6.1 THE STABILITY OF ATOMIC NUCLEI

Until the 1870s, science considered the atom to be the smallest unit of matter, and to be unchanging and unchangeable. Thomson’s discovery of the electron - the first known sub-atomic particle - began the great upheaval in the theory of atomic structure. Some twenty years later, another supposed ‘truth’ was to be shattered.

In 1896, a French scientist, Henri Becquerel, found that some new photographic plates left in a drawer in the dark had somehow been exposed. He was certain that they could not have been accidentally exposed to light. The dark exposed ‘blob’ on the plate was the same shape as a piece of uranium ore left in the drawer on top of the plates. Becquerel’s best explanation was that the uranium ore was somehow emitting radiation (but not visible light) and had therefore exposed the photographic plates. This was subsequently proven to be the case, though at the time, it could be explained what was actually happening.

The phenomenon was given the name radioactivity: the spontaneous emission of radiation by atomic nuclei (this definition later extended to include as a result of the breakdown of the atomic nucleus). You will see later that in this context, radiation refers to not only true electromagnetic radiation, but also atomic particles. Nuclei which exhibited this phenomenon were called known as radioisotopes.

Rutherford’s gold foil experiment, which proved that atoms were mostly empty space, used the emission of a radioisotope. He was able to prove that the emissions were not radiation, but in fact, the nuclei of helium. If this was the case, then radioactivity must involve the breakdown of the nuclei of the radioisotope. The process of spontaneous nuclear decomposition is known as radioactive decay. The decaying nuclei is known as the parent nuclide, and the product is called the daughter nuclide. The products of radioactive decay may be unstable themselves, and decay further.

Not all atomic nuclei are radioactive. A number of observations have been made about the stability of nuclei:

- all known nuclei with atomic numbers greater than 83 are unstable,
- nuclei with even numbers of protons and neutrons are more stable,
- nuclei with 2, 8, 20, 50, 80 or 126 protons or neutrons (not total) are more common than nuclei one more or less than these so-called magic numbers (apparently the equivalent of the stable electron numbers 2, 10, 18, 36, 54 or 86); for example, there are 16 stable nuclei with 50 neutrons or 50 protons, but only 3 with 49 or 51 of either particle, and
- when neutron numbers are plotted against atomic number for the stable nuclei, it is found that all fall close, known as the belt of stability

Figure 6.1 illustrates the belt of stability. Isotopes in the greyed region on the are stable, those in region A have too many neutrons, those in region B too few neutrons, and those in C simply too heavy.

6.2 NUCLEAR DECAY PROCESSES

Nuclear equations deal only with the nuclear particles (not with the circling electrons or the various chemical states that the element may find itself in). Thus, a uranium atom always decays in the same way, regardless of whether it is in metallic uranium, a compound or in solution as a free or complexed ion.
Writing Nuclear Equations

To balance a nuclear equation, the mass number and atomic number of each component is important. The atomic number is the number of protons in the atom, the mass number is the number of protons and neutrons. It is the sum of these which is in balance between products and reactants.

EXAMPLE

In a nuclear reactor, uranium-235 absorbs a neutron, and then fragments, producing three more neutrons, krypton-91 and one other product. What is it?

A nuclear equation is balanced when the mass numbers of reactants and products are equal, and similarly for the atomic numbers. The equation for the known nuclei is:

\[ ^{235}_{92}U + ^{1}_0n \rightarrow ^{91}_{36}Kr + 3^1_0n + ? \]

Reactants
Mass number: \(235 + 1 = 236\)
Atomic number: \(92 + 0 = 92\)

Products
Mass number: \(236 = 91 + (3 \times 1) + ?\)
Atomic number: \(92 = 36 + (3 \times 0) + ?\)

The mass and atomic number for the unknown product are 142 and 56, respectively. This means the product is **barium-142**, when the periodic table is checked.
PRACTICE QUESTIONS
1. Neutron absorption by uranium-235 can also result in a different fragmentation, where tellurium-137 and zirconium-97. How many neutrons are produced in this process?

Modes Of Natural Nuclear Decay
Study of the products of the decay of many radioisotopes has shown that five basic decay modes exist. Of these, three occur from naturally radioactive nuclei (α, β, γ), while the other two (positron emission and electron capture) are more likely to occur in man-made radioisotopes.

**Alpha (α)**
An alpha particle is a group of two protons and two neutrons ejected together from the nucleus of an atom which has been too heavy to be stable. Given its composition, it should be obvious that an alpha particle is a helium-4 atom, stripped of its electrons.

Uranium-238 is an example of an alpha emitter.

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

**Beta (β)**
Beta emission involves the production of a high-energy electron from the nucleus (denoted by the symbol β⁻). This may seem rather unusual since there are no electrons in the nucleus. The beta particle is produced by the breakdown of a neutron in nuclei which have too many neutrons.

\[ {}^{1}\text{n}_{0} \rightarrow {}^{1}\text{p}_{1} + {}^{0}\text{e}^{-} \]

Notice the symbol for a β⁻ particle: \( {}^{0}\text{e}^{-} \). This is because an electron is electrically the inverse of a proton. The above equation shows that β⁻ emission is the result of the transformation of a neutron into a proton. The mass number of the nucleus remains unaltered, but the atomic number increases by one.

Iodine-131 is a typical β⁻ emitter.

\[ {}^{131}\text{I}_{53} \rightarrow {}^{131}\text{Xe}_{54} + {}^{0}\text{e}^{-} \]

**Gamma (γ)**
Gamma radiation is pure electromagnetic radiation: no mass, no charge, only energy. It accompanies almost all α and β decay processes as the result of the daughter nuclide being energetically excited and unstable. Certain daughter nuclides remain energetically unstable for an extended period, and thus appear to undergo a process of only gamma emission. Such isotopes are known as metastable and indicated with the letter m or an asterisk. The process is often called isomeric transition (IT).

\[ {}^{85}\text{Kr}_{36}^{m} \rightarrow {}^{85}\text{Kr}_{36} + \gamma \]

The gamma radiation emitted during a particular decay is of definite and fixed energy values. This can be used as an identification tool, or for quantitative analysis.

**Positron Emission (β⁺)**
Positrons are positive electrons - that is, they have the same mass and absolute charge as an electron, but that charge is positive. Positron emission is given the symbol β⁺, and occurs when a nucleus has too many protons. As shown below, a proton decays to form a neutron and a positron. The mass number of the nucleus remains unchanged, but the atomic number decreases by one.
6. Nuclear Chemistry

\[ ^{1}p_{1} \longrightarrow ^{1}n_{0} + ^{0}e_{1} \]

Platinum-187 is a positron emitter.

\[ ^{187}Pt_{78} \longrightarrow ^{187}Ir_{77} + ^{0}e_{1} \]

The net result is a nucleus with one less proton and one more neutron. Positrons would have similar penetration properties to electrons, were it not for the annihilation reaction that occurs when the two meet. Positrons are anti-matter: when they collide with their "normal" matter equivalent, the result is the complete destruction of both particles and the release of considerable amounts of energy.

**Electron Capture**

Electron capture occurs in the same type of nuclei as positron emission: excess protons. Such nuclei will often exhibit both forms of decay. The captured electron will come from the inner shell of orbiting electrons.

\[ ^{1}p_{1} + ^{0}e_{-1} \longrightarrow ^{1}n_{0} \]

Calcium-41 is a nucleus that undergoes electron capture.

\[ ^{41}Ca_{20} + ^{0}e_{-1} \longrightarrow ^{41}K_{19} \]

---

**PRACTICE QUESTIONS**

2. Determine the product of the following decay processes.
   (a) the alpha decay of astatine-198
   (b) the \( \beta^- \) decay of fluorine-18
   (c) positron emission by yttrium-84
   (d) electron capture by neodymium-141
   (e) isomeric transition of chlorine-38*

3. What decay processes are occurring in the following reactions:
   (a) iodine-118 decays to tellurium-118
   (b) plutonium-244 decays to uranium-240
   (c) aluminium-29 decays to silicon-29
   (d) polonium-200 to lead-196
   (e) holmium-157 to dysprosium-157
   (f) magnesium-29 to aluminium-29

4. Uranium-238 undergoes a long sequence of decays before attaining a stable nucleus, lead-206. How many \( \alpha \) and \( \beta^- \) decay processes have occurred?

5. Construct a table for the five modes of decay, which summarises the change in atomic and mass number of the radioisotope as it decays.

---

**Radioactive Decay Series**

Some of the heavy radioisotopes cannot become stable by one decay process. As has been observed, no stable nuclei exist with atomic numbers greater than 83. Uranium isotopes (atomic number 92), for example, must undergo a series of \( \alpha \) and \( \beta^- \) decays before reaching a stable nucleus. Figure 6.2 shows the decay series for uranium-235, which ends with lead-207 (atomic number 82).

**Properties Of Decay Particles**

The three types of radioactive emissions from natural radioisotopes differ widely in their properties - energy, charge, ability to penetrate matter. Table 6.1 and Figure 6.3 summarise some of these variations. The difference in the ability of the various types of matter to block the emission is essentially due to density (number of atoms that the emission can be absorbed by), rather than chemical composition.
6.3 ARTIFICIAL NUCLEAR PROCESSES

Rutherford was the first scientist to perform an artificial nuclear reaction, when he caused the high speed collision of alpha particles with a target nucleus, nitrogen-14. The result was the formation of oxygen-17 and hydrogen-1. Such a process has been repeated many times with different target nuclei and colliding species. There are three possible outcomes of these artificial nuclear reactions, depending on the relative size of the colliding species and the target nucleus.

**Nuclear Transmutation**

This involves the collision of a small species, generally a neutron or alpha particle with a large target nucleus. The result is a nucleus which is increased in mass (and possibly) atomic number. The daughter nuclide probably will be radioactive. However, this process is used to manufacture most of the radioisotopes used in medicine and other fields (see Section 6.?). An example is the production of copper-64 by bombarding stable copper-63 with neutrons.

\[ ^{63}\text{Cu}_{29} + {\text{n}}_0 \longrightarrow ^{64}\text{Cu}_{29} \]

All the elements with atomic numbers greater than 92 have been produced in this way. Most of these so-called **transuranium elements** are not known in nature, but can be produced in nuclear reactors, called cyclotrons, which accelerate the particles to very high speeds to increase the chances of a successful collision and reaction.
Nuclear Fission

The collision of neutrons with heavy nuclei can cause the splitting of the nuclei. Such a reaction is the basis for the nuclear power industry and the early atomic bombs. Various uranium isotopes can be fragmented by neutron bombardment, with the release of huge amounts of energy and more neutrons. The neutrons produced can then go on to cause fission of more nuclei, and so on. This is known as a chain reaction.

\[ ^{235}\text{U} + \text{n} \rightarrow ^{94}\text{Sr} + ^{139}\text{Xe} + 3\text{n} \]

If only a small amount if the reacting nuclei are present, enough of the neutrons produced will be lost to the surroundings, and the reaction does not continue. If, however, an amount exceeding the critical mass is present, the process is self-sustaining and potentially uncontrollable. Such a requirement is met in an atomic bomb, where two masses of the material, each below the critical mass, are brought together. In a nuclear power plant, the process is controlled (or moderated) by neutron-absorbing material.

The products of nuclear fission reactions are frequently high radioactive themselves, and this problem is the main reason why so much controversy surrounds the operation and building of nuclear power plants.

Nuclear Fusion

The combination of two smaller nuclei to form a larger one is what drives our sun.

\[ ^{3}\text{He} + ^{1}\text{H} \rightarrow ^{4}\text{He} + ^{0}\text{e} \]

The fusion of nuclei of hydrogen-1 and helium-3, among others, release huge amounts of energy, and produce relatively safe products, unlike those from nuclear fission. It is for these reasons that nuclear fusion is a potential energy source of the future. There is one very stumbling block, however, in the implementation of a feasible fusion power plant: to combine fuse nuclei requires temperatures of many millions of degrees. Achieving and maintaining such a temperature is exceptionally difficult and not energetically efficient at this time.

6.4 Nuclear Decay Rates

Radioactive decay does not occur instantly. If it did, many of the problems associated with radioactive wastes would not be significant. Unlike normal chemical reactions, the rate of nuclear decay is independent of temperature, pressure and chemical form of the nuclei. However, like many chemical reactions, nuclear decay occurs at a rate proportional to the number of reactants (nuclei) present. This is known as a first order reaction, and is given by the following equation.

\[ \text{Rate} = \lambda N \quad \text{Eqn 6.1} \]

where N is the number of nuclei at any time and \( \lambda \) a constant determined by the isotope itself, known as the decay constant. If the exact rate can be determined (known as the activity) then the number of nuclei of that particular type can be calculated, since \( \lambda \) is known for all isotopes.

**EXAMPLE**

A 1.00 g sample of sand has an activity due to thorium-234 of \( 5.5 \times 10^6 \) counts per second. Given that the decay constant for this isotope is \( 3.2 \times 10^{-7} \) per second, calculate the ug/kg of thorium in the sample.

Number of nuclei = \( \frac{\text{Rate}}{\lambda} \)

= \( \frac{5.5 \times 10^6/3.2 \times 10^{-7}}{} \)

= \( 1.72 \times 10^{13} \) atoms

= \( 2.86 \times 10^{-11} \) moles

Mass of thorium = \( 6.69 \times 10^{-3} \) micrograms

= \( 6.69 \text{ ug/kg} \) in sand
PRACTICE QUESTION
6. A 2.000 g sample of contaminated soil has an activity due to strontium-90 of $1.38 \times 10^4$ counts over a 24 hour period. Its decay constant is $1.1 \times 10^{-9}$ per second. Calculate the level of Sr-90 in the soil.

As a function of time, the above equation can be modified to an exponential relationship as in Equation 6.2.

\[ N = N_0 e^{-\lambda t} \]

where $N_0$ is the number of nuclei originally present.

The decay constant varies greatly from one isotope to another. Some isotopes decay completely within milliseconds, others may take millions of years before most of the nuclei have decayed to a stable form. The most common way of measuring the decay rate of a radioisotope is not the decay constant, whose values are difficult to visualise, but rather the **half-life**: the time taken for half the number of nuclei to decay. Figure 6.4 shows the change in the number of nuclei of a radioisotope over time. Table 6.2 gives examples of the half-lives of some radioisotopes.

![Graph showing the decay of a radioisotope over time.](image)

**FIGURE 6.4** Decay of radioisotope

**TABLE 6.2** Half-lives of some radioisotopes

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Half-life</th>
<th>Isotope</th>
<th>Half-life</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon-14</td>
<td>5730 yrs</td>
<td>Gold-177</td>
<td>1.3 s</td>
</tr>
<tr>
<td>Copper-64</td>
<td>12.7 hrs</td>
<td>Magnesium-27</td>
<td>9.5 mins</td>
</tr>
<tr>
<td>Strontium-90</td>
<td>29 yrs</td>
<td>Thorium-219</td>
<td>1 us</td>
</tr>
<tr>
<td>Uranium-238</td>
<td>$4.46 \times 10^6$ yrs</td>
<td>Iodine-131</td>
<td>8 days</td>
</tr>
<tr>
<td>Plutonium-242</td>
<td>$3.76 \times 10^8$ yrs</td>
<td>Oxygen-13</td>
<td>8.3 ms</td>
</tr>
</tbody>
</table>
EXAMPLES
1. What proportion of copper-64 remains 2 days after production in a medical isotope reactor?

2 days (48 hours) is approximately 4 half-lives of Cu-64. Therefore, one-quarter of the original amount would remain.

2. To the nearest number of half-lives, how long would it require for 99% of plutonium-242 to decay?

99% decay means that 1% remains. In half-life terms, 1/100 is closest to 1/128, which is the fraction that would remain after 7 half-lives - 2.63 million years.

PRACTICE QUESTIONS
4. Using the same method as for the half-life example in your notes, determine the relationship between time and decay constant for 99% decay of a radioisotope.
5. How many half-lives would pass before the following changes occur:
   (a) 15 g of barium-137 decays to 0.94 g
   (b) the counts of a sample of strontium-90 decrease from 12030 to 90
6. Calculate how much is left of the following radioisotopes after the following time periods (round to the nearest half-life):
   (a) 12 g of calcium-39 after 12 seconds (t ½ = 860 ms)
   (b) 1.4 g of bromine-82 after 6 days (t ½ = 35.3 hours)
   (c) 750 mg of silver-117 if the counts drop from 2567 to 80

The half-life of a radioisotope is measured, not by counting the number of isotopes left (difficult to say the least), or by analysing the concentration of the element (firstly because there may be other isotopes of the same element present, and secondly, the number of radioactive nuclei is decreasing continually), but by measuring the rate of emission of the particular particle produced by the nuclear decay.

Figure 6.5 shows the change in count rate for a sample of palladium-111 over a period of time. The half-life (t ½) is equal to the time taken for the counts to decrease by half. The measurement of the half-life by this method is made inaccurate by the need to draw a curve through the data.

An alternative method relying on the same data can be used to provide a straight line graph. Since the decay is an exponential one, it can be shown that the natural logarithm of the counts is directly proportional to the time, and that (ignoring the negative sign) the slope of the graph is equal to the decay constant, \( \lambda \), **The half-life is equal to \(-0.693/(\text{slope of log graph})\).**

EXAMPLE

Determine the half-life of palladium-111, using the direct and logarithmic methods, given the graphs in Figures 6.5a and b.

From Figure 6.5a, the time when the counts were 2000 was just under 12 minutes, and the time when these counts had halved from 2000 to 1000 was just under 32 minutes. Therefore, the half-life is around 20 minutes.

The straight line has a slope of \(-0.034 \text{ min}^{-1}\). The half-life is equal to \(-0.693 \div \text{slope}\). Therefore, the half-life is **20.4 minutes** (logarithm method).
6. Nuclear Chemistry

FIGURE 6.5 Radioactive decay of palladium-111 (a) direct plot of counts (b) logarithmic plot

PRACTICE QUESTIONS

7. Determine the half-life an isotope, given the following data.

<table>
<thead>
<tr>
<th>Time (secs)</th>
<th>Counts</th>
<th>Time (s)</th>
<th>Counts</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>3923</td>
<td>90</td>
<td>980</td>
</tr>
<tr>
<td>30</td>
<td>2472</td>
<td>120</td>
<td>620</td>
</tr>
<tr>
<td>60</td>
<td>1560</td>
<td>150</td>
<td>400</td>
</tr>
</tbody>
</table>

8. Injected into the human body, iodine-131 is used as a tracer to follow the buildup of the element in the thyroid. The sample needs to have an activity of 60 uCi before injection. If the sample when received by the hospital has an activity of 2 mCi, how long before the sample is useless.

9. What fraction of strontium-90 will remain in the soil around Chernobyl power plant since the explosion of the reactor in 1986?
10. Uranium-238 decays to form the stable isotope lead-206. In rocks without the more common isotope of lead, 208, the relative masses of U-238 and Pb-206 can be used to date the rocks. A sample of rock is found to contain 0.257 mg of lead-206 for every mg of uranium-238. How old is the rock?

**PRACTICAL WORK**
You are required to measure the half-life of a radioisotope. No details of the procedure can be provided here because of the difference in equipment and radioisotopes that will be used in different locations.

It is recommended that the isotope have a half life of 2-10 minutes, and that the total counting period should be at least eight half-lives.

A suggested counting pattern (used for a 2½ minute half-life isotope) is:
- count for 30 seconds
- pause for 15 second (to reset timer and counter)
- repeat 20 minutes

Q1. Plot the corrected counts against time (in seconds). Draw a smooth curve through the data points. Estimate the half-life by the interval method, using at least three measurements.
Q2. Plot ln(corrected counts) against time. Calculate the slope of the graph and determine the half-life.
Q3. Compare the two values of half-life that you have determined with the literature value. Account for variations.
Q4. Write an equation for the decay process.
Q5. Show why the graph ln(count rate) versus time is a straight line

**6.5 MEASUREMENT OF RADIOACTIVITY**

The emission of particles or radiation from decaying nuclei is measured as a result of the ionising or energetic nature of the emission.

Alpha and beta particles in particular, being ionised themselves, produce ions as they collide with matter. If these ions are free to move, as in gas molecules or atoms, then they can be collected by applying a high voltage between positive and negative electrodes. The most common detector is known as a Geiger counter and a simplified diagram is shown below in Figure 6.6.

![Typical Geiger counter](image)

**FIGURE 6.6** Typical Geiger counter
Gamma radiation generally doesn’t produce sufficient ion density in gases, because of the low probability of interaction. It is measured in a different way, which we won’t cover here.

**Units of measurements associated with radiation**

While measuring radioactivity is a relatively simple task, as described above, the method by which the results are reported varies. It depends on the reason for the measurement: determining the amount of radiation emitted (activity) the energy absorbed by a particular object (exposure) or the biological effect of the radiation (damage).

The simplest measurement is the **intensity** of radiation, which is a measure of the activity of the source, i.e. the number of decays per second. This is not a measure of what is emitted, but rather the process that produces the radiation. It is measured in the units of Becquerels (Bq): 1 Becquerel is equal to one radioactive decay per second. Normal amounts of radioactive isotopes will be measured in millions or billions of becquerels. The Becquerel has replaced the Curie (Ci). \( 3.7 \times 10^{10} \text{Bq} \) are equal to 1 Ci.

The **energy** of the radiation absorbed by matter is potentially dangerous, and thus is an important measure. The amount of radiation that delivers one joule of energy to one kilogram of matter is known as the Gray (Gy). This is dependent on the energy of the radiation. Gamma rays are very high energy radiation. The previous unit was the rad. 100 rads are equivalent to 1 Gy.

To gauge the **potential biological damage** caused by radiation, a final measure known as the Sievert (Sv) is used. The number of Grays is multiplied by a factor which takes into account the type of tissue affected and the type of particles involved. Alpha particles do far more damage than gamma radiation: 1 Gy of alpha particles does 15 times as much damage as 1 Gy of gamma radiation. However, as a result of the correction factor, 1 Sv of alpha particles does exactly the same amount of damage as 1 Sv of gamma radiation. As a general guide, normal human exposure from background radiation is about 2 mSv.

### 6.6 Applications Of Radioactive Nuclei

Radioisotopes are used in such fields as medicine, archaeology, metallurgy, chemistry, geography, geology and biology. It is outside the scope of this topic to examine all of these. Two applications will be examined: neutron activation analysis (NAA) and radiochemical dating.

Neutron activation analysis involves the bombardment of a non-radioactive sample with a stream of neutrons (generated in a nuclear reactor). Some atoms will absorb the neutron, resulting in an unstable nucleus and therefore subsequent nuclear decay, which can be measured, both in terms of its intensity (quantitative analysis) and the energy of the gamma rays (qualitative analysis).

For example, a sample containing copper, which has two common isotopes - 63 and 65 - is irradiated with neutrons (often for up to 24 hours, depending on expected concentrations). A small fraction of the copper atoms (among others) will absorb a neutron, giving the following transformations:

\[
\begin{align*}
^{63}\text{Cu}_{29} + \gamma \text{n}_0 & \rightarrow ^{64}\text{Cu}_{29} \\
^{65}\text{Cu}_{29} + \gamma \text{n}_0 & \rightarrow ^{66}\text{Cu}_{29}
\end{align*}
\]

Copper-64 and copper-66 are unstable and decay via \( \beta^- \) emission and produce characteristic energy \( \gamma \) radiation also. It is the intensity of the \( \gamma \) radiation that is measured and related via standards to the concentration of copper in the sample.

As may be obvious from the earlier calculation regarding numbers of atoms in a radioactive sample, NAA is a very sensitive technique, capable of analysing nanogram levels of certain elements. It has the advantage of not being limited to metals as are certain other trace analytical species.

Radiochemical dating uses naturally occurring radioactive isotopes to establish the age of materials, such as rocks, fossils and man-made ancient artefacts. Carbon-14, which has a half-life of 5730 years, is very commonly used to determine the age of carbonaceous material.
Carbon-14 is produced in the upper atmosphere by the action of neutrons in cosmic radiation on nitrogen atoms.

\[ ^{14}N + ^{1}n \rightarrow ^{14}C + ^{1}p \]

Carbon-14 forms carbon dioxide, which is then used in the photosynthetic process within plants to make carbon compounds. Animals eat plants containing these compounds. Thus, all living organisms contain carbon-12 and carbon-14. The radioactive isotope slowly decays in the organism's tissues (by $\beta$ decay). It is, however, continuously replaced by more carbon-14. Therefore, while the organism is alive, its ratio of carbon-12 to carbon-14 remains unchanged. After the organism dies, it does not incorporate any new carbon, and thus the ratio of C-14/C-12 decreases with time, according to the half-life of C-14. Carbon dating is reliable for carbon-containing matter of ages between 200 and 50,000 years.

Measuring the ratio of atoms of each isotope of carbon is not easy. The simplest method burns the carbon-containing matter and the resulting carbon dioxide is trapped, and its radioactivity and mass measured. It is known that the equilibrium activity of carbon-14 in the atmosphere is 15.3 disintegrations per minute per gram of carbon. Thus, the activity of the sample gives an indication of the number of half-lives of carbon-14 have passed since the organism died. Obviously, very low decay rates are being measured, so long counting times (typically 24 hours) are used.

**EXAMPLE**

A sample of charcoal from the fireplace of a prehistoric village is combusted and the carbon dioxide produced is trapped. Its activity is measured, giving a corrected count value of 5510. The mass of carbon dioxide was 3.67 g. What was the age of the village?

The mass of carbon in 3.67 g of CO$_2$ is $(12 \div 44) \times 3.67 = 1.0$ g. The decay rate is $5510 \div (24 \times 60) = 4$. This is $\frac{1}{4}$ of the equilibrium value, and therefore, 2 half-lives - **11,460 years** - have passed since the timber was burnt in the village.

**PRACTICE QUESTION**

11. A sample of wood from an archaeological dig is burnt and the activity of the CO$_2$ produced (mass 3.340 g) measured over 2 days to be 8010 counts. How old is the sample?

12. Describe how a gamma emitter such as sodium-24 (half-life 15 hours) could be used to check for leaks in underground pipes. Explain why it would be more effective than an $\alpha$- or $\beta$-emitting isotope.

**6.7 HEALTH ASPECTS**

The dangers associated with radioactive isotopes are due to their ability to produce radiation or particles which ionise and fragment molecules within our bodies. Thus, the alpha, beta and gamma particles & radiation are included in the general classification of ionising radiation. X-rays are also classified as this. As radiation moves through matter it gives up its energy to the molecules it encounters, thereby leaving a trail of ions and molecular debris. The chemicals produced in this fashion are highly reactive and are able to further disrupt the normal operation of the body's cells. This may result in medical problems for that organism or their subsequent offspring, if the genetic material, DNA, is damaged.

The damage caused by a radiation source outside the body depends on the penetrating ability of the radiation. Gamma rays are particularly dangerous because they penetrate human tissue very effectively. Therefore, their potential for damage is not limited to the skin, unlike
that of alpha and beta particles whose limited penetration power means that the former is all absorbed the outer layer of skin, and the latter by the outer 1 cm of tissue. However, if an alpha or beta emitter enters the body in some way, then the potential for damage to the body’s cells is substantial. For example, strontium-90, a major β-emitting radioisotope produced in nuclear power reactors, when ingested is absorbed into the bones of the body, because of its chemical similarity to calcium. Therefore, it could do substantial damage to the bone marrow as it decays, resulting in possible bone cancer or leukemia.

PRACTICE QUESTION
13. Explain why a spill of the radioisotope iron-53 (half-life 9 minutes) in a river would be not be considered particularly dangerous, but that of iron-59 (half-life 45 days) would be a cause for concern.

WHAT YOU NEED TO BE ABLE TO DO

- write and balance nuclear equations
- write equations for the α, β and γ modes of decay
- describe the properties of α, β and γ radiation
- describe artificially-induced nuclear processes
- carry out measurement and calculations associated with half-life
- describe the meaning of measurement units of activity, dosage and exposure associated with radiation
- describe applications of radioisotopes
- understand the safety aspects associated with ionising radiation