Chapter 2
A Short History of Nuclear Physics to the Mid-1930s

Until the late 1930s, the study of radioactivity and nuclear physics were relatively low-profile academic research fields whose applications were limited primarily to medical treatments such as radiation therapies for cancers. But within a very few months between late 1938 and mid-1939, some investigators began to realize that one of the quieter provinces of pure research could become a geopolitical game-changer of immense military potential. How did this transformation come to be?

To set the stage for a description of the development of nuclear weapons, it is helpful to understand a lengthy progression of underlying background discoveries. Our understanding that atoms comprise protons, neutrons, and electrons; that nuclei of various elements occur in different isotopic forms; that some elements are radioactive; and that nuclear weapons can be made are now facts of common knowledge: not many people know, or ever wonder how, their predecessors divined such knowledge. The purpose of this chapter is to give an overview of how the scientific community came to these understandings.

In developing such an overview there are many facts to be considered, and they interconnect so tightly that it can be difficult to decide how to order the presentation. I use a largely chronological ordering, but slavishly maintaining a chronological description would be awkward in that full understanding of some phenomena took decades to develop. A dramatic example of this is that neutrons were not discovered until 1932, a full twenty years after the existence of nuclei had been established (Fig. 2.1). Consequently, there are points where I abandon the chronological approach for the sake of coherence or to digress on some background material. Also, there are instances where a concept introduced in one section is revisited in a later one for fuller elaboration. Readers are urged to consider the sections of this chapter as linked units; treat them as a whole, and re-read them as necessary.

The discovery of the neutron in 1932 is considered such a pivotal event that nuclear physicists divide the history of their discipline into two eras: that time before awareness of the neutron, and that time after. In keeping with this, the first section of this chapter covers, in a number of subsections, important developments from the discovery of radioactivity in 1896 up to 1931. Sections 2.2 through 2.4

take the story from 1932 through the discoveries of the neutron and artificially-induced radioactivity, and Enrico Fermi's neutron-bombardment experiments of the mid-1930s. Section 2.5, which can be considered optional, fills in some technical details that are skirted in Sect. 2.1.4.
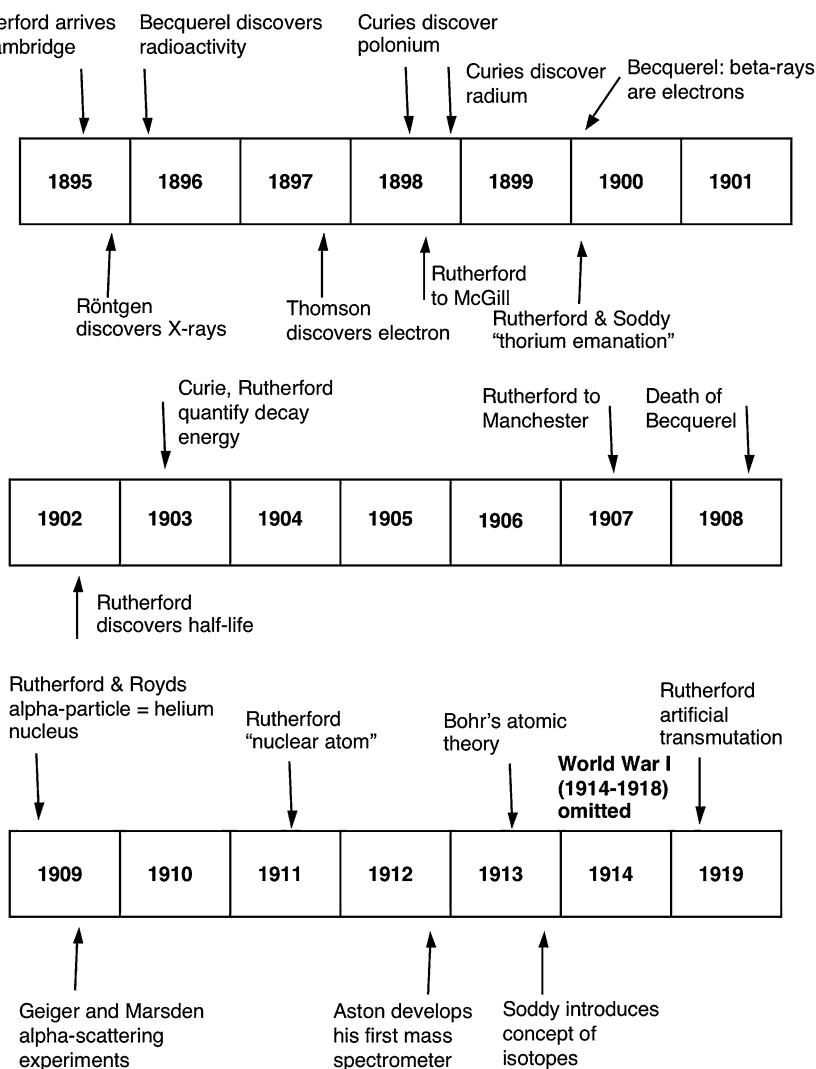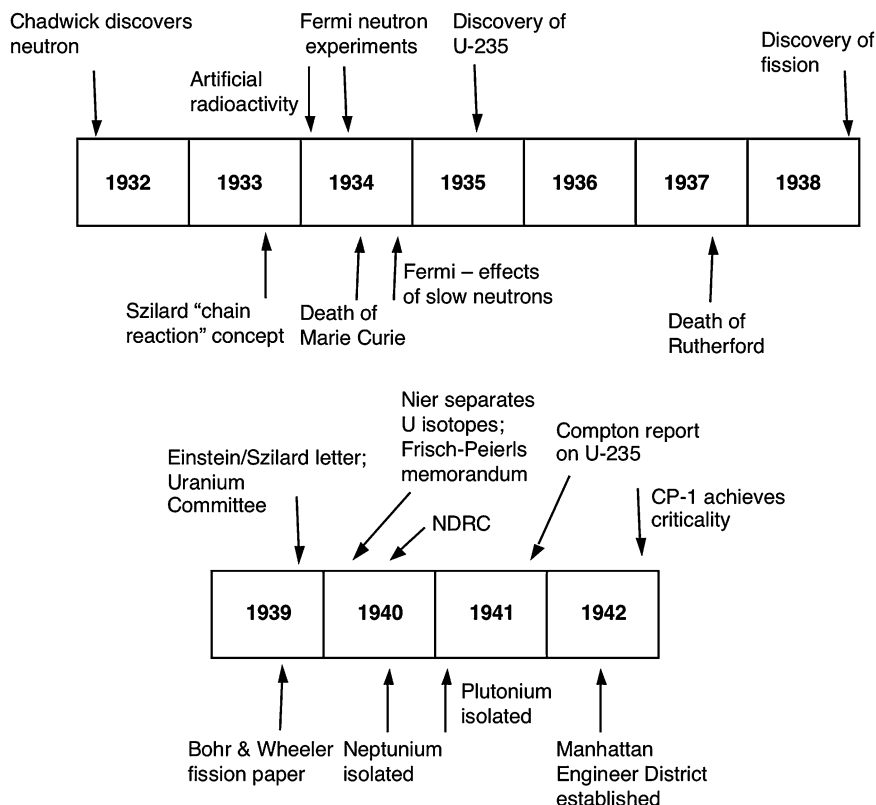

Fig. 2.1 Chronology of early nuclear physics

Chadwick discovers neutron

Fermi neutron experiments

Discovery of U-235

Discovery of fission

Artificial radioactivity

| 1932 | 1933 | 1934 | 1935 | 1936 | 1937 | 1938 |

Szilard "chain reaction" concept

Death of Marie Curie

Fermi – effects of slow neutrons

Death of Rutherford

Nier separates U isotopes; Frisch-Peierls memorandum

Compton report on U-235

Einstein/Szilard letter; Uranium Committee

NDRC

CP-1 achieves criticality

| 1939 | 1940 | 1941 | 1942 |

Bohr & Wheeler fission paper

Neptunium isolated

Plutonium isolated

Manhattan Engineer District established

Fig. 2.1 continued

2.1 Radioactivity, Nuclei, and Transmutations: Developments to 1932

The era of "modern" physics is usually considered to have begun in late 1895, when Wilhelm Conrad Röntgen, working in Germany, accidentally discovered X-rays. Röntgen discovered that not only could his mysterious rays pass through objects such as his hand, but they also ionized air when they passed through it; this was the first known example of what we now call "ionizing radiation". A part of Röntgen's discovery involved X-rays illuminating a phosphorescent screen, a fact which caught the attention of Antoine Henri Becquerel, who lived in France. Becquerel was an expert in the phenomenon of phosphorescence, where a material emits light in response to illumination by light of another color. Becquerel wondered if phosphorescent materials such as uranium salts might be induced to emit X-rays if they were exposed to sunlight. While this supposition was wrong, investigating it led Becquerel, in February 1896, to the accidental discovery of
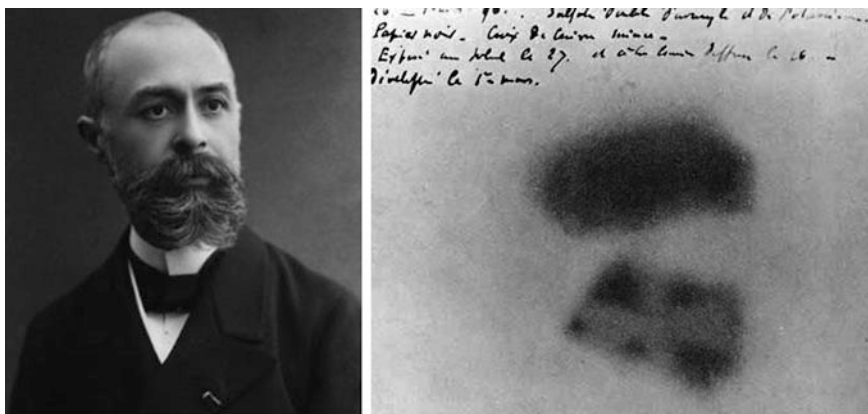
Fig. 2.2 Henri Becquerel (1852–1908) and the first image created by "Becquerel rays" emitted by uranium salts placed on a wrapped photographic plate. In the lower part of the plate a Maltese cross was placed between the plate and the lump of uranium ore. *Sources* http://upload.wikimedia.org/wikipedia/commons/a/a3/Henri_Becquerel.jpg; http://upload.wikimedia.org/wikipedia/commons/1/1e/Becquerel_plate.jpg

radioactivity. Becquerel observed that samples of uranium ores left on top of wrapped photographic plates would expose the plates even in the *absence* of any external illumination; that is, the exposures seemed to be created by the uranium itself. When the plates were unwrapped and developed, an image of the samples would be apparent (Fig. 2.2). Nuclear physics as a scientific discipline originated with this discovery.

We now attribute these exposures to the action of so-called "alpha" and "beta" particles emitted by nuclei of uranium and other heavy elements as they decay naturally to more stable elements; the nature of these particles is elaborated in the sections that follow. Some of the decay timescales are fleeting, perhaps only minutes, while others are inconceivably long, hundreds of millions or billions of years. In the latter event it is only because there are so many trillions upon trillions of individual atoms in even a small lump of ore that enough are likely to decay within a span of seconds or minutes to leave an image on a film or trigger a Geiger counter. (The operation of a Geiger counter is described at the end of Sect. 2.1.5.) A third form of such emission, "gamma rays," was discovered by French chemist Paul Villard in 1900. Gamma-rays are photons, just like those entering your eyes as you read this, but of energies about a million times greater than visible-light photons.

By the time of Becquerel's death in 1908, the field of research he had opened was producing developments which would lead, by about 1920, to humanity's first true scientific understanding of the structure of the most fundamental constituents of matter: atoms and their nuclei. Remarkably, just less than a half-century would pass between the discovery of radioactivity and the development of nuclear weapons.

2.1.1 Marie Curie: Polonium, Radium, and Radioactivity

Becquerel's work came to the attention of Marie Sklodowski, a native of Poland who had graduated from the Sorbonne (part of the University of Paris) with a degree in physical science in 1893; the following year she would add a degree in mathematics. In 1895 she married Pierre Curie, a physicist at the Paris School of Physics and Chemistry (Fig. 2.3). Seeking a subject for a doctoral thesis, Marie turned to Becquerel's work, a subject about which not a great deal had been published. She set up a laboratory in her husband's School, and began work in late 1897.

Becquerel had reported that the energetic "rays" emitted by uranium could ionize air as they passed through it; in modern parlance the rays collide with molecules in the air and cause them to lose electrons. Pierre Curie and his brother had developed a device known as an electrometer for detecting minute electrical currents. Making use of this device, Marie found that the amount of electricity generated was directly proportional to the amount of uranium in a sample. Testing other materials, she found that the heavy element thorium also emitted Becquerel rays (a fact discovered independently by Gerhard Schmidt in Germany), although not as many per gram per second as did uranium. Further work, however, revealed that samples of pitchblende ore, a blackish material rich in uranium oxides, emitted more Becquerel rays than could be accounted for solely by the quantity of uranium that they contained. Drawing the conclusion that there must be some other "active element" present in pitchblende, Curie began the laborious task of chemically isolating it from the tons of ore she had available. By this time, Pierre had abandoned his own research on the properties of crystals in order to join Marie in her work.

Spectroscopic analysis of the active substance proved that it was a new, previously unknown element. Christening their find "polonium" in honor of Marie's
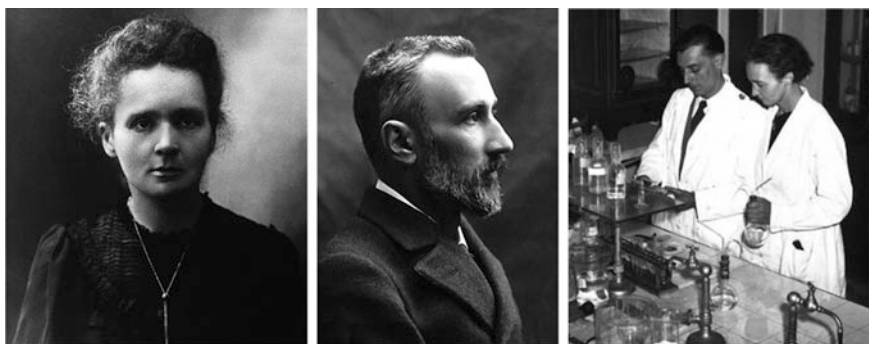

Fig. 2.3   Marie (1867–1934) and Pierre (1859–1906) Curie; Right Irène (1897–1956) and Frédéric Joliot-Curie (1900–1958) in 1935. *Sources* http://commons.wikimedia.org/wiki/File:Mariecurie. jpg; http://commons.wikimedia.org/wiki/File:PierreCurie.jpg; http://commons.wikimedia.org/ wiki/File:Irène_et_Frédéric_Joliot-Curie_1935.jpg

native country, they published their discovery in July, 1898, in the weekly proceedings of the French Academy of Sciences. That paper introduced two new words to the scientific community: "radioactivity" to designate whatever process deep within atoms was giving rise to Becquerel's ionizing rays, and "radioelement" to any element that possessed the property of doing so. The term "radioisotope" is now more commonly used in place of "radioelement," as not all of the individual isotopes of elements that exhibit radioactivity are themselves radioactive.

In December, 1898, the Curies announced that they had found a second radioactive substance, which they dubbed "radium." By the spring of 1902, after starting with ten tons of pitchblende ore, they had isolated a mere tenth of a gram of radium, which was enough for definite spectroscopic confirmation of its status as a new element. In the summer of 1903 Marie defended her thesis, "Researches on Radioactive Substances," and received her doctorate from the Sorbonne. In the fall of that year the Curies would be awarded half of the 1903 Nobel Prize for Physics; Henri Becquerel received the other half.

2.1.2 Ernest Rutherford: Alpha, Beta, and Half-Life

In the fall of 1895, Ernest Rutherford (Fig. 2.4), a New Zealand native, arrived at the Cavendish Laboratory of Cambridge University in England on a postgraduate scholarship. The Director of the Laboratory was Joseph John "J. J." Thomson, who in the fall of 1897 was credited with discovering the electron, the fundamental, negatively-charged particles of matter which account for the volumes of atoms. It is rearrangements of the outermost electrons of atoms which cause the chemical reactions by which, for example, we digest meals to provide ourselves
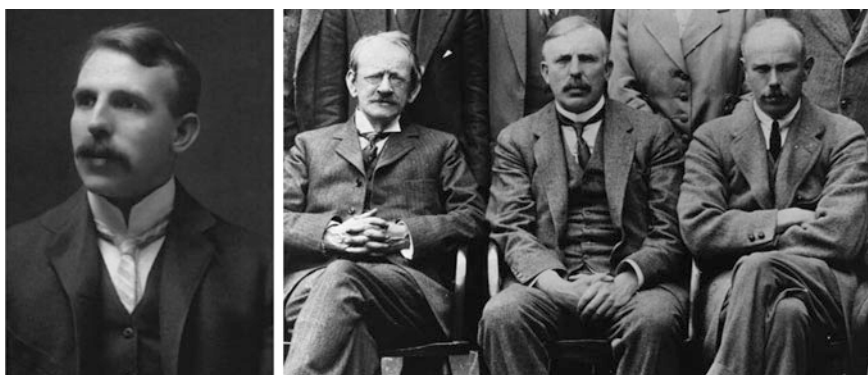

Fig. 2.4 Left Ernest Rutherford (1871–1937) about 1910. Right Seated left to right in this 1921 photo are J. J. Thomson (1856–1940), Rutherford, and Francis Aston (1877–1945), inventor of the mass spectrograph (Sect. 2.1.4). *Sources* http://commons.wikimedia.org/wiki/File:Ernest_Rutherford.jpg; AIP Emilio Segre Visual Archives, Gift of C.J. Peterson

with the energy we use to do useful work such as the preparation of book manuscripts.

Rutherford's intrinsic intelligence, capacity for sheer hard work, and unparalleled physical insight combined with propitious timing to set him on a path to become one of history's great nuclear pioneers. Soon after Rutherford arrived in Cambridge, Röntgen discovered X-rays. As a student, Rutherford had developed considerable experience with electrical devices, and the Cavendish Laboratory was well-equipped with Thomson's "cathode ray tubes," the core apparatus for generating X-rays. Rutherford soon began studying their ionizing properties. When the discovery of radioactivity was announced, it was natural for him to turn his attention to this new ionizing phenomenon.

Rutherford discovered that he could attenuate some of the uranium activity by wrapping the samples in thin aluminum foils; adding more layers of foil decreased the activity. Rutherford deduced that there appeared to be two types of radiation present, which he termed "alpha" and "beta." Alpha-rays could be stopped easily by a thin layer of foil or a few sheets of paper, but beta-rays were more penetrating. Henri Becquerel later showed that both types could be deflected by a magnetic field, but in opposite directions and by differing amounts. This meant that the rays must be electrically charged; alphas proved to be positive, and betas negative. Becquerel also later proved that beta rays were identical to electrons. Alpha-rays were much less affected by a magnet, which meant that they must be much more massive than electrons (further details on this point appear in Sect. 2.1.4).

In the fall of 1898, Rutherford completed his studies at Cambridge, and moved to McGill University in Montreal, Canada, where he had been appointed as the McDonald Professor of Physics. Over the next three decades he continued his radioactivity research, both at McGill and later back in England. This research would contribute to a series of groundbreaking discoveries in the area of atomic structure, and would earn him the 1908 Nobel Prize for Chemistry.

Rutherford's first major discovery at McGill occurred in 1900, when he found that, upon emitting its radiation, thorium simultaneously emitted a product which he termed "emanation." The emanation was also radioactive, and, when isolated, its radioactivity was observed to decline in a geometrical progression with time. Specifically, the activity decreased by a factor of one-half for every minute of time that elapsed. Rutherford had discovered the property of *radioactive half-life*, the quintessential natural exponential decay process.

As an example, suppose that at "time zero" you have 1,000 atoms of some isotope that has a half-life of 10 days. You can then state that 500 of them will have decayed after 10 days. You cannot predict *which* of the 500 will have decayed, however. Over the following 10 days a further 250 of the original remaining atoms will decay, and so on. Remarkably, the probability that a given atom will decay in some specified interval of time is completely independent of how long it has managed to avoid decaying; in the subatomic world, age is *not* a factor in the probability of continued longevity.

The following paragraphs examine the mathematics of half-life. Readers who wish to skip this material should proceed to the paragraph following Eq. (2.8).

If the number of nuclei of some radioactive species at an arbitrarily-designated starting time $t = 0$ is N_O, then the number that remain after time t has elapsed can be written as

$$N(t) = N_O e^{-\lambda t}, \tag{2.1}$$

where λ is the so-called decay constant of the species. If $t_{1/2}$ is the half-life of the species for some mode of decay (alpha, beta, …), λ is given by

$$\lambda = \frac{\ln 2}{t_{1/2}}. \tag{2.2}$$

Since half-lives run from tiny fractions of a second to billions of years, there is no preferred unit for them; one must be careful to put t and $t_{1/2}$ in the same units when doing calculations.

What is measured in the outside world is the rate of decays R, which is determined by taking the derivative of (2.1):

$$R(t) = \frac{dN(t)}{dt} = -\lambda N_O e^{-\lambda t} = -\lambda N(t). \tag{2.3}$$

The meaning of the negative sign is that the number of nuclei of the original species steadily decreases as time goes on; what is customarily quoted is the *absolute value* of $R(t)$.

Since one is more likely to know the mass of material m that one is working with rather than the number of atoms, it is helpful to have a way of relating these two quantities. This is given by

$$N = \frac{m N_A}{A}, \tag{2.4}$$

where N_A is Avogadro's number and A is the molecular weight of the species. Tradition is to quote A in grams per mole, which means that m must be expressed in grams.

Marie and Pierre Curie adopted the rate of decay of a freshly-isolated one-gram sample of radium-226 as a standard for comparing radioactivity rates of different substances. This isotope, which has a half-life of 1,599 years, is a rather prodigious emitter of alpha-particles. With $A = 226.025$ g/mol,

$$N_O = \frac{m N_A}{A} = \frac{(1 \text{ g})(6.022 \times 10^{23} \text{nuclei}/mol)}{(226.025 \text{ g}/mol)} = 2.664 \times 10^{21} \text{ nuclei.} \tag{2.5}$$

To compute the decay rate in nuclei per second, convert 1,599 years to seconds; 1 year $= 3.156 \times 10^7$ s. Hence 1,599 years $= 5.046 \times 10^{10}$ s, and the initial decay rate will be

$$R_O = \lambda N_O = \frac{(\ln 2)(2.664 \times 10^{21} \text{ nuclei})}{(5.046 \times 10^{10} \text{ s})} = 3.66 \times 10^{10} \text{ nuclei}/s. \tag{2.6}$$

This rate of activity, slightly rounded, is now known as the *Curie*, abbreviated *Ci*:

$$1 \, Ci = 3.7 \times 10^{10} \text{decay/s.} \tag{2.7}$$

This is a large number, but a gram of radium contains some 10^{21} atoms and so will maintain its activity for a long time; even after several years it will possess essentially the same rate of activity it began with.

In many situations, a Curie is too large a unit of activity for practical use, so in technical papers one often encounters *millicuries* (one-thousandth of a Curie, *mCi*) or *microcuries* (one-millionth of a Curie; *μCi*). Household smoke detectors contain about 1 *μCi* of radioactive material (37,000 decays per second) which ionizes a small volume of air around a sensor in order to aid in the detection of smoke particles. A more modern unit of activity is the *Becquerel* (*Bq*); one *Bq* is equal to one decay per second. In this unit, a smoke detector would be rated as having an activity of 37 kiloBecquerels (*kBq*). If you would like to try a quick exercise, imagine that you have 1 *kilogram* of plutonium-239, which has $A = 239.05$ g/mol and a half-life for alpha-decay of 24,100 years. You should be able to prove that the decay rate would be 62 *Ci*, a substantial number. We will see in Chap. 7 that decay rates are an important consideration in nuclear weapons engineering.

To better understand the comment above about decay probability being independent of age, consider the following argument. If the number of undecayed nuclei at some time is $N(t)$, then (2.3) tells us that in the subsequent dt seconds the number that will decay is $dN = \lambda N(t) \, dt$. The probability $P(t, dt)$ that any given nucleus will decay during these dt seconds is therefore the number that do decay, divided by the number that was available at the start of the interval:

$$P(t, dt) = \frac{\lambda N(t) \, dt}{N(t)} = \lambda \, dt. \tag{2.8}$$

As claimed, $P(t, dt)$ is independent of the time t that had elapsed before the interval considered.

Rutherford sought to identify what element the thorium "emanation" actually was, and to do so teamed up with Frederick Soddy, a young Demonstrator in Chemistry (Fig. 2.5). They had expected the emanation to be some form of thorium, but Soddy's analysis revealed that it behaved like an inert gas. This suggested that the thorium was spontaneously transmuting itself into another element, a conclusion that would prove to be one of the pivotal discoveries of twentieth-century physics. Soddy initially thought that the emanation was argon (element 18), but it would later come to be recognized as radon (element 86). Various isotopes of thorium, radium, and actinium decay to various isotopes of radon, which themselves subsequently decay.

In the 1920s half-life came to be understood as a quantum–mechanical process that is a manifestation of the wave-nature of particles at the atomic level; it is a purely probabilistic phenomenon. This more sophisticated understanding has no bearing on the issues discussed in this book, however: for our purposes, we can

regard radioactive decay as an empirical phenomenon described by the mathematics developed above.

2.1.3 Units of Energy in Nuclear Physics and the Energy of Radioactive Decay

In this and the following section we break with chronological progression to fill in some background physics on the units of energy conventionally used in nuclear physics, and the history of the discovery of isotopes. We will return to Rutherford in Sect. 2.1.5.

In the rare circumstances when people consider the quantities of energy that they consume or produce, the unit of measure involved will likely be something such as the kilowatt-hours that appear on an electric bill or the food-calories on a nutrition label. Science students will be familiar with units such as Joules and physical calories (1 cal = 4.187 J). The *food* calorie appearing on nutrition labels is equivalent to 1,000 physical calories, a so-called *kilocalorie*. The food calorie was introduced because the physical calorie used by physicists and chemists is inconveniently small for everyday use.

The words *energy* and *power* are often confused in common usage. Power is the *rate* at when energy is created or used. For physicists, the standard unit of power is

the Watt, which is equivalent to producing (or consuming) one Joule of energy per second. A kilowatt (kW) is 1,000 W, or 1,000 J/s. A kilowatt-hour (kWh) is 1,000 W times one hour, that is, 1,000 J/s times 3,600 s, or 3.6 million Joules. A 60-W bulb left on for one hour will consume (60 J/s)(3,600 s) = 216,000 J, or 0.06 kWh. If electricity costs 10 cents per kWh, your bill for that hour will be six-tenths of one cent. The cost would still be a bargain at twice the price, so you can keep reading.

When dealing with processes that happen at the level of individual atoms, however, calories, Joules, and Watts are all far too large to be convenient; one would be dealing with exceedingly tiny fractions of them in even very energetic reactions. To address this, physicists who study atomic processes developed a handier unit of energy: the so-called *electron-Volt*. One electron-Volt is equivalent to a mere 1.602×10^{-19} J. This oddly-named quantity, abbreviated eV, actually has a very sound basis in fundamental physics. You can skip this sentence if you are unfamiliar with electrical units, but for those in the know, an eV is technically defined as the kinetic energy acquired by a single electron when it is accelerated through a potential difference of one Volt. As an everyday example, the electrons supplied by a 1.5-V battery each emerge with 1.5 eV of kinetic energy. A common 9-V battery consists of six 1.5-V batteries connected in series, so their electrons emerge with 9 eV of energy. On an atom-by-atom basis, chemical reactions involve energies of a few eV. For example, when dynamite is detonated, the energy released is equivalent to 9.9 eV per molecule.

Nuclear reactions are much more energetic than chemical ones, typically involving energies of *millions* of electron-volts (MeV). We will see many reactions involving MeVs in this book. If a nuclear reaction liberates 1 MeV per atom involved (nucleus, really) while a chemical reaction liberates 10 eV per atom involved, the ratio of the nuclear to chemical energy releases will be 100,000. This begins to give you a hint as to the compelling power of nuclear weapons. An "ordinary" bomb that contains 1,000 pounds of chemical explosive could be replaced with a nuclear bomb that utilizes only 1/100 of a pound of a nuclear explosive, presuming that the weapons detonate with equal efficiency. Thousands of *tons* of conventional explosive can be replaced with a few tens of *kilograms* of nuclear explosive. Nuclear fission weapons like those used at Hiroshima and Nagasaki involved reactions which liberated about 200 MeV per reaction, so a nuclear explosion in which even only a small amount of the "explosive" actually reacts (e.g., one kilogram) can be incredibly devastating.

It did not take physicists long to appreciate that natural radioactivity was accompanied with substantial energy releases. In 1903, Pierre Curie and a collaborator, A. Laborde, found that just one gram of radium released on the order of 100 physical calories of heat energy per hour. Rutherford and Soddy were also on the same track. In a May, 1903, paper titled "Radioactive Change," they wrote that (expressed in modern units) "the total energy of radiation during the disintegration of one gram of radium cannot be less than 10^8 calories and may be between 10^9 and 10^{10} calories. The union of hydrogen and oxygen liberates approximately 4×10^3 calories per gram of water produced, and this reaction sets

free more energy for a given weight than any other chemical change known. The energy of radioactive change must therefore be at least 20,000 times, and may be a million times, as great as the energy of any molecular change." Another statistic Rutherford was fond of quoting was that a single gram of radium emitted enough energy during its life to raise a 500-t weight a mile high.

The moral of these numbers is that nuclear reactions liberate vastly more energy per reaction than any chemical reaction. As Rutherford and Soddy wrote: "All these considerations point to the conclusion that the energy latent in the atom must be enormous compared with that rendered free in ordinary chemical change." That enormity would have profound consequences.

2.1.4 Isotopes, Mass Spectroscopy, and the Mass Defect

In this section I give a brief history of the discovery of isotopes, and a description of the notations now used to designate them. This material is somewhat technical, but the concept of isotopy is an important one that will run throughout this book.

In modern terminology, an element's location in the periodic table is dictated by the number of protons in the nuclei of its atoms. This is known as the *atomic number*, and is designated by the letter Z. Atoms are usually electrically neutral, so the atomic number also specifies an atom's normal complement of electrons. Chemical reactions involve exchanges of so-called valence electrons, which are the outermost electrons of atoms. Quantum physics shows us that the number of electrons in an atom, and hence the number of protons in its nucleus, accounts for its chemical properties. The periodic table as it is published in chemistry texts is deliberately arranged so that elements with similar chemical properties (identical numbers of valence electrons) appear in the same column of the table.

The number of neutrons in a nucleus is designated by the letter N, and the total number of neutrons plus protons is designated by the letter A: $A = N + Z$. A is known as the *mass number*, and also as the *nucleon number*; the term *nucleon* means either a proton or a neutron. By specifying Z and A, we specify a given isotope. Be careful: A is also used to designate the *atomic weight* of an element (or isotope) in grams per mole. The atomic weight and nucleon number of an isotope are always close, but the difference between them is important. The nucleon number is always an integer, but the atomic weight will have decimals. For example, the nucleon number of uranium-235 is 235, but the atomic weight of that species is 235.0439 g/mol. The term *nuclide* is also sometimes encountered, and is completely synonymous with *isotope*.

The general form for isotope notation is

$$_Z^A X. \tag{2.9}$$

In this expression, X is the symbol for the element involved. The subscript is always the atomic number, and the superscript is always the mass number. For

example, the oxygen that you are breathing while reading this passage consists of three stable isotopes: $^{16}_{8}$O, $^{17}_{8}$O, and $^{18}_{8}$O. All oxygen atoms have eight protons in their nuclei, but either eight, nine, or ten neutrons. These nuclides are also referred to as oxygen-16 (O-16), oxygen-17 (O-17), and oxygen-18 (O-18). By far the most common isotope of oxygen is the first one: 99.757 % of naturally-occurring oxygen is O-16, with only 0.038 and 0.205 % being O-17 and O-18, respectively. Three isotopes that will prove very important in the story of the Manhattan Project are uranium-235, uranium-238, and plutonium-239: $^{235}_{92}$U, $^{238}_{92}$U, and $^{239}_{94}$Pu.

The concepts of atomic number and isotopy developed over many years. The foundations of modern atomic theory can be traced back to 1803, when English chemist John Dalton put forth a hypothesis that all atoms of a given element are identical to each other and equal in weight. An important development in Dalton's time came about when chemical evidence indicated that the masses of atoms of various elements seemed to be very nearly equal to integer multiples of the mass of hydrogen atoms. This notion was formally hypothesized about 1815 by English physician and chemist William Prout, who postulated that all heavier elements are aggregates of hydrogen atoms. He called the hydrogen atom a "protyle," a fore-runner of Ernest Rutherford's "proton." Parts of both Dalton's and Prout's hypotheses would be verified, but other aspects required modification. In partic-ular, something looked suspicious about Prout's idea from the outset, as some elements had atomic weights that were *not* close to integer multiples of that of hydrogen. For example, chlorine atoms seemed to weigh 35.5 times as much as hydrogen atoms.

The concept of isotopy first arose from evidence gathered in studies of natural radioactive decay chains (Sect. 2.1.6). Substances that appeared in different decay chains through different modes of decay often seemed to have similar properties, but could not be separated from each other by chemical means. The term "isotope" was introduced in 1913 by Frederick Soddy, who had taken a position at the University of Glasgow. Soddy argued that the decay-chain evidence suggested that "the net positive charge of the nucleus is the number of the place which the element occupies in the periodic table". Basing his hypothesis on the then-current idea that the electrically neutral mass in nuclei was a combination of protons and electrons, Soddy went on to state that the "algebraic sum of the positive and negative charges in the nucleus, when the arithmetical sum is different, gives what I call "isotopes" or "isotopic elements," because they occupy the same place in the periodic table." The root "iso" comes from the Greek word "isos," meaning "equal," and the *p* in *tope* serves as a reminder that it is the number of protons which is the same in all isotopes of a given element. In the same paper, Soddy also developed an ingenious argument to show that the electrons emitted in beta-decay had to be coming from *within the nucleus*, not from the "orbital" electrons.

True understanding of the nature and consequences of isotopy came with the invention of *mass spectroscopy*, an instrumental technique for making extremely precise measurements of atomic masses. In his 1897 work, J. J. Thomson mea-sured the ratio of the electrical charge carried by electrons to their mass by using

electric and magnetic fields to deflect them and track their trajectories. In 1907, Thomson modified his apparatus to investigate the properties of positively-charged (ionized) atoms, and so developed the first "mass spectrometer." In this device, electric and magnetic fields were configured to force ionized atoms to travel along separate, unique parabolic-shaped trajectories which depended on the ions' masses. The separate trajectories could be recorded on a photographic film for later analysis.

In 1909, Thomson acquired an assistant, Francis Aston, a gifted instrument-maker (Fig. 2.6). Aston improved Thomson's instrument, and, in November, 1912, obtained evidence for the presence of two isotopes of neon, of mass numbers 20 and 22 (taking hydrogen to be of mass unity). The atomic weight of neon was known to be 20.2. Aston reasoned that this number could be explained if the two isotopes were present in a ratio of 9:1, as is now known to be the case. (There is a third isotope of neon, of mass 21, but it comprises only 0.3 % of natural neon.) Aston tried to separate the two neon isotopes using a technique known as *diffusion*. As described in Chap. 1, this refers to the passage of atoms through a porous membrane. Aston passed neon through clay tobacco pipes, and did achieve a small degree of enrichment.

Following a position involving aircraft research during World War I, Aston returned to Cambridge, and in 1919 he built his own mass spectrometer which incorporated some improvements over Thomson's design. In a series of papers published from late that year through the spring of 1920, he presented his first results obtained with the new instrument. These included a verification of the two

Fig. 2.6 Francis Aston (1877–1945) *Source* http://commons.wikimedia.org/wiki/File:Francis_William_Aston.jpg
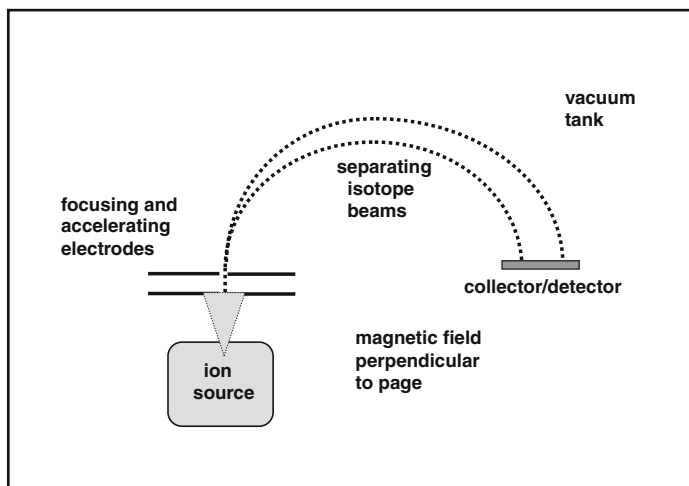
Fig. 2.7 Principle of mass spectroscopy. Positive ions are accelerated by an electric field and then directed into a magnetic field which emerges perpendicularly from the page. Ions of different mass will follow different circular trajectories, with those of greater mass having larger orbital radii

previously-detected neon isotopes, and a demonstration that chlorine comprised a mixture of isotopes of masses 35 and 37 in an abundance ratio of about 3:1. In later years (1927 and 1937), Aston developed improved instruments, his so-called second and third mass spectrometers.

The principle of Aston's mass spectrometer is sketched very simplistically in Fig. 2.7. Inside a vacuum chamber, the sample to be investigated is heated in a small oven. The heating will ionize the atoms, some of which will escape through a narrow slit. The ionized atoms are then accelerated by an electric field, and directed into a region of space where a magnetic field of strength B is present. The magnetic field is arranged to be perpendicular to the plane of travel of the positively-charged ions, that is, perpendicular to the plane of the page in Fig. 2.7; the electrical coils or magnet poles for creating the field are not sketched in the diagram. The magnetic field gives rise to an effect known as the Lorentz Force Law, which causes the ions to move in circular trajectories; an ion of mass m and net charge q that is moving with speed v will enter into a circular orbit of radius $r = mv/qB$.

If all ions are ionized to the same degree and have the same speed, heavier ones will be deflected somewhat less than lighter ones; that is, they will have larger-radius orbits. Only two different mass-streams are sketched in Fig. 2.7; there will be one stream for each mass-species present. The streams will be maximally separated after one-half of an orbit, where they can be collected on a film. Present-day models incorporate electronic detectors which can feed data to a computer for immediate analysis.
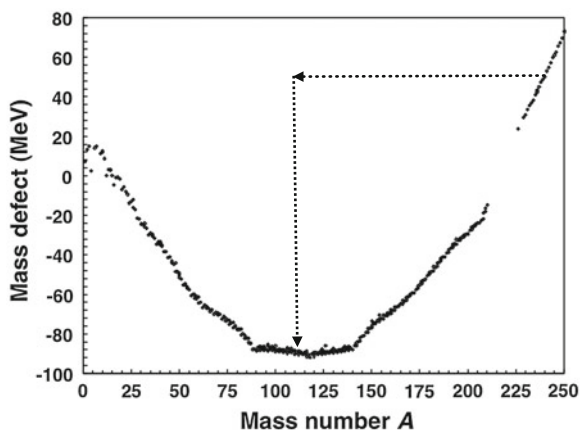
During his career, Aston discovered over 200 naturally-occurring isotopes, including uranium-238. Surprisingly, he does not have an element named after him, but he did receive the 1922 Nobel Prize for Chemistry. Aston's work showed that John Dalton's 1803 conjecture had been partially correct: atoms of the same element behave identically as far as their chemistry is concerned, but the presence of isotopy means that not all atoms of the same element have the same weight. Similarly, Aston found that Prout's conjecture that the masses of all atoms were integer multiples of that of hydrogen, if one substitutes "isotopes" for "atoms," was also *very nearly* true. But that *very nearly* proved to involve some very important physics.

What is meant by *very nearly* here? As an example, consider the common form of iron, Fe-56, nuclei of which contain 26 protons and 30 neutrons. Had Prout been correct, the mass of an iron-56 atom should be 56 "mass units," if one neglects the very tiny contribution of the electrons. (A technical aside: 56 electrons would weigh about 1.4 % of the mass of a proton. We are also assuming, for sake of simplicity, that protons and neutrons each weigh one "mass unit"; neutrons are about 0.1 % heavier than protons.) Mass spectroscopy can measure the masses of atoms to remarkable precision; the actual weight of an iron-56 atom is 55.934937 atomic mass units (see Sect. 2.5 for a definition of the atomic mass unit). The discrepancy of $55.934937 - 56 = -0.065063$ mass units, what Aston called the "mass defect," is significant, amounting to about 6.5 % of the mass of a proton. This mass defect effect proved to be systemic across the periodic table: *all stable atoms are less massive than one would predict on the basis of Prout's whole-number hypothesis*. Iron has a fairly large mass defect, but by no means the largest known (Fig. 2.8). The mass-defect is *not* an artifact of protons and neutrons having slightly different masses; if one laboriously adds up the masses of all of the constituents of atoms, the defects are still present. The unavoidable conclusion is that when protons and neutrons assemble themselves into nuclei, they give up some of their mass in doing so.

Physicists now quote mass defects in terms of equivalent energy in MeVs, thanks to $E = mc^2$. One mass unit is equivalent to 931.4 MeV, so the iron-56 mass defect amounts to just over 60 MeV. Because this is a mass *defect*, it is formally cited as a negative number, -60.6 MeV. The capital Greek letter delta (as in "Defect") is now used to designate such quantities: $\Delta = -60.6$ MeV.

Where does the mass *go* when nature assembles nuclei? Empirically, nuclei somehow have to hold themselves together against the immense mutual repulsive Coulomb forces of their constituent protons; some sort of nuclear "glue" must be present. To physicists, this "glue" is known synonymously as the "strong force" or as "binding energy," and is presumed to be the "missing" mass transformed into some sort of attractive energy. The greater the magnitude of the mass defect, the more stable will be the nucleus involved. Figure 2.8 shows a graph of the mass defects of 350 nuclides that are stable or have half-lives greater than 100 years, as a function of mass number A. The deep valley centered at $A \sim 120$ attests to the great stability of elements in the middle part of the periodic table; negative values of Δ connote intrinsic stability. The gap between $A \sim 210$ and 230 is due to the

Fig. 2.8 Mass defect in MeV versus mass number A for 350 nuclides with half-lives ≥100 years; $1 \leq A \leq 250$. The dashed arrows indicate a splitting of a heavy nucleus ($A \sim 240$) into two nuclides of $A \sim 120$. The energy released would be $\sim (40) - 2(-90) \sim 220$ MeV

fact that there are no long-lived isotopes of elements between bismuth ($Z = 83$) and thorium ($Z = 90$). Isotopes with $A > 230$ could be said to have a "mass surplus." Consistent with the idea that *negative* Δ-values connote stability, all such *positive* Δ-valued nuclei eventually decay.

The two forgoing paragraphs actually muddle the concepts of mass defect and binding energy for sake of simplifying the description. Strictly, these are separate (but related) quantities. At a qualitative level, the details of the technical distinctions between them do not really add to the central concept that "lost mass" transforms to "binding energy." For sake of completeness, however, further details are discussed in Sect. 2.5, which can be considered optional.

Figure 2.8 can be used to estimate the energy released in hypothetical nuclear reactions. This is discussed in greater detail in Sect. 2.1.6, but the essence is straightforward: Add up the Δ-values of all of the input reactants (be careful with negative signs!), and then subtract from that result the sum of the Δ-values of the output products. The arrows in Fig. 2.8 show an example: a hypothetical splitting of a nucleus with $A \sim 240$ into two nuclei of $A \sim 120$. The input Δ-value is $\sim +40$ MeV. The sum of the output Δ-values is approximately $(-90$ MeV$) + (-90$ MeV$) \sim -180$ MeV. The difference between these is $\sim (+40) - (-180) \sim +220$ MeV, a large amount of energy even by nuclear standards. In late 1938 it was discovered that reactions like this are very real possibilities indeed.

There exist 266 apparently permanently stable, naturally occurring isotopes of the various elements, and about a hundred more "quasi-stable" ones with half-lives of a hundred years or greater. A compact way of representing all these nuclides is to plot each one as a point on a graph where the *x*-axis represents the number of neutrons, and the *y*-axis the number of protons. All isotopes of a given element will then lie on a horizontal line, since the number of protons in all nuclei of a given element is the same. This is shown in Fig. 2.9 for the 350 stable and quasi-stable nuclei of Fig. 2.8. Clearly, stable nuclei follow a very well-defined $Z(N)$ trend. Nature provides nuclei with neutrons to hold them together against the
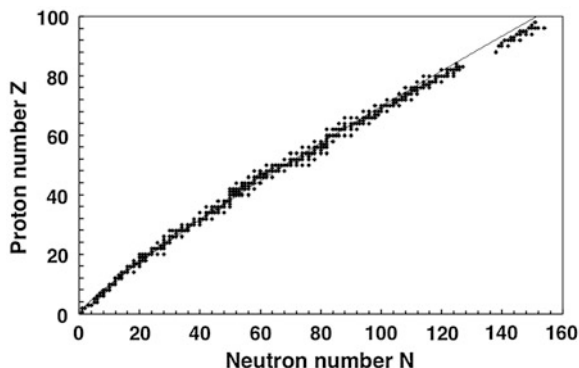
Fig. 2.9 Proton number Z versus neutron number N for 350 isotopes with half-lives $\geq$100 years, showing the narrow "band of stability" for nuclides. The trendline is described by the equation $Z \sim 1.264 \, N^{0.87}$. As discussed in Sect. 2.1.6, "neutron-rich" nuclei would lie below the band of points and decay to stability along leftward-diagonally-upwards trajectories by β^- decay. Conversely, "neutron-poor" nuclei would lie above the band of points and decay along rightward-diagonally-downwards trajectories by β^+ decay; see Fig. 2.12

mutual repulsion of their protons, but she is economical in doing so. Mass represents energy ($E = mc^2$), and Nature is evidently unwilling to invest more mass-energy to stabilize nuclei than is strictly necessary.

Note also that the points in the graph curve off to the right; this indicates that the vast majority of nuclei, except for a very few at the bottom-left of the graph, contain more neutrons than protons; this effect is known as the *neutron excess*. We will revisit such plots in Sect. 2.1.6.

We now return for one section to pick up the story of early nuclear physics, after which will follow another tutorial on nuclear reactions.

2.1.5 Alpha Particles and the Nuclear Atom

In the spring of 1907, Rutherford returned to England to take a position at Manchester University. When he arrived there, he made a list of promising research projects, one of which was to pin down the precise nature of alpha particles. Based on experiments where the number of alphas emitted by a sample of radium had been counted and the charge carried by each had been determined, he had begun to suspect that they were ionized helium nuclei. However, he needed to trap a sample of alphas for confirming spectroscopic analysis. Working with student Thomas Royds, Rutherford accomplished this with one of his typically elegant experiments.

In the Rutherford-Royds experiment, a sample of radon gas was trapped in a very thin-walled glass tube, which was itself surrounded by a thicker-walled tube.

The space between the two tubes was evacuated, and the radon was allowed to decay for a week. The energetic radon alphas could easily penetrate through the 1/100-mm thick wall of the inner tube. During their flights they would pick up electrons, become neutralized, and then become trapped in the space between the tubes. The neutralized alphas were then drawn off for analysis, and clearly showed a helium spectrum. Rutherford and Royds published their finding in 1909. In the notation described in the preceding section, alpha particles are identical to helium-4 nuclei: $_2^4$He.

The discovery for which Rutherford is most famous is that atoms have nuclei; this also had its beginnings in 1909. One of the projects on Rutherford's to-do list was to investigate how alpha particles "scattered" from atoms when they (the alphas) were directed through a thin metal foil. At the time, the prevailing notion of the structure of an atom was of a cloud of positive electrical material within which were embedded negatively-charged electrons. Thomson had determined that electrons weighed about 1/1,800 as much as a hydrogen atom; since hydrogen was the lightest element, it seemed logical to presume that electrons were small in comparison to their host atoms. This picture has been likened to a pudding, with electrons playing the role of raisins inside the body of the pudding. Another line of atomic structure evidence came from the chemistry community. From the bulk densities of elements and their atomic weights, it could be estimated that individual atoms behaved as if they were a few Ångstroms in diameter (1 Å $= 10^{-10}$ m; see Exercise 2.1). The few Ångstroms presumably represented the size of the overall cloud of positive material.

Rutherford had been experimenting with the passage of alpha-particles through metal foils since his earliest days of radioactivity research, and all of his experience indicated that the vast majority of alphas were deflected by only a very few degrees from straight-line paths as they barreled their way through a layer of foil. This observation was in line with theoretical expectations. Thomson had calculated that the combination of the size of a positively-charged atomic sphere and the kinetic energy of an incoming alpha (itself also presumably a few Ångstroms in size) would be such that the alpha would typically suffer only a small deflection from its initial trajectory. Deflections of a few degrees would be rare, and a deflection of 90° was expected to be so improbable as to never have any reasonable chance of being observed. In the Thomson atomic model, a collision between an alpha and an atom should not be imagined as like that between two billiard balls, but rather more like two diffuse clouds of positive electricity passing through each other. The alphas would presumably strike a number of electrons during the collision, but the effect of the electrons' attractive force on the alphas would be negligible due to the vast difference in their masses, a factor of nearly 8,000. Electrons played no part in Rutherford's work.

Rutherford was working with Hans Geiger (Fig. 2.10) of Geiger counter fame, who was looking for a project to occupy an undergraduate student, Ernest Marsden, another New Zealand native. Rutherford suggested that Geiger and Marsden check to see if they could observe any large-angle deflections of alphas

when they passed through a thin gold foil, fully expecting a negative result. Gold was used because it could be pressed into a thin foil only about a thousand atoms thick. To Geiger and Marsden's surprise, a few alphas, about one in every 8,000, were *bounced backward* toward the direction from which they came. The number of such reflections was small, but was orders of magnitude more than what was expected on the basis of Thomson's model. Rutherford was later quoted as saying that the result was "almost as incredible as if you had fired a 15-in. shell at a piece of tissue paper and it came back and hit you." Geiger and Marsden published their anomalous result in July, 1909. The work of detecting the scattered alpha-particles was excruciating. A Geiger counter could have been used to *detect* the alphas, but they had to be *seen* to get detailed information on their direction of travel. This was done by having the scattered alphas strike a phosphorescent screen; a small flash of light (a "scintillation") would be emitted, and could be counted by an observer working in a darkened room. Geiger and Marsden counted thousands of such scintillations.

So unexpected was Geiger and Marsden's result that it took Rutherford the better part of 18 months to infer what it meant. The conclusion he came to was that the positive electrical material within atoms must be confined to much smaller volumes than had been thought to be the case. The alpha-particles (themselves also nuclei) had to be similarly minute; only in this way could the electrical force experienced by an incoming alpha be intense enough to achieve the necessary repulsion to turn it back if it should by chance strike a target nucleus head-on; the vast majority of alpha nuclei sailed through the foil, missing gold nuclei by wide margins. The compaction of the positive charge required to explain the scattering experiments was stunning: down to a size of about 1/100,000 of an Ångstrom. But, atoms as a whole still behaved *in bulk* as if they were a few Ångstroms in diameter. Both numbers were experimentally secure and had to be accommodated.

This, then, was the origin of our picture of atoms as miniature solar systems: very small, positively-charged "nuclei" surrounded by orbiting electrons at distances out to a few Ångstroms. This configuration is now known as the "Rutherford atom."
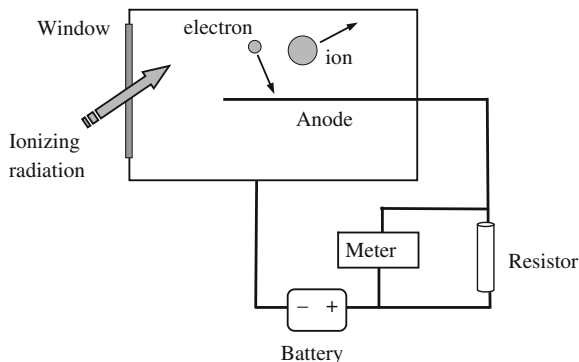
A sense of the scale of Rutherford's atom can be had by thinking of the lone proton that forms the nucleus of an ordinary hydrogen atom as scaled up to being two millimeters in diameter, about the size of an uncooked grain of rice. If this enlarged proton is placed at the center of a football field, the diameter of the lowest-energy electron orbit (that which comes closest to the nucleus) would reach to about the goal lines. In giving us nuclei and being credited with the discovery of the positively-charged protons that they contain, Rutherford bequeathed us atoms that are largely empty space.

The first public announcement of this new model of atomic structure seems to have been made on March 7, 1911, when Rutherford addressed the Manchester Literary and Philosophical Society; this date is often cited as the birthdate of the nuclear atom. The formal scientific publication came in July, and directly influenced Niels Bohr's famous atomic model which was published two years later. Rutherford's nucleus paper is a masterpiece of fusion of experimental evidence and theoretical reasoning. After showing that the Thomson model could not possibly generate the observed angular distribution of alpha scatterings, he demonstrated that the nuclear "point-mass" model gave predictions in accord with the data. Rutherford did not use the term "nucleus" in his paper; that nomenclature seems to have been introduced by Cambridge astronomer John Nicholson in a paper published in November, 1911. The term "proton" was not introduced until June, 1920, but was coined by Rutherford himself.

With the understanding that scattering events were the results of such nuclear collisions, Rutherford's analysis could be applied to other elements in the sense of using an observed scattering distribution to infer how many fundamental "protonic" charges the element possessed; this helped to place elements in their proper locations in the periodic table. Elements had theretofore been defined by their atomic weights (A), but it was the work of researchers such as Rutherford, Soddy, Geiger, and Marsden which showed that it is an element's atomic number (Z) that dictates its chemical identity.

The atomic weights of elements were still important, however, and very much the seat of a mystery. Together, chemical and scattering evidence indicated that the atomic weights of atoms seemed to be proportional to their number of protonic charges. Specifically, atoms of all elements weighed about twice as much or more as could be accounted for on the basis of their numbers of protons. For some time, this extra mass was thought to be due to additional protons in the nucleus which for some reason contained electrons within themselves, an electrically neutral combination. This would give net-charge nuclei consistent with the scattering experiments, while explaining measured atomic weights. By the mid-1920s, however, this proposal was becoming untenable: the Uncertainty Principle of quantum mechanics ruled against the possibility of containing electrons within so small a volume as a single proton, or even a whole nucleus. For many years before

Fig. 2.11 Schematic drawing of a Geiger-Müller Counter

its discovery, Rutherford speculated that there existed a third fundamental constituent of atoms, the neutron. As described in Sect. 2.2, he would live to see his suspicion proven by one of his own students. That atoms are built of electrons orbiting nuclei comprised of protons and neutrons is due very much to Rutherford and his collaborators and students.

Having mentioned Hans Geiger, it is worthwhile to describe briefly the operation of his eponymous counter, as its use will turn up in other contexts in this book. The original version of the Geiger counter was invented by Geiger and Rutherford in 1908. In 1928, Geiger and a student, Walther Müller, made some improvements on the design, and these devices are now properly known as Geiger-Müller counters.

Geiger-Müller counters operate by detecting ionizing radiation, that is, particles that ionize atoms in a sample of gas through which they pass. Fundamentally, the counter consists of a metal case which is closed at one end and which has a thin plastic "window" at the other end (Fig. 2.11). Inside the tube is an inert gas, usually helium. The case and a metallic central anode are connected to a battery which makes (in the Figure) the tube negative and the anode positive. Energetic particles such as alpha or beta rays penetrate through the window, and ionize atoms of the inert gas. The liberated electrons will be attracted to the anode, while the ionized atoms are attracted to the case. The net effect is to create an electrical current, which is shunted by a resistor to pass through a meter. By incorporating a speaker in the electronics, the current can be converted to the familiar "clicks" one hears on news reports and in movies.

2.1.6 Reaction Notations, Q-Values, Alpha and Beta Decay, and Decay Chains

In this and the following section we again break with chronological order to fill in some background on notations for nuclear reactions, and the details of alpha and

beta-decays and naturally-occurring decay schemes. We will pick up with more of Rutherford's work in Sect. 2.1.7. The present section is somewhat technical, but involves no computations; the concept of a Q-value will prove very important later on.

The notation for writing a nuclear reaction is very similar to used that for describing a chemical reaction. *Reactants* or *input nuclides* are written on the left side of a rightward-pointing arrow, and *products* or *output nuclides* are placed on the other side, like this:
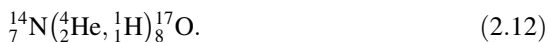
$$reactants \rightarrow products. \tag{2.10}$$

Decades of experimental evidence indicate that there are two rules that are *always* obeyed in nuclear reactions:

(i) The total number of input nucleons must equal the total number of output nucleons. *The numbers of protons and neutrons may (and usually do) change, but their sum must be conserved.*

(ii) Total electric charge must be conserved. Protons count as one unit of positive charge. Beta decays involve nuclei which create within themselves and then eject either an electron or a positively-charged particle with the same mass as an electron, a so-called *positron*. The charges of these ejected particles must be taken into account in ensuring charge conservation (negative or positive one unit), but they are not considered to be nucleons and so are not counted when applying rule (i). Positrons are also known as beta-positive (β^+) particles, while ordinary electrons are also known as beta-negative particles (β^-).

As an example of a typical reaction, here is one that will be discussed in more detail in Sect. 2.1.7: alpha-bombardment of nitrogen to produce hydrogen and oxygen:

$$_2^4\text{He} + _7^{14}\text{N} \rightarrow _1^1\text{H} + _8^{17}\text{O}. \tag{2.11}$$

Verification that both rules are followed can be seen in that (i) $4 + 14 = 1 + 17$, and (ii) $2 + 7 = 1 + 8$. In this type of bombardment reaction, the notational convention is to write the lighter incoming reactant first on the left side, followed by the target nucleus. Note that a hydrogen nucleus, $_1^1\text{H}$, is simply a proton. A proton is sometimes written as just "p", but this book will always employ the more explicit $_1^1\text{H}$ notation. On the output side, the lighter product is usually written first. In a "four-body" reaction such as this, a more compact shorthand notation that puts the target nucleus first is sometimes employed:

$$_7^{14}\text{N}\left(_2^4\text{He}, _1^1\text{H}\right)_8^{17}\text{O}. \tag{2.12}$$

In this format, the convention can be summarized as:

$$\text{target (projectile, light product) heavy product} \tag{2.13}$$

In any reaction where the input and input reactants are different, experiments show that *mass is not conserved*. That is, the sum of the input masses will be different from the sum of the output masses. Mass can either be created or lost; what happens depends on the nuclides involved. The physical interpretation of this relates to Einstein's famous $E = mc^2$ equation. If mass is lost (sum of output masses < sum of input masses), the lost mass will appear as kinetic energy of the output products. If mass is gained (sum of output masses > sum of input masses), energy must be drawn from somewhere to create the mass gained, and the only source available is the kinetic energy of the "bombarding" input reactant. Nuclear physicists always express the mass gain or loss in units of energy equivalent, almost always in MeV. Such energy gains or losses are termed *Q-values*. If $Q > 0$, kinetic energy is created by consuming input-reactant mass, whereas if $Q < 0$, input-particle kinetic energy has been consumed to create additional output mass. The technical definition of Q is

$$Q = (\text{sum of input masses}) - (\text{sum of product masses}), \qquad (2.14)$$

quoted in units of equivalent energy (one atomic mass unit = 931.4 MeV). When applied to computing the energy consumed or liberated in a reaction, this definition gives the same result as the graphical method illustrated in Fig. 2.8.

The alpha-nitrogen reaction above has $Q = -1.19$ MeV. To approach the nucleus and initiate the reaction, the incoming alpha-particle must possess at least this much kinetic energy. In fact, the alpha must possess considerably more than 1.19 MeV of kinetic energy due to an effect that is not accounted for in computing the Q-value alone, the so-called *Coulomb barrier*; this is discussed in Sect. 2.1.8.
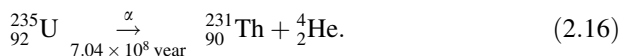
Returning for a moment to mass spectroscopy, the development of means to measure precise masses for isotopes was a crucial step forward in the progress of nuclear physics. With precise masses and knowledge of Einstein's $E = mc^2$ equivalence, the energy liberated or consumed in reactions could be predicted. Measurements of the kinetic energies of reaction products would then serve as checks on the mass values. Conversely, for a reaction where the mass or identity of some of the particles involved was not clear, measurements of the kinetic energies could be used to infer what was happening. On reflecting on these connections, you might wonder how Rutherford measured such kinetic energies; after all, tracking a nucleus is obviously not the same as using a radar gun to measure the speed of a car or a baseball. Experimenters had to rely on proxy measurements such as how far a particle traveled through air or a stack of thin metal foils before being brought to a stop. If precise mass defects are known from mass spectroscopy, the energy liberated (or consumed) in a reaction can be computed, and the numbers can be used to calibrate a range-versus-energy relationship. This combination of theory, experimental technique, and instrumental development is an excellent example of scientific cross-fertilization.

2.1.6.1 Alpha Decay

Ernest Rutherford decoded alpha-decay as a nucleus spontaneously transmuting itself to a more stable mass-energy configuration by ejecting a helium nucleus. In doing so, the original nucleus loses two protons and two neutrons, which means that it ends up two places down in atomic number on the periodic table and has four fewer nucleons in total. Alpha-emission is a common decay mechanism in heavy elements, and can be written in the arrowed notation as
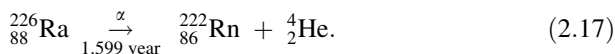
$$\ _{Z}^{A}X \ \xrightarrow{\alpha} \ _{Z-2}^{A-4}Y \ + \ _{2}^{4}\text{He}. \tag{2.15}$$

Here, X designates the element corresponding to the original nucleus, and Y that of the "daughter product" nucleus. Sometimes the half-life is written below the arrow; for example, the alpha-decay of uranium-235 can be written as

$$\ _{92}^{235}\text{U} \ \xrightarrow[7.04 \times 10^{8} \text{ year}]{\alpha} \ _{90}^{231}\text{Th} + \ _{2}^{4}\text{He}. \tag{2.16}$$

As always, electrical charge and nucleon number are conserved. In such decays, the total mass of the output products is always less than that of the input particles: Nature spontaneously seeks a lower mass-energy configuration ($Q > 0$). The energy release in alpha-decays is typically $Q \sim 5 - 10$ MeV, the majority of which appears as kinetic energy of the alpha-particle itself. Helium nuclei tend to be ubiquitous in nuclear reactions as they have a large mass defect and so are very stable.

As a tool to induce nuclear reactions, the Curies and Rutherford often utilized alpha particles emitted in radium decay:

$$\ _{88}^{226}\text{Ra} \ \xrightarrow[1,599 \text{ year}]{\alpha} \ _{86}^{222}\text{Rn} \ + \ _{2}^{4}\text{He}. \tag{2.17}$$

This decay has a Q-value of +4.87 MeV, which explains how the $_{7}^{14}\text{N}\left(_{2}^{4}\text{He}, _{1}^{1}\text{H}\right)_{8}^{17}\text{O}$ reaction described above can be made to happen.

2.1.6.2 Beta Decay

Two types of beta decay occur naturally. Look back at Fig. 2.9, which illustrates the narrow "band of stability" of long-lived isotopes. If an isotope should find itself with too many neutrons for the number of protons that it possesses (or, equivalently, too few protons for its number of neutrons), it will lie to the right of the band of points. Conversely, should it have too few neutrons for the number of protons that it possesses (or too many protons for its number of neutrons), it will lie to the left of the band of points.

Suppose that a nucleus is too neutron-rich for its number of protons. Purely empirically, it has been found that Nature deals with this by having a neutron spontaneously decay into a proton. But this, by itself, would represent a net creation of electric charge, and hence a violation of charge conservation. So, a

negative electron is created in the bargain to render no net charge created. Nucleon number is conserved; remember that electrons do not count as nucleons. The electron is also known as a β^- particle, and the reaction can be symbolized as $n \to p + e^-$, or, equivalently, $n \to p + \beta^-$. The number of neutrons drops by one while the number of protons grows by one, so the number of nucleons is unchanged. The overall effect is

$$ {}^A_Z X \xrightarrow{\beta^-} {}^A_{Z+1} Y + {}^0_{-1} e^-. \tag{2.18} $$

Note that a "nucleon-like" notation has been appended to the electron to help keep track of the charge and nucleon numbers. *The result of β^- decay is to move a nucleus up one place in the periodic table*. It was Henri Becquerel who showed, in 1900, that the negatively-charged beta-rays being observed in such decays were identical in their properties to Thomson's electrons.
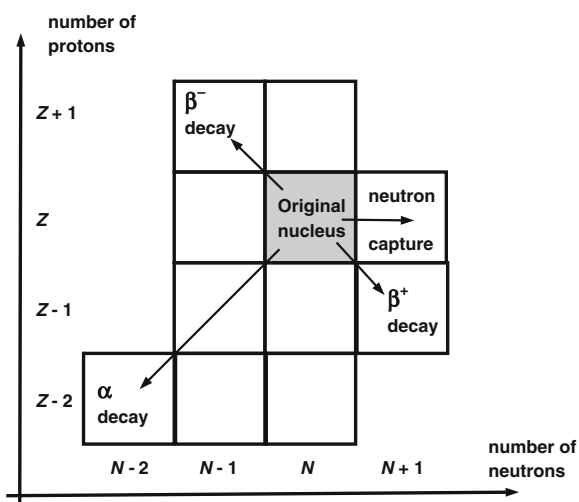
If a nucleus is neutron-poor, a proton will spontaneously decay into a neutron. But this would represent a loss of one unit of charge, so Nature creates a *positron*—an anti-electron—to maintain the charge balance: $p \to n + e^+$, or $p \to n + \beta^+$. Here the overall effect is

$$ {}^A_Z X \xrightarrow{\beta^+} {}^A_{Z-1} Y + {}^0_1 e^+. \tag{2.19} $$

The result of β^+ decay is to move a nucleus down one place in the periodic table.

As shown in Fig. 2.12, decay mechanisms can be represented graphically in the (Z, N) grid format of Fig. 2.9. Also shown in Fig. 2.12 is the effect of *neutron capture*. In this process, a nucleus absorbs an incoming neutron to become a heavier isotope of itself; this will be important in later discussions of fission and the synthesis of plutonium.
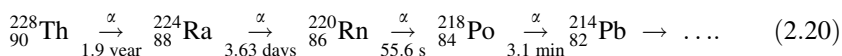

Fig. 2.12 Decay and neutron-capture transmutation trajectories of an original nucleus of Z protons and N neutrons. This (Z, N) arrangement is as that in Fig. 2.9

2.1.6.3 Natural Decay Schemes

The work of the Curies and Rutherford and their various collaborators culminated in the understanding that three lengthy decay sequences occur spontaneously in Nature. All three begin with an isotope of thorium or uranium, and terminate with three different isotopes of lead. These are illustrated in Fig. 2.13, which is of the same form as Fig. 2.12.

As an example of the use of decay-chain notation, consider again Rutherford's one-minute thorium "emanation" of Sect. 2.1.2. The observed one-minute half-life probably represented the decay of radon-220:

$$\underset{90}{228}\text{Th} \underset{1.9\ \text{year}}{\overset{\alpha}{\rightarrow}} \underset{88}{224}\text{Ra} \underset{3.63\ \text{days}}{\overset{\alpha}{\rightarrow}} \underset{86}{220}\text{Rn} \underset{55.6\ \text{s}}{\overset{\alpha}{\rightarrow}} \underset{84}{218}\text{Po} \underset{3.1\ \text{min}}{\overset{\alpha}{\rightarrow}} \underset{82}{214}\text{Pb} \rightarrow \ldots \qquad (2.20)$$

The first step in this chain, thorium-228, is itself a decay product of uranium. The identification of Rn-220 as Rutherford's thorium emanation is strong, but not absolutely secure: many heavy-element half-lives are on the order of a minute; as described in a McGill University website, it is not clear exactly what isotopes Rutherford detected, as the concept of isotopy had not yet been established in 1900.

The work accomplished by the world's first generation of radiochemists was staggering. Uranium and thorium ores contain constantly varying amounts of a number of isotopes which are created by decays of heavier parent isotopes, and which themselves decay to lighter daughter products until they arrive at stable
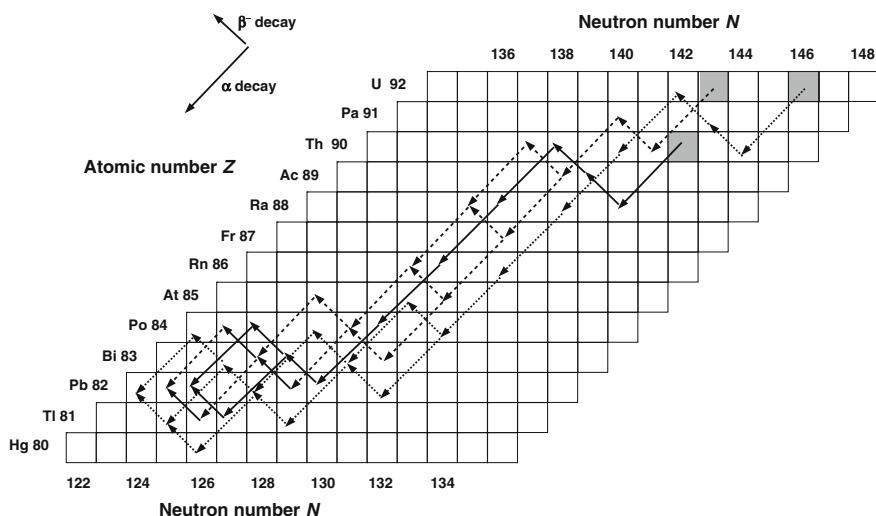

Fig. 2.13 Natural radioactive decay sequences. Nuclei of uranium-238, uranium-235, and thorium-232 all decay to isotopes of lead (Pb-206, 207, and 208, respectively) via sequences of alpha and beta decays. These sequences are respectively indicted by the chains of *dotted*, *dashed*, and *solid lines*

neutron/proton configurations. Only by isolating samples of individual elements and subjecting them to mass-spectroscopic analyses could individual isotopes be characterized.
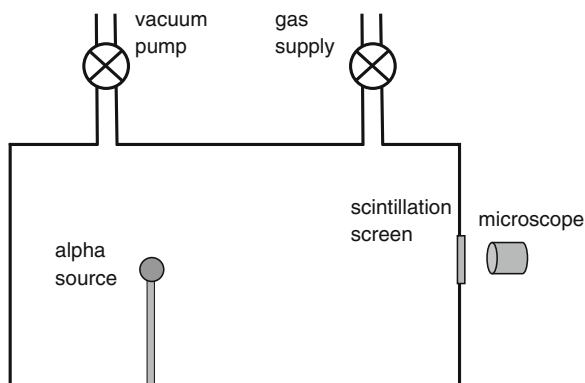
2.1.7 Artificial Transmutation

Rutherford's last great discovery came in 1919. This was his realization that it was possible to set up experimental situations wherein atoms of a given element could be *transmuted* into those of another, when bombarded by nuclei of yet a third. The idea of elemental transmutation was not new; after all, this is precisely what happens in natural alpha and beta-decays. What was new was the realization that transmutations could be induced by human intervention.

   The work that led to this discovery began around 1915, and was carried out by Ernest Marsden. As part of an experimental program involving measurements of reaction energies, Marsden bombarded hydrogen atoms with alpha-particles produced by the decay of samples of radon gas contained in a small glass vials. A hydrogen nucleus would receive a significant kick from a collision with an alpha-particle and be set into motion at high speed. These experiments were done by sealing the alpha source and hydrogen gas inside a small chamber, as sketched in Fig. 2.14. At one end of the chamber was a small scintillation screen which could be viewed through a microscope, as had been done in the alpha-scattering experiments. By placing thin metal foils just behind the screen, Marsden could determine the ranges, and hence the energies, of the struck protons. So far, there is nothing unusual here; these experiments were routine work that involved the use of known laws of conservation of energy and momentum to cross-check and interpret measurements.

   Breakthroughs favor an attentive and experienced mind, and Marsden's was ready. His crucial observation was to notice that when the experimental chamber was evacuated, the radon source itself seemed to give rise to scintillations like

Fig. 2.14 Sketch of Rutherford's apparatus for the discovery of artificial transmutation

those from hydrogen, *even though there was no hydrogen in the chamber*. The implication seemed to be that hydrogen was arising in radioactive decay, an occurrence that had never before been observed. Marsden returned to New Zealand in 1915, and Rutherford, heavily occupied with research for the British Admiralty, could manage only occasional experiments until World War I came to an end in late 1918. In 1919, he turned to investigating Marsden's unexpected radium/hydrogen observation, and was rewarded with yet another pivotal discovery.
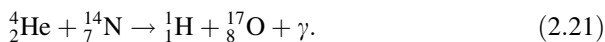
Look again at Fig. 2.14. Rutherford placed a source of alpha particles within a small brass chamber which could be evacuated and then filled with a gas with which he wished to experiment. As Rutherford reported in his June, 1919, discovery paper, he set out to investigate the phenomenon that "a metal source, coated with a deposit of radium-C [bismuth-214], always gives rise to a number of scintillations on a zinc sulphide screen far beyond the range of the α particles. The swift atoms causing these scintillations carry a positive charge and are deflected by a magnetic field, and have about the same range and energy as the swift H atoms produced by the passage of α particles through hydrogen. These 'natural' scintillations are believed to be due mainly to swift H atoms from the radioactive source, but it is difficult to decide whether they are expelled from the radioactive source itself or are due to the action of α particles on occluded hydrogen."

Rutherford proceeded by investigating various possibilities as to the origin of the hydrogen scintillations. No vacuum pump is ever perfect; some residual air would always remain in the chamber no matter how thoroughly it had been pumped down. While hydrogen is normally a very minute component of air (about half a part per million), more could be present if the air contained water vapor. Suspecting that the alpha particles might be striking residual hydrogen-bearing water molecules, Rutherford began by introducing dried oxygen and carbon dioxide into the chamber, observing, as he expected, that the number of scintillations decreased. Surprisingly, however, when he admitted *dry air* into the chamber, the number of hydrogen-like scintillations *increased*. This suggested that hydrogen was arising not from the radium-C itself, but from some interaction of the alpha particles with air.

The major constituents of air are nitrogen and oxygen; having eliminated oxygen, Rutherford inferred that nitrogen might be involved. On admitting pure nitrogen into the chamber, the number of scintillations increased yet again. As a final test that hydrogen was not somehow arising from the radioactive source itself, he found that on placing thin metal foils close to the radioactive source, the scintillations persisted, but their range was reduced in accordance with what would be expected if the alpha particles were traveling through the foils before striking nitrogen atoms; the scintillations were evidently arising from within the volume of the chamber. As Rutherford wrote, "it is difficult to avoid the conclusion that the long-range atoms arising from collision of α particles with nitrogen are … probably atoms of hydrogen …. If this be the case, we must conclude that the nitrogen atom is disintegrated under the intense forces developed in a close collision with a

swift α particle, and that the hydrogen atom which is liberated formed a constituent part of the nitrogen nucleus".
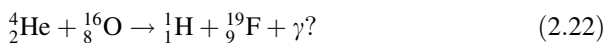
In modern notation, the reaction is written as

$$_2^4\text{He} + _7^{14}\text{N} \rightarrow _1^1\text{H} + _8^{17}\text{O} + \gamma. \tag{2.21}$$

The "γ" here indicates that this reaction also releases a gamma-ray. The gamma-ray plays no role in the interpretation of Rutherford's experiment through conservation of charge and mass numbers, but is included here for sake of completeness; it will play a role in the discussion of the discovery of the neutron in Sect. 2.2.

Because atoms are mostly empty space, only about one alpha per hundred thousand induces such a reaction. Nuclear physicists speak of the *yield* of a reaction, which is the fraction of incident particles that cause a reaction. Yield values on the order of 10^{-5} in alpha-induced reactions are not uncommon.

Why does such a reaction not take place with oxygen, say,

$$_2^4\text{He} + _8^{16}\text{O} \rightarrow _1^1\text{H} + _9^{19}\text{F} + \gamma? \tag{2.22}$$

The Q-value of this reaction is -8.1 MeV. Radium-C alphas have energies of about 5.5 MeV, which is not enough to make the reaction happen. In the case of nitrogen, the Q-value is about -1.2 MeV, so the alphas are sufficiently energetic to make that reaction occur.

Rutherford and Marsden's discovery opened yet another experimental venue: Could alpha-particles induce transmutations in any other elements? What products could be created? What yields were involved? As discussed in the next section, however, there was a serious natural limitation to further experiments.

Later in 1919, Rutherford moved from Manchester University to Cambridge University to fill the position of Director of the Cavendish Laboratory, which had become vacant upon the retirement of J. J. Thomson. Rutherford would remain at Cambridge until his death in October, 1937, nurturing another generation of nuclear experimentalists. He died just 2 years before the discovery of nuclear fission, which would lead, in a few more years, to the development and use of nuclear weapons.

2.1.8 The Coulomb Barrier and Particle Accelerators

Consider again Rutherford's alpha-bombardment of nitrogen, the first artificial transmutation of an element (neglecting the gamma-ray):

$$_2^4\text{He} + _7^{14}\text{N} \rightarrow _1^1\text{H} + _8^{17}\text{O}. \tag{2.23}$$

Neglecting the fact that this reaction has a negative Q-value, a simple interpretation of this equation is that if you were to mix helium and nitrogen, say at

room-temperature conditions, hydrogen and oxygen would result spontaneously. But even if Q were positive, this would not happen because of an effect that is not accounted for in merely writing down the reaction or in computing the Q-value: the so-called "Coulomb barrier" problem.

Electrical charges of the same sign repel each other. This effect is known as "the Coulomb force" after French physicist Charles-Augustin de Coulomb, who performed some of the first quantitative experiments with electrical forces in the late 1700s. Because of the Coulomb force, nitrogen nuclei will repel incoming alpha-particles; only if an alpha has sufficiently great kinetic energy will it be able to closely approach a nitrogen nucleus. Essentially, the two have to collide before stronger but shorter-range "nuclear forces" between nucleons that effect transmutations can come into play. The requisite amount of kinetic energy that the incoming nucleus must possess to achieve a collision is called the "Coulomb barrier."

For an alpha-particle striking a nitrogen nucleus, the barrier amounts to about 4.2 MeV, a fairly substantial amount of energy. An atom or molecule at room temperature will typically possess only a fraction of an eV of kinetic energy (about 0.025 eV on average), not nearly enough to initiate the reaction. Rutherford was able to induce the nitrogen transmutation because his radium-C alphas possessed over 5 MeV of kinetic energy.

The following material examines the physics of the Coulomb barrier. Readers who wish to skip this material should proceed to the paragraph following Eq. (2.31).

Figure 2.15 sketches two nuclei that are undergoing a collision. One of them, of atomic number Z_1, is presumed to be bombarding a fixed target nucleus of atomic number Z_2. Since the charge on each proton is the same magnitude as the electron charge "e," the charges within the nuclei will be $+eZ_1$ and $+eZ_2$.
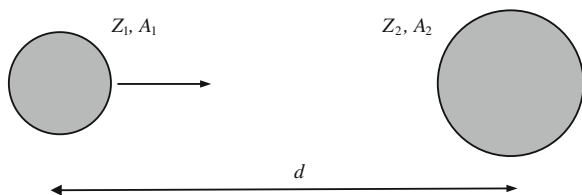
According to Coulomb's law, if the centers of the two nuclei are distance d apart, the system will possess a potential energy PE given by

$$PE = \frac{(eZ_1)(eZ_2)}{4\pi\varepsilon_0 d} = \frac{e^2 Z_1 Z_2}{4\pi\varepsilon_0 d}, \tag{2.24}$$

where ε_O is a physical constant, 8.8544×10^{-12} C^2/(J-m).

To effect a collision, the incoming nucleus needs to approach to a distance d which is equal to the sum of the radii of the two nuclei. To achieve such an approach, the incoming nucleus must start with an amount of kinetic energy which

Fig. 2.15 Colliding nuclei. The nucleus on the right is presumed to be fixed while the one on the left approaches

Z_1, A_1 Z_2, A_2

d

is at least equal to the potential energy of the system at the moment of contact in order that it will not have been brought to a halt beforehand by the Coulomb repulsion. Equation (2.24), when evaluated for the value of d corresponding to the two nuclei just touching, gives the Coulomb barrier. We thus need a general way to evaluate (2.24) for when two nuclei are just touching.

Empirically, scattering experiments show that the radii of nuclei can be expressed in terms of their mass numbers according as the expression

$$radius \sim a_0 A^{1/3}, \tag{2.25}$$

where $a_0 \sim 1.2 \times 10^{-15}$ m. Designating the mass numbers as A_1 and A_2, we have

$$Coulomb\ barrier \sim \left(\frac{e^2}{4\,\pi\,\varepsilon_0 a_0} \right) \frac{Z_1 Z_2}{\left(A_1^{1/3} + A_2^{1/3} \right)}. \tag{2.26}$$

The value of e is 1.6022×10^{-19} C; substituting this, ε_0, and $a_0 \sim 1.2 \times 10^{-15}$ m into (2.26) gives the bracketed factor as

$$\frac{e^2}{4\,\pi\,\varepsilon_0 a_0} \ = \ 1.9226 \ \text{x} \ 10^{-13}\ \text{J}. \tag{2.27}$$

This can be expressed more conveniently in terms of MeV; $1\ \text{MeV} = 1.6022 \times 10^{-13}$ J:

$$\frac{e^2}{4\,\pi\,\varepsilon_0 a_0} = 1.2\ \text{MeV}. \tag{2.28}$$

We can then write (2.26) as

$$Coulomb\ barrier \sim \frac{1.2(Z_1 Z_2)}{\left(A_1^{1/3} + A_2^{1/3} \right)}\ \text{MeV}. \tag{2.29}$$

For an alpha-particle striking a nitrogen nucleus, this gives, as claimed above,

$$Coulomb\ barrier \sim \frac{1.2(2 \times 7)}{(4^{1/3} + 14^{1/3})} \sim \frac{16.8}{(1.587 + 2.410)} \sim 4.2\ \text{MeV}. \tag{2.30}$$

Now imagine trying to induce a reaction by having alpha-particles strike nuclei of uranium-235. The experiment would be hopeless if you are using an alpha whose kinetic energy is of the typical 5–10 MeV decay energy:

$$Coulomb\ barrier \sim \frac{1.2(2 \times 92)}{(4^{1/3} + 235^{1/3})} \sim \frac{220.8}{(1.587 + 6.171)} \sim 28.5\ \text{MeV}. \tag{2.31}$$

If one is using alphas created in natural decays, it is practical to carry out bombardment experiments with target elements only up to $Z \sim 20$. By the mid-1920s this was becoming a serious problem: researchers were literally running out of elements to experiment with. The curiosity-driven desire to bombard heavier
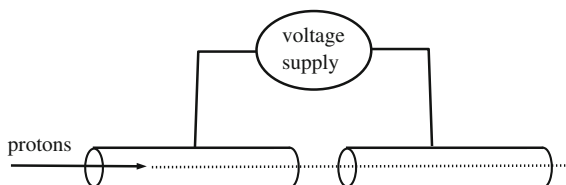
Fig. 2.16 Left Rolf Wideröe (1902–1996). Right Ernest Lawrence (1901–1958). Sources AIP Emilio Segre Visual Archives; http://commons.wikimedia.org/wiki/File:Ernest_Orlando_Lawrence.jpg

elements thus generated a technological challenge: Was there any way that the alpha (or other) particles could be accelerated once they had been emitted by their parent nuclei? It was this challenge that gave birth to the first generation of particle accelerators.

The first practical particle acceleration scheme was published by Norwegian native Rolf Wideröe (Fig. 2.16) in a German electrical engineering journal in 1928. The essence of Wideröe's proposal is pictured in Fig. 2.17. Two hollow metal cylinders are placed end-to-end and connected to a source of variable-polarity voltage. This means that the cylinders can be made positively or negatively charged, and the charges can be switched as desired. A stream of protons (say) is directed into the leftmost cylinder, which is initially negatively charged. This will attract the protons, which will speed up as they pass through the cylinder.

Just as the bunch of protons emerges from the first cylinder, the voltage polarity is switched, making the left cylinder positive and the right one negative. The protons then get a push from the first cylinder while being pulled into the second one, which further accelerates them. By placing a number of such units back-to-back, substantial accelerations can be achieved; this is the principle of a *linear accelerator*. Obviously, many of the incoming particles will be lost by crashing into the side of a cylinder or because their speed does not match the frequency of the polarity shifts of the voltage supplies; only a small number will emerge from the last cylinder. But the point here is not necessarily efficiency; it is to generate *some* high-speed particles which could surmount the Coulomb barriers of heavy target nuclei. The longest linear accelerator in the world is now the Stanford Linear

Fig. 2.17 Wideröe's linear
acceleration scheme

Accelerator in California, which can accelerate electrons to 50 *billion* electron-
volts of kinetic energy over a distance of 3.2 km (2 miles).

Wideröe's work came to the attention of Ernest Orlando Lawrence (Fig. 2.16),
an experimental physicist at the University of California at Berkeley. Lawrence
and collaborator David Sloan built a Wideröe device, which by late 1930 they had
used to accelerate mercury ions to kinetic energies of 90,000 eV. While experi-
menting with the Wideröe design, however, Lawrence had an inspiration that was
to have profound consequences. He desired to achieve higher energies, but was
daunted by the idea of building an accelerator that would be meters in length. How
could the device be made more compact?

In Sect. 2.1.4 a description was given of how Francis Aston utilized the Lorentz
force caused by a magnetic field to separate ions of different masses in his mass
spectrometer. Lawrence's new device, which he called a *cyclotron*, also made use of
this force law, but in a way that simultaneously incorporated Wideröe's alternating-
voltage acceleration scheme.

Lawrence's cyclotron is sketched in Fig. 2.18, which is taken from his appli-
cation for a patent on the device. Here the voltage supply is connected to two
D-shaped metal tanks placed back-to back; they are known to cyclotron engineers
as "Dees." The entire assembly must be placed within a surrounding vacuum tank
to avoid deflective effects of collisions of the accelerated particles with air
molecules.

The source of the ions (usually positive) is placed between the Dees. In the
diagram, the ions are initially directed toward the upper Dee, which is set to carry a
negative charge to attract them. If the voltage polarity is not changed and there is
nothing to otherwise deflect the ions, they would crash into the edge of the Dee.
But Lawrence knew from Aston's work that if the assembly were placed between
the poles of a magnet (with the magnetic field again emerging from the page), the
Lorentz force would try to make the ions move in circular paths. The net result of
the combination of the ions' acceleration toward the charged Dee and the Lorentz
force is that the ions move in outward-spiraling trajectories. If the magnetic field is
strong, the spiral pattern will be "tight", and the ions will get nowhere near the
edge of the Dee in their first orbit. As ions leave the upper Dee, the polarity is
switched in order to attract them to the lower Dee. Switching and acceleration
continues (for microseconds only) until the ions strike a target at the periphery of
one of the Dees.

Lawrence and graduate student Nils Edlefsen first reported on the cyclotron
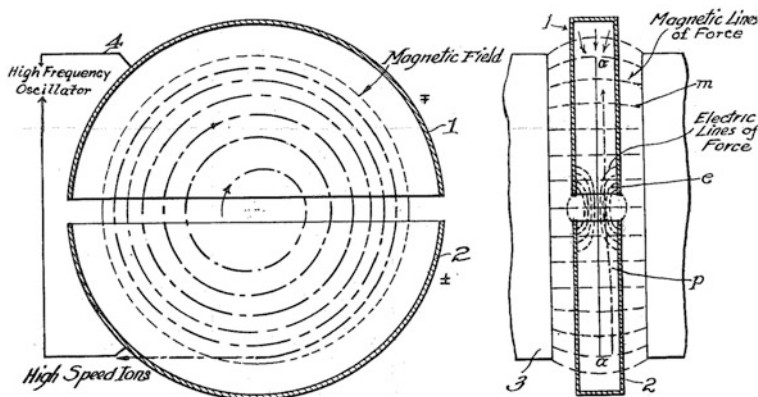concept at a meeting of the American Association for the Advancement for

Fig. 2.18 Schematic illustration of Lawrence's cyclotron concept in top and side view, from his patent application. *Source* http://commons.wikimedia.org/wiki/File:Cyclotron_patent.png

Science held in September, 1930, but they had no results available at that time. By May, 1931, Lawrence had a 4.5-in. diameter device in operation (Fig. 2.19); he and student M. Stanley Livingston reported at an American Physical Society meeting that they were able to accelerate hydrogen molecule-ions (H_2^+) to energies of 80,000 eV using only a 2,000 V power supply. Later the same year, Lawrence achieved MeV energies with an eleven-inch cyclotron. By 1932 he had constructed a 27-in. device which achieved an energy of 3.6 MeV (Fig. 2.20), but had bigger plans yet. Lawrence was as adept at fundraising as he was at electrical engineering, and by 1937 had constructed a 37-in. model capable of accelerating deuterons (nuclei of "heavy hydrogen," $_1^2H$) to energies of 8 MeV. By 1939 he had brought into operation a 60-in. model that required a 220-t magnet, and which could accelerate deuterons to 16 MeV. In 1942 he brought online his 184-in. diameter cyclotron, which is still operating and can accelerate various types of particles to energies exceeding 100 MeV. Along the way, Lawrence established the University of California Radiation Laboratory ("Rad Lab"), which is now the Lawrence Berkeley National Laboratory (LBNL).
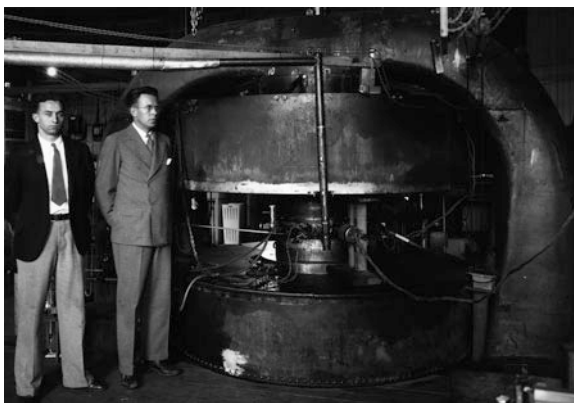
Particle accelerators allowed experimenters to surmount the Coulomb barrier and so open up a broad range of energies and targets to experimentation. Lawrence's ingenuity earned him the 1939 Nobel Prize for Physics, and a variant of his cyclotron concept would play a significant role in the Manhattan Project. Today's giant accelerators at the Fermi National Accelerator Laboratory (Fermilab) and the European Organization for Nuclear Research (CERN) are the descendants of Wideröe's and Lawrence's pioneering efforts, and still use electric and magnetic fields to accelerate and direct particles.

Lawrence's development of the cyclotron occurred just as a pivotal discovery was unfolding in Europe: the existence of the neutron. This is the topic of the next section.

Fig. 2.19 Left Lawrence's original 4.5-in. cyclotron. Middle Lawrence at the controls of his later 184-in. cyclotron. Right Lawrence, Glenn Seaborg (1912–1999), and Robert Oppenheimer (1904–1967). *Sources* Lawrence Berkeley National Laboratory, courtesy AIP Emilio Segre Visual Archives
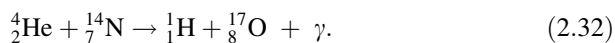
Fig. 2.20 M. Stanley Livingston and Ernest O. Lawrence at the Berkeley 27-in. cyclotron. *Source* Lawrence Berkeley National Laboratory, courtesy AIP Emilio Segre Visual Archives

2.2 Discovery of the Neutron

The discovery of the neutron in early 1932 by Ernest Rutherford's protégé, James Chadwick, was a critical turning point in the history of nuclear physics. Within two years, Enrico Fermi would generate artificially-induced radioactivity by neutron bombardment, and five years after that, Otto Hahn, Fritz Strassmann, and Lise Meitner would discover neutron-induced uranium fission. The latter would lead directly to the *Little Boy* uranium-fission bomb, while Fermi's work would lead to reactors to produce plutonium for the *Trinity* and *Fat Man* bombs.

The experiments which led to the discovery of the neutron were first reported in 1930 by Walther Bothe (Fig. 2.21) and his student, Herbert Becker, who were working in Germany. Their research involved studying the gamma radiation which is produced when light elements such as magnesium and aluminum are bombarded by energetic alpha-particles. In such reactions, the alpha particles often interact with a target nucleus to yield a proton and a gamma-ray, as Ernest Rutherford had found when he first achieved an artificially-induced nuclear transmutation:

$$^4_2\text{He} + ^{14}_7\text{N} \rightarrow ^1_1\text{H} + ^{17}_8\text{O} \; + \; \gamma. \tag{2.32}$$

The mystery began when Bothe and Becker found that boron, lithium, and particularly beryllium gave evidence of gamma emission under alpha bombardment, *but with no accompanying protons being emitted*. A key point here is that they were certain that some sort of energetic but electrically neutral "penetrating radiation" was being emitted; this radiation could penetrate foils of metal but could not be deflected by a magnetic field as charged particles would be. Gamma-rays were the only electrically neutral form of penetrating radiation known at the time, so it was natural for them to interpret their results as evidence of gamma-ray emission despite the anomalous lack of protons.
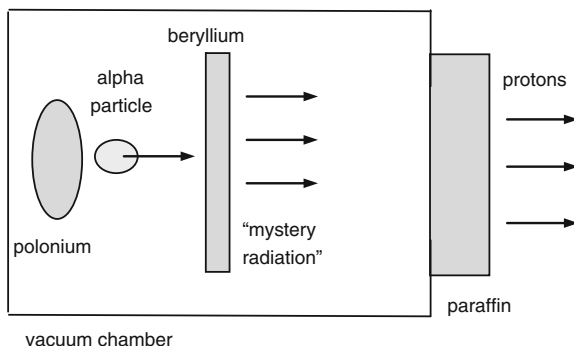
Bothe and Becker's beryllium result was picked up by the Paris-based husband-and-wife team of Frédéric Joliot and Irène Curie (the daughter of Pierre and Marie; Fig. 2.3), hereafter referred to as the Joliot-Curies. In January, 1932, they reported that the presumed gamma-ray "beryllium radiation" was capable of knocking protons out of a layer of paraffin wax that had been put in its path. The situation is shown schematically in Fig. 2.22, where the supposed gamma-rays are labeled as "mystery radiation."

At Cambridge, this interpretation struck Chadwick as untenable. He had searched for neutrons for many years with no success, and suspected that Bothe and Becker and the Joliot-Curies had stumbled upon them. He immediately set about to reproduce, re-analyze, and extend their work. In his recreation of the Joliot-Curies' work, Chadwick's experimental setup involved polonium (the alpha source)

Fig. 2.21 Left Walter Bothe (1891–1957); Right James Chadwick (1891–1974). *Sources* Original drawing by Norman Feather, courtesy AIP Emilio Segre Visual Archives; http://commons.wikimedia.org/wiki/File:Chadwick.jpg

Fig. 2.22 The "beryllium radiation" experiment of Bothe, Becker, the Joliot-Curies, and Chadwick

deposited on a silver disk 1 cm in diameter placed close to a disk of pure beryllium 2 cm in diameter, with both enclosed in a small vessel which could be evacuated. In comparison to the gargantuan particle accelerators of today, these experiments were literally table-top nuclear physics.

Let us first assume that Bothe and Becker and the Joliot-Curies were correct in their interpretation that α-bombardment of beryllium creates gamma-rays. To account for the lack of protons created in the bombardment, the Joliot-Curies hypothesized that the reaction was

$$\ce{^4_2He + ^9_4Be -> ^{13}_6C} + \gamma. \tag{2.33}$$

The Q-value of this reaction is 10.65 MeV. Polonium decay yields alpha particles with kinetic energies of about 5.3 MeV, so the emergent γ-ray can have at most an energy of about 16 MeV. A more detailed analysis which accounts for the energy and momentum transmitted to the carbon atom shows that the energy of the gamma ray comes out to be about 14.6 MeV. The 14.6-MeV gamma-rays then strike protons in the paraffin, setting them into motion. Upon reproducing the experiment, Chadwick found that the struck protons would emerge with maximum kinetic energies of about 5.7 MeV.

The problem, Chadwick realized, was that if a proton was to be accelerated to this amount of energy by being struck by a gamma-ray, conservation of energy and momentum demanded that the gamma-ray would have to possess about 54 MeV of energy, nearly four times what it could have! This strikingly high energy demand is a consequence of the fact that photons do not possess mass. Relativity theory shows that massless particles do carry momentum, but much less than a "material" particle of the same kinetic energy; only an extremely energetic gamma-ray can kick a proton to a kinetic energy of several MeV. Analyzing a collision between a photon and a material particle involves relativistic mass-energy and momentum conservation; details can be found in Reed (2007). The results of such an analysis show that if a target nucleus of rest-energy E_t (that is, mc^2 equivalent energy) is to be accelerated to kinetic energy K_t by being struck head-on by a photon of energy E_γ which then recoils backwards (this transfers maximum momentum to the struck nucleus), then the energy of the photon must be
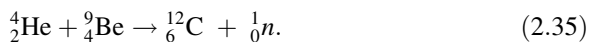
$$E_\gamma = \frac{1}{2}\left[K_t + \sqrt{2\,E_t\,K_t}\,\right]. \tag{2.34}$$

For a proton, $E_t \sim 938$ MeV; with $K_t \sim 5.7$ MeV, the value of E_γ works out to about 54 MeV, as claimed above. Remarkably, the Joliot-Curies had realized that this discrepancy was a weak point in their interpretation, but attributed it to the difficulty of accurately measuring the energy of their "gamma rays." Another clue that led Chadwick to suspect a material particle as opposed to a high-energy photon was that the "beryllium radiation" was more intense in the forward direction than in the backward direction; if the radiation was photonic, it should have been of equal intensity in all directions.

Before invoking a mechanism involving a (hypothetical) neutron, Chadwick devised a further test to investigate the remote possibility that 54-MeV gamma-rays could be being created in the α-Be collision. In addition to having the "beryllium radiation" strike protons, he also arranged for it to strike a sample of nitrogen gas. If struck by such a photon, a nucleus of nitrogen should acquire a kinetic energy of about 450 keV [A nitrogen nucleus has a rest energy of about 13,000 MeV; check the consistency of these numbers with Eq. (2.34)]. From prior experience, Chadwick knew that when an energetic particle travels through air it produces ions, with about 35 eV required to produce a single ionization, which yields one pair of ions. A 450 keV nitrogen nucleus should thus generate some (450 keV/35 eV) $\sim 13{,}000$ ion pairs. Upon performing the experiment, however, he found that some 30,000–40,000 ion pairs would typically be produced, which implied kinetic energies of about 1.1–1.4 MeV for the recoiling nitrogen nuclei. Such numbers would in turn require the nitrogen nuclei to have been struck by gamma-rays of energy up to ~ 90 MeV, a value completely inconsistent with the ~ 54 MeV indicated by the proton experiment. Upon letting the supposed gamma-rays strike heavier and heavier target nuclei, Chadwick found that "if the recoil atoms are to be explained by collision with a quantum, we must assume a larger and larger energy for the quantum as the mass of the struck atom increases." The absurdity of this situation led him to write that, "It is evident that we must either relinquish the application of conservation of energy and momentum in these collisions or adopt another hypothesis about the nature of the radiation."

After refuting the Joliot-Curies' interpretation, Chadwick provided a more physically realistic one. This was that if the protons in the paraffin were in reality being struck by neutral *material* particles of mass equal or closely similar to that of a proton, then the kinetic energy of the striking particles need only be on the order of the kinetic energy that the protons acquired in the collision. As an everyday example, think of a head-on collision between two equal-mass billiard balls: the incoming one stops, and the struck one is set into motion with the speed that the incoming one had. This is the point at which the neutron makes it debut.

Chadwick hypothesized that instead of the Joliot-Curie reaction, the α-Be collision leads to the production of carbon and a neutron via the reaction

$$_{2}^{4}\mathrm{He} + _{4}^{9}\mathrm{Be} \rightarrow {}_{6}^{12}\mathrm{C} \; + \; {}_{0}^{1}n. \tag{2.35}$$

$_{0}^{1}n$ denotes a neutron: it carries no electric charge but it does count as one nucleon. In this interpretation, a ^{12}C atom is produced as opposed to the Joliot-Curies' proposed ^{13}C. Since the "beryllium radiation" was known to be electrically neutral, Chadwick could not invoke a charged particle such as a proton or electron to explain the reaction. Hypothesizing that the neutron's mass was similar to that of a proton (he was thinking of neutrons as being electrically neutral combinations of single protons and single electrons), Chadwick was able to show that the kinetic energy of the ejected neutron would be about 10.9 MeV. A subsequent neutron/proton collision will be like a billiard-ball collision, so it is entirely plausible that a neutron which begins with about 11 MeV of kinetic energy would be sufficiently energetic to accelerate a proton to a kinetic energy of 5.7 MeV, even after the neutron battered its way out of the beryllium target and through the window of the vacuum vessel on its way to the paraffin. As a check on his hypothesis, Chadwick calculated that a neutron of kinetic energy 5.7 MeV striking a nitrogen nucleus should set the latter into motion with a kinetic energy of about 1.4 MeV, which was precisely what he had measured in the ion-pair experiment!

Further experiments with other target substances showed similarly consistent results. Chadwick estimated the mass of the neutron as between 1.005 and 1.008 atomic mass units; the modern figure is 1.00866. The accuracy he obtained with equipment which would now be regarded as primitive is nothing short of awe-inspiring. Chadwick reported his discovery in two papers. The first, titled "Possible Existence of a Neutron," was dated February 17, 1932, and was published in the February 27 edition of *Nature*. An extensive follow-up analysis dated May 10 was published in the June 1 edition of the *Proceedings of the Royal Society of London*. Chadwick was awarded the 1935 Nobel Prize in Physics for his discovery. While later experiments showed that the neutron is a fundamental particle in its own right (as opposed to being a proton/electron composite), that development does not affect the above analysis.

Why is the discovery of the neutron regarded as such a pivotal event in the history of nuclear physics? The reason is that neutrons do not experience any electrical forces, so they experience no Coulomb barrier. With neutrons, experimenters now had a way of producing particles that could be used to bombard nuclei without being repelled by them, no matter what the kinetic energy of the neutron or the atomic number of the target nucleus. It was not long before such experiments were taken up. Neutrons would prove to be the gateway to reactors and bombs, but, at the time, Chadwick anticipated neither development. In the February 29, 1932, edition of the *New York Times*, he is quoted as stating that, "I am afraid neutrons will not be of any use to any one."

About 18 months after Chadwick's dismissal of the value of neutrons, an idea did arise as to a possible application for them: As links in the progression of a nuclear chain reaction. This notion seems to have occurred inspirationally to a Hungarian-born engineer, physicist, and inventor named Leo Szilard (Fig. 2.23), a personal friend and sometimes collaborator of Albert Einstein.

Fig. 2.23 Leo Szilard (1898–1964) *Source* http://commons.wikimedia.org/wiki/File:Leo_Szilard.jpg

Szilard was living in London in the fall of 1933, and happened to read a description of a meeting of the British Association for the Advancement of Science published in the September 12 edition of the London *Times*. In an article describing an address to the meeting by Rutherford on the prospects for reactions that might be induced by accelerated protons, the *Times* quoted Rutherford as stating that, "We might in these processes obtain very much more energy than the proton supplied, but on the average we could not expect to obtain energy in this way. It was a very poor and inefficient way of producing energy, and anyone who looked for a source of power in the transformation of the atoms was talking moonshine." Historian of science John Jenkin has pointed out that Rutherford's private thoughts on the matter may have been very different, however. Some years before World War II, Rutherford evidently advised a high government official that he had a hunch that nuclear energy might one day have a decisive effect on war.

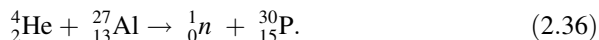Szilard reflected on Rutherford's remarks while later strolling the streets of London. From a 1963 interview:

Pronouncements of experts to the effect that something cannot be done have always irritated me. That day as I was walking down Southampton Row and was stopped for a traffic light, I was pondering whether Lord Rutherford might not prove to be wrong. As the light changed to green and I crossed the street, it suddenly occurred to me that if we could find an element which is split by neutrons and which would emit two neutrons when it absorbed one neutron, such an element, if assembled in sufficiently large mass, could sustain a nuclear chain reaction, liberate energy on an industrial scale, and construct atomic bombs. The thought that this might be possible became an obsession with me. It led me to go into nuclear physics, a field in which I had not worked before, and the thought stayed with me.

It did not take Szilard long to get up to speed in his new area. Envisioning a chain reaction as a source of power and possibly as an explosive, he filed for patents on the idea in the spring and summer of 1934. His British patent number 630,726, "Improvements in or relating to the Transmutation of Chemical Elements," was issued on July 4, 1934 (curiously, the date of Marie Curie's death), and referred specifically to being able to produce an explosion given a sufficient mass of material. To keep the idea secret, Szilard assigned the patent to the British Admiralty in February, 1936. The patent was reassigned to him after the war, and finally published in 1949.

2.3 Artificially-Induced Radioactivity

Irène and Frédéric Joliot-Curie must have been deeply disappointed at their failure to detect the neutron in early 1932, but scored a success almost exactly two years later when they discovered that normally stable nuclei could be induced to become radioactive upon alpha-particle bombardment. In early 1934, the Joliot-Curies were performing some follow-up experiments involving bombarding thin foils of aluminum with alpha-particles emitted by decay of polonium, the same source of alphas used in the neutron-discovery reaction. To their surprise, their Geiger counter continued to register a signal after the source of the alpha particles was removed. The signal decayed with a half-life of about 3 min. Performing the experiment in a magnetic field led them to conclude that positrons were being emitted, that is, that β^+ decays were occurring.

They proposed a two-stage reaction to explain their observations. First was formation of phosphorous-30 by alpha-capture and neutron emission:

$$_2^4\text{He} + \, _{13}^{27}\text{Al} \rightarrow \, _0^1 n \, + \, _{15}^{30}\text{P}. \tag{2.36}$$

The phosphorous-30 nucleus subsequently undergoes positron decay to silicon; the modern value for the half-life is 2.5 min (the emitted beta-particle is omitted here; it is the decay product that is important):

$$_{15}^{30}\text{P} \, \xrightarrow[2.5\,\text{min}]{\beta^+} \, _{14}^{30}\text{Si}. \tag{2.37}$$

To be certain of their interpretation, the Joliot-Curies dissolved the bombarded aluminum in acid; the small amount of phosphorous created could be separated and chemically identified as such. That the radioactivity "carried with" the separated phosphorous and not the aluminum verified their suspicion. Bombardment of boron and magnesium showed similar effects. They first observed the effect on January 11, 1934, and reported it in the January 15 edition of the journal of the French Academy of Sciences; an English version appeared in the February 10

edition of *Nature*. The discovery of artificially-induced radioactivity opened up the whole field of synthesizing short-lived isotopes for medical treatments. Emilio Segrè, one of Enrico Fermi's students, described this development as one of the most important discoveries of the century.

Induced radioactivity had almost been discovered in California, where Ernest Lawrence's cyclotron operators often noticed that their detectors kept registering a signal after the cyclotron had been shut down following bombardment experiments. Thinking that the detectors were misbehaving, they arranged circuitry to shut them down simultaneously with the cyclotron. The history of nuclear physics, particularly events surrounding the discovery of fission, is replete with such missed chances.

2.4 Enrico Fermi and Neutron-Induced Radioactivity

Surprisingly, neither the Joliot-Curies nor James Chadwick particularly experimented with using neutrons as bombarding particles. Norman Feather, one of Chadwick's collaborators, did carry out some experiments with light elements, and found that neutrons would disintegrate nitrogen nuclei to produce an alpha-particle and a boron nucleus:

$$^{1}_{0}n + {}^{14}_{7}\text{N} \rightarrow {}^{4}_{2}\text{He} + {}^{11}_{5}\text{B}. \tag{2.38}$$

The same type of reaction also occurs with elements such as oxygen, fluorine, and neon, but apparently neither British nor French researchers carried out experiments with heavy-element targets.

The idea of systematically using neutrons as bombarding particles did occur to a physicist at the University of Rome, Enrico Fermi (Fig. 2.24). Fermi had established himself as a first-rate theoretical physicist at a young age, publishing his first paper while still a student. As a postdoctoral student with quantum mechanist Max Born, Fermi had prepared an important review article on relativity theory while in his early twenties, and a few years later made seminal contributions to statistical mechanics. At the young age of 26 he was appointed to a full professorship at the University of Rome, and in late 1933 he developed a quantum–mechanically-based theory of beta decay. He was to prove equally gifted as a nuclear experimentalist.

The reticence of Chadwick and the Joliot-Curies to carry out neutron-bombardment experiments may seem strange, but was understandable in view of the low yields expected. Chadwick estimated that he produced only about 30 neutrons for every million alpha-particles emitted by his sample of polonium. If the neutrons interacted with target nuclei with similarly low yields, virtually nothing could be expected to result. Otto Frisch, one of the co-interpreters of fission, later remarked that, "I remember that my reaction and probably that of many others was that Fermi's was a silly experiment because neutrons were so
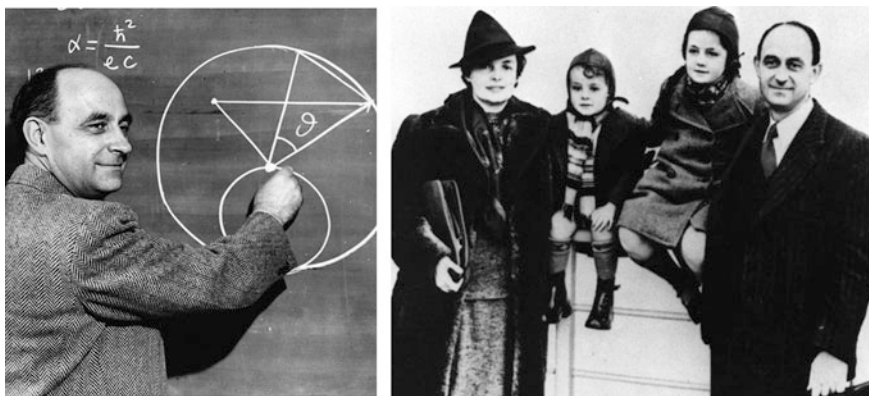
Fig. 2.24 Left Enrico Fermi (1901–1954). Right The Fermi family (Laura, Giulio, Nella and Enrico) arrive in America, January, 1939. *Sources* University of Chicago, courtesy AIP Emilio Segre Visual Archives; AIP Emilio Segre Visual Archives, Wheeler Collection

much fewer than alpha particles." But this overlooked the fact that neutrons would not experience a Coulomb barrier.

Fermi desired to break into nuclear experimentation, and saw his opening in this under-exploited possibility. He began work in the spring of 1934 with a group of students and collaborators including Edoardo Amaldi, Franco Rasetti, chemist Oscar D'Agostino, and Emilio Segrè (Fig. 2.25), who would later write an very engaging biography of Fermi, titled *Enrico Fermi*: *Physicist.*
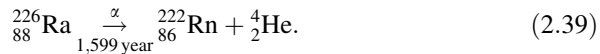
In the early 2000s a group of Italian historians, Giovanni Acocella, Francesco Guerra, Matteo Leone, and Nadia Robotti found Fermi's original laboratory notebooks (and some of his neutron sources!) from the spring of 1934, so there is now available a very detailed record of his work. Much of the material in this section is adapted from their analysis of Fermi's notes.

Fermi's first challenge was to secure a strong neutron source. In this sense he was fortunate in that his laboratory was located in the same building as the Physical Laboratory of the Institute of Public Health, which was charged with
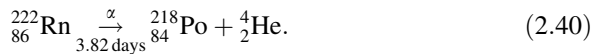
Fig. 2.25 Some of Fermi's collaborators. Left to right Oscar D'Agostino (1901–1975), Emilio Segrè (1905–1989), Edoardo Amaldi (1908–1989), and Franco Rasetti (1901–2001). *Source* Agenzia Giornalistica Fotovedo, courtesy AIP Emilio Segre Visual Archives

controlling radioactive substances in Italy. The Laboratory held many radium sources that had been used for cancer treatments, and Fermi used them as a source of radon gas. When mixed with powdered beryllium, the radon gave rise to a copious supply of neutrons. Radon is produced in the decay of radium,

$$\ce{^{226}_{88}Ra ->[\alpha][1,599\,year] ^{222}_{86}Rn + ^4_2He}. \tag{2.39}$$

The radon daughter product has a very short half-life, which means a correspondingly great flux of alpha-particles from the decay

$$\ce{^{222}_{86}Rn ->[\alpha][3.82\,days] ^{218}_{84}Po + ^4_2He}. \tag{2.40}$$
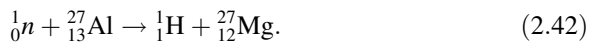
After being harvested from the decaying radium, the radon gas was captured in inch-long glass vials which contained powdered beryllium. The radon-produced alphas in (2.40) then gave rise to neutrons via the same reaction that Bothe and Becker, the Joliot-Curies, and Chadwick had experimented with:

$$\ce{^4_2He + ^9_4Be -> ^{12}_6C + ^1_0n}. \tag{2.41}$$

This series of reactions yields neutrons with energies of up to about 10 MeV, more than energetic enough to escape through the thin walls of the glass vials and so bombard a sample of a target element. Fermi estimated that his sources yielded about 100,000 neutrons per second. Because the neutrons generated by his radon-beryllium sources tended to be emitted in all directions, Fermi usually formed samples of the target elements to be investigated into cylinders which could be placed around the sources in order to achieve maximum exposure. The cylinders were made large enough so that after being irradiated they could be slipped around a small handmade Geiger counter.
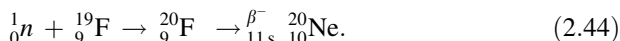
Fermi's goal was to see if he could induce artificial radioactivity with neutron bombardment. Possibly anxious to see if he could induce *heavy-element* radioactivity, his first target was the heavy element platinum (atomic number 78). Fifteen minutes of irradiation gave no discernible signal. Perhaps inspired by the Joliot-Curies' experience, he then turned to aluminum. Here he did succeed, and found a different half-life than they had. The reaction involved ejection of a proton from the bombarded aluminum, leaving behind magnesium,

$$\ce{^1_0n + ^{27}_{13}Al -> ^1_1H + ^{27}_{12}Mg}. \tag{2.42}$$

The magnesium beta-decays back to aluminum with a half-life of about 10 min:

$$\ce{^{27}_{12}Mg ->[\beta^-][9.5\,min] ^{27}_{13}Al}. \tag{2.43}$$
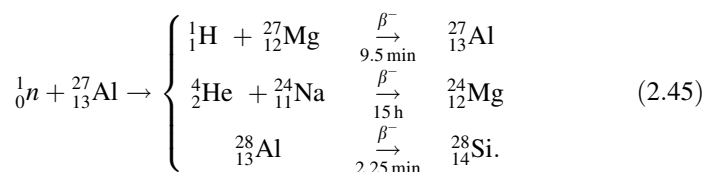
After aluminum, Fermi tried lead, but with negative results. His next attempt was with fluorine, irradiation of which produced a very short-lived heavier isotope of that element:

$$\frac{1}{0}n + \frac{19}{9}F \rightarrow \frac{20}{9}F \xrightarrow[11\,s]{\beta^-} \frac{20}{10}Ne. \tag{2.44}$$
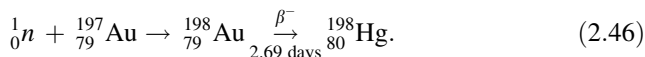
Guerra and Robotti have pinpointed the date of Fermi's first success with aluminum as having occurred on March 20, 1934. Fermi announced his discovery five days later in the official journal of the Italian National Research Council, and an English-language report dated April 10 appeared in the May 19 edition of *Nature*. By late April, the Rome group had performed experiments on about 30 elements, 22 of which yielded positive results, including the four medium-weight elements antimony ($Z = 51$), iodine (53), barium (56), and lanthanum (57).

Fermi and his co-workers found that, as a rule, light elements exhibited three reaction *channels*: a proton or an alpha could be ejected, or the element might simply capture the neutron to become a heavier isotope of itself and then subsequently decay. In all three cases, the products would undergo β^- decay. Aluminum is typical in this regard:
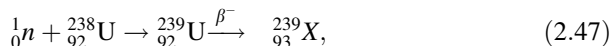
$$\frac{1}{0}n + \frac{27}{13}Al \rightarrow \begin{cases} \frac{1}{1}H + \frac{27}{12}Mg & \xrightarrow[9.5\,min]{\beta^-} & \frac{27}{13}Al \\[2mm] \frac{4}{2}He + \frac{24}{11}Na & \xrightarrow[15\,h]{\beta^-} & \frac{24}{12}Mg \\[2mm] \frac{28}{13}Al & \xrightarrow[2.25\,min]{\beta^-} & \frac{28}{14}Si. \end{cases} \tag{2.45}$$

With a *heavy*-element target, the result is typically the latter of the above channels. Gold is characteristic in this regard:

$$\frac{1}{0}n + \frac{197}{79}Au \rightarrow \frac{198}{79}Au \xrightarrow[2.69\,days]{\beta^-} \frac{198}{80}Hg. \tag{2.46}$$

By the early summer of 1934, Fermi had prepared improved sources, which he estimated were yielding about a million neutrons per second. Based on work with these new sources, he published a stunning result in the June 16, 1934, edition of *Nature*: that his group was producing *transuranic* elements, that is, ones with atomic numbers greater than that of uranium. Since uranium was the heaviest-known element, this meant that they believed that they were synthesizing new elements. If true, this would be a remarkable development.

Fermi's radical assertion was based on the fact that uranium could be activated to produce beta-decay upon neutron bombardment. The results were complex, however, with evidence for half-lives of 10 s, 40 s, 13 min, and at least two more half-lives of up to one day. Whether this was a chain of decays or some sort of parallel sequence was unknown. Whatever sequence was occurring, however, the initial step was presumably the formation of a heavy isotope of uranium, followed by a beta-decay as in the gold reaction above:

$$\frac{1}{0}n + \frac{238}{92}U \rightarrow \frac{239}{92}U \xrightarrow{\beta^-} \frac{239}{93}X, \tag{2.47}$$

where X denotes a new, transuranic element.

The 13-min decay was convenient to work with, and the Rome group managed to separate chemically its decay product from the bombarded uranium. Analysis showed that the decay product did not appear to be any of the elements between lead ($Z = 82$) and uranium. Since no natural or artificial transmutation had ever been observed to change the identity of a target element by more than one or two places in the periodic table, it would have seemed perfectly plausible to assume that a new element was being created.

To isolate the product of the 13-min activity, Fermi and his group began with manganese dioxide as a chemical carrier. The rationale for this was that if element 93 were actually being created, it was expected that it would fall in the same column of the periodic table as manganese ($Z = 25$), and so the two should have similar chemistry (see Fig. 3.2). The Romans' analysis came in for criticism, however, from a German scientist, Ida Noddack (Fig. 2.26). Noddack was a well-regarded chemist who in 1925 had participated in the discovery of rhenium; she would be nominated for a Nobel Prize on three occasions. In a paper published in September, 1934, Noddack criticized Fermi on the grounds that numerous elements were known to precipitate with manganese dioxide, and that he should have checked for the possibility that elements of lower atomic numbers than that of lead were being produced. In what would prove to be a prescient comment, Noddack remarked that, "When heavy nuclei are bombarded by neutrons, it is conceivable that the nucleus breaks up into several large fragments, which would of course be isotopes of known elements but would not be neighbors of the irradiated element." Noddack's "breaking up" is now known as "fission."

Noddack was ahead of her time in suggesting that heavy nuclei might fission. Ironically, Fermi was probably both inducing fissions *and* creating transuranic elements. Nuclei of the most common isotope of uranium, U-238 (>99 % of natural uranium), are fissile when bombarded by the very fast neutrons that Fermi was using, but when struck by slow neutrons tend to capture them and subsequently decay to neptunium and plutonium. These processes will be discussed at length in subsequent sections.
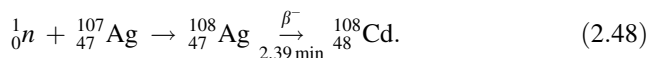
That Noddack's idea was not taken seriously has sometimes been construed as an example of blatant sexism. But the reasons were much more prosaic. She offered no supporting calculations of the energetics of such a proposed splitting, and years of experience with nuclear reactions had always yielded products that were near the bombarded elements in atomic number. Nobody had any reason to anticipate such a splitting. Otto Frisch thought Noddack's paper was "carping criticism." In any case, by the summer of 1934, Fermi's group had developed an improved rhenium-based chemical analysis of the 13-min uranium activation which appeared to strengthen the transuranic interpretation.

Fermi's next discovery would prove pivotal to the eventual development of plutonium-based nuclear weapons. In the fall of 1934, his group decided that they needed to more precisely quantify their assessments of activities induced in various elements; previously they had assigned only qualitative "strong-medium-

weak" designations. As a standard of activity, they settled on a 2.4-min half-life
induced in silver:

$$\,_0^1 n \, + \, _{47}^{107}\mathrm{Ag} \; \rightarrow \; _{47}^{108}\mathrm{Ag} \; \xrightarrow[2.39\,\mathrm{min}]{\beta^-} \; _{48}^{108}\mathrm{Cd}. \tag{2.48}$$

However, they soon ran into a problem: the activity induced in silver seemed to
depend on where in the laboratory the sample was irradiated. In particular, silver
irradiated on a wooden table became much more active then when irradiated on a
marble-topped one. To try to discern what was happening, a series of calibration
experiments was undertaken, some of which involved investigating the effects of
"filtering" neutrons by interposing layers of lead between the neutron source and
the target sample.

Fermi made the key breakthrough on October 22, 1934: "One day, as I came
into the laboratory, it occurred to me that I should examine the effect of placing a
piece of lead before the incident neutrons. Instead of my usual custom, I took great
pains to have the piece of lead precisely machined. I was clearly dissatisfied with
something; I tried every excuse to postpone putting the piece of lead in its place.
When finally, with some reluctance, I was going to put it in its place, I said to
myself: "No, I don't want this piece of lead here; what I want is a piece of
paraffin." It was just like that with no advance warning, no conscious prior rea-
soning. I immediately took some odd piece of paraffin and placed it where the
piece of lead was to have been."

To Fermi's surprise, the presence of the paraffin caused the level of induced radioactivity to increase. Further experimentation showed that the effect was characteristic of filtering materials which contained hydrogen; paraffin and water were most effective. Within a few hours of the discovery, Fermi developed a working hypothesis: That by being slowed by collisions with hydrogen nuclei, the neutrons would have more time in the vicinity of target nuclei to induce a reaction. Neutrons and protons have essentially identical masses, and, as with a billiard-ball collision, a head-on strike would essentially bring a neutron to a stop. Since atoms always have random motions due to being at a temperature that is above absolute zero, the incoming neutrons will never be brought to dead stops, but in practice only a few centimeters of paraffin or water are needed to bring them to an average speed characteristic of the temperature of the slowing medium. This process is now called "thermalization." Nuclear physicists define "thermal" neutrons as having kinetic energy equivalent to a temperature of 298 K, or 77 °F—not much warmer than the average daily temperature in Rome in October. The speed of a thermal neutron is about 2,200 m/s, and the corresponding kinetic energy is about 0.025 eV, much less than the ∼ 10 MeV of Fermi's radon-beryllium neutrons. Thermal neutrons are also known as "slow" neutrons; those of MeV-scale kinetic energies are, for obvious reasons, termed "fast." The water or paraffin is now known as a "moderator"; graphite (crystallized carbon) also works well in this respect. Fermi's wooden lab bench, by virtue of its water content, was a more effective moderator than was his marble-topped one.

Be sure to understand what is meant by "fast" and "slow" neutrons. When uranium is bombarded by neutrons, what happens depends very critically on the kinetic energies of the neutrons. Fast and slow neutrons lie at the heart of why nuclear reactors and bombs function differently, and why a bomb requires "enriched" uranium to function. This is a complex topic with a number of interconnecting aspects; the following chapter is devoted to a detailed analysis of the ramifications of this fast-versus-slow issue.

Following Fermi's serendipitous discovery, his group began re-investigating all elements which they had previously subjected to *fast* (energetic) neutron bombardment. Extensive results were reported in a paper published in the spring of 1935. For some target elements, the effect was dramatic: activity in vanadium and silver were increased by factors of 40 and 30, respectively, over that achieved by unmoderated neutrons. Uranium also showed increased activation, by a factor of about 1.6.

Fermi's hypothesis that slower neutrons have a greater chance of inducing a reaction is now quantified in the concept of a reaction *cross-section*. This is a measure of the cross-sectional area that a target nucleus effectively presents to a bombarding particle that results in a given reaction. Because of a quantum–mechanical effect known as the de Broglie wavelength, a target nucleus will *appear* larger to a slower bombarding particle than to a faster one, sometimes by factors of hundreds. Each possible reaction channel for a target nucleus will have its own characteristic run of cross-section as a function of bombarding-particle energy.

Cross-sections are designated with the Greek letter sigma (σ), equivalent to the English letter "s," which serves as a reminder that they have units of surface area. The fundamental unit of cross-section is the "barn"; 1 bn = 10^{-28} m^2. This miniscule number is characteristic of the geometric cross-sectional area of nuclei, which is given approximately in terms of the mass number by the empirical relationship

$$Geometric \; cross-section \sim 0.0452 \, A^{2/3} \; (barns). \tag{2.49}$$

As an example, Fig. 2.27 shows the "radiative capture cross-section" for aluminum-27 when bombarded by neutrons of energies from 10^{-11} to 10 MeV. (Al-27 is the one stable isotope of that element.) In this reaction, the aluminum absorbs the neutron, sheds some energy via a gamma-ray, and eventually decays to silicon via the last branch of the three-channel reaction in (2.45); both "capture" and "radiation" occur, hence the name of the cross-section. Both scales of the graph are logarithmic; this is done in order to accommodate a wide range of energies and cross-sections. Thermal neutrons have log (Energy) ~ -7.6 when the energy is measured in MeV. For an Al-27 nucleus, the geometric cross-section is about 0.407 barns, or log(area) = -3.91. For fast neutrons, reaction cross-sections are typically of the size of the geometrical cross-sectional area of the target nucleus.

The spikes in Fig. 2.27 are known as "resonance capture lines." Just as atomic orbital electrons can be excited to different energy levels, so can the protons and neutrons within nuclei; resonance energies correspond to the bombarding particles having just the right energies to excite nucleons to higher energy levels. As the number of nucleons grows, so does the complexity of the structure of the resonance spikes; for a more dramatic example, see the graph for the uranium-238 neutron capture cross-section in Fig. 3.11.

Fermi was awarded the 1938 Nobel Prize for Physics for his demonstration of the existence of new radioactive elements produced by neutron irradiation. Fermi's

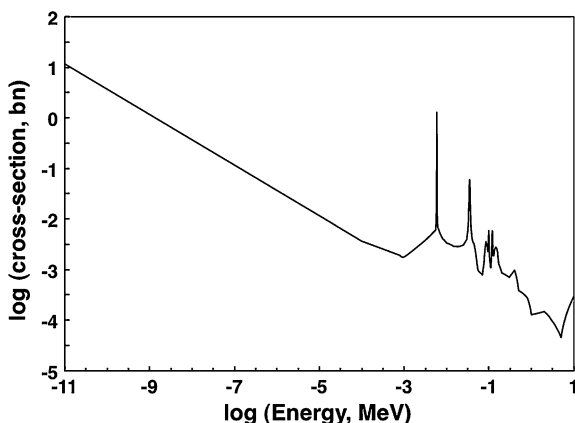**Fig. 2.27** Radiative capture cross-section for neutrons on aluminum-27

Fig. 2.28 Arthur Dempster (1886–1950). *Source* http://commons.wikimedia.org/wiki/File:Arthur_Jeffrey_Dempster_-_Portrait.jpg

wife and children were Jewish, and he and his family used the excuse of the trip to Stockholm to escape the rapidly deteriorating fascist political situation in Italy by subsequently emigrating to America, where he had arranged for a position at Columbia University. The American branch of the Fermi family was established on January 2, 1939.

Before proceeding to the story of the discovery of nuclear fission, a brief but important intervening discovery needs to be mentioned. This is that uranium possesses a second, much less abundant isotope than the U-238 that Fermi had assumed was the sole form of that element. In 1931, Francis Aston had run uranium hexafluoride through his mass spectrometer and concluded that only an isotope of mass number 238 was present. In the summer of 1935, Arthur Dempster of the University of Chicago (Fig. 2.28) discovered evidence for a lighter isotope of mass number 235. Dempster estimated U-235 to be present to an extent of less than one percent of the abundance of its sister isotope of mass 238. Within a few years that one percent would prove very important indeed.

2.5 Another Look at Mass Defect and Binding Energy (Optional)

In Sect. 2.1.4, the concepts of mass defect and binding energy were treated as interchangeable. They are, however, separate but related quantities. The strict definition of binding energy is described in this section, which can be considered optional.

Atomic masses are measured in terms of *mass units*. Abbreviated simply as *u*, the mass unit is defined as one-twelfth of the mass of a neutral carbon-12 atom. Chemists will know the mass unit as a "Dalton," and older readers will be more familiar with the term *atomic mass unit* (*amu*). The numerical value is

$$1\,u = 1.660539 \times 10^{-27}\,\text{kg}. \tag{2.50}$$

The masses of the proton, neutron, and electron in mass units are

$$m_p = 1.00727646677\,u \tag{2.51}$$

$$m_n = 1.00866491597\,u \tag{2.52}$$

$$m_e = 5.4857990943 \times 10^{-4}\,u \tag{2.53}$$

An important conversion factor here is that the energy equivalent of one mass unit is 931.494 MeV; this comes from $E = mc^2$. Give this value the symbol ε:

$$\varepsilon = 931.494\,\text{MeV}. \tag{2.54}$$

A neutral atom comprises Z protons, Z electrons, and N neutrons. In the notation of (2.51)–(2.53), one would naively expect the mass of the assembled atom to be equal to $Z(m_p + m_e) + N(m_n)$ mass units. However, naturally-occurring "assembled" neutral atoms are always lighter than what this argument predicts. The *binding energy E_B* of an atom is defined as the difference between the naively-predicted mass and the "true" measured mass m_U of the assembled atom (in mass units), all expressed as an energy equivalent:

$$E_B = \left[Z\left(m_p + m_e\right) + N(m_n) - m_U \right]\varepsilon. \tag{2.55}$$

Substituting (2.51)–(2.54) into (2.55) gives the binding energy as

$$E_B = \left[938.783\,Z + 939.565\,N - 931.494\,m_U\right]\text{MeV}. \tag{2.56}$$
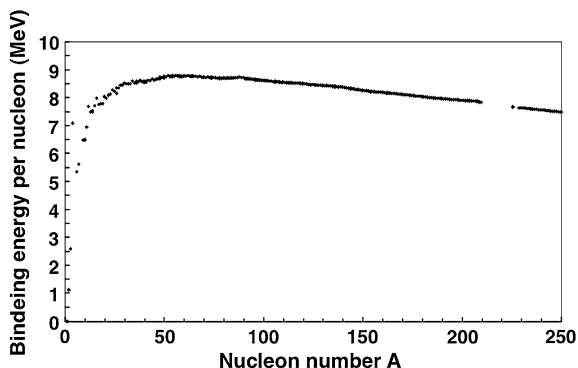
As an example, for iron-56 ($m_U = 55.934937$):

$$E_B = 938.783(26) + 939.565(30) - 931.494(55.934937) = 492.2\,\text{MeV}.$$

To correct a misleading statement in Sect. 2.1.4, it is this binding energy that holds nuclei together, not the mass-defect; a positive mass defect does not by itself connote instability. Stable atoms will have positive E_B values, but, conversely, a positive E_B value does not necessarily denote intrinsic stability. For example, uranium-235 has $E_B = 1{,}784$ MeV, but is unstable against alpha-decay, which is fundamentally a quantum-mechanical effect that cannot be understood on the basis of energy considerations alone.

For heavy elements, E_B values are large. To display them graphically it is more convenient to plot the binding emery *per nucleon*, E_B/A, versus the nucleon number A. This is shown in Fig. 2.29 for the same 350 nuclides as in Fig. 2.8. This plot immediately tells us that for $A > \sim 25$, each nucleon in a nucleus is "glued" into the structure of the nucleus to the extent of about 8 MeV/nucleon. This plot is known as "the curve of binding energy." For Fe-56, $E_B/A = 8.79$ MeV/nucleon.

Fig. 2.29 Binding energy curve for 350 stable and quasi-stable nuclei

Exercises

2.1 Consider an element with atomic weight A grams per mole and density ρ grams per cubic centimeter. If atoms are imagined to be hard spheres of radius R packed edge-to edge, each atom will effectively occupy a cube of volume $8R^3$. Show that R can be expressed approximately as

$$R \sim 0.59(A/\rho)^{1/3}\text{Å}.$$

Apply this result to lithium; $(\rho, A) = (0.534 \text{ g/cm}^3, 7 \text{ g/mol})$, and uranium; $(\rho, A) = (18.95 \text{ g/cm}^3, 238 \text{ g/mol})$. [Ans: $R \sim 1.4$ Å in each case]

2.2 Show that a 100 food-calorie snack is equivalent to about 2.6×10^{18} MeV.

2.3 Take radium to have an atomic weight of 226 g/mol. The energy of each radium alpha-decay is 4.78 MeV. If all of the atoms in one gram of radium decay and all of that energy could be used to lift a mass m to a height $h = 1$ mile $= 1,609$ m, how much mass could be so lifted (hint: mgh)? You will find Rutherford's estimate of 500 t to be optimistic, but the answer, 129,000 kg $\sim$ 143 t, is still impressive.

2.4 Consider one gram of freshly-isolated radium-226 ($t_{1/2} = 1,599$ year). If each alpha-decay liberates 4.78 MeV, how much energy will be emitted in the first year after the sample is isolated? [Ans: 884 kJ]

2.5 A serving of a sports drink contains 50 mg of potassium to help athletes restore their electrolyte levels. However, one naturally-occurring isotope of potassium, K-40, is a beta-decayer with a half-life of 1.25 billion years. This isotope is present to the level of 1.17 % in natural potassium. If the average atomic weight of potassium is 39.089 g/mol, what level of beta-activity will you consume with one serving—at least until you excrete it? [Ans: 158 decay/s]

2.6 Given the empirical relationship R $\sim a_O A^{1/3}$ between nucleon number and nuclear radius with $a_O = 1.2 \times 10^{-15}$ m, verify Eq. (2.49) for the geometric cross-section of a nucleus.

2.7 In a particle accelerator, it is desired to fire calcium atoms, (Z, A) = (20, 40), into a stationary uranium target, (Z, A) = (92, 238), in an effort to synthesize nuclei of high atomic number. What Coulomb barrier will have to be overcome? [Ans: 230 MeV]

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