus splits into two smaller nuclei of roughly equal size, plus a few left-over neutrons. This process is called **nuclear fission**. Although it is relatively rare, it has the following crucial property: the left-over neutrons can trigger the fission of other, neighboring nuclei, leading to a **chain reaction** in which a large number of nuclei all undergo fission in a relatively short time. This is the principle of all nuclear reactors, and also of the original "atomic" bomb. The most common fissionable isotopes are uranium-235 (which makes up a little under one percent of all naturally occurring uranium) and plutonium-239 (which can be manufactured by bombarding uranium-238 with neutrons).

While fission splits very heavy nuclei into lighter ones, the opposite process, combining light nuclei to form heavier ones, can also release energy. This process is called **nuclear fusion**. The most important example is the series of nuclear reactions that power the sun. These reactions occur in several steps, but the net result is

$$4\,{}^1_1\mathrm{H} + 2\,{}^{\;\,0}_{-1}\mathrm{e} \longrightarrow {}^4_2\mathrm{He}, \tag{7.3}$$

the combination of four hydrogen nuclei (protons) with two electrons to form a single helium-4 nucleus. Fusion reactions inside some other stars can even extract energy by combining helium nuclei to form carbon and oxygen, then combining these to form heavier nuclei, all the way up to iron-56, the most stable nucleus of all. Still heavier nuclei can be formed by other nuclear reactions, but these reactions require energy input, rather than releasing nuclear energy.

Exercise 7.1. Determine the identity of "X" in each of the following nuclear reactions, referring to a periodic table as necessary. Note that α represents an alpha particle (helium-4 nucleus), and β represents a "beta particle" (an electron).

(a) $^{239}\mathrm{Pu} \longrightarrow \mathrm{X} + \alpha$ (an example of alpha decay)

(b) $^{3}\mathrm{H} \longrightarrow \mathrm{X} + \beta$ (an example of beta decay)

(c) $^{239}\mathrm{Pu} + \mathrm{n} \longrightarrow {}^{92}\mathrm{Kr} + \mathrm{X} + 2\mathrm{n}$ (an example of fission)

(d) $^{4}\mathrm{He} + {}^{4}\mathrm{He} + {}^{4}\mathrm{He} \longrightarrow \mathrm{X}$ (an example of fusion)