Pearson

20.4: The Valley of Stability: Predicting the Type of Radioactivity

So far, we have described various different types of radioactivity. But what causes a particular nuclide to be radioactive in the first place? And why do some nuclides decay via alpha decay, while others decay via beta decay or positron emission? Nuclear properties—like all properties—depend on structure, but in this case the relevant structure is not that of an atom or molecule but of the nucleus. The particles that compose the nucleus—protons and neutrons—occupy energy levels that are similar to the energy levels occupied by electrons. A full examination of nuclear structure is beyond the scope of this text, but we can examine a couple of simple factors that influence the stability of the nucleus and the nature of its decay.

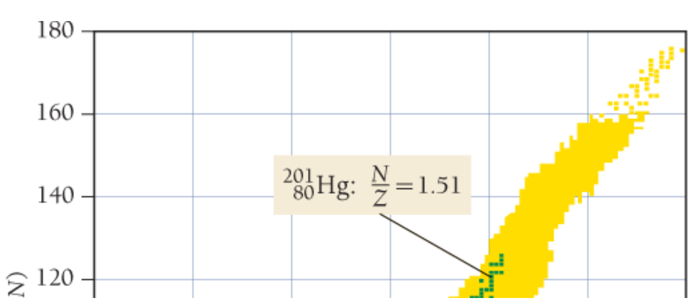
A nucleus is a collection of **nucleons**—protons (positively charged) and neutrons (uncharged). We know that positively charged particles such as protons repel one another. So what binds the nucleus together? A fundamental force of physics known as the **strong force** binds the nucleus together. All nucleons are attracted to one another by the strong force. However, the strong force acts only at very short distances. We can think of the stability of a nucleus as a balance between the *repulsive* coloumbic force among protons and the *attractive* strong force among all nucleons.

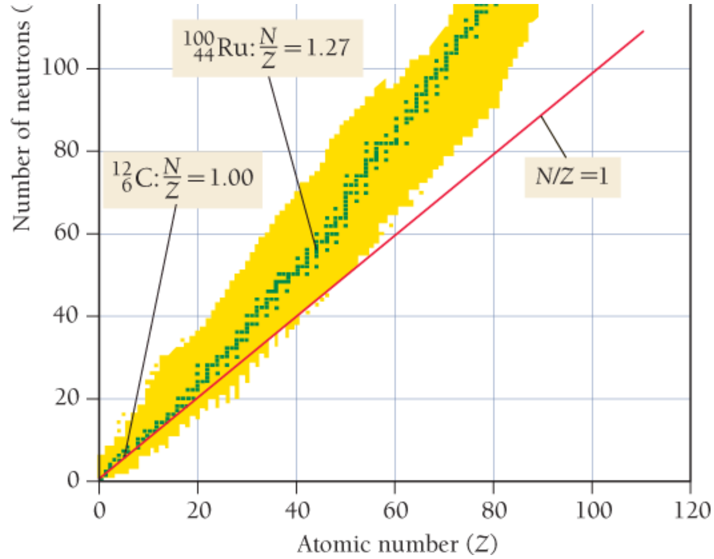
The neutrons in a nucleus, therefore, play an important role in stabilizing the nucleus because they attract other nucleons (through the strong force) but lack the repulsive force associated with positive charge. It might seem that adding more neutrons would *always* lead to greater stability, so that the more neutrons the better. This is not the case, however, because as we just discussed protons and neutrons occupy energy levels in a nucleus. As you add more neutrons, they must occupy increasingly higher-energy levels within the nucleus. At some point, the energy payback from the strong force is not enough to compensate for the high-energy state that the neutrons must occupy.

An important number in determining nuclear stability is the *ratio* of neutrons to protons (N/Z). Figure 20.5 shows a plot of the number of neutrons versus the number of protons for all known nuclei. The yellow dots represent unstable nuclei and the green dots represent stable nuclei. The region of the graph with the green dots (stable nuclei) is known as the *valley* (or *island*) *of stability*. Notice that for the lighter elements, the N/Z ratio of stable nuclei is about 1 (equal numbers of neutrons and protons). For example, the most abundant isotope of carbon ($Z = 6$) is carbon-12, which contains six protons and six neutrons. However, beyond about $Z = 20$, the N/Z ratio of stable nuclei begins to get larger. For example, at $Z = 44$, stable nuclei have an N/Z ratio of about 1.27 and at $Z = 80$, the N/Z ratio reaches about 1.5. Above $Z = 83$, stable nuclei do not exist—bismuth ($Z = 83$) is the heaviest element with stable (nonradioactive) isotopes.

Figure 20.5 Stable and Unstable Nuclei

A plot of N (the number of neutrons) versus Z (the number of protons) for all known nuclei. The green dots represent stable nuclei, and the yellow dots represent unstable nuclei. An unstable nucleus with an N/Z ratio that is too high tends to undergo beta decay. An unstable nucleus with an N/Z ratio that is too low tends to undergo positron emission or electron capture.





The type of radioactivity emitted by a nuclide depends in part on the N/Z ratio.

***N/Z* too high:** Nuclides that lie above the valley of stability have too many neutrons and tend to convert neutrons to protons via beta decay. The process of undergoing beta decay moves the nuclide down in the plot in Figure 20.5 and closer to (or into) the valley of stability.

***N/Z* too low:** Nuclides that lie below the valley of stability have too many protons and tend to convert protons to neutrons via positron emission or electron capture. This moves the nuclide up in the plot in Figure 20.5 and closer to (or into) the valley of stability.

One way to decide whether a particular nuclide has an N/Z that is too high, too low, or about right is to consult Figure 20.5. Those nuclides that lie within the valley of stability are stable. Alternatively, we can also compare the mass number of the nuclide to the atomic mass listed in the periodic table for the corresponding element. The atomic mass is an average of the masses of the most stable nuclides for an element (which is why they occur naturally) and thus represents an N/Z that is about right. For example, suppose we want to evaluate N/Z for Ru-112. Ruthenium has an atomic mass of 101.07, so we know that the nuclide with a mass number of 112 must contain too many neutrons and therefore its N/Z is too high. Example 20.3 demonstrates how to apply these considerations in predicting the mode of decay for a nucleus.

Example 20.3 Predicting the Type of Radioactive Decay

Predict whether each nuclide is more likely to decay via beta decay or positron emission.

a. Mg-28
b. Mg-22
c. Mo-102

SOLUTION

a. Magnesium-28 has 16 neutrons and 12 protons, so $N/Z = 1.33$. However, for $Z = 12$, you can see from Figure 20.5 that stable nuclei should have an N/Z of about 1. Alternatively, you can see from the periodic table that the atomic mass of magnesium is 24.31. Therefore, a nuclide with a mass number of 28 is too heavy to be stable because the N/Z ratio is too high, so Mg-28 undergoes *beta decay*, resulting in the conversion of a neutron to a proton.

b. Magnesium-22 has 10 neutrons and 12 protons, so $N/Z = 0.83$ (too low). Alternatively, you can see from the periodic table that the atomic mass of magnesium is 24.31. A nuclide with a mass number of 22 is too light; the N/Z ratio is too low. Therefore, Mg-22 undergoes *positron emission*, resulting in the is too light; the N/Z ratio is too low. Therefore, Mg-22 undergoes *positron emission*, resulting in the conversion of a proton to a neutron. (Electron capture would accomplish the same thing as positron emission, but in Mg-22, positron emission is the only decay mode observed.)

c. Molybdenum-102 has 60 neutrons and 42 protons, so $N/Z = 1.43$. For $Z = 42$, you can see from Figure 20.5 that stable nuclei should have an N/Z ratio of about 1.3. Alternatively, you can see from the periodic table that the atomic mass of molybdenum is 95.94. A nuclide with a mass number of 102 is too heavy to be stable; the N/Z ratio is too high. Therefore, Mo-102 undergoes *beta decay*, resulting in the conversion of a neutron to a proton.

FOR PRACTICE 20.3 Predict whether each nuclide is more likely to decay via beta decay or positron emission.

a. Pb-192
b. Pb-212
c. Xe-114

Magic Numbers

In addition to the N/Z ratio, the *actual number* of protons and neutrons also affects the stability of the nucleus. Table 20.2 shows the number of nuclei with different possible combinations of even or odd nucleons. Notice that a large number of stable nuclides have both an even number of protons and an even number of neutrons. Only five stable nuclides have an odd number of protons and an odd number of neutrons.

Table 20.2 **Number of Stable Nuclides with Even and Odd Numbers of Nucleons**

Z	*N*	Number of Nuclides
Even	Even	157
Even	Odd	53
Odd	Even	50
Odd	Odd	5

The reason for this is related to how nucleons occupy energy levels within the nucleus. Just as atoms with certain numbers of electrons are uniquely stable (in particular, the number of electrons associated with the noble gases: 2, 10, 18, 36, 54, etc.), so nuclei with certain numbers of nucleons (N or $Z = 2, 8, 20, 28, 50, 82$, and $N = 126$), are uniquely stable. These numbers are often referred to as **magic numbers**. Nuclei containing a magic number of protons or neutrons are particularly stable. Note that the magic numbers are even; this accounts in part for the abundance of stable nuclides with even numbers of nucleons. Moreover, nucleons also have a tendency to pair together (much as electrons pair). This tendency and the resulting stability of paired nucleons also contribute to the abundance of stable nuclides with even numbers of nucleons.

Radioactive Decay Series

Atoms with $Z > 83$ are radioactive and decay in one or more steps involving primarily alpha and beta decay (with some gamma decay to carry away excess energy). For example, uranium (atomic number 92) is the heaviest naturally occurring element. Its most common isotope is U-238, an alpha emitter that decays to Th-234.

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

The daughter nuclide, Th-234, is itself radioactive—it is a beta emitter that decays to Pa-234.

$$^{234}_{90}\mathrm{Th} \rightarrow {}^{234}_{91}\mathrm{Pa} + {}^{0}_{-1}\mathrm{e}$$

Protactinium-234 is also radioactive, decaying to U-234 via beta emission. Radioactive decay continues until a stable nuclide, Pb-206, is reached. Figure 20.6 illustrates the entire uranium-238 decay series.

Figure 20.6 The Uranium-238 Radioactive Decay Series

Uranium-238 decays via a series of steps ending in Pb-206, a stable element. Each diagonal line to the left represents an alpha decay, and each diagonal line to the right represents a beta decay.