Atomic number, atomic mass, and isotopes

Fundamental properties of atoms including atomic number and atomic mass. The atomic number is the number of protons in an atom, and isotopes have the same atomic number but differ in the number of neutrons.

Introduction

Radioactivity pops up fairly often in the news. For instance, you might have read about it in discussions of nuclear energy, the Fukushima reactor tragedy, or the development of nuclear weapons. It also shows up in popular culture: many superheroes' origin stories involve radiation exposure, for instance—or, in the case of Spider-Man, a bite from a radioactive spider. But what exactly does it mean for something to be radioactive?

Radioactivity is actually a property of an atom. Radioactive atoms have unstable nuclei, and they will eventually release subatomic particles to become more stable, giving off energy—radiation—in the process. Often, elements come in both radioactive and nonradioactive versions that differ in the number of neutrons they contain. These different versions of elements are called isotopes, and small quantities of radioactive isotopes often occur in nature. For instance, a small amount of carbon exists in the atmosphere as radioactive carbon-14, and the amount of carbon-14 found in fossils allows paleontologists to determine their age.

In this article, we'll look in more detail at the subatomic particles that different atoms contain as well as what makes an isotope radioactive.

Atomic number, atomic mass, and relative atomic mass

Atoms of each element contain a characteristic number of protons. In fact, the number of protons determines what atom we are looking at (e.g., all atoms with six protons are carbon atoms); the number of protons in an atom is called the **atomic number**. In contrast, the number of neutrons for a given element can vary. Forms of the same atom that differ only in their number of neutrons are called **isotopes**. Together, the number of protons and the number of neutrons determine an element's **mass number**: mass number = protons + neutrons. If you want to calculate how many neutrons an atom has, you can simply subtract the number of protons, or atomic number, from the mass number.

A property closely related to an atom's mass number is its **atomic mass**. The atomic mass of a single atom is simply its total mass and is typically expressed in atomic mass units or amu. By definition, an atom of carbon with six neutrons, carbon-12, has an atomic mass of 12 amu. Other atoms don't generally have round-number atomic masses for reasons that are a little beyond the scope of this article. In general, though, an atom's atomic mass will be very close to its mass number, but will have some deviation in the decimal places.

Since an element's isotopes have different atomic masses, scientists may also determine the **relative atomic mass**—sometimes called the **atomic weight**—for an element. The relative atomic mass is an average of the atomic masses of all the different isotopes in a sample, with each isotope's contribution to the average determined by how big a fraction of the sample it makes up. The relative atomic masses given in periodic table entries—like the one for hydrogen, below—are calculated for all the naturally occurring isotopes of each element, weighted by the abundance of those isotopes on earth. Extraterrestrial objects, like asteroids or meteors, might have very different isotope abundances.

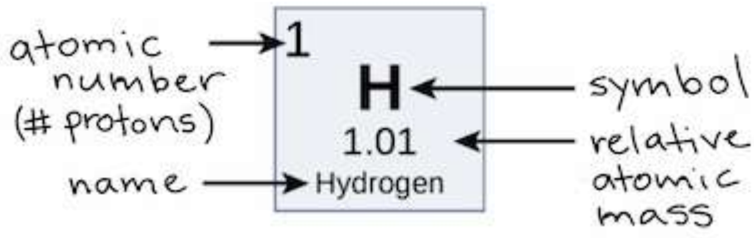


Image showing the "anatomy" of a periodic table entry. At the upper left is the atomic number, or number of protons. In the middle is the letter symbol for the element (e.g., H). Below is the relative atomic mass, as calculated for the isotopes found naturally on Earth. At the very bottom is the name of the element (e.g., hydrogen).

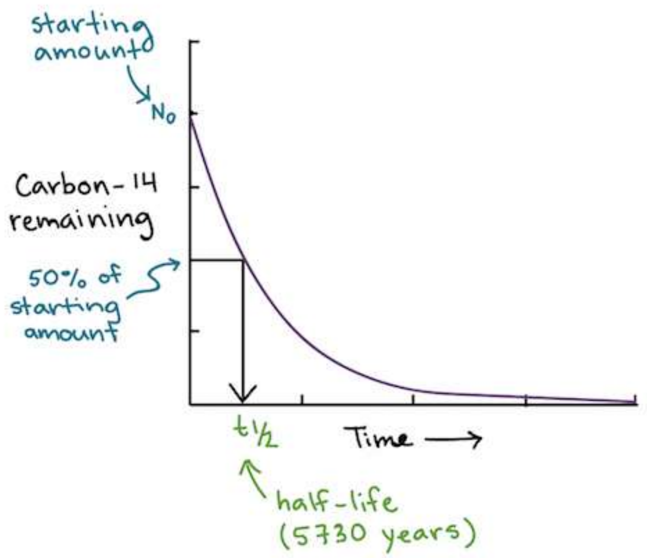
Image credit: modified from OpenStax CNX Biology

Isotopes and radioactive decay

As mentioned above, isotopes are different forms of an element that have the same number of protons but different numbers of neutrons. Many elements—such as carbon, potassium, and uranium—have multiple naturally occurring isotopes. Carbon-12 contains six protons, six neutrons, and six electrons; therefore, it has a mass number of 12 (six protons plus six neutrons). Carbon-14 contains six protons, eight neutrons, and six electrons; its mass number is 14 (six protons plus eight neutrons). These two alternate forms of carbon are isotopes.

Some isotopes are stable, but others can emit, or kick out, subatomic particles to reach a more stable, lower-energy, configuration. Such isotopes are called **radioisotopes**, and the process in which they release particles and energy is known as **decay**. Radioactive decay can cause a change in the number of protons in the nucleus; when this happens, the identity of the atom changes (e.g., carbon-14 decaying to nitrogen-14).

Radioactive decay is a random but exponential process, and an isotope's **half-life** is the period over which half of the material will decay to a different, relatively stable product. The ratio of the original isotope to its decay product and to stable isotopes changes in a predictable way; this predictability allows the relative abundance of the isotope to be used as a clock that measures the time from the incorporation of the isotope (e.g., into a fossil) to the present.



Graph of radioactive decay of carbon-14. The amount of carbon-14 decreases exponentially with time. The time at which half of the original carbon-14 has decayed—and half still remains—is designated as t 1/2. This time is also known as the half-life of the radioisotope and, for carbon-14, is equal to 5730 years.

Image credit: modified from CK-12 Biology

For example, carbon is normally present in the atmosphere in the form of gases like carbon dioxide, and it exists in three isotopic forms: carbon-12 and carbon-13, which are stable, and carbon-14, which is radioactive. These forms of carbon are found in the atmosphere in relatively constant proportions, with carbon-12 as the major form at about 99%, carbon-13 as a minor form at about 1%, and carbon-14 present only in tiny amounts1^11start superscript, 1, end superscript. As plants pull carbon dioxide from the air to make sugars, the relative amount of carbon-14 in their tissues will be equal to the concentration of carbon-14 in the atmosphere. As animals eat the plants, or eat other animals that ate plants, the concentrations of carbon-14 in their bodies will also match the atmospheric concentration. When an organism dies, it stops taking in carbon-14, so the ratio of carbon-14 to carbon-12 in its remains, such as fossilized bones, will decline as carbon-14 decays gradually to nitrogen-14[2].

After a half-life of approximately 5,730 years, half of the carbon-14 that was initially present will have been converted to nitrogen-14. This property can be used to date formerly living objects such as old bones or wood. By comparing the ratio of carbon-14 to carbon-12 concentrations in an object to the same ratio in the atmosphere, equivalent to the starting concentration for the object, the fraction of the isotope that has not yet decayed can be determined. On the basis of this fraction, the age of the material can be calculated with accuracy if it is not much older than about 50,000 years. Other elements have isotopes with different half lives, and can thus be used to measure age on different timescales. For example, potassium-40 has a half-life of 1.25 billion years, and uranium-235 has a half-life of about 700 million years and has been used to measure the age of moon rocks[2].