Radioactive Chemistry

13

13.0 CHAPTER PREVIEW

In this chapter we will:

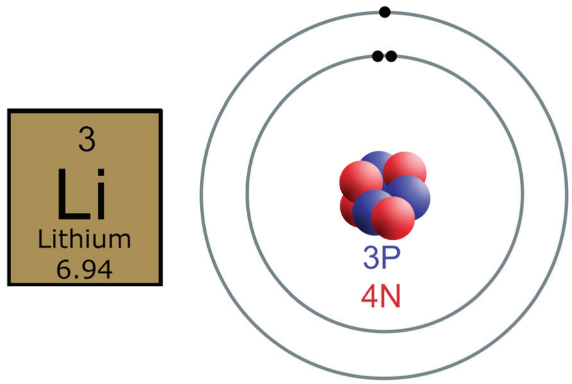
- Introduce new terms related to nuclear chemistry and radioactive elements.
- Learn what makes a nucleus unstable (radioactive).
- Study the nuclear decay processes of alpha, beta and gamma radiation.
- Understand half-life and how it works.
- Discuss a couple basic applications of nuclear chemistry in everyday life.

13.1 INTRODUCTION

Like a lot of concepts, sometimes it takes a little bit to explain something that is brand new, and this is true of our next topic—radioactivity. You have probably heard the term "radioactive," but may not know exactly what it means. In this chapter, we're going to dig into the basics of "radioactivity" with our study of **nuclear chemistry**, the study of radioactive elements. "**Radioactive**" simply means that an element's nucleus is unstable. We will learn what radioactivity is, why it occurs and some other cool stuff about **nuclear decay**, or what happens to the nucleus of an atom that is radioactive.

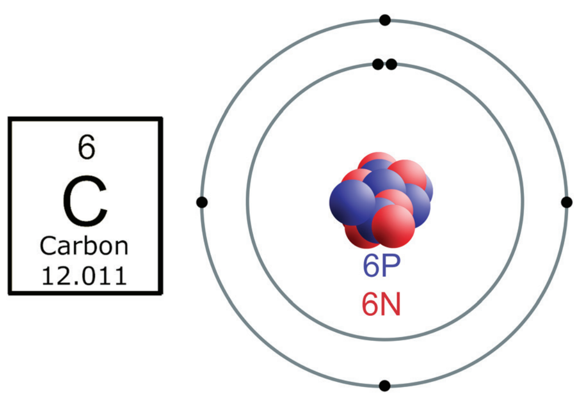
13.2 OUR FRIEND, THE PERIODIC TABLE AND ITS RELATIONSHIP TO RADIOACTIVITY

As usual, we need our old friend the Periodic Table to get us started. Specifically, let's briefly review what we have learned about the atom and the subatomic nuclear particles since that is where radioactivity occurs. So far, our study has focused on the atoms of "stable" elements, which means that the subatomic nuclear particles—the protons and neutrons—are there to stay, so to speak. Protons and neutrons in non-radioactive elements stay in the nucleus, which we can see using lithium as an example. Lithium has 3 protons and 4 neutrons in its nucleus, and those protons and neutrons aren't going anywhere. They stay right in the nucleus.



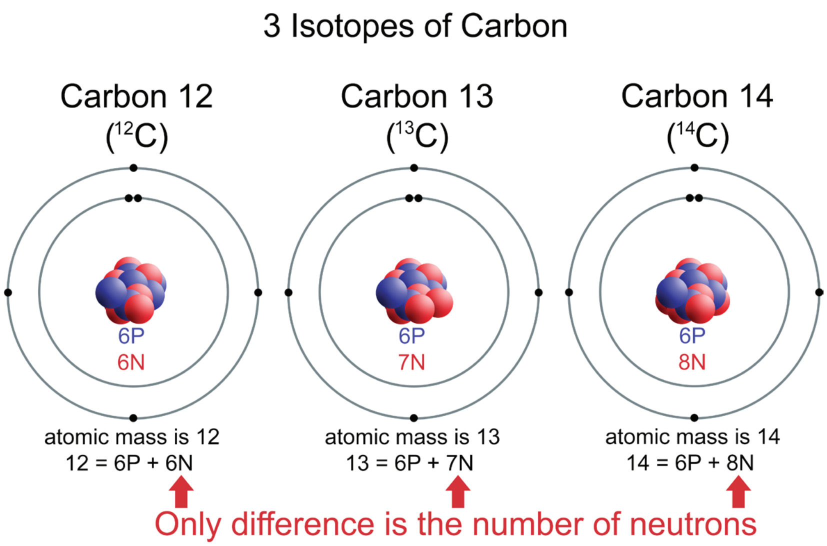
No matter how long we observe this nucleus and no matter how may lithium atoms we have, they will always have 3 protons and 4 neutrons because they have the right combination of protons and neutrons so they are stable. It may seem obvious, that the nuclear subatomic particles don't go anywhere, but for radioactive elements, being unstable means that protons and neutrons do go somewhere—they come out of the nucleus—or they change from one to the other! The process of a nucleus releasing subatomic particles, or of one changing into another, is "radioactivity." Therefore, we say that the nuclei of radioactive atoms are unstable, one moment it has a certain number of protons and neutrons, and the next moment, it doesn't.

We can begin to understand the process of radioactivity by considering the carbon atom:



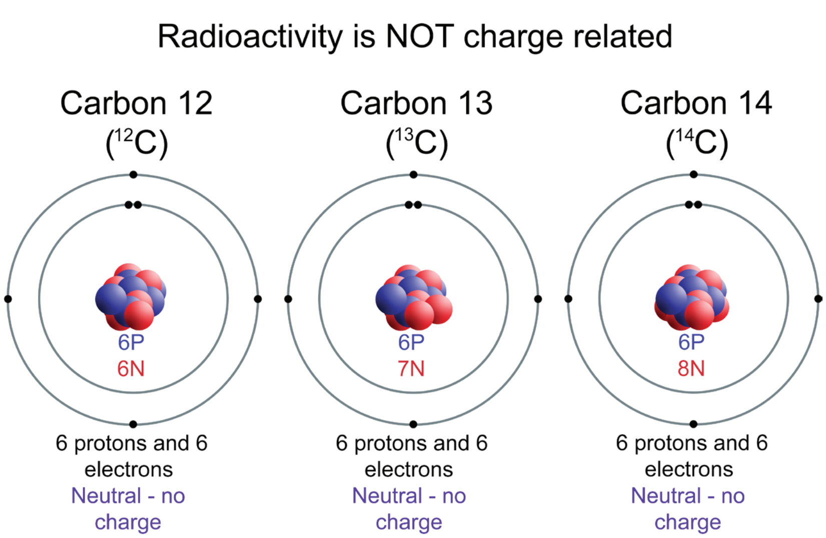
This is the most common form of carbon, which has 6 protons and 6 neutrons in its nucleus. A million carbon atoms that right now have 6 protons and 6 neutrons in their nuclei will still have 6 protons and 6 neutrons in their nuclei 57,300 years from now (we will see why I chose this seemingly strange number in a few pages) because carbon with an atomic mass of 12 is a stable atom with a stable nucleus.

But, recall way back to **Section 4.9** and **Figure 4.9.1** when we learned about isotopes and discussed how to calculate atomic mass. To refresh your memory, an isotope is two or more forms of the same element. Isotopes have the same number of protons, but different numbers of neutrons. Remember, an element is defined by number of protons in its nucleus. For example, here are 3 isotopes of carbon. All have 6 protons. Any element that has 6 protons is carbon. However, they have different numbers of neutrons:



These three common isotopes of carbon are known as **carbon 12**, **carbon 13**, and **carbon 14**. Those numbers come from the atomic masses. Carbon 12 has 6 neutrons and 6 protons, so its atomic mass is 12. Carbon 13 has 6 protons and 7 neutrons, so its atomic mass is 13 and carbon 14 has 6 protons and 8 neutrons for an atomic mass of 14. I will digress briefly to mention that there are two standard ways to denote isotopes. We can either write the element's symbol with its atomic mass after it—"C12" for carbon 12, "C13" for carbon 13, and "C14" for carbon 14—or we can write the atomic mass as a superscript before the atomic symbol—^{12}C for carbon 12, ^{13}C for carbon 13 and ^{14}C for carbon 14. I will use the "^{12}C" method.

It is a common misunderstanding that radioactivity occurs because of a difference in charge between the isotopes, but that is not the case. Charge on an atom is determined by protons and electrons and since the three isotopes of carbon have the same numbers of positively-charged protons and negatively charged electrons, they have no overall charge. The only difference between ^{12}C, ^{13}C and ^{14}C is in the number of neutrons, which have no charge; therefore, there is no charge difference between the three isotopes.



^{12}C is not radioactive and is the most common form of carbon found on earth; about 99% of all carbon is ^{12}C. ^{13}C is also stable and accounts for slightly under 1% of the carbon on earth. The number of protons and neutrons in the nuclei of ^{13}C and ^{12}C do not change over time; they are stable.

However, ^{14}C, which accounts for a tiny amount of carbon on earth, is not stable; it is radioactive. When an element is not stable, that means that the number of neutrons and/or protons changes over time. Therefore, if we had 1 million ^{14}C atoms and counted the protons and neutrons in their nuclei 10,000 years from now, we would find that there were no longer 1 million ^{14}C atoms. We would still have 1 million total carbon atoms, and some would still be ^{14}C, but they would not all be ^{14}C because their nuclei are unstable and that causes them to change over time. We will talk specifically about what happens to them in a bit, but for now, the important concept to understand is that radioactive nuclei are unstable and so their numbers of protons and neutrons change over time. I don't mean to beat a dead horse, but this is what it means to be unstable, or "radioactive."

There are a total of about 50 naturally occurring radioactive isotopes (and about a thousand man-made ones), and we will learn about some of them, why they are radioactive and what happens as a result of their radioactivity.

13.3 WHY ARE SOME ISOTOPES STABLE AND SOME UNSTABLE?

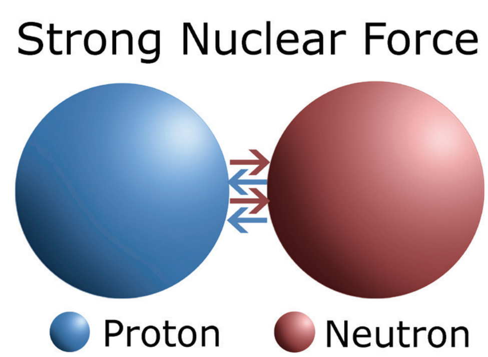
The reason for radioactivity takes us back to our discussion of the strong force. With all of those positively-charged protons packed into the nucleus, all sitting there repelling each other, the strong force is a non-charge-based force that develops between protons and neutrons to hold the nucleus together. The strong force is strong enough to counteract the repulsive power of those positive charges and keep the protons from flying out of the nucleus. The strong force does not make the repulsive force of the protons go away, though. It acts more like glue and holds the protons in the nucleus despite them repelling each other.

A stable nucleus is stable because it has the right "mix" of neutrons and protons to develop the necessary strong force to create a stable nucleus. If there are too many, or too few, neutrons, then the strong force will not be as strong as it needs to be to create a stable nucleus and the atom will be radioactive. That is to say there is a "sweet spot" in terms of the ratio of neutrons to protons.

Figure 13.3.1

STRONG NUCLEAR FORCE REVISITED

We first learned of the strong nuclear force in Chapter 3, but it is relevant here, too. The strong nuclear force, or "strong force" for short, is an attractive force between neutrons and protons. Remember that protons are positively charged (and particles with the same charge repel one another), so normally you'd expect all the positive charges packed into the nucleus to repel each other causing the nucleus to fly apart. However, since that doesn't happen, there must be some force that can overcome that repulsive power, and there is—the strong force. Further, since neutrons have no charge, the strong force is not based upon charge. So, even though there are many positive charges packed into the nucleus, all pushing each other away, and the neutrons have no charge at all, the strong force—the attraction that exists between neutrons and protons—is powerful enough to overcome the proton's positive electrical charges and keep the protons firmly in the nucleus.



The ratio is known as the **neutron-proton** ratio, abbreviated the N:P ratio, or just N:P. Only hydrogen, which has a N:P of 0 (no neutrons and 1 protons) has a N:P less than 1 and is stable. Otherwise, the minimum N:P ratio for a stable isotope is 1.0 (1 neutron for every 1 proton, 1:1). Isotopes that have a N:P ratio of less than 1.0 are unstable; therefore, radioactive. Once you get to a N:P ratio of 1.0 and higher, it is difficult to make a blanket statement of "all elements with a N:P ratio of such-and-such are stable," so we will assess the stability on a case by case basis, but the following general statements are true:

- For elements with atomic numbers 1–20, the optimal N:P ratio is 1.0 (1 neutron for every 1 proton).
- The elements with atomic numbers 21 and higher have a N:P of 1.14–1.5 to be stable.

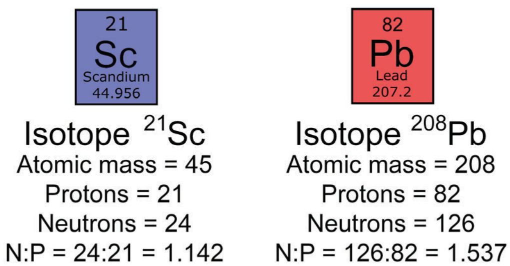


Figure 13.3.2

N:P The Highs and Lows

There are wide ranges of N:P ratios for an element to be stable. For the elements atomic number 21 and higher, the lowest N:P ratio for a stable element is for scandium (^{21}Sc) at 1.142 and the highest N:P ratio for a stable element is 1.537 for the lead isotope ^{208}Pb. By the way, ratios are commonly expressed as a single number by dividing the first number of the ratio by the second number. So, for scandium, 24 neutrons to 21 protons is a ratio of 24:21 and 24 ÷ 21 = 1.142, so we say that N:P is 1.142. For lead, the ratio is 126N:82P and 126 ÷ 82 = 1.537, so we say that neutron to proton ratio is 1.537.

Let's look a little more at that sweet spot concept for the N:P. When there are more than 20 protons packed into a nucleus, the optimal number of neutrons to generate an adequate strong force is more than a 1 to 1 ratio of neutrons to protons. However, there is also an upper limit to the "mismatch" ratio between the number of neutrons and protons, which is around 1.5-ish, to generate a stable nucleus. When the N:P gets over 1.537 for any element, all isotopes become unstable and are radioactive. Besides hydrogen, all stable isotopes have a N:P of 1.0–1.537; however, not all isotopes with a N:P between 1.0 and 1.537 are stable. In general, the N:P increases as the atomic number gets larger, with the max value of 1.537 representing the highest N:P ratio for a stable atom.

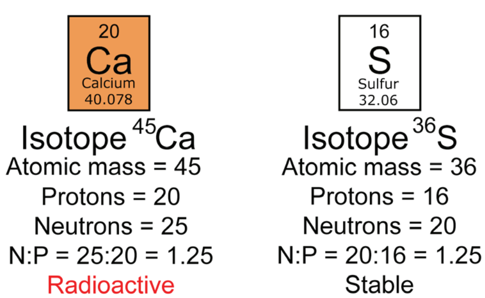


Figure 13.3.3

Not All Isotopes Respond the Same Way to the Same N:P Ratio

The relationship between the N:P and the stability of isotopes is highly dependent on the particular element such that you can't make the blanket statement, "All isotopes with a N:P of such-and-such-value are stable." In other words, you can't make a generalized statement that a certain isotope with a certain N:P is stable just because an isotope of another element is stable with that same N:P. For example, the ^{45}Ca isotope has a N:P of 1.25 and it is radioactive, but sulfur has an isotope, ^{36}S, which also has a N:P of 1.25 but it is stable. N:P ratios and the stability of isotopes do not generalize out across elements. The effects of different ratios of neutrons to protons on the stability of the nucleus are specific for each element's isotopes. You are in no way expected to be able to predict if an atom is stable, or not, based only upon its N:P, unless I ask you a N:P that we KNOW is unstable. For example, if I asked, "True or False: a fictional element, elementium, has a N:P of 0.872 and so it is unstable." I would expect that you would answer that "True" because all elements, except hydrogen, are unstable with N:P less than 1.