Nuclear Chemistry

The AP exam requires you to know about nuclear equations, half-lives, radioactivity, and chemical applications of nuclear properties. This chapter begins with a brief review of the history of the nucleus and how we came to know about it and then moves into the required topics.

ROAD MAP

- *Properties of the Nucleus*
- *Nuclear Equations*
- *Half-lives*
- *Radioactivity*
- *Applications of Nuclear Chemistry*

THE DISCOVERY OF RADIOACTIVITY

A simple, but working, definition of *radioactivity* is "the spontaneous decay of particles from the nucleus of an atom." Henri Becquerel first discovered radiation in 1896 while doing research on the fluorescence of different materials. One day, he set a sample of uranium ore in a drawer atop some unexposed photographic plates. Upon later developing the film, he discovered that the film had been exposed in the location where the ore had been sitting. He concluded that there must be some high-energy emissions emerging from the material. Becquerel did not wish to pursue this separate line of research, so he passed the work on to one of his graduate students, Marie Curie, and her husband, Pierre. The Curies painstakingly worked with large samples of the uranium ore to isolate the material responsible for the emissions (the material they isolated was radium).

RUTHERFORD DISCOVERS DIFFERENT TYPES OF RADIATION

The next major discoveries about radiation came from Ernest Rutherford. These experiments came before his famous "gold-foil" experiment. He discovered that an electrical field affected the emissions from radioactive material. By placing a sample of material near two charged plates (similar to the design of Thomson's cathode ray studies), he was able to observe the behavior of the radioactive emissions in an electric field. He accomplished this by firing a sample of radioactive material in a thin stream between two charged plates. He placed a photographic plate on the opposite side of the charged plates to observe the deflections of any particles.

From the deflections, he determined that some of the particles were positively charged, some were negatively charged, and some were not charged at all. He discovered that the positively charged particles were much more massive than the others, and when combined with electrons, they formed helium atoms. He

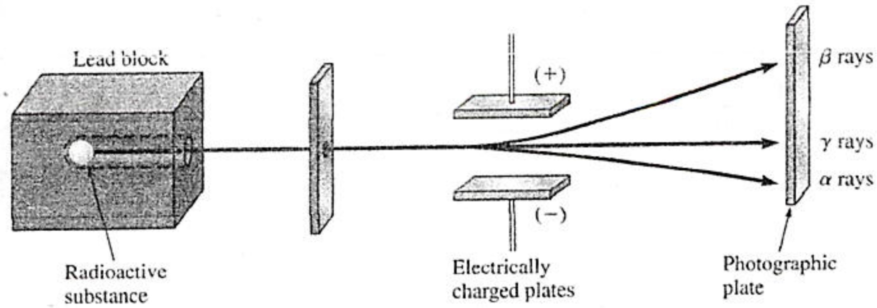

Figure 6.1

called these particles **alpha particles** (α- particles), and today we know them to be helium nuclei. The negatively charged particles were lighter and faster-moving than alpha particles and behaved like cathode rays. Rutherford called these **beta particles** (β-particles), and today we know them to be electrons. The third, neutral emission was determined to be extremely high-energy radiation, unaffected by an electrical field. These emissions became known as **gamma rays** (γ rays).

Alpha particles are relatively large and slow-moving and are easily stopped. Beta particles are much smaller, much faster, and about 100 times more penetrating than alpha particles. Gamma rays (often referred to as gamma photons) are massless and extremely fast and possess high energy. They are by far the most penetrating radiation—about 100 times greater than a beta particle, or 10,000 times more penetrating than an alpha particle.

The remainder of this chapter will be devoted to the process of radioactive decay.

DIFFERENT TYPES OF RADIOACTIVE EMISSIONS

There are five main types of emissions: alpha emission, beta emission, positron emission, electron capture, and gamma emission. Four of these produce changes in the elements undergoing decay, and the end result is a more stable atomic structure.

Alpha Emissions

These emissions result in the release of an alpha particle from the atom. Recall that an α-particle is a helium nucleus. The result in alpha decay is the atom's atomic number decreasing by two and the mass number decreasing by four. An example of an α-decay is

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

You Try It!

Fill in the missing isotope in the reaction that follows:

$$^{211}_{83}\mathrm{Bi} \rightarrow {}^{4}_{2}\mathrm{He} + \underline{\hspace{2cm}}$$

Answer: $^{207}_{81}\mathrm{Tl}$

Beta Emission

Although there are two types of β-particles (β+ and β-), the former is usually referred to as a **positron**, so we'll refer to only the β- particle as a beta particle. In a beta emission, a beta particle is ejected from the atom. A beta particle has all of the properties of an electron (virtually massless, negative charge), yet it is created by the conversion of a neutron in the nucleus to a proton and an electron (beta particle). The proton remains in the nucleus, and the beta particle is ejected from the nucleus. An example of a beta emission is

$$^{227}_{89}\mathrm{Ac} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{227}_{90}\mathrm{Th}$$

Notice that since the number of nucleons in the atom does not change, the mass number remains unchanged. However, the gain of a proton increases the atomic number by one (and consequently changes the element).

You Try It!

Fill in the missing isotope in the reaction that follows:

$$^{209}_{82}\mathrm{Pb} \rightarrow {}^{0}_{-1}\mathrm{e} + \underline{\hspace{2cm}}$$

Answer: $^{209}_{83}\mathrm{Bi}$

Positron Emission

Positron emissions are also known as β+ emissions. The positron is known as an **antiparticle**. Antiparticles are the exact opposites of particles. The positron is the antiparticle to an electron and is represented by the symbol $^{0}_{+1}\mathrm{e}$. The electron has virtually no mass and a charge of negative one (relative to a proton). A positron has virtually no mass and a charge of positive one. When an electron and its antiparticle, the positron, collide, they disintegrate and their matter is converted entirely into energy in the form of two gamma rays. In a positron emission, a proton in the nucleus is converted into a neutron and a positron. The neutron remains in the nucleus, and the positron is ejected. The life span of the positron is very brief since it will disintegrate upon collision with an electron. An example of a positron emission can be seen in the example showing the breakdown of carbon-11:

$$^{11}_{6}\mathrm{C} \rightarrow {}^{0}_{+1}\mathrm{e} + {}^{11}_{5}\mathrm{B}$$

Notice that the number of nucleons doesn't change here, either. As a result, the mass number of the atom does not change. However, the conversion of a proton to a neutron decreases the atomic number by one (and changes the element).

You Try It!

Fill in the missing isotope in the reaction that follows:

$$^{58}_{29}Cu \rightarrow {}^{0}_{+1}e + \_\_\_\_\_\_$$

Answer: $^{58}_{28}Ni$

Electron Capture

The fourth type of emission is called electron capture. In this process, an inner shell electron is pulled into the nucleus, and when this occurs, the electron combines with a proton to form a neutron. In electron capture reactions, the atomic number decreases by one, the mass number remains the same, and the element changes. One difference in this type of reaction is that the electron is written to the left of the arrow to show that it is consumed, rather than produced, in the process. An example of electron capture can be seen in the following reaction:

$$^{201}_{80}Hg + {}^{0}_{-1}e \rightarrow {}^{201}_{79}Au$$

You Try It!

Fill in the missing isotope in the reaction that follows:

$$^{68}_{32}Ge + {}^{0}_{-1}e \rightarrow \_\_\_\_\_\_$$

Answer: $^{68}_{31}Ga$

Gamma Radiation

The fifth type of radioactive emission, gamma radiation, does not result in a change in the properties of the atoms. As a result, they are usually omitted from nuclear equations. Gamma emissions often accompany other alpha or beta reactions—any decay that has an excess of energy that is released. For example, when a positron collides with an electron, two gamma rays are emitted, a phenomenon usually referred to as **annihilation radiation.**

RADIOACTIVE DECAY

There are a variety of reasons why radioactive decay occurs, but the primary reason is increased nuclear stability. To briefly review, attractive forces must overcome electrostatic repulsions between the like-charged protons. These attractive forces are provided by gluons (but you needn't concern yourself with knowing about gluons). Neutrons seem to play a role in this attractive process as well, both by attracting neighboring nucleons (including protons) and by "diluting" the electrostatic repulsions between the protons, by spreading them out a bit. Small atoms tend to have stable structures with equal numbers of neutrons and protons. As atoms get larger, the number of neutrons exceeds the number of protons. There are also two other properties that seem to determine stability of atomic nuclei. These are based on a model of nuclear stability known as the shell model of the nucleus. Just as there are stable electron configurations, there seem to be stable configurations of nucleons. These two properties fall into two categories:

1. **Magic numbers—There are certain numbers of protons and neutrons that are found in the most stable nuclei.** These are known as the magic numbers for protons and neutrons. For protons, the numbers are 2, 8, 20, 28, 50, and 82. For neutrons, the numbers are 2, 8, 20, 28, 50, 82, and 126.
2. **Nuclei with even numbers of protons AND neutrons are more stable than atoms with odd numbers of neutrons.** It is believed that perhaps proton-proton pairs and neutron-neutron pairs form stable relationships, much like electron-electron pairs in molecules. Above atomic number 83 (elements 84 and up), all elements are radioactive.

THE BAND OF STABILITY

Figure 6.2 shows the pattern of stable nuclides of the elements. The shaded area of the graph indicates stable isotopes and is known as the belt (or band) of stability. Most radioactive isotopes are located outside this region. The figure may also help you to see that there are three main situations that determine the types of decay that elements are likely to undergo.

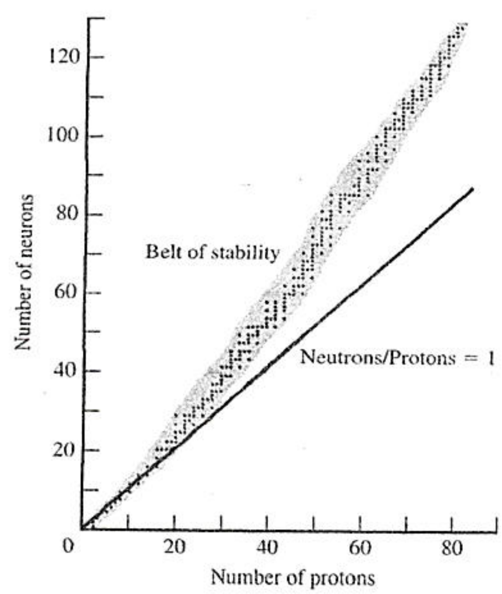

Figure 6.2

1. **Atoms whose nuclei are above the band of stability (high neutron-to-proton ratio) can lower their numbers of neutrons by undergoing beta emissions.** The typical pattern for these is that the mass number (number of neutrons + number of protons) is greater than the atomic weight. Remember that beta emissions convert neutrons into protons and beta particles.
2. **Atoms whose nuclei are below the band of stability (low neutron-to-proton ratio) can raise their numbers of neutrons by undergoing positron emissions, or electron capture.** The typical pattern here is that the mass number is less than the atomic weight. Remember that both processes involve the conversion of a proton into a neutron.
3. **Atoms with atomic numbers higher than 84 are too large to remain stable.** The easiest way to decrease size is to undergo alpha emission. Remember that alpha emissions eliminate two protons and two neutrons.

A large sample of radioactive material will spontaneously decay. Over time, the amount of the original sample will decrease and the amounts of products will increase. In the next section, you will look at the different ways this decay is described.

NOTE
These are just general guidelines, and they do not work for all elements.

Half-Life

A common way to describe the rate at which radioactive decay is occurring is a measurement known as the **half-life**. Half-life is defined as the time necessary for one half of a radioactive sample to undergo decay into new elements. Different isotopes have different decay rates. Some are as long or longer than 4.5 billion years (Uranium-238) to as short as 10 microseconds (astatine-215). It is not necessary to understand the factors that contribute to the length of the half-life, but you are expected to be able to perform various calculations involving half-life. There is also no way to really predict when an atom will undergo a single decay, but it is possible to observe large amounts of decay and come up with an average rate. If the amount of a radioactive substance is measured over time and the results are plotted, the resulting graph is known as a decay curve. Figure 6.3 shows the decay pattern more clearly.

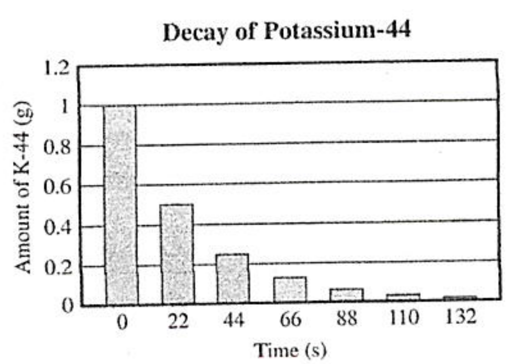

Figure 6.3

From a conceptual approach, we can look at half-life as the time it takes for ½ of a sample to decay into some other substance. For instance, if we start out with 1.0 g of a radioactive sample, after one half-life has elapsed, we will be left with only 0.5 g of the original material. After two half-lives, we will have 0.25 g. After 3, 0.125 g. As you can see, this could go on for some time, but it is generally accepted that after about 10 half-lives have elapsed, there is a negligible amount of the original radioactive material left.
Most of the problems on past AP exams that contain half-lives are relatively simple to solve, using either conceptual or mathematical approaches. On the exam, you are not provided with any equations related to nuclear chemistry. Therefore, any calculations you will have to make should be fairly simplistic and easy to solve using a few simple rules. From a conceptual perspective, this that means half-life problems can be solved by repeatedly cutting the starting amount in half. For example, after one half-life, a sample will have ½ the number of radioactive nuclei that it started with. After two half-lives, the sample will have (½)(½), or $(\frac{1}{2})^2$, times the number of radioactive nuclei left. After three half-lives, the sample will have (½)(½)(½), or $(\frac{1}{2})^3$, times the number of radioactive nuclei left. If you haven't already spotted the pattern here, it is that the amount of sample left after time *t* will be

$$N_t = N_0(\tfrac{1}{2})^n \qquad \text{(Equation 6.1)}$$

where *n* = the number of half-lives that have elapsed during time interval *t*. This can also be rewritten as

$$\frac{N_t}{N_0} = (\tfrac{1}{2})^n \qquad \text{(Equation 6.2)}$$

which is the same thing as saying

$$\text{Fraction remaining} = (\tfrac{1}{2})^n \qquad \text{(Equation 6.3)}$$

Sample: If you start with 64 g of a material with a half-life of 10 years, how much will be left at the end of 40 years?

Answer: Conceptually, you can work through this step-by-step:

At end of	You will have left
10 years	32 g
20 years	16 g
30 years	8 g
40 years	4 g

Or you can use the formula

$$40 \text{ years} = 4 \text{ half-lives, so}$$
$$N_t = N_0(\tfrac{1}{2})^n = (64\text{ g})(\tfrac{1}{2})^4 = (64\text{ g})(1/16) = 4\text{ g}$$

The conceptual approach is particularly effective when solving problems that have half-lives that are whole number values. For more complex problems, we need to use some ideas borrowed from chemical kinetics . Radioactive decay can be described as a first order processes, which means it can be described with the following equation:

(Equation 6.4)

$$Rate = kN_t$$

where k is a constant, known as the decay constant, and N is the number of radioactive nuclei in the sample at time t. Note that the rate is proportional to the size of the sample.

By applying some basic calculus, we can integrate this equation into a new formula (it is not necessary for you to know how this was derived):

(Equations 6.5, 6.6)

$$\ln\frac{N_t}{N_0} = -kt \quad \text{or} \quad \log\frac{N_t}{N_0} = \frac{-kt}{2.303}$$

where t is the time interval, k is the decay constant, N_0 is the initial number of nuclei, and N_t is the number of nuclei after time t. Mass values can be substituted for the N values since the mass will be proportional to the number of nuclei.

If you are going to be considering the half-life of a substance, then the quantities of nuclei will be related by the expression $N_t = \frac{1}{2} N_0$, or $2N_t = N_0$. Remember, after one half-life, the number of nuclei of a particular radioisotope will have decreased by half. This relationship will allow us to simplify the above expressions through substitution (again, it is not necessary for you to know the derivation) into the following expression:

(Equation 6.7)

$$k = \frac{0.693}{t^{\frac{1}{2}}} \quad \text{or} \quad t_{1/2} = \frac{0.693}{k} \quad \text{(the number } 0.693 = \ln 2\text{)}$$

Sample: The half-life of Iodine-125 is 60 days. If the original sample had a mass of 50.0 grams, how much is left after 360 days?

Answer: If you compare the half-life to the time elapsed, you will see that the number of half-lives that elapse during the time interval is a whole number (60/360 = 6). The easiest way to approach this problem, then, is to use equation 6.1:

$$N_t = N_0(1/2)^n$$

Given information

$N_0 = 50.0$ g
$n = 6$ half-lives
$N_t = ?$

We can now substitute these values into the equation to solve for N_t

$N_t = 50.0\text{ g}(1/2)^6 = 50.0\text{ g }(1/64) = 0.781\text{ g}$

You Try It!

Titanium-51 has a half-life of 6 minutes. If 98 g of the material are obtained, how many grams of the sample will remain after 1 hour?

Answer: 0.096 g

Sample: Iodine-131, a radioactive isotope of iodine, has a half-life of 8.07 days. A lab worker discovers that a sample of iodine-131 has been sitting on a shelf for 7 days. What fraction of the original nuclei is still present after 7 days?

Answer: To solve this problem, you will want to begin with the equation 6.6:

$$\log \frac{N_t}{N_0} = \frac{-kt}{2.303}$$

We will use this equation because we are concerned with the ratio of the initial to the final sample. We don't know k, so we can substitute the expression

$k = \frac{0.693}{t^{\frac{1}{2}}}$ for k in the equation. Doing so will produce the new expression:

$$\log \frac{N_t}{N_0} = \frac{-0.693t}{2.303t^{\frac{1}{2}}} = \frac{-0.693\ (7.0\text{ days})}{2.303\ (8.07\text{ days})} = 0.2610$$

Therefore, we need to rearrange the equation so that

$$\frac{N_t}{N_0} = 10^{-0.2610} = 0.548;\ 0.548 \text{ X } 100\% = 55\%$$

We also could have used equation 6.5:

$$\ln \frac{N_t}{N_0} = -kt;\ \text{substituting } k = \frac{0.693}{t^{\frac{1}{2}}} \text{ we obtain } \ln \frac{N_t}{N_0} = -\frac{0.693}{t^{\frac{1}{2}}}\text{t}$$

$$\ln \frac{N_t}{N_0} = \frac{0.693}{8.07\text{ days}}(7\text{ days}) = -0.601$$

$$\frac{N_t}{N_0} = e^{-.601} = 0.548 = 55\%$$

This is about what we predicted, so chances are we've done it correctly!

TIP

Before beginning, it helps to consider what you are doing. If you know that the half-life is about 8 days and the sample is 7 days old, you know there should be just more than half of the sample present in the sample. When you finish your work, check to make sure your answer is close to your estimate.

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You Try It!

The half-life of phosphorus-30 is 2.5 min. What fraction of phosphous-30 nuclides would remain after 14 min?

Answer: 2.1%

Sample: If a 1.0×10^{-3} g sample of technetium-99 has a decay rate of 6.3×10^5 nuclei s^{-1}, what is its decay constant?

Answer: We want to use the formula: $Rate = kN_t$

The first thing we will need to do is rearrange the equation to solve for k:

$$k = \frac{rate}{N_t}$$

We know the rate, but we still need to determine the number of nuclei present at time t (we only know the mass of the sample). To do this, we will need to determine how many technetium-99 atoms are present in $1.0 \text{ x} \times 10^{-3}$ g of the substance). This is a mole-conversion problem (see Chapter 13).

$$1.0 \times 10^{-3} \text{g } ^{99}_{43}\text{Tc} \times \frac{1 \text{ mol } ^{99}_{43}\text{Tc}}{99 \text{ g } ^{99}_{43}\text{Tc}} \times \frac{6.02 \times 10^{23} \ ^{99}_{43}\text{Tc nuclei}}{1 \text{ mol } ^{99}_{43}\text{Tc}}$$

$$= 6.1 \times 10^{18} \ ^{99}_{43}\text{Tc nuclei}$$

Now, we have all of the information we need to solve the problem:

$$k = \frac{rate}{N_t} = \frac{6.3 \times 10^5 \text{ nuclei / s}}{6.1 \times 10^{18} \text{nuclei}} = 1.0 \times 10^{-13} \text{ s}^{-1}$$

You Try It!

A $2.8 \text{ x} \times 10^{-6}$ g collection of plutonium-238 is decaying at a rate of 1.8×10^6 disintegrations per second. What is the decay constant (k) of plutonium-238 in s^{-1}?

Answer: $2.5 \times 10^{-10} \text{ s}^{-1}$

Sample: What is the half-life of technetium-99?

Answer: We know from problem 3 that the decay constant of Tc-99 is $1.0 \times 10^{-13} \text{ s}^{-1}$. We also know from equation 6.7 that

$$t_{1/2} = \frac{0.693}{k}$$

Therefore, by substitution, we will get

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{1.0 \times 10^{-13} \text{s}^{-1}} = 6.9 \times 10^{12} \text{ s}$$

Depending on your needs, you might need to convert this into more suitable units. For instance,

$$6.9 \times 10^{12}\ \text{s} \times \frac{1\ \text{min}}{60\ \text{s}} \times \frac{1\ \text{h}}{60\ \text{min}} \times \frac{1\ \text{day}}{24\ \text{h}} = 7.99 \times 10^{7}\ \text{days}$$

This still isn't a very useful amount since it is so large, so we can convert it further into years:

$$7.99 \times 10^{7}\ \text{days} \times \frac{1\ \text{yr}}{365.25\ \text{days}} = 2.2 \times 10^{5}\ \text{years}$$

You Try It!

A sample of curium-243 was produced for transport to a lab. When it was first produced, the activity of the sample was measured at 3012 dps (disintegrations per second). After 1.00 year, the lab needed to use the sample. The lead scientist measured the activity as 2921 dps. What is the half-life of curium-243?

Answer: 22.6 years

Radioactive Isotope Dating

Because radioactive isotopes seem to decay at very constant rates, they can be used as "clocks." One of the first radioactive dating techniques involved the use of the radioisotope carbon-14. Carbon-14 is produced in the upper atmosphere when neutrons (produced by cosmic rays from space) collide with nitrogen-14 molecules in the reaction shown below:

$$^{14}_{7}\text{N} + ^{1}_{0}\text{n} \rightarrow ^{14}_{6}\text{C} + ^{1}_{1}\text{H}$$

Carbon-14 incorporates into molecules just as Carbon-12 (ordinary carbon), including CO_2 gas. This carbon dioxide becomes incorporated into plants (and subsequently animals). It is assumed that the ratio of C-14 to C-12 within an organism is similar to the ratio in the atmosphere, about $1{:}10^{12}$ (it is believed to have been at this ratio for about 50,000 years). When a living organism dies, carbon no longer cycles into or out of it. Since Carbon-12 does not decay and Carbon-14 does, the change in ratio of C-12 to C-14 can be used to estimate the time of death of living organisms. There are other radioisotopes that are used for radioactive dating that include inorganic compounds in nonliving things, but the mechanisms for obtaining dates are similar. We will only examine C-14 dating.

The carbon-14 atom has a half-life of 5730 years and decays at a rate of 15.3 disintegrations per minute per gram of total carbon in living organisms. What this tells us is that if an organisms dies, the carbon-14 in it will decay at a rate of 15.3 disintegrations per minute per gram. As more carbon-14 decays, this rate will decrease. After one half-life has elapsed (5730 years), only half of the original C-14 will be left. This means that the total decay rate will be 7.65 disintegrations/minute/gram. After two half-lives, the disintegration will have dropped to 3.825, etc. Let's see how this information can be used to determine the date of an object.

Sample: How old is a piece of ancient wood that is giving off beta emissions from carbon-14 at the rate of 1.9 disintegrations/minute/gram?

Answer: We know that the rate of disintegrations is proportional to the number of nuclei. Therefore, we can substitute the activity for the *N* values in equation 6.6:

$$\ln \frac{N_t}{N_0} = -kt; \text{ substituting } k = \frac{0.693}{t^{\frac{1}{2}}} \text{ we obtain } \ln \frac{N_t}{N_0} = -\frac{0.693}{t^{\frac{1}{2}}} t$$

$$\ln \frac{1.9 dis/\min/g}{15.3 dis/\min/g} = -\frac{0.693}{t_{1/2}} t$$

$$-\frac{t_{1/2}}{0.693} \ln \frac{1.9 dis/\min/g}{15.3 dis/\min/g} = t = -\frac{5730\text{ y}}{0.693} \ln \frac{1.9 dis/\min/g}{15.3 dis/\min/g} = 1.7 \times 10^4 \text{y}$$

This object is about 17, 000 years old (rounded for significant digits).

You Try It!

A wooden artifact found in a mummy's tomb is found to have 9.4 disintegrations/min/g. How old is the artifact?

Answer: 4030 years old (rounded)

Transuranium Elements

In almost all of the previous examples, we have looked at nuclear reactions that occur by spontaneous decay. There are other types of nuclear reactions that can occur, known as **transmutation reactions**. These reactions can be induced by forcing a reaction between the nucleus of an element and nuclear particles (such as neutrons), or nuclei. Ernest Rutherford carried out the first transmutation by bombarding Nitrogen-14 nuclei with alpha particles. This resulted in the production of oxygen-17 and a proton, as shown below:

$${}^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$$

This process has been used to produce countless isotopes, including many radioactive isotopes. In addition, it has allowed scientists to produce elements with atomic numbers that are higher than that of the largest naturally occurring element, uranium. These elements are known as transuranium elements. In 1940, E. M. McMillan and P.H. Abelson of the

University of California Berkeley produced the first transuranium element, neptunium (Np, Z=93), by bombarding uranium-238 with neutrons. The nuclei that captured the neutrons were converted to uranium-239, which decayed into neptunium-239 during a beta emission. The reaction is shown below:

Step 1: $^{238}_{92}U + ^{1}_{0}n \rightarrow ^{239}_{92}U$

Step 2: $^{239}_{92}U \rightarrow ^{239}_{93}Np + ^{0}_{-1}e$

NUCLEAR REACTIONS

Energy Relationships

Nuclear decay allows a nucleus to form products with lower energy. However, according to Albert Einstein's famous equation $E = mc^2$, any change in energy must be accompanied by a corresponding change in mass. In chemical reactions, this change in mass is negligible (though it does occur), but it is much more pronounced in nuclear reactions. If E changes by a certain amount, ΔE, then mass will have to change proportionally by an amount Δm (c being constant at all times). Thus,

(Equation 6.8)

$$\Delta E = (\Delta m)c^2$$

If we examine a specific reaction, we can see how this works. Take for instance,

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

The masses of each of the particles in the equation are

$$^{238}_{92}U = 238.0003 \text{ g mol}^{-1}, \quad ^{234}_{90}Th = 233.9942 \text{ g mol}^{-1}, \quad \text{and } ^{4}_{2}He = 4.00150 \text{ g mol}^{-1}$$

The change in mass can be determined by subtracting the mass of the reactant (parent nuclei—in this case, uranium-238), from the combined masses of the products.

$$\Delta m = (233.9942 \text{ g} + 4.00150 \text{ g}) - 238.0003 \text{ g} = -0.0046 \text{ g}$$

To determine the energy change, we can now use Einstein's equation and solve for ΔE. Before we can do this, we must convert the mass to kilograms (because the energy unit of joules requires the mass in kilograms).

$$\Delta E = -4.6 \times 10^{-6} \text{ kg}(3.0 \times 10^{8} \text{ m/s})^2 = -4.14 \times 10^{11} \text{ J}$$

This is an enormous amount of energy!

You will not be expected to perform these calculations on the AP exam, but you should appreciate from a conceptual level that there is an enormous amount of energy released during nuclear reactions. In addition, you should understand that the amount of energy released in nuclear reactions is much larger than that released in chemical reactions. The main reason is that during chemical reactions, the only energy released during a reaction will result from electrostatic forces between the protons and electrons in the atom. During nuclear reactions, the energy result is dependent on the energies associated with the strong nuclear force, which is many orders of magnitude larger than electrostatic forces. On the atomic level, we can see the effects of these strong nuclear forces when we look at the phenomena of binding energies and mass defect.

Binding Energy

It turns out that the mass of an individual atom is always less than the sum of its parts. That is, if you add up the masses of all the components of an atom, you will not get the total mass of the atom. As an example, let's look at oxygen-16.

Oxygen-16 contains 8 protons and 8 neutrons. Therefore, we would expect the mass to equal

mass of 8 protons +	mass of 8 neutrons =	total mass
8(1.0078252 amu) +	8(1.0086652 amu) =	16.1319232 amu
However, the actual mass of oxygen-16 is		15.9949150 amu

So, what happened to the 0.1370082 amu that should have been present? It turns out that when the neutrons and protons come together to form the nucleus, they form a more stable entity. This means that energy is released. When that energy is released, there must be a corresponding loss of mass. The more energy that is released, the more stable the nucleus is. In order to break apart a nucleus, you would have to add that much energy. The amount of energy you must add to break a nucleus into its constituent neutrons and protons is known as the **binding energy**. The difference in mass between the expected and actual masses is known as the **mass defect**.

It is not necessary to be able to perform these calculations for the AP exam. However, it is very important that you understand the underlying idea that large amounts of energy are released when atomic nuclei are broken apart. It is also important to understand that the difference in mass between the components of a nucleus and the actual mass of the nucleus can be accounted for by a change in the energy state of those components. The nature of that relationship is captured in Einstein's equation $E = mc^2$.

Nuclear Fission

We have seen so far that the process of heavy nuclei splitting apart (known as **fission**) is highly exothermic. In addition, the joining of lighter nuclei (known as **fusion**) is also a highly exothermic process. Fission is typically accomplished by bombarding heavy nuclei with slow-moving neutrons. Once the neutron is absorbed, the resulting unstable nucleus breaks into smaller nuclei. One of the most well known fissions involves the splitting of uranium-235, shown in the reaction below:

$$^{235}_{92}\text{U} + {}^{1}_{0}\text{n} \rightarrow {}^{139}_{56}\text{Ba} + {}^{94}_{36}\text{Kr} + 3\,{}^{1}_{0}\text{n}$$

The neutrons produced by this fission reaction can potentially collide with other U-235 nuclei. The likelihood of the extra neutrons striking other nuclei increases as the mass of the sample increases. At a characteristic mass, the neutrons are assured to collide with U-235 nuclei, and as a result, a chain reaction begins. In this chain reaction, the neutrons from one fission will strike other nuclei and cause additional fission reactions. The mass at which a self-sustaining chain reaction will occur is known as the **critical mass**. Fission reactions are responsible for the production of nuclear power and for the design of nuclear weapons.

Nuclear Fusion

In a fusion reaction, light nuclei will combine to form heavier ones. While this process is quite commonplace on stars (including our sun), it is very difficult to accomplish in a laboratory setting. In order to fuse nuclei (such as hydrogen), extremely high temperatures are necessary (around 100 million degrees Celsius) to overcome the repulsive forces between nuclei. These temperatures are very difficult to achieve and to maintain long enough to achieve the reaction.

Summary: Nuclear Chemistry

- Atomic nuclei are composed of neutrons and protons.
- For each nuclei, there are some isotopes that are more stable than others. The stability of each nucleus is determined by the ratio of neutrons to protons. The belt of stability can be used to estimate the stability of any given nucleus.
- To gain stability, neutrons undergo decay reactions: alpha emission, beta emission, positron emission, and electron capture are possible.
- Although it is not possible to predict when a single decay will occur, the overall rate of decay for any isotope is relatively consistent.
- The rate of decay can be used to determine the half-life of an isotope; that is the time during which one half of a radioactive sample is converted into a different material.
- The half-lives of certain materials, like carbon-14, can be used to determine the ages of artifacts.
- The sum of the masses of the individual components of an atom is greater than the actual mass of the atom. This is due to the mass that is converted into energy as the nucleons bind together.
- Nuclear fission is a process whereby unstable nuclei are bombarded by neutrons in order to split them into smaller nuclei.
- Nuclear fusion is a process where small nuclei are forced together with an extremely large amount of energy in an effort to join them into a larger nucleus.