CHAPTER 11 – NUCLEAR CHEMISTRY

Introduction

Most of chemistry focuses on the changes in the electronic structure of the atoms and molecules because it is those changes that result in bond breaking and bond formation (i.e., in chemical reactivity). In this, our final chapter, we examine reactions that involve changes in the nucleus. This branch of chemistry is called ***nuclear chemistry*** or radiochemistry.

11.1 The Nucleus

Introduction

All of the positive charge (protons) of an atom is concentrated in a very small volume called the nucleus. The repulsive force between the protons is very large, so the force responsible for holding the nucleus together is also very large. In this section, we discuss the origin of that force and the characteristics of a nucleus that dictate whether or not it is stable.

Objectives

- Determine the number of protons and neutrons in a nucleus given its symbol or atomic number and its mass number.
- Determine the atomic weight of an element from the masses and natural abundance of its isotopes.
- Calculate the mass defect of a nucleus from the number of protons and neutrons it contains and the mass of the nucleus.
- Convert between mass and energy.
- Determine the binding energy per nucleon of a nucleus and predict which of several nuclei is most stable.

11.1-1. Terms

There are three major subatomic particles: electrons, protons, and neutrons. Their masses and charges are summarized in Table 11.1. Notice that electrons and protons carry a net charge but neutrons are neutral. Also, the mass of the neutron and the proton are each very close to 1 amu ($M_m \sim 1$ g/mol) while the mass of an electron is much smaller. Because neutrons and protons reside within the nucleus, they are referred to as ***nucleons***. The number of protons in the nucleus is given by the ***atomic number***, Z, while the number of nucleons (protons plus neutrons) in the nucleus is given by the ***mass number***, A. The symbol N will be used to denote the number of neutrons in the nucleus. Thus, we can write

$$A = Z + N \qquad \text{The Mass Number} \tag{11.1}$$

The mass of a neutron and a proton are each 1.0 amu, so the mass number is the integer that is closest to the mass of the nucleus. For example, a carbon atom that contains 6 protons and 6 neutrons has a mass number of 12, and its mass is 12.0 amu.

Particle	**Mass (amu)**	**Charge**
electron	0.000549	−1
proton	1.00728	+1
neutron	1.00867	0

Table 11.1: Major Subatomic Particles

- An amu is an atomic mass unit. The mass of an atom expressed in amu is numerically equal to its molar mass.
- The charge is given in terms of the fundamental unit of charge, which is 1.6e−19 C.

11.1-2. Isotopes

The atomic number is the number that characterizes the atom. Two atoms with different atomic numbers are atoms of different elements, but atoms of the same atomic number are atoms of the same element, even if they have different masses. Atoms with the same atomic number but different mass numbers are called ***isotopes***. Thus, isotopes of the same element have the same number of protons but have a different number of neutrons (N). Isotopes are distinguished by indicating their mass number as superscripts in front of the symbol of the element. For example, ^{13}C (read "carbon-13"), is an isotope of carbon that has seven neutrons ($N = A - Z = 13 - 6 = 7$). The atomic mass of ^{13}C is 13.003 amu.

The atomic mass scale is based on the assignment of the mass of a carbon-12 atom (^{12}C), which is defined as exactly 12 amu. The reason that the molar mass of naturally occurring carbon is 12.011 g/mol and not 12.000 g/mol is that the molar mass of an element is the mass-weighted average of the masses of all of its naturally occurring isotopes. Naturally occurring carbon is a mixture that is 98.9% ^{12}C and 1.1% ^{13}C, so a mole of carbon contains 0.989 mol ^{12}C and 0.011 mol ^{13}C and has a mass of

$$(0.989 \text{ mol } ^{12}\text{C})(12.000 \text{ g/mol}) + (0.011 \text{ mol } ^{13}\text{C})(13.003 \text{ g/mol}) = 12.011 \text{ g}$$

Thus, any element whose atomic mass is not nearly an integer must have more than one naturally occurring isotope.

11.1-3. Isotope Exercise

EXERCISE 11.1:

Determine the molar mass of magnesium given the masses and abundances of the three isotopes. Express all masses to 0.01 g.

Isotope	**Mass**	**Abundance**
^{24}Mg	23.9850	78.70%
^{25}Mg	24.9858	10.13%
^{26}Mg	25.9826	11.17%

Mass of ^{24}Mg in one mole of magnesium = ______________ g

Mass of ^{25}Mg in one mole of magnesium = ______________ g

Mass of ^{26}Mg in one mole of magnesium = ______________ g

Molar mass of magnesium = ______________ g/mol

Nuclear Stability

11.1-4. Binding Energy

The ***binding energy*** of a nucleus is the energy required to separate the nucleus into its nucleons. The binding energy of a ^{12}C nucleus is the energy change for the following process:

$$^{12}\text{C} \rightarrow 6 \text{ p } + 6 \text{ n} \qquad \Delta E = 8.90 \times 10^{09} \text{ kJ}$$

Just as it is the energy per bond rather than the atomization energy dictates molecular stability, it is the energy per nucleon not the total binding energy that dictates the nuclear stability. The binding energy per nucleon of a ^{12}C nucleus is

$$\frac{8.90 \times 10^{09}\ \text{kJ}}{12\ \text{nucleons}} = 7.42 \times 10^{08}\ \text{kJ/nucleon}$$

As shown in Figure 11.1, nuclei with mass numbers close to that of iron ($A = 56$) are the most stable.

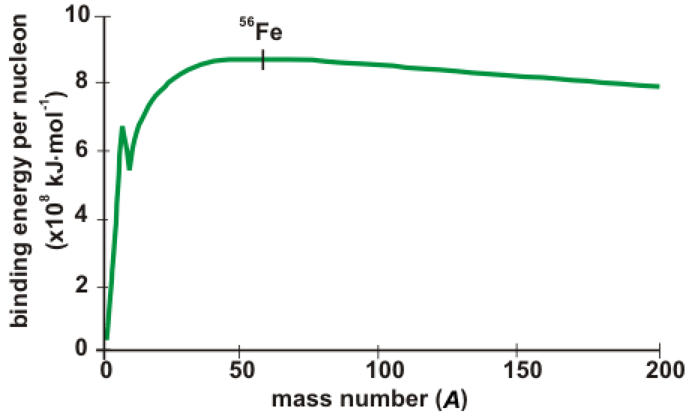

Figure 11.1: Binding Energy per Nucleon Versus Mass Number The binding energy per nucleon (nuclear stability) reaches a maximum for nuclei with mass numbers in the range of 50–60.

11.1-5. Mass Defect

Consider the process where a ^{12}C nucleus disintegrates into its nucleons: $^{12}C \rightarrow 6\ p + 6\ n$. The mass of a ^{12}C nucleus is the mass of the atom (exactly 12 amu) less the mass of the six electrons, so

$$\begin{aligned}
\text{mass of } ^{12}\text{C nucleus} &= \text{mass of atom} - \text{mass of electrons} \\
&= 12.0000 \text{ amu } - (6 \text{ e})(0.000549 \text{ amu/e}) \\
&= 11.9967 \text{ amu}
\end{aligned}$$

$$\begin{aligned}
\text{mass of } ^{12}\text{C nucleons} &= (6 \text{ p})(1.00728 \text{ amu/p}) + (6 \text{ n})(1.00867 \text{ amu/n}) \\
&= 12.0957 \text{ amu}
\end{aligned}$$

- The initial and final masses of the process are not equal, and the mass difference is called the ***mass defect***, Δm.

For the process $^{12}C \rightarrow 6\ p + 6\ n$, $\Delta m = 12.0957 - 11.9967 = 0.0990$ amu, which is 0.0990 g/mol or 9.90×10^{-5} kg/mol.

$$\begin{aligned}
\Delta m &= \text{final mass } - \text{ initial mass} \\
&= \text{mass of product } - \text{ mass of reactant}
\end{aligned} \qquad \text{Mass Defect} \qquad (11.2)$$

11.1-6. Mass-Energy

The origin of mass defect can be understood in terms of Einstein's famous equation that relates mass and energy.

$$E = mc^2 \qquad \text{Equivalence of Mass and Energy} \qquad (11.3)$$

Or, in terms of changes in energy due to changes in mass:

$$\Delta E = \Delta mc^2 \qquad \text{Equivalence of Mass and Energy Change} \qquad (11.4)$$

Equation 11.3 and Equation 11.4 show the equivalence of mass and energy, and the term ***mass-energy*** is sometimes used to express the equivalence. Indeed, the law of conservation of mass and the law of conservation of energy are combined into the law of conservation of mass-energy:

- The total mass-energy of the universe is constant.

The binding energy of a nucleus is determined from its mass defect and the application of Equation 11.4. However, a joule is a kg·m^2·s^{-2}, so, in order to obtain ΔE in joules,

- Δm must be expressed in kg
- $c = 2.998 \times 10^8$ m·s^{-1}

11.1-7. Binding Energy Exercise

EXERCISE 11.2:

The mass defect of a ^{12}C nucleus is 0.0990 amu. What is its binding energy?

Binding energy = ______________ J/mol

11.1-8. Binding Energy per Nucleon Exercise

EXERCISE 11.3:

Determine the binding energy per nucleon for a ^{64}Zn nucleus (atomic mass = 63.9291 amu).

Particle Masses (amu)	
electron	0.000549
proton	1.00728
neutron	1.00867

Number of nucleons

protons = ______________

neutrons = ______________

Mass of nucleons

protons = ______________ amu

neutrons = ______________ amu

Mass of nucleus = ______________ amu

Mass defect = ______________ amu

Binding energy = ______________ J/mol

Binding energy per nucleon = ______________ J/mol-nucleon

11.1-9. Nuclear Stability Exercise

EXERCISE 11.4:

Determine the binding energy per nucleon for ^{56}Fe ($Z = 26$) and ^{209}Bi ($Z = 83$) nuclei. The atomic masses are ^{56}Fe = 55.9349 amu and ^{209}Bi = 208.9804 amu.

Particle Masses (amu)	
electron	0.000549
proton	1.00728
neutron	1.00867

	Fe	Bi	
Mass of nucleons			amu
Mass of nucleus			amu
Mass defect			amu
Binding energy			J/mol
Binding energy/nucleon			J/mol-nucleon

Which nucleus is more stable?

^{56}Fe
^{209}Bi

11.2 Nuclear Reactions and Radioactivity

Introduction

Most nuclei found in nature are stable, and those that are not are said to be ***radioactive***. ***Radioactive nuclei*** spontaneously emit particles and electromagnetic radiation to change into other, more stable, nuclei. Radioactive nuclei are also called ***radioisotopes***. All of the first 83 elements except technetium ($Z = 43$) have at least one stable nucleus. However, the ^{209}Bi nucleus is the heaviest stable nucleus. Furthermore, many of the elements that have stable nuclei also have radioisotopes. In this section, we examine the different types of radioactive decay and present some observations that help us predict the mode of decay that a particular radioisotope is likely to undergo. We begin with a discussion about how nuclear reactions are written.

Objectives

- Define the terms "radioactive" and "radioisotopes."
- Identify a missing particle in a nuclear reaction.
- Identify the decay particles by name and symbol.
- Predict the probable mode of decay of an unstable nucleus.

Writing Nuclear Reactions

11.2-1. Particles

As in chemical reactions, nuclear reactions involve balancing both mass and charge. In a chemical reaction, the charge is given explicitly on each ion, but in a nuclear reaction, the charge is the charge on the nucleus, and that is given by the atomic number. Thus, the atomic number is included with the symbol in nuclear reactions to aid in charge balance. For example, element X with a mass number A and an atomic number Z is represented as

$$^{A}_{Z}X$$

and chlorine-37 is

$$^{37}_{17}\mathrm{Cl}$$

The following table lists the names and symbols of several small particles that are encountered in many nuclear reactions.

Particle	Second Name	Symbol	Comment
proton		$^{1}_{1}\mathrm{p}$	mass = 1, charge = 1
neutron		$^{1}_{0}\mathrm{n}$	mass = 1, charge = 0
electron	beta particle	$^{0}_{-1}\mathrm{e} = ^{0}_{-1}\beta = \beta^-$	mass = 0, charge = −1
positron		$^{0}_{+1}\mathrm{e} = ^{0}_{+1}\beta = \beta^+$	mass = 0, charge = +1, a positively charged electron
helium nucleus	***alpha particle***	$^{4}_{2}\mathrm{He} = ^{4}_{2}\alpha = \alpha$	mass = 4, charge = +2, the heaviest common particle

Table 11.2: Common Particles in Nuclear Reactions

11.2-2. Balancing a Nuclear Equation

In a nuclear equation, the masses are represented by the mass numbers and the charges by the atomic numbers:

- **charge balance**: The sum of the atomic numbers (Z) must be the same on both sides of the equation.
- **mass balance**: The sum of the atomic masses (A) must be the same on both sides of the equation.

11.2-3. Identifying a Product Exercise

<u>EXERCISE 11.5</u>:

Determine the atomic mass, number, and symbol of the unknown particle, X, in each reaction.

Do not include A or Z in the symbol, and use the following symbols for common particles:

- beta = e-
- positron = e+
- alpha = He

Reaction	A	Z	Symbol
$^{69}_{30}\mathrm{Zn} \rightarrow\ ^{69}_{31}\mathrm{Ga} + X$			
$^{27}_{13}\mathrm{Al} + \alpha \rightarrow n + X$			
$^{65}_{30}\mathrm{Zn} \rightarrow\ ^{65}_{29}\mathrm{Cu} + X$			
$^{59}_{26}\mathrm{Fe} \rightarrow \beta + X$			
$^{213}_{83}\mathrm{Bi} \rightarrow\ \alpha + X$			

Nuclear Decay

11.2-4. Trends in Nuclear Stability

Nuclear forces are not understood well enough to allow us to predict whether a nucleus is stable or not. Neutrons play an important role in holding the nucleus together, and every stable nucleus (except ^{1}H and ^{3}He) contains at least one neutron per proton. Figure 11.2 shows the number of protons and neutrons in the stable nuclei, which lie in a narrow band, referred to as the ***band of stability*** (shown as the blue line in Figure 11.2). Only one neutron per proton is sufficient for the lighter elements. However, the number of neutrons exceeds the number of protons in the stable nuclei of the larger elements. The following two rules summarize two empirical observations about nuclear stability that indicate the importance of the neutron to proton ratio and the size of the nucleus

- **Neutron/proton ratios**: The neutron/proton ratio remains near one through the third period ($Z = 18$), then it begins to rise reaching a maximum of 1.52 for ^{209}Bi $\left(N/Z = \dfrac{(209-83)}{83} = 1.52\right)$, the heaviest stable nucleus. The N/Z ratio is high to the left of the band of stability, so elements in that region tend to be β emitters as a β emmission converts a neutron into a proton ($\mathrm{n} \rightarrow \mathrm{p} + \mathrm{e}$), which lowers N/Z. The N/Z ratio is low to the right of the band of stability, so elements in this region tend to emit positrons or capture electrons, both of which convert protons to neutrons, which increases N/Z.
- **Total number of protons**: There are no stable nuclei with atomic numbers greater than 83. All elements with $Z > 83$ are radioactive. Heavy elements tend to emit α particles because they are the most massive particles.

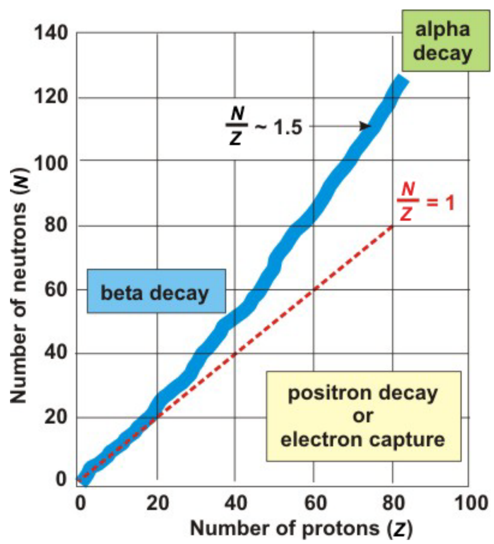

Figure 11.2: N/Z for Stable Isotopes The number of protons and neutrons in the stable nuclei is shown as the blue line, which represents the band of stability.

11.2-5. Alpha Decay

Alpha decay is the loss of an alpha particle. The loss reduces the mass number by four and the atomic number by two.

- **The alpha particle is the most massive particle of the common decay particles, so alpha decay is the most common mode of decay for the heavy nuclei.**

For example, ^{238}U undergoes α-decay to ^{234}Th:

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

α-decay is not limited to the heavier nuclei, but it is found in only a few of the lighter elements. ^{8}Be is the lightest element to undergo alpha decay:

$$^{8}_{4}\mathrm{Be} \rightarrow 2\, ^{4}_{2}\mathrm{He}$$

11.2-6. Beta Decay

Beta decay is the ejection of an electron by the nucleus. It results in an increase of one in the atomic number. The ejected electron is produced by the disintegration of a neutron, n $\rightarrow$ p + e. Because beta decay results from the conversion of a neutron into a proton, it decreases the neutron/proton ratio. As such,

- **β-decay reduces the proton/neutron ratio, so it is the common mode of decay for those nuclei lying above the belt of stability.**

For example, the neutron/proton ratio of ^{14}C is $8/6 = 1.3$, which is well above the value of 1.0 found for stable nuclei of the first three periods. Consequently, ^{14}C undergoes β-decay to a stable ^{14}N nucleus with $N/Z = 7/7 = 1.0$:

$$^{14}_{6}\mathrm{C} \rightarrow\ ^{14}_{7}\mathrm{N} +\ ^{0}_{-1}\mathrm{e}$$

α-decay, the most common decay among the heavy elements, is the loss of two protons and two neutrons, which increases N/Z slightly. Thus, successive α-decays produce isotopes with unfavorable N/Z ratios. Consequently, some heavy nuclei formed by α-decay undergo β-decay in order to maintain N/Z 1.5.^{234}Th, formed from the α-decay of ^{238}U, is a heavy nucleus and might be expected to undergo α-decay, but it also has a very high neutron/proton ratio of $\dfrac{(234-90)}{90} = 1.60$.

- **Reducing a high neutron/proton ratio is usually favored over reducing the mass in heavy nuclei.**

Consequently, ^{234}Th undergoes β-decay to ^{234}Pa, which has a N/Z ratio of 1.57:

$$^{234}_{90}\mathrm{Th} \rightarrow\ ^{234}_{91}\mathrm{Pa} +\ ^{0}_{-1}\mathrm{e}$$

11.2-7. Positron Decay

Positron decay is the emission of a positron and has the opposite effect of β-decay. It reduces the atomic number as it converts a proton into a neutron.

$$^{1}_{1}\mathrm{p} \rightarrow\ ^{1}_{0}\mathrm{n} +\ ^{0}_{+1}\mathrm{e}$$

- **Positron decay reduces the atomic number by one, so it is a common mode of decay for nuclei below the belt of stability.**

Positron emission of ^{13}N produces ^{13}C, which results in an increase of N/Z from 0.86 to 1.2:

$$^{13}_{7}\mathrm{N} \rightarrow\ ^{13}_{6}\mathrm{C} +^{0}_{+1}\mathrm{e}$$

A positron is the antimatter analog of the electron because it is identical to the electron in every respect except charge. Occasionally, the emitted positron collides with an orbital electron. The collision results in the annihilation of the two particles, so all of their mass is converted into a large amount of energy, which is released in the form of gamma radiation:

$$\beta^- + \beta^+ \rightarrow \gamma$$

Gamma radiation is high energy electromagnetic radiation and has no mass or charge.

11.2-8. Electron Capture

Electron capture (EC) is the capture by the nucleus of an electron from an inner-shell orbital. EC, like positron emission, increases the neutron/proton ratio by converting a proton into a neutron,

$$^{1}_{1}\mathrm{p} +^{0}_{-1}\mathrm{e} \rightarrow\ ^{1}_{0}\mathrm{n}$$

- **EC reduces the neutron/proton ratio so it is a common decay used by nuclei below the belt of stability.**

For example, ^{7}Be ($N/Z = 0.75$) undergoes electron capture to become ^{7}Li ($N/Z = 1.3$) :

$$^{7}_{4}\text{Be} + ^{0}_{-1}\text{e} \rightarrow \ ^{7}_{3}\text{Li}$$

11.2-9. Summary

Use the following rules to predict the mode of decay of a nucleus:

- Heavy nuclei ($Z > 83$) undergo alpha decay if their N/Z ratios are not too high.
- Nuclei with N/Z ratios above the band of stability undergo beta decay.
- Nuclei with N/Z ratios below the band of stability undergo either positron decay or electron capture.
- Reducing a high N/Z ratio (beta decay) is usually favored over reducing nuclear mass (alpha decay).

The above rules are summarized in Figure 11.2, which shows the common modes of decay in the regions where they are common.

11.2-10. Predicting the Mode of Decay Exercise

EXERCISE 11.6:

(a) The most abundant isotope of krypton is ^{84}Kr. Predict the mode of decay of the radioactive nucleus ^{76}Kr and determine the identity of the heavy product.

The mode of decay:
- alpha
- beta
- positron or EC

Heavy Product:

$A =$ ______________

symbol: ______________

(b) Cobalt occurs naturally as ^{59}Co. Predict the mode of decay of the radioactive nucleus ^{62}Co and determine the identity of the heavy product.

The mode of decay:
- alpha
- beta
- positron or EC

Heavy Product:

$A =$ ______________

symbol: ______________

(c) Predict the mode of decay of the radioactive nucleus ^{33}Cl and determine the identity of the heavy product.

The mode of decay:
- alpha
- beta
- positron or EC

Heavy Product:

$A =$ ______________

symbol: ______________

(d) Predict the mode of decay of the radioactive nucleus ^{220}Fr and determine the identity of the heavy product.

The mode of decay:
- alpha
- beta
- positron or EC

Heavy Product:

$A =$ ______________

symbol: ______________

11.3 Kinetics of Radioactivity

Introduction

Radioactive decay is an elementary process involving only one particle, so it follows first order kinetics. In this section, we examine the rate equation and the half-life of radioactive nuclei and then show how nuclear decay can be used to determine the age of materials.

Objectives

- Determine the time required for a given fraction of a radioactive material to disappear given the half-life or rate constant for the decay.
- Determine the age of an organic material given its ^{13}C rate of decay, the rate of ^{13}C decay in living organisms, and the half-life of ^{13}C.

11.3-1. The Rate Law

The first order decay of a nucleus is explained by the integrated rate law for first order kinetics (Equation 10.4), which is reproduced below:

$$\ln\left(\frac{[\mathrm{A}]}{[\mathrm{A}]_0}\right) = -kt$$

[A] and $[\mathrm{A}]_0$ are the concentrations of A at time t and the beginning of the measurement, respectively. Molar concentration is moles per liter, so $[\mathrm{A}] = \frac{n}{V}$ and $[\mathrm{A}]_0 = \frac{n_0}{V}$. The volumes cancel in the ratio, so $\frac{[\mathrm{A}]}{[\mathrm{A}]_0} = \frac{n}{n_0}$, where $n = \frac{m}{\mathrm{M_m}}$ and $n_0 = \frac{m_0}{\mathrm{M_m}}$. The molar masses cancel in the ratio of moles, so the ratio can be expressed as a ratio of moles or masses.

$$\ln\left(\frac{n}{n_0}\right) = \ln\left(\frac{m}{m_0}\right) = -kt \qquad \text{Logarithmic Form of the Integrated Rate Law for Nuclear Decay} \qquad (11.5)$$

Equation 11.5 can be expressed as an exponential instead of a logarithm as follows:

$$n = n_0 e^{-kt} \qquad \text{Exponential Form of the Integrated Rate Law for Nuclear Decay} \qquad (11.6)$$

Radioactive decay is exponential. Radioactive decays are usually characterized by their half-lives rather than their rate constants. Equation 10.6, which relates the half-life to the first order rate constant, is reproduced below:

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

11.3-2. Nuclear Decay Rate Exercise

EXERCISE 11.7:

Magnesium-23 undergoes positron decay. What is the product of the decay?

symbol: ______________

17.9% of a ^{23}Mg remained after 30.0 seconds. What is its half-life?

$t_{1/2} =$ ______________ s

How long would it take for 99.90% of a sample to disintegrate?

$t =$ ______________ s

Radioactive Dating

11.3-3. Introduction

Radioactive dating is the process of determining the age of an object from its radioactive components. It is based on Equation 11.5, which indicates that the time required for some known initial amount of radioisotope to decay to another known amount can be determined if the rate constant (half-life) for the decay is known. Determining the amount of radioisotope present in the object today is straightforward, and we outline the approximations used in two techniques to obtain the initial amounts. One technique is used for historical time scales and the other for geological time scales.

11.3-4. Carbon-14 Dating

Historical ages are frequently determined with carbon-14 dating, which is based on the fact that there is a constant exchange of carbon containing compounds between living organisms and the atmosphere. Atmospheric CO_2 is used in photosynthesis to produce organic compounds that are ingested by animals, and the carbon that was in the CO_2 becomes incorporated into the compounds in the organism. The organism returns some of the carbon back to the atmosphere in the form of CO_2 to continue the cycle. A small fraction of the carbon is in the form of radioactive ^{14}C, which is formed in the upper atmosphere by the following reaction:

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

^{14}C then undergoes beta decay with a half-life of 5730 years ($k = 1.21 \times 10^{-4}$ yr^{-1}).

$$^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e}$$

The two competing processes have resulted in an equilibrium ^{14}C:^{12}C ratio in the atmosphere of about 1:10^{12}, which is also maintained in all living organisms. Consequently, all living organisms show ^{14}C radioactivity of 15.3 disintegrations per minute per gram of carbon (d$\cdot$min$^{-1}\cdot$g^{-1}). However, when the organism dies, it no longer replaces the decaying ^{14}C, and the disintegration rate drops. The age of a material can be estimated by measuring the rate of ^{14}C disintegration by the following method:

1 Determine the ratio $\dfrac{n}{n_0}$ as the ratio of the ^{14}C disintegration rates.

$$\frac{n}{n_0} = \frac{\text{observed rate of disintegration}}{\text{initial rate of disintegration}} = \frac{\text{observed rate of disintegration}}{15.3\ \mathrm{d\cdot min^{-1}\cdot g^{-1}}}$$

2 Use Equation 11.5 with the above ratio and the known rate constant for the decay to determine its age (t).

$$t = -\frac{\ln(n/n_0)}{1.21\mathrm{e}{-04}\ \mathrm{yr^{-1}}}$$

11.3-5. Carbon Dating Exercise

EXERCISE 11.8:

A piece of charred bone found in the ruins of an American Indian village has a ^{14}C disintegration rate of 9.2 d·min^{-1}·g^{-1}. Determine the approximate age of the bone. The rate of decay in living species is 15.3 d·min^{-1}·g^{-1}.

age = ______________ years

11.3-6. Dating Geological Times

Carbon dating can be used to estimate the age of materials that are up to 50,000 years old. The rate of ^{13}C decay for older objects is too slow to be measured. Thus, when a geological age is required, a radioisotope with a much longer half-life must be used. One method used to determine the age of rocks is based on the decay of ^{238}U to ^{206}Pb, a process with a half-life of 4.5 × 10^9 yr. In this method, it is assumed that all of the ^{206}Pb found in the rock originated from ^{238}U, which is represented by the following:

$$\frac{n}{n_0} = \frac{\text{mol } ^{206}\text{Pb in sample}}{(\text{mol } ^{206}\text{Pb} + \text{mol } ^{238}\text{U}) \text{ in sample}}$$

This presumes that none of the lead was in the rock when it was formed, which is an acceptable assumption if there is not much of the more abundant ^{208}Pb present in the sample.

11.4 Nuclear Radiation and Living Tissue

Introduction

When visible light is absorbed by a molecule, the electrons can be excited into excited states, but they soon find their way back to the ground state. Thus, the visible light excited the molecule; but, because the electron was not lost, the radiation is said to be ***nonionizing radiation***. Radio and TV waves, microwaves, and visible light are some examples of nonionizing radiation. However, when radiation of greater energy interacts with the molecule, the electron can be lost to produce an ion. This kind of radiation is said to be ***ionizing radiation***. Ionizing radiation includes alpha and beta particles, gamma rays, and x-rays. Unlike nonionizing radiation, ionizing radiation can be very harmful to living tissue.

Objectives

- Distinguish between ionizing and nonionizing radiation.
- Compare the penetrating power of α-particles, β-particles and γ-rays.

11.4-1. Some Examples

In order for ionizing radiation to be harmful, it must encounter the tissue. Thus, ionizing radiation produced in an experiment conducted in a laboratory next door would have to pass through at least one wall and your clothing before it could harm you. The ability of radiation to pass through material is called its penetrating power. The penetrating power decreases as the mass and charge of the particle increases. Alpha particles are both highly charged and massive, which results in very poor penetrating power. Alpha particles are stopped by a piece of paper or by the layer of dead skin cells covering the body. They can be very damaging to internal organs, but they must be ingested or inhaled to do so.

Approximately 40% of the background radiation to which humans are exposed is produced by radon, which is formed during the decay of ^{238}U to ^{206}Pb. The other members of this decay pathway are also radioactive, but they are solids and remain in the rock. Radon, on the other hand, is a noble gas and can escape from the rock and into our homes. It is a source of alpha particles that has been attributed to up to 10% of lung cancer deaths. As a gas, radon is readily inhaled and, after inhalation, the resulting alpha particles bombard the lung tissue. ^{218}Po is also an alpha emitter ($t_{1/2}$ = 3 minutes), but it is a solid and is not exhaled. Thus, exposure to radon can produce a constant bombardment of lung tissue by alpha particles, which damages the growth-regulation mechanism of the cells, causing the uncontrolled cell reproduction we call cancer.

Beta particles are not as highly charged and not nearly as massive as alpha particles. Consequently, they have greater penetrating power. However, even beta particles are stopped by a sheet of metal or wood. Beta particles can cause damage to the skin and the surface of organs, but they also do their worst damage if ingested or inhaled. Gamma rays are photons and have excellent penetrating power because they have neither charge nor mass. Dense materials like lead or concrete are required to stop gamma rays. Recall that gamma rays are used to carry excess energy away from a nuclear reaction. Consequently, many radioisotopes emit gamma rays. ^{60}Co is a gamma emitter that is used in cancer treatment by bombarding the tumor with gamma rays to destroy the cancerous cells.

11.5 Nuclear Fission

Introduction

Fission reactions are extremely exothermic and are the basis for nuclear power plants (controlled fission) and weaponry (uncontrolled fission). In this section, we examine both the process and its uses.

Objectives

- Describe nuclear fission and chain reactions.
- Define critical mass and explain its origin.
- Explain how fission is controlled in a nuclear reactor.
- Describe what is meant by "meltdown."

11.5-1. Chain Reactions

Nuclear fission is the process of splitting a large nucleus into smaller nuclei. The most common example is the fission of ^{235}U, which uses a neutron to start the reaction, but the reaction then produces three more neutrons.

$$^{235}_{92}\mathrm{U} + \,^{1}_{0}\mathrm{n} \rightarrow \,^{92}_{36}\mathrm{Kr} + \,^{141}_{56}\mathrm{Ba} + 3\,^{1}_{0}\mathrm{n}$$

Reactions like the fission of ^{235}U, in which one of the products of the reaction initiates further reaction, are called ***chain reactions***. Table 11.3 shows the number of product particles produced in the fission of ^{235}U and Figure 11.3 is a schematic of the reaction. In general, 3^n neutrons are produced in the n^{th} step. Thus, in the 10^{th} step, 3^{10} or 59,049 neutrons are produced.

Step	^{141}Ba	^{92}Kr	Neutrons
1	1	1	3
2	3	3	9
3	9	9	27
4	27	27	81

Table 11.3

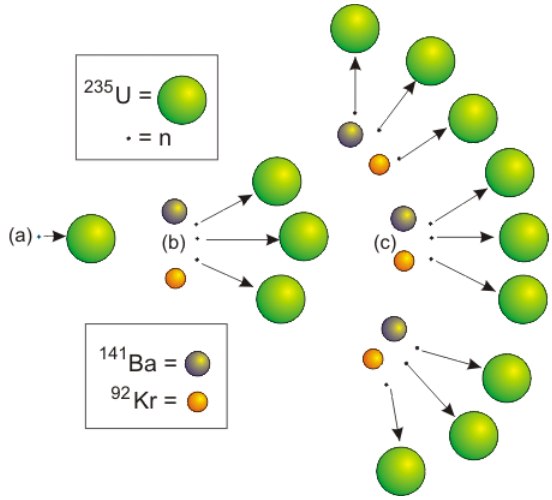

Figure 11.3: Chain Reaction

11.5-2. Rate of Fission Reaction

The ^{235}U fission reaction involves a bimolecular collision between a neutron and a ^{235}U nucleus, so the rate of this elementary reaction is proportional to the product of the two concentrations:

$$\text{rate} = k[\text{n}][^{235}\text{U}] \qquad \text{Rate of Fission} \tag{11.7}$$

where [n] is the concentration of neutrons. As the reaction proceeds, the concentration of neutrons increases faster than the concentration of ^{235}U decreases, which causes the rate of the reaction to increase. Furthermore, each step of the reaction produces three times the energy of the previous step. If it is not controlled, the chain reaction results in an explosion as a vast amount of energy is released in a very short period of time.

Equation 11.7 indicates that the rate of fission can be reduced by reducing either [n] or [^{235}U]. ^{235}U does not undergo a chain reaction in nature because both concentrations are low. The natural abundance of ^{235}U in uranium ore is only 0.7%, which means that [^{235}U] is low. Indeed, the uranium must be enriched to levels of around 4% if it is to serve as a nuclear fuel.

11.5-3. Energy Released by Fission Exercise

EXERCISE 11.9:

Fission reactions have large mass defects, and the large amounts of energy they give off makes them useful in energy production and weaponry. Determine how much energy is released by the fission of 1.00 g of ^{235}U in the following fission reaction:

$$^{235}_{92}\text{U} + {}^{1}_{0}\text{n} \rightarrow {}^{92}_{36}\text{Kr} + {}^{141}_{56}\text{Ba} + 3\,{}^{1}_{0}\text{n}$$

The masses are:

U = 235.0439
Kr = 91.9263
Ba = 140.9144
n = 1.0087

Determine the mass defect for the fission of one mole of ^{235}U.

mass of products = ______________ g   mass of reactants = ______________ g

Δm = ______________ g

Determine ΔE for the fission of one mole of ^{235}U.

$\Delta E =$ _______________ J

How much heat is released in the fission of one gram of ^{235}U?

$\Delta E =$ _______________ J

11.5-4. Critical Mass

Even enriched uranium does not get out of control if the sample size is kept small because many of the neutrons produced in the fission process are near the surface and escape the sample without colliding with other ^{235}U nuclei. However, as the sample size increases, the fraction of neutrons initiating fission increases. The minimum mass of uranium required to maintain a chain reaction is called the ***critical mass***. At the critical mass, one neutron from each fission encounters a uranium nucleus. Masses that are less than the critical mass are said to be subcritical. Subcritical masses cannot sustain a chain reaction because less than one neutron per fission initiates a subsequent fission. Masses in excess of the critical mass are called supercritical. In a supercritical mass, most of the neutrons initiate further reaction. The critical mass of ^{235}U, which depends upon its purity, the shape of the sample, and the energy of the neutrons, ranges from about 15 kg to over 50 kg.

11.5-5. Atomic Bomb

The atomic bomb is an example of uncontrolled fission. The design of the first bomb, shown schematically in Figure 11.4, is quite simple. It is transported with the fissionable uranium divided into two sections, each with a subcritical mass and located at the opposite ends of a large gun barrel. A chemical explosive, TNT, is used to send one subcritical mass into the other. The combined mass of the two samples exceeds the critical mass, and an uncontrolled chain reaction is initiated. The first bomb dropped on Japan at the end of World War II produced an explosion equivalent to 19,000 tons of TNT.

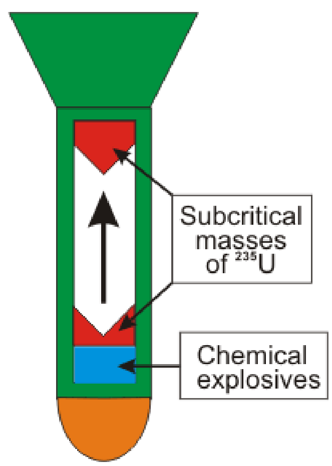

Figure 11.4: Schematic of an Atomic Bomb Chemical explosive (TNT) is used to drive one subcritical mass into another. If the sum of the two subcritical masses exceeds the critical mass, an uncontrolled chain reaction is initiated.

11.5-6. Nuclear Reactor

A nuclear reactor is a controlled chain reaction. A schematic representation of a nuclear reactor is shown in Figure 11.5. Enriched ^{235}U in the form UO_2 is contained in fuel rods that are tubes made of a zirconium alloy. A constant rate of reaction is maintained by varying the height of the control rods, which function by absorbing neutrons. When there is new fuel present, the rods are lowered to capture a greater number of neutrons, but as the fuel is consumed, the rods are raised to increase the number of neutrons available to initiate fission. Heat generated by the nuclear reaction is carried out of the reactor core by high-pressure water (300 °C, 2250 psi) in the primary water loop. Over 30,000 gal/min can flow through this loop in a large reactor. The heat is used to boil water in a steam generator. The escaping steam in a secondary water loop drives a turbine connected to a generator. The steam leaving the turbine is condensed and cooled in the condenser with cooling water from a lake or river.

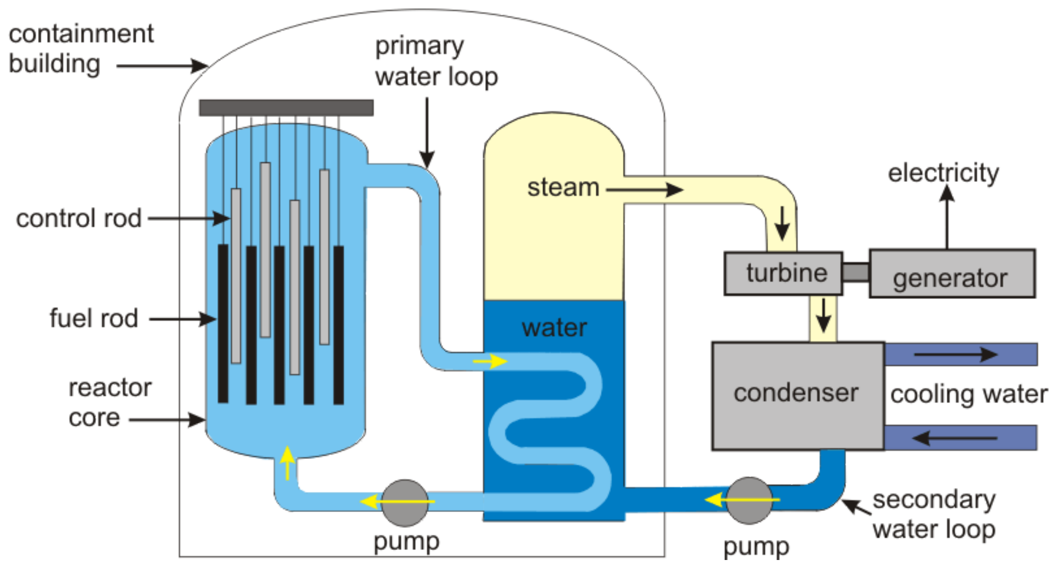

Figure 11.5: Schematic of a Nuclear Power Plant

11.5-7. Concerns

The fuel in a nuclear plant cannot explode like an atomic bomb, but if the reaction gets out of control, the reactor can experience a meltdown. The worst nuclear disaster occurred at Chernobyl in the Ukraine in 1986. Operators disabled the safety system to carry out some tests. During the tests, the reactor cooled and nearly shut down, so, to avoid a costly shutdown, they removed most of the control rods. In the absence of the control rods and with the safety system disabled, the reactor heated beyond safe limits. The excess heat boiled the superheated water and melted the fuel rods, which then mixed with the superheated water. High-pressure steam generated by boiling the superheated water blew off the top of the reactor, and spread the radioactive fuel into the atmosphere. A malfunction of the cooling system was also responsible for the Three Mile Island accident in 1979, but no explosion accompanied that partial meltdown and only a very small amount of radiation was released.

Nuclear reactors are built with many levels of safeguards that have proved effective in preventing accidents except in the case of gross operator error. However, there is one other problem presented by the use of nuclear power. The major concern surrounding nuclear power today is nuclear waste disposal. Not all of the radioactive fuel in the fuel rod can be consumed, and many of the products of the fission reactions are radioactive with long half-lives. Two problems arise: where do you store this radioactive waste, and how do you get it there? Nobody wants to live near a nuclear waste site, and there is major opposition to the transport of radioactive material along our highways and railways. However, a national repository for radioactive waste has been developed at Yucca Mountain, Nevada.

11.6 Nuclear Fusion

Introduction

In ***nuclear fusion***, two lighter atoms combine, or fuse, to form a heavier atom. It is the process that powers the sun and other stars.

Objectives

- Describe nuclear fusion and the problems associated with controlling it.

11.6-1. Deuterium Plus Tritium

As in fission, some of the mass of the fusing nuclei is converted into energy. The most studied fusion reaction is the fusion of deuterium (^{2}H) with tritium (^{3}H) to form helium and a neutron:

$$^{2}_{1}\mathrm{H} + \, ^{3}_{1}\mathrm{H} \rightarrow \, ^{4}_{2}\mathrm{He} + \, ^{1}_{0}\mathrm{n} \qquad \Delta E = -1.7\mathrm{e}09 \text{ kJ}$$

Even with a natural abundance of only 0.015%, deuterium is a readily available isotope because it is present in all water. Tritium atoms can be prepared by bombarding lithium atoms with the neutrons released in the above reaction:

$$^{6}_{3}\mathrm{Li} + \, ^{1}_{0}\mathrm{n} \rightarrow \, ^{3}_{1}\mathrm{H} + \, ^{4}_{2}\mathrm{He} \qquad \Delta E = -4.6\mathrm{e}08 \text{ kJ}$$

The fusion of deuterium and tritium offers almost limitless energy.

11.6-2. Problems

The reason we do not have fusion power plants is that the activation energy for a fusion reaction is enormous. The potential energy of two nuclei as a function of the distance between them rises very sharply at distances less than the bond length. The rise in energy is due to the repulsion between the two positively charged nuclei; but, in order for fusion to occur, this high repulsion energy must be overcome. Consequently, extremely high temperatures are required to bring about fusion. For this reason, fusion reactions are also called ***thermonuclear***. Instead of a critical mass that must be exceeded, fusion reactions have temperatures that must be exceeded. The fusion of deuterium and tritium has the lowest threshold temperature for any fusion reaction, a mere 40,000,000 K! The uncontrolled fusion of deuterium and tritium is called a hydrogen bomb. The threshold temperatures required for the fusion in a hydrogen bomb are achieved by first detonating a fission bomb!

In order to achieve controlled fusion, the nuclei not only have to have sufficient energy to fuse, they must also be held together long enough for fusion to occur. As we shall see in the next section, stars use enormous gravitational fields to both heat the nuclei and to confine them. Scientists on Earth are trying two techniques to produce fusion in the laboratory. In magnetic confinement, the nuclei are confined by a strong magnetic field and heated by powerful microwaves. In inertial confinement, a pellet of frozen hydrogen is compressed and heated by an intense energy beam so quickly that fusion occurs before the atoms can fly apart. Fusion has been achieved in the laboratory, but the nuclei fly apart before a self-sustained reaction can be initiated. Consequently, more energy is pumped into the system than is extracted from it. However, it is expected that fusion reactions that produce more energy than they consume will soon be achieved.

11.7 Origin of the Heavy Elements

Introduction

Nature has mastered fusion in nuclear reactors called stars, and the by-products of these thermonuclear reactions are the elements that populate the periodic table. The universe is comprised mostly of hydrogen, and the story of how the heavier elements came into being is illuminating.

Objectives

- Explain where and how elements are formed.

11.7-1. Birth of a Star

Hydrogen atoms in space are attracted to one another by gravitational forces. As the number of atoms that are attracted to one another increases, the gravitational forces between the atoms also increase, causing the system to begin to collapse. As the body of hydrogen atoms collapses, the pressure at the center begins to build, and the increase in pressure results in an increase in temperature. If there is sufficient mass, the system continues to collapse until the temperature reaches about 4×10^{7} K, at which point the density is about 100 g/cm^{3}. At this temperature, the protons begin to fuse, and a star is born. Further collapse of the star is offset by the enormous energy released by the fusion process, and the star stabilizes as long as the fuel lasts. The overall reaction is:

$$4\,{}^1_1\mathrm{H} \rightarrow\ {}^4_2\mathrm{He} + 2\,{}^0_1\mathrm{e} + 2\,\gamma$$

11.7-2. Red Giant

After about 10% of the hydrogen has been consumed, the core again begins to collapse. When the temperature reaches about 2×10^8 K and the density is around 10,000 g/cm^3, ^{4}He begins to burn:

$$3\,{}^4_2\mathrm{He} \rightarrow {}^{12}_6\,\mathrm{C}$$

The energy released by burning helium expands the hydrogen into a sphere over a hundred times larger than the original star. At this point, the star is called a red giant. When the concentration of ^{12}C gets sufficiently high, it begins to burn and produce other elements.

$$\begin{aligned}
&{}^{12}\mathrm{C} + {}^4\mathrm{He} \rightarrow {}^{14}\mathrm{N} + {}^2\mathrm{H}\\
&{}^{12}\mathrm{C} + {}^4\mathrm{He} \rightarrow {}^{16}\mathrm{O}\\
&{}^{12}\mathrm{C} + {}^{12}\mathrm{C} \rightarrow {}^{24}\mathrm{Mg}\\
&{}^{12}\mathrm{C} + {}^{16}\mathrm{O} \rightarrow {}^{28}\mathrm{Si}\\
&{}^{12}\mathrm{C} + {}^{12}\mathrm{C} \rightarrow {}^{23}\mathrm{Na} + {}^1\mathrm{H}
\end{aligned}$$

11.7-3. White Dwarf

Further collapse and heating produces elements up to ^{56}Fe. Reactions of this type are highly exothermic, but reactions to form elements heavier than ^{56}Fe are endothermic and are produced by neutron capture, which is a very slow process. Thus, once a star contains mostly ^{56}Fe, there is no further nuclear fuel and the star collapses to a **white dwarf** with a radius similar to Earth's and a density of 10^4 to 10^8 g/cm^3. This is the fate that awaits our sun.

11.7-4. Neutron Star

However, if the star is large enough, the collapse continues to even greater densities and temperatures of 4×10^9 K, where many neutron releasing reactions are initiated:

$$^{56}\mathrm{Fe} + \text{ energy } \rightarrow 13\ {}^4\mathrm{He} + 4\ {}^1\mathrm{n}$$

This final collapse occurs in minutes or less with the release of immense amounts of energy and neutrons. The elements in the outer shell of the star absorb many neutrons almost simultaneously and very large masses ($A = 238$) are achieved. The shell is then blown off at near the speed of light in a supernova, leaving a core of many solar masses, a diameter ~10 km, and a density of 10^{14} g/cm^3. At such pressures, electrons are captured by the protons to form neutrons. Eventually, the core consists of nothing but neutrons and is called a **neutron star**. It is interesting to realize that all of the atoms that are heavier than iron were formed in supernovas, which makes a gold necklace all the more interesting.

11.8 Exercises and Solutions

Links to view either the end-of-chapter exercises or the solutions to the odd exercises are available in the HTML version of the chapter.