Although all atoms of an element have the same number of protons, individual atoms may have different numbers of neutrons. These differing atoms are called isotopes. All atoms of chlorine (Cl) have 17 protons, but there are chlorine isotopes with 15 to 23 neutrons. Only two chlorine isotopes exist in significant amounts in nature: those with 18 neutrons (75.53% of all chlorine atoms found in nature), and those with 20 neutrons (24.47%). To write the symbol for an isotope, place the atomic number as a subscript and the mass number (protons plus neutrons) as a superscript to the left of the atomic symbol. The symbols for the two naturally occurring isotopes of chlorine are written as follows: \(^{35}_{17}\text{Cl}\) and \(^{37}_{17}\text{Cl}\). The subscript is somewhat unnecessary, because all atoms of chlorine have 17 protons; isotope symbols are usually written without the subscript, as in \(^{35}\text{Cl}\) and \(^{37}\text{Cl}\). In discussing these isotopes, the terms chlorine-35 and chlorine-37 are used to differentiate between them. In general, for an atom to be stable, it must have more neutrons than protons. Nuclei with too many of either kind of fundamental particle are unstable, and break down radioactively.

Example:\(\PageIndex{1}\):

How many protons, neutrons, and electrons are there in an atom of uranium-238? Write the symbol for this isotope.

**SOLUTION**

The atomic number of uranium (see periodic table) is 92, and the mass number of the isotope is given as 238. Therefore, it has 92 protons, 92 electrons, and 238 — 92 : 146 neutrons. Its symbol is \(^{238}_{92}\text{U}\) (or \(^{238}\text{U}\)).

The total mass of an atom is called its atomic weight, and is the approximate sum of the masses of its constituent protons, neutrons, and electrons. When protons, neutrons, and electrons combine to form an atom, some of their mass is converted to energy and is given off. (This is the source of energy in nuclear fusion reactions. Because the atom cannot be broken down into its fundamental particles unless the energy for the missing mass is supplied from outside it, this energy is called the **binding energy** of the nucleus.)

Example:\(\PageIndex{2}\):

Calculate the mass that is lost when an atom of carbon-12 is formed from protons, electrons, and neutrons.

**SOLUTION**

Because the atomic number of every carbon atom is 6, carbon-12 has 6 protons and therefore 6 electrons. To find the number of neutrons, we subtract the number of protons from the mass number: 12 — 6 = 6 neutrons. The data in Table 1-1 can be used to calculate the total mass of these particles:

- **Protons:** \(6 \times 1.00728 \text{ amu} = 6.04368 \text{ u}\)
- **Neutrons:** \(6 \times 1.00867 \text{ amu} = 6.05202 \text{ u}\)
- **Electrons:** \(6 \times 0.00055 \text{ amu} = 0.00330 \text{ u}\)
- **Total particle mass:** \(12.09900 \text{ u}\)

However, by the definition of the scale of atomic mass units, the mass of one carbon-12 atom is exactly 12 amu. Therefore, 0.0990 u of mass has disappeared in the process of building the atom from its particles.
Each isotope of an element is characterized by an atomic number (the number of protons), a mass number (the total number of protons and neutrons), and an atomic weight (mass of atom in atomic mass units). Because mass losses upon formation of an atom are small, the mass number is usually the same as the atomic weight rounded to the nearest integer (for example, the atomic weight of chlorine-37 is 36.966, which is rounded to 37). If there are several isotopes of an element in nature, then the experimentally observed atomic weight (the natural atomic weight) is the weighted average of the isotope weights. The average is weighted according to the percent abundance of the isotopes. Chlorine occurs in nature as 75.53% chlorine-35 (34.97 u) and 24.47% chlorine-37 (36.97 u), so the weighted average of the isotope weights is \((0.7553 \times 34.97 \text{ u}) + (0.2447 \times 36.97 \text{ u}) = 35.46 \text{ u}\). The atomic weights found in periodic tables are all weighted averages of the isotopes occurring in nature, and these are the figures used for the remainder of this article, except when discussing one isotope specifically. In general, all isotopes of an element behave the same way chemically. Their behaviors differ with regard to mass-sensitive properties such as diffusion rates.

Example \(\PageIndex{3}\):

Magnesium (Mg) has three significant natural isotopes: 78.70% of all magnesium atoms have an atomic weight of 23.985 u, 10.13% have an atomic weight of 24.986 u, and 11.17% have an atomic weight of 25.983 u. How many protons and neutrons are present in each of these three isotopes? How are the symbols for each isotope written? Finally, what is the weighted average of the atomic weights?

**SOLUTION**

There are 12 protons in all magnesium isotopes. The isotope whose atomic weight is 23.985 u has a mass number of 24 (protons and neutrons), so 24 - 12 protons gives 12 neutrons. The symbol for this isotope is \(^{24}\text{Mg}\). Similarly, the isotope whose atomic weight is 24.986 amu has a mass number of 25, 13 neutrons, and \(^{25}\text{Mg}\) as a symbol. The third isotope (25.983 amu) has a mass number of 26, 14 neutrons, and \(^{26}\text{Mg}\) as a symbol. The average atomic weight is calculated as follows:

\[
(0.7870 \times 23.985) + (0.1013 \times 24.986) + (0.1117 \times 25.983) = 24.31 \text{ u}
\]

Example \(\PageIndex{4}\): Boron

Boron has two naturally occurring isotopes, \(^{10}\text{B}\) and \(^{11}\text{B}\). In nature, 80.22% of its atoms are \(^{11}\text{B}\), with atomic weight 11.009 u. From the natural atomic weight, calculate the atomic weight of the \(^{10}\text{B}\) isotope.

**SOLUTION**

If 80.22% of all boron atoms are \(^{11}\text{B}\), then 100.00 — 80.22, or 19.78%, are the unknown isotope. In the periodic table the atomic weight of boron is found to be 10.81 u. We can use \(W\) to represent the unknown atomic weight in our calculation:

\[
(0.8022 \times 11.009) + (0.1978 \times W) = 10.81 \text{ (natural atomic weight)}
\]

\[
W = \frac{10.81 - 8.831}{0.1978} = 10.01 \text{ u}
\]
Contributors

• Dickerson, Richard E. and Gray, Harry B. and Haight, Gilbert P (1979) *Chemical principles.*