CHAPTER 1

Nuclear Systematics

Isotope geology is concerned with the measurement and interpretation of the variations of the isotope composition of certain elements in natural materials. These variations are the result of two quite different processes:

1. the spontaneous decay of the nuclei of certain atoms to form stable nuclei of other elements and the accumulation of these radiogenic daughter atoms in the minerals in which they formed and
2. the enrichment (or depletion) of certain stable atoms of elements of low atomic number in the products of chemical reactions as a result of changes in state such as evaporation and condensation of water and during physical processes such as diffusion.

The interpretation of the changes in the isotope compositions of the affected elements has become a powerful source of information in all branches of the earth sciences. The use of this investigative tool requires a thorough understanding of geological, hydrological, biospheric, and atmospheric processes occurring on the Earth and elsewhere in the solar system. In addition, the interpretation of isotopic data requires knowledge of the relevant principles of atomic physics, physical chemistry, and biochemistry. The decay of unstable atoms is accompanied by the emission of nuclear particles and radiant energy, which together constitute the phenomenon of radioactivity. The discovery of this process near the end of the nineteenth century was a milestone in the history physics and greatly increased our understanding of the Earth.

1.1 DISCOVERY OF RADIOACTIVITY

The rise of geology as a science is commonly associated with the work of James Hutton in Scotland. He emphasized the importance of very slow but continuously acting processes that shape the surface of the Earth. This idea conflicted with Catastrophism and foreshadowed the concept of Uniformitarianism, developed by Hutton in his book *Theory of the Earth*, published in 1785. His principal point was that the same geological processes occurring at the present time have shaped the history of the Earth in the past and will continue to do so in the future. He stated that he could find “no vestige of a beginning—no prospect of an end” for the Earth. Hutton’s conclusion regarding the history of the Earth was not well received by his contemporaries. However, as time passed, geologists accepted the principle of Uniformitarianism, including the conviction that very long periods of time are required for the deposition of sedimentary rocks whose accumulated thickness amounts to many miles. In 1830, Charles Lyell published the first volumes of his *Principles of Geology*. By the middle of the nineteenth century, geologists seemed to be secure in their conviction that the Earth was indeed very old and that long periods of time are required for the deposition of the great thickness of sedimentary rocks that had been mapped in the field.

The apparent antiquity of the Earth and the principle of Uniformitarianism were unexpectedly attacked by William Thomson, better known as Lord Kelvin (Burchfield, 1975). Thomson was Britain's most prominent physicist during the second half of the nineteenth century. His invasion into geology profoundly influenced geological opinion regarding the age of the Earth for about 50 years. Between 1862 and 1899 Thomson published a number of papers in which he set a series of limits on the possible age of the Earth. His calculations were based on considerations of the luminosity of the Sun (Thomson, 1862), the cooling history of the Earth, and the effect of lunar tides on the rate of rotation of the Earth. He initially concluded that the Earth could not be much more than 100 million years old. In subsequent papers, he further reduced the age of the Earth. In 1897, Lord Kelvin (he was raised to peerage in 1892) delivered his famous lecture, "The Age of the Earth as an Abode Fitted for Life" (Thomson, 1899), in which he narrowed the possible age of the Earth to between 20 and 40 million years.

These and earlier estimates of the age of the Earth by Lord Kelvin and others were a serious embarrassment to geologists. Kelvin's arguments seemed to be irrefutable, and yet they were inconsistent with the evidence as interpreted by geologists on the basis of Uniformitarianism. Ironically, one year before Lord Kelvin presented his famous lecture, the French physicist Henri Becquerel (1896) had announced the discovery of radioactivity. Only a few years later it was recognized that the disintegration of radioactive elements is an exothermic process. Therefore, the natural radioactivity of rocks produces heat, and the Earth is not merely a cooling body, as Lord Kelvin had assumed in his calculation.

Becquerel's discoveries attracted the attention of several young scientists, among them Marie (Manya) Sklodowska who came to Paris in 1891 from her native Poland to study at the Sorbonne. On July 25, 1895, she married Pierre Curie, a physics professor at the Sorbonne. After Becquerel reported his discoveries regarding salts of uranium, Marie Curie decided to devote her doctoral dissertation to a systematic search to determine whether other elements and their compounds emit similar radiation (Curie, 1898). Her work was rewarded when she discovered that thorium is also an active emitter of penetrating radiation. Turning to natural uranium and thorium minerals, she noticed that these materials are far more active than the pure salts of these elements. This important observation suggested to her that natural uranium ore, such as pitchblende, should contain more powerful emitters of radiation than uranium. For this reason, Marie and Pierre Curie requested a quantity of uranium ore from the mines of Joachimsthal in Czechoslovakia and, in 1898, began a systematic effort to find the powerful emitter whose presence she had postulated. The search eventually led to the discovery of two new active elements, which they named polonium and radium. Marie Curie coined the word "radioactivity" on the basis of the emissions of radium. In 1903, the Curies shared the Nobel Prize for physics with Henri Becquerel for the discovery of radioactivity.

1.2 INTERNAL STRUCTURE OF ATOMS

Every atom contains a small, positively charged nucleus in which most of its mass is concentrated. The nucleus is surrounded by a cloud of electrons that are in motion around it. In a neutral atom, the negative charges of the electrons exactly balance the total positive charge of the nucleus. The diameters of atoms are of the order of 10^{-8} centimeters (cm) and are conveniently expressed in angstrom units ($1\ \text{Å} = 10^{-8}$ cm). The nuclei of atoms are about 10,000 times smaller than that and have diameters of 10^{-12} cm, or 10^{-4} Å. The density of nuclear matter is about 100 million tons per cubic centimeter. The nucleus contains a large number of different elementary particles that interact with each other and are organized into complex patterns within the nucleus. It will suffice for the time being to introduce only two of these, the proton (p) and the neutron (n), which are collectively referred to as nucleons. Protons and neutrons can be regarded as the main building blocks of the nucleus because they account for its mass and electrical charge. Briefly stated, a proton is a particle having a positive charge that is equal in magnitude but opposite in polarity to the charge of an electron. Neutrons have a slightly larger mass than protons and carry no electrical charge. Extranuclear neutrons are unstable and decay spontaneously to form protons and electrons with a "halflife" of 10.6 min. The other principal components of atoms are the electrons, which swarm around the nucleus. Electrons at rest have a small mass (1/1836.1 that of hydrogen atoms) and a negative electrical charge. The number of extranuclear electrons in a neutral atom is equal to the number of protons. The protons in the nucleus of an atom therefore determine how many electrons that atom can have when it is electrically neutral. The number of electrons and their distribution about the nucleus in turn determine the chemical properties of that atom.

1.2a Nuclear Systematics

The composition of atoms is described by specifying the number of protons and neutrons that are present in the nucleus. The number of protons (Z) is called the *atomic number* and the number of neutrons (N) is the *neutron number*. The atomic number Z also indicates the number of extranuclear electrons in a neutral atom. The sum of protons and neutrons in the nucleus of an atom is the *mass number* (A). The composition of the nucleus of an atom is represented by the simple relationship

$$A = Z + N \tag{1.1}$$

Another word for atom that is widely used is *nuclide*. The composition of any nuclide can be represented by means of a shorthand notation consisting of the chemical symbol of the element, the mass number written as a superscript, and the atomic number written as a subscript. For example, ${}^{14}_{6}\mathrm{C}$ identifies the nuclide as an atom of carbon having 6 protons (therefore 6 electrons in a neutral atom) and a total of 14 nucleons. Equation 1.1 indicates that the nucleus of this nuclide contains $14 - 6 = 8$ neutrons. Similarly, ${}^{23}_{11}\mathrm{Na}$ is a sodium atom having 11 protons and $23 - 11 = 12$ neutrons. Actually, it is redundant to specify Z when the chemical symbol is used. For this reason, the subscript (Z) is sometimes omitted in informal usage.

A great deal of information about nuclides can be shown on a diagram in which each nuclide is represented by a square in coordinates Z and N. Figure 1.1 is a part of such a chart of the nuclides. Each element on this chart is represented by several nuclides having different neutron numbers arranged in a horizontal row. Atoms which have the same Z but different values of N are called *isotopes*. The isotopes of an element have identical chemical properties and differ only in their masses. Nuclides that occupy vertical columns on the chart of the nuclides have the same value of N but different values of Z and are called *isotones*. Isotones are therefore atoms of different elements. The chart also contains nuclides that occupy diagonal rows. These have the same value of A and are called *isobars*. Isobars have different values of Z and N and are therefore atoms of different elements. However, because they contain the same number of nucleons, they have similar but not identical masses.

1.2b Atomic Weights of Elements

The masses of atoms are too small to be conveniently expressed in grams. For this reason, the *atomic mass unit* (amu) is defined as one-twelfth of the mass of ${}^{12}_{6}\mathrm{C}$. In other words, the mass of ${}^{12}_{6}\mathrm{C}$ is arbitrarily fixed at 12.00... amu, and the masses of all other nuclides and subatomic particles are expressed by comparison to that of ${}^{12}_{6}\mathrm{C}$. The masses of the isotopes of the elements have been measured by mass spectrometry and are known with great precision and accuracy.

The total number of different nuclides is close to 2500, but only 270 of these are stable, including long-lived radioactive isotopes that still occur naturally because of their slow rate of decay. The stable nuclides, along with a small number of naturally occurring long-lived unstable nuclides, make up the elements in the periodic table. Many elements have two or more naturally occurring isotopes, some have only one, and two elements (technetium and promethium) have none. These

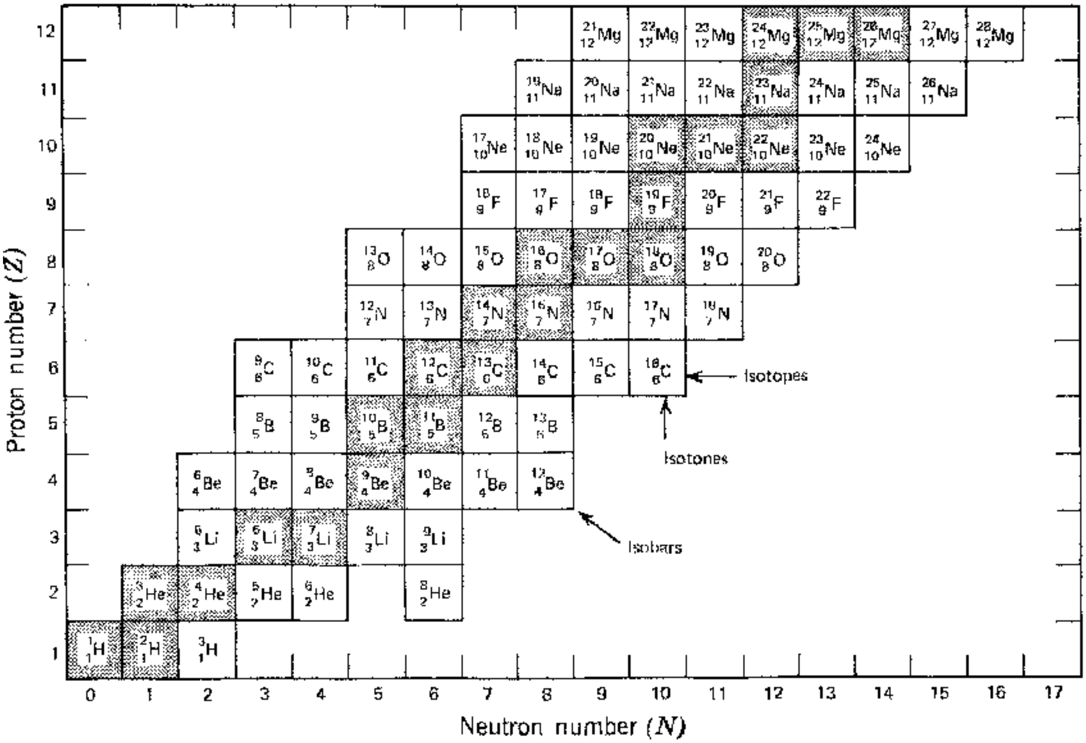

FIGURE 1.1 Partial chart of the nuclides. Each square represents a particular nuclide defined in terms of number of protons (Z) and neutrons (N) that make up its nucleus. The shaded squares represent stable atoms, whereas the white squares are the unstable or radioactive nuclides. Isotopes are atoms having the same Z but different values of N. Isotones have the same N but different values of Z. Isobars have the same A but different values of Z and N. Isotopes are atoms of the same element and therefore have identical chemical properties.

two elements therefore do not occur naturally on the Earth. However, they have been identified in the optical spectra of certain stars where they are synthesized by nuclear reactions.

The relative proportions of the naturally occurring isotopes of an element are expressed in terms of percent by number. For example, the statement that the isotopic abundance of $^{85}_{37}\mathrm{Rb}$ is 72.15 percent means that in a sample of 10,000 Rb atoms 7215 are the isotope $^{85}_{37}\mathrm{Rb}$. When the masses of the naturally occurring isotopes of an element and their abundances are known, the atomic weight of that element can be calculated. The atomic weight of an element is the sum of the masses of its naturally occurring isotopes weighted in accordance with the abundance of each isotope expressed as a decimal fraction. For example, the atomic weight of chlorine (Cl) is calculated from the masses and abundances of its two naturally occurring isotopes:

Isotope	Mass × Abundance
$^{35}_{17}\mathrm{Cl}$	$34.96885 \times 0.7577 = 26.4958$
$^{37}_{17}\mathrm{Cl}$	$36.96590 \times 0.2423 = \underline{\ 8.9568}$
	Atomic weight = 35.4526 amu

The abundances of the naturally occurring isotopes of the elements and their measured masses are listed in tables such as those of the *Handbook of Chemistry and Physics* (Lide and Frederikse, 1995).

Although the atomic weights of the elements are expressed in atomic mass units, it is convenient to define the *gram atomic weight*, or *mole*, which is the atomic weight of an element in grams. One mole of an atom or a compound contains a fixed number of atoms or molecules, respectively. The number of atoms or molecules in one mole is given by Avogadro's number, which is equal to 6.022045×10^{23} atoms or molecules per mole.

1.2c Binding Energy of Nucleus

The definition of the atomic mass unit provides an opportunity to calculate the mass of a particular nuclide by adding the masses of protons + electrons ($M_{\mathrm{H}} = 1.00782503$ amu) and of the neutrons ($M_{\mathrm{n}} = 1.00866491$ amu) of which it is composed. These calculated masses are consistently greater than the measured masses. It appears, therefore, that the mass of an atom is less than the sum of its parts. This phenomenon is an important clue to an understanding of the nature of the atomic nucleus. The explanation of the observed *mass defect* is that some of the mass of the nuclear particles is converted into *binding energy* that holds the nucleus together. The binding energy (E_B) is calculated by means of Einstein's equation:

$$E_B = \Delta m\, c^2 \tag{1.2}$$

where Δm is the mass defect and c is the speed of light in a vacuum ($2.99792458 \times 10^{10}$ cm/s).

The calculation of the binding energy requires a review of the relationship between units of mass and energy. The basic unit of energy in the cgs system (centimeter, gram, second) is the erg. However, the amount of energy released by a nuclear reaction involving a single atom is only a small fraction of one erg. For this reason, the *electron volt* (eV) is defined as the energy acquired by any charged particle carrying a unit electronic charge when it is acted upon by a potential difference of one volt. One electron volt is equivalent to 1.60210×10^{-12} erg. It is convenient to define two additional units, the kiloelectron volt (keV) and the million electron volt (MeV), where 1 keV = 10^3 eV and 1 MeV = 10^6 eV.

The conversion of the amu into grams follows from the definition of the atomic mass unit:

$$1 \text{ amu} = \frac{1}{12} \times \frac{12.000}{A} = \frac{1}{A}\text{gram}$$

where A is Avogadro's number. The amount of energy equivalent to 1 amu is obtained from Equation 1.2 by substituting $1/A$ for the mass of 1 amu in grams and by converting ergs to million electron volts by means of the appropriate conversion factor given above:

$$E = \frac{(2.99792458 \times 10^{10})^2}{6.022045 \times 10^{23} \times 1.60210 \times 10^{-12} \times 10^6}$$
$$= 931.5 \text{ MeV/amu}$$

The result of this calculation indicates that 1 amu of mass is equivalent to 931.5 MeV of energy. The binding energy of a nucleus (E_B) in million electron volts can now be calculated from the mass defect (Δm) in amu by means of the equation

$$E_B = 931.5\ \Delta m \tag{1.3}$$

For example, the theoretical mass of ${}^{27}_{13}\mathrm{Al}$ is $13 \times 1.00782503 + 14 \times 1.00866491 = 27.22303413$ amu. Its measured mass is only 26.981538 amu. Therefore, the mass defect and binding energy of ${}^{27}_{13}\mathrm{Al}$ are

$$\Delta m = 0.24149613 \text{ amu}$$
$$E_B = \frac{0.24149613 \times 931.5}{27} = 8.332 \text{ MeV/nucleon}$$

The binding energies per nucleon of the atoms of most elements have values ranging from about 7.5 to 8.8 MeV. The binding energy per nucleon rises slightly with increasing mass number and reaches a maximum value for ${}^{56}_{26}\mathrm{Fe}$. Thereafter, the binding energies decline slowly with increasing mass number. The binding energies of the atoms of H, He, Li, and Be are lower than the binding energies of the other elements.

1.2d Nuclear Stability and Abundance

It is reasonable to expect that a relationship exists between the stability of the nucleus of an atom and the abundance of that atom in nature. Conversely, one can use the observed abundances of the elements in the solar system and the abundances of their naturally occurring isotopes to derive information about the apparent stabilities of different kinds of atoms.

Given that only about 270 of nearly 2500 known nuclides are stable, nuclear stability is the exception rather than the rule. The point is illustrated in Figure 1.2, which is a schematic plot of the nuclides in coordinates of N and Z. In this diagram, the stable nuclides (shown in black) form a band flanked on both sides by unstable nuclides (shown in white). The existence of such a region of stability indicates that only those nuclei are stable in which Z and N are nearly equal. Actually, the ratio of N to Z increases from 1 to about 3 with increasing values of A. Only ${}^{1}_{1}\mathrm{H}$ and ${}^{3}_{2}\mathrm{He}$ have fewer neutrons than protons.

Another interesting observation is that most of the stable nuclides have even numbers of protons and neutrons. Stable nuclides with even Z and odd N or vice versa are much less common, whereas nuclides having odd Z and odd N are rare. These facts are shown in Table 1.1, where long-lived radioactive nuclides having halflives greater than 500×10^6 years are considered to be stable (Walker et al., 1989a). The relations in Table 1.1 suggest that the nucleons within the nucleus are arranged into regular patterns consisting of even numbers of protons and neutrons.

A careful study of this phenomenon combined with theoretical considerations has led to the concept of magic numbers for Z and N. Nuclides having magic proton numbers or magic neutron numbers or both are unusually stable, as indicated by their greater abundance or, in the case of unstable nuclides, by their slower decay rates. The magic numbers for Z and N are 2, 8, 10, 20, 28, 50, 82, and 126.

Calcium in Figure 1.3 is a good example of the effect of the magic number. It has a magic proton number (20) and six stable isotopes whose mass numbers are 40, 42, 43, 44, 46, and 48. All but one of these stable isotopes have even neutron numbers. The nucleus of ${}^{40}_{20}\mathrm{Ca}$ is doubly magic because it also has a magic number of neutrons ($N = 20$). Its isotopic abundance is 96.94 percent. The nucleus of ${}^{48}_{20}\mathrm{Ca}$ is also doubly magic ($N = 28$), but its abundance is only 0.19 percent, which is nevertheless remarkably high considering that it is some distance from the region of stability. Figure 1.3 also demonstrates that the halflives of the unstable (radioactive) isotopes of Ca increase from $A = 35$ to $A = 41$ and then decline from $A = 45$ to $A = 53$. The halflife of an unstable isotope is the time required for one-half of the atoms of that isotope to decay. In the context of Figure 1.3, the halflife is a measure of the relative stability of the unstable Ca isotopes: The longer the halflife, the more "stable" the nucleus. Therefore, Figure 1.3 illustrates the phenomenon that the stable isotopes of a given element are confined to a limited range of their mass numbers and that the radioactive isotopes at greater and smaller values of A become increasingly unstable (i.e., have progressively shorter halflives).

The chart of the nuclides (Figures 1.1 and 1.2) as well as the isotopes of Ca in Figure 1.3 reveal that most known nuclides are not stable but decompose spontaneously until they achieve a stable nuclear configuration. These are the so-called radioactive nuclides, or radionuclides. The spontaneous transformations that occur in their nuclei give rise to the phenomenon of *radioactivity*. Most of the known radionuclides do not occur naturally because their decay rates are rapid compared with the age of the solar system. However, they can be produced artificially by means of nuclear reactions in the laboratory. Radionuclides occur in nature for several reasons:

1. They have not yet completely decayed because their decay rates are very slow (${}^{238}_{92}\mathrm{U}$, ${}^{235}_{92}\mathrm{U}$, ${}^{232}_{90}\mathrm{Th}$, ${}^{87}_{37}\mathrm{Rb}$, ${}^{40}_{19}\mathrm{K}$, and others).
2. They are produced by the decay of long-lived, naturally occurring, radioactive parents (${}^{234}_{92}\mathrm{U}$, ${}^{230}_{90}\mathrm{Th}$, ${}^{226}_{88}\mathrm{Ra}$, and others).
3. They are produced by nuclear reactions occurring in nature (${}^{14}_{6}\mathrm{C}$, ${}^{10}_{4}\mathrm{Be}$, ${}^{32}_{14}\mathrm{Si}$, and others).

There is also a fourth group of radionuclides that can now be found in nature because they have been produced artificially, mainly as a result of the operation of nuclear fission reactors and by the testing of explosive fission and fusion devices. The dispersal of these radionuclides into the atmosphere

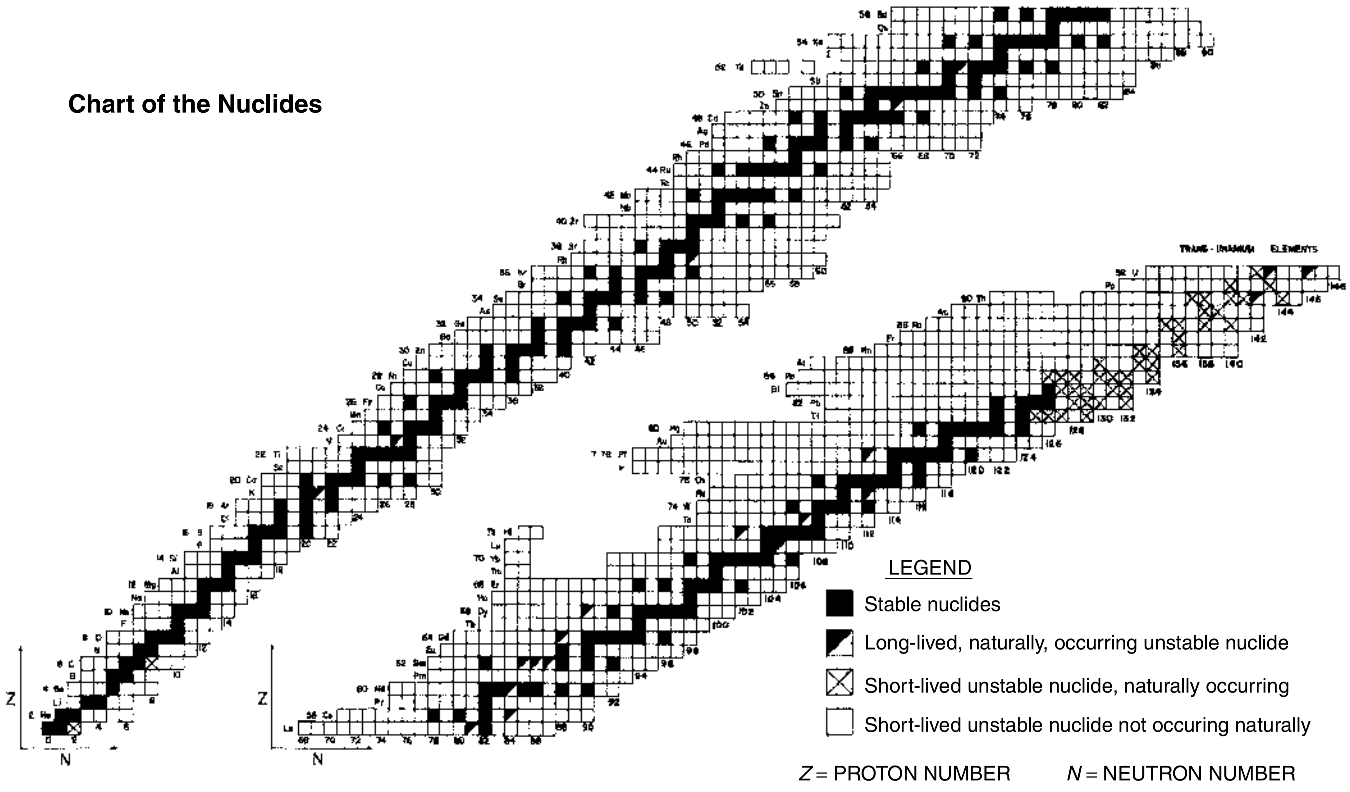

FIGURE 1.2 Chart of the nuclides in coordinates of Z and N. The diagram does not show all of the short-lived radioactive nuclides. After Walker et al. (1989b).

Table 1.1. Distribution of Stable Nuclides Depending on "Evenness" or "Oddness" of A, Z, and N

A	Z	N	Number of Stable Nuclides
Even	Even	Even	161
Odd	Even	Odd	55
Odd	Odd	Even	50
Even	Odd	Odd	4
Total number of stable nuclides			270

Source: Walker et al., 1989a.

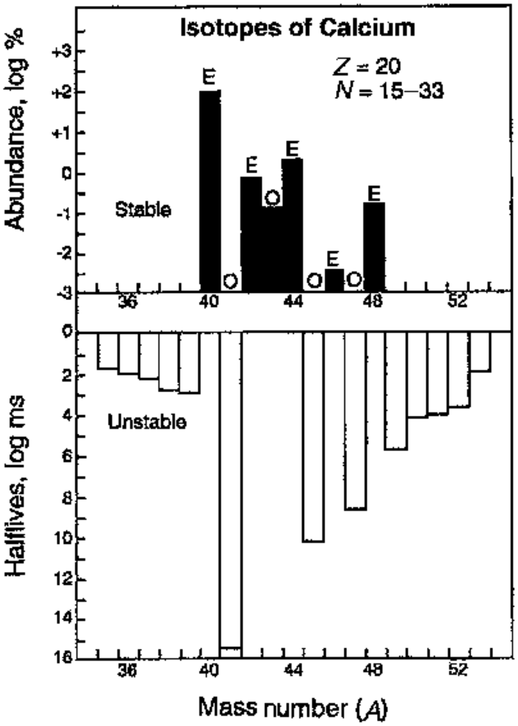

FIGURE 1.3 Stable and unstable (radioactive) isotopes of calcium. The abundances of the stable isotopes (black) are expressed as the logarithm to the base ten of the percent by number, whereas the relative stability of the unstable isotopes is represented by the logarithm to the base ten of their halflives expressed in milliseconds. The stable isotopes have even (E) mass numbers, except for $^{43}_{20}$Ca, whose abundance is less than that of its neighbors at $A = 42$ and $A = 44$. The gaps between the stable isotopes are occupied by unstable isotopes having odd (O) values of A. The halflives of the unstable isotopes decline with increasing "distance" from the central region of stability. Data from Walker et al. (1989b).

results in "fallout" and may cause contamination of food crops and drinking water. The safe disposal of radioactive waste products is an increasingly serious problem that will become more acute in the future as the quantity of radioactive "waste" increases (Winograd, 1981; Brookins, 1984; Faure, 1998).

The abundances of the chemical elements in the solar system given in Table 1.2 and Figure 1.4 are expressed in terms of the number of atoms of each element per 10^6 atoms of Si. They were compiled by Anders and Grevesse (1989) from spectroscopic analyses of sunlight and from chemical analyses of meteorites. The logarithm to the base 10 of the elemental abundances in Figure 1.4 forms a "sawtooth" pattern with increasing atomic number because the elements with even Z are more abundant than the adjacent elements that have odd values of Z. This statement is known as the Oddo–Harkins rule after the codiscoverers of this phenomenon.

The pattern of variation displayed in Figure 1.4 permits several additional observations concerning the abundances of the chemical elements in the solar system:

1. The elements H ($Z = 1$) and He ($Z = 2$) are by far the most abundant in the solar system.
2 The abundances of Li ($Z = 3$), Be ($Z = 4$), and B ($Z = 5$) are anomalously low.
3. The abundances of the elements whose atomic numbers are greater than 6 (C) decrease exponentially with increasing values of Z.
4. Iron ($Z = 26$) is exceptionally abundant, whereas the abundance of F ($Z = 9$) is less than expected.
5. Technetium (Tc) and promethium (Pm) do not occur naturally in the solar system because all of their isotopes are unstable and have short halflives.

Table 1.2. Abundance of Elements in the Solar System (atoms per 10^6 Si atoms)

Z	Element	Abundance
1	H	2.79×10^{10}
2	He	2.72×10^{9}
3	Li	5.71×10^{1}
4	Be	7.3×10^{-1}
5	B	2.12×10^{1}
6	C	1.01×10^{7}
7	N	3.13×10^{6}
8	O	2.38×10^{7}
9	F	8.43×10^{2}
10	Ne	3.44×10^{6}
11	Na	5.74×10^{4}
12	Mg	1.074×10^{6}
13	Al	8.49×10^{4}
14	Si	1.00×10^{6}
15	P	1.04×10^{4}
16	S	5.15×10^{5}
17	Cl	5.24×10^{3}
18	Ar	1.01×10^{5}
19	K	3.770×10^{3}
20	Ca	6.11×10^{4}
21	Sc	3.42×10^{1}
22	Ti	2.400×10^{3}
23	V	2.93×10^{2}
24	Cr	1.35×10^{4}
25	Mn	9.50×10^{3}
26	Fe	9.00×10^{5}
27	Co	2.250×10^{3}
28	Ni	4.93×10^{4}
29	Cu	5.22×10^{2}
30	Zn	1.260×10^{3}
31	Ga	3.78×10^{1}
32	Ge	1.19×10^{2}
33	As	6.56
34	Se	6.21×10^{1}
35	Br	1.18×10^{1}
36	Kr	4.5×10^{1}
37	Rb	7.09
38	Sr	2.35×10^{1}
39	Y	4.64
40	Zr	1.14×10^{1}
41	Nb	0.698×10^{-1}
42	Mo	2.55
43	Tc	—
44	Ru	1.86
45	Rh	3.44×10^{-1}
46	Pd	1.39
47	Ag	4.86×10^{-1}
48	Cd	1.61
49	In	1.84×10^{-1}
50	Sn	3.82
51	Sb	3.09×10^{-1}
52	Te	4.81
53	I	9.0×10^{-1}
54	Xe	4.7
55	Cs	3.72×10^{-1}
56	Ba	4.49
57	La	4.460×10^{-1}
58	Ce	1.136
59	Pr	1.669×10^{-1}
60	Nd	8.279×10^{-1}
61	Pm	—
62	Sm	2.582×10^{-1}
63	Eu	9.73×10^{-2}
64	Gd	3.30×10^{-1}
65	Tb	6.03×10^{-2}
66	Dy	3.42×10^{-1}
67	Ho	8.89×10^{-2}
68	Er	2.508×10^{-1}
69	Tm	3.78×10^{-2}
70	Yb	2.479×10^{-1}
71	Lu	3.67×10^{-2}
72	Hf	1.54×10^{-1}
73	Ta	2.07×10^{-2}
74	W	1.33×10^{-1}
75	Re	5.17×10^{-2}
76	Os	6.75×10^{-1}
77	Ir	6.1×10^{-1}
78	Pt	1.34
79	Au	1.87×10^{-1}
80	Hg	3.4×10^{-1}
81	Tl	1.84×10^{-1}
82	Pb	3.15
83	Bi	1.44×10^{-1}
90	Th	3.35×10^{-2}
92	U	9.00×10^{-3}

Source: Anders and Grevesse, 1989.

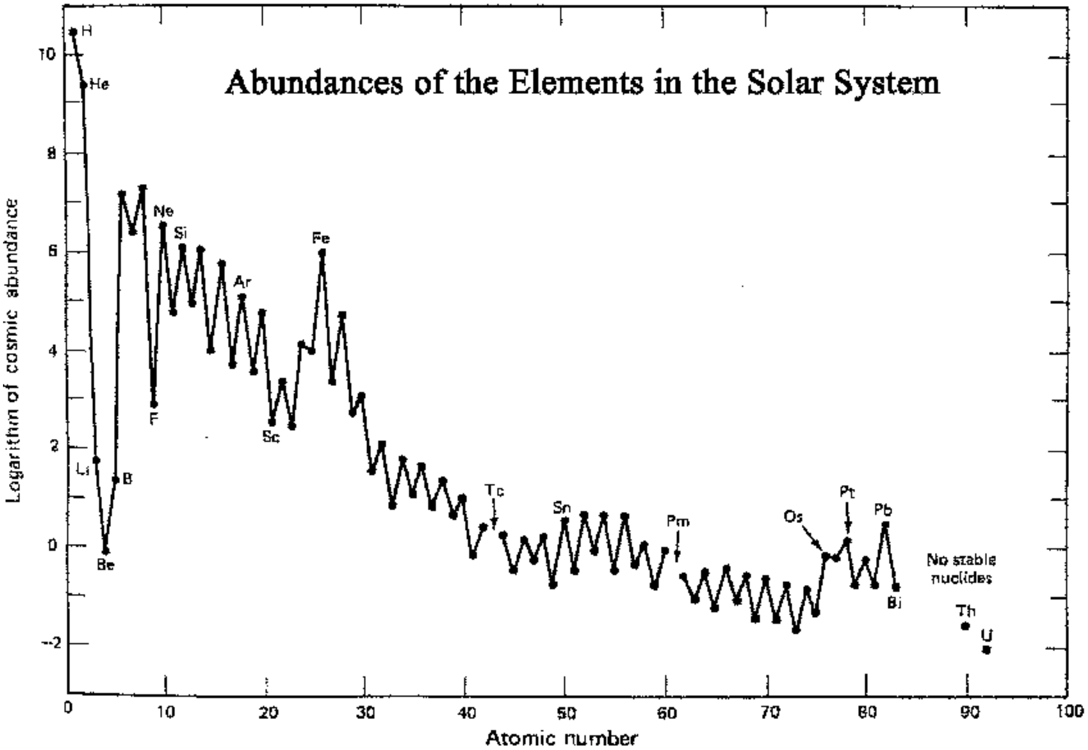

FIGURE 1.4 Abundances of the elements in the solar system versus their atomic number. The abundances are expressed as the logarithm to the base 10 of the number of atoms of each element relative to 10^6 atoms of silicon. Data are listed in Table 1.2. After Anders and Grevesse (1989).

6. The abundance of Pb ($Z = 82$) is somewhat greater than expected because its isotopes are the stable end products of the decay of the isotopes of U and Th.
7. Uranium has the lowest abundance of any element in the solar system, except as noted in 5 above.
8. The abundances of elements that lie between Bi ($Z = 83$) and Th ($Z = 90$) (e.g., Ra) and at $Z = 91$ are exceedingly low because these elements are the unstable daughters of the decay of the naturally occurring isotopes of U and Th.

The observed pattern of abundances of the chemical elements constrains the process of nucleosynthesis in the ancestral stars, which contributed matter to the solar nebula from which the Sun and the planets of the solar system formed. The nuclear reactions that occur in stars during their evolution and during their final explosions as supernovas produced all of the known nuclides (except ^1_1H), including both the stable and the unstable isotopes of the elements. Consequently, the Sun, the Earth, and all of the other planets and their satellites in the solar system are composed of the comparatively small number of stable and long-lived radioactive nuclides that have survived to the present time.

1.3 ORIGIN OF THE ELEMENTS

The nuclear reactions in the interiors of stars start at temperatures in excess of 10×10^6K with fusion of H nuclei (protons and deuterons) to form nuclei of He. The Sun has been generating the energy it radiates into space by the H fusion reaction, which means that it is producing only helium and deuterium. Therefore, all of the other elements in the Sun originated from ancestral stars whose terminal explosions formed the cloud of gas and dust in interstellar space from which the Sun and the planets of the solar system formed 4.6×10^9 years ago.

The fusion of H nuclei to form He takes place by the proton–proton chain, which includes ${}^2_1\text{H}$ and ${}^3_2\text{He}$ as intermediate products:

$$ {}^1_1\text{H} + {}^1_1\text{H} \rightarrow {}^2_1\text{H} + \beta^+ + \nu + 0.422\ \text{MeV} \quad (1.4)$$

$$ \beta^+ + \text{e}^- \rightarrow 1.02\ \text{MeV} \quad \text{(annihilation)} \quad (1.5)$$

$$ {}^2_1\text{H} + {}^1_1\text{H} \rightarrow {}^3_2\text{He} + \gamma + 5.493\ \text{MeV} \quad (1.6)$$

$$ {}^3_2\text{He} + {}^3_2\text{He} \rightarrow {}^4_2\text{He} + 2\ {}^1_1\text{H} + 12.859\ \text{MeV} \quad (1.7)$$

where β^+ is a positron (positively charged electron) and ν is a neutrino. The stable intermediate products ${}^2_1\text{H}$ and ${}^3_2\text{He}$ are consumed in reaction 1.6 and 1.7 to form the stable nucleus of ${}^4_2\text{He}$, which is identical to the alpha-particle (α-particle) named by Ernest Rutherford in 1899. The total amount of energy released by the synthesis of one helium nucleus is 19.794 MeV. Therefore, the formation of 1 mol of ${}^4_2\text{He}$ by the proton–proton chain releases an amount of heat (Q) equal to

$$\begin{aligned} Q &= 19.794 \times 6.022045 \times 10^{23} \\ &\quad \times 1.60210 \times 10^{-6} \times 10^{-7}\ \text{J} \\ &= 190.97 \times 10^{10}\ \text{J/mol} \end{aligned}$$

where 1 MeV $= 1.60210 \times 10^{-6}$ erg, 1 erg $= 10^7$ joule (J), and Avogadro's number $= 6.022045 \times 10^{23}$ atoms per mole.

Stars like the Sun that inherited ${}^{12}_6\text{C}$ and other nuclides from a previous cycle of stellar evolution and nucleosynthesis also synthesize ${}^4_2\text{He}$ by the so-called CNO cycle, in which ${}^{12}_6\text{C}$ acts as a catalyst (Bethe, 1939, 1968):

$$ {}^{12}_6\text{C} + {}^1_1\text{H} \rightarrow {}^{13}_7\text{N} + \gamma \quad (1.8)$$

$$ {}^{13}_7\text{N} \rightarrow {}^{13}_6\text{C} + \beta^+ + \nu \quad \text{(positron decay)} \quad (1.9)$$

$$ {}^{13}_6\text{C} + {}^1_1\text{H} \rightarrow {}^{14}_7\text{N} + \gamma \quad (1.10)$$

$$ {}^{14}_7\text{N} + {}^1_1\text{H} \rightarrow {}^{15}_8\text{O} + \gamma \quad (1.11)$$

$$ {}^{15}_8\text{O} \rightarrow {}^{15}_7\text{N} + \beta^+ + \nu \quad \text{(positron decay)} \quad (1.12)$$

$$ {}^{15}_7\text{N} + {}^1_1\text{H} \rightarrow {}^{12}_6\text{C} + {}^4_2\text{He} \quad (1.13)$$

The nucleus of ${}^{12}_6\text{C}$ is released and is available for another iteration of the CNO cycle. In this way, the CNO cycle converts four protons into one ${}^4_2\text{He}$ nucleus, thereby releasing the same amount of heat energy as the proton–proton chain.

The entire sequence of nuclear reactions by means of which stars generate energy at different stages in their evolution was revealed in a famous paper by Burbidge et al. (1957). A critical step in the nucleosynthesis process is the formation of Li, Be, and B, whose abundances in the solar system (Figure 1.4) are anomalously low by more than five orders of magnitude compared to C ($Z = 6$). The "gap" in the pattern of abundances of the elements is bridged by the so-called triple-α process:

$$ {}^4_2\text{He} + {}^4_2\text{He} \rightarrow {}^8_4\text{Be} \quad \text{(very unstable)} \quad (1.14)$$

$$ {}^8_4\text{Be} + {}^4_2\text{He} \rightarrow {}^{12}_6\text{C} \quad (1.15)$$

The nucleus of ${}^8_4\text{Be}$ has a halflife of only 7×10^{-17} s, which requires that it must capture an α-particle almost immediately after its formation (hence the name triple-α process). The low probability of this reaction is overcome by the high density and high temperature of He in the cores of stars in the red giant stage of their evolution. The triple-α process has a profound effect on the evolution of stars and makes possible the synthesis of all chemical elements having atomic numbers of 6 and higher.

The nucleosynthesis of the chemical elements in stars progresses from ${}^{12}_6\text{C}$ by means of fusion of nuclei having Z values of 6, 8, and 10 as well as by capture of α-particles, protons, and neutrons. Many nuclides having atomic numbers greater than 26 (Fe) form by neutron capture on two different timescales. In the s-process (slow timescale) the radioactive products of neutron capture decay to a stable isotope of another element before the next neutron capture occurs. In the r-process (rapid timescale), neutron capture occurs rapidly before the product nuclides have time to decay. The nuclear reactions by means of which the naturally occurring isotopes of the elements were synthesized in stars were identified by Anders and Grevesse (1989).

1.4 SUMMARY

The chemical elements were synthesized by nuclear reactions in stars starting with the fusion of four hydrogen nuclei to form helium. This is the process by means of which the Sun has been producing the energy it radiates into space since its formation from the solar nebula more than 4.6×10^9 years ago. The solar nebula originally contained about 2500 different nuclides, most of which were unstable and have since decayed to form stable products. Only 270 nuclides have survived to the present either because they are stable or because they decay very slowly.

The stable and unstable nuclides are conveniently inventoried in the chart of the nuclides in coordinates of atomic number (Z) and neutron number (N). The chart reveals that the stable isotopes of the elements occur in a band that extends diagonally across it with the unstable nuclides arranged on either side of it. The stable nuclides predominantly have even values of Z and N, whereas only a small number of stable nuclides have odd values of both Z and N. The even–odd criterion also applies to the abundances of the stable isotopes of a given element and is recognized by the Oddo–Harkins rule, which applies to the abundances of the stable isotopes as well as to the abundances of the chemical elements in the solar system.

The abundances and masses of the naturally occurring isotopes of the elements determine their atomic weights and hence their gram-atomic weights or moles. The transformation of moles to numbers of atoms or molecules is based on Avogadro's law, which plays an important role not only in physical chemistry but also in the derivation of the law of radioactivity.

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