review

Nuclear Transmutations

A nuclear reaction in which a nucleus transforms into another nucleus as it is **struck by a neutron or by another nucleus**. Such nuclear reactions can be used to synthesize radioisotopes not discovered in nature.

Non-spontaneously nuclear reactions

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

$$^{58}_{26}Fe + ^{1}_{0}n \longrightarrow ^{59}_{26}Fe \longrightarrow ^{59}_{27}Co + ^{0}_{-1}e \longrightarrow ^{60}_{27}Co$$

Synthetic cobalt-60 is radioactive with a half-life of 5.27 years.
It **emits gamma rays with energies of 1.33 MeV** and is used in cancer radiation therapy. (…as also known dirty weapon, **Cobalt bomb,** which would contaminate large areas with ^{60}Co nuclear fallout, rendering them uninhabitable.)

Transmutation Occurring in Nature

Cosmic rays continually bombard the Earth's atmosphere. This bombardment causes many atoms in the upper atmosphere to transmute. These transmutations result in many protons and neutrons being "sprayed out" into the environment. Most of the protons are stopped as they collide with the atoms of the upper atmosphere. These protons strip electrons from the atoms they collide with and thus become hydrogen atoms. The neutrons, however, continue for longer distances because they have no electric charge and therefore do not interact electrically with matter. Eventually, many of them collide with atomic nuclei in the lower atmosphere. A nitrogen atom that captures a neutron, for instance, becomes an isotope of carbon by emitting a proton:

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

$$\boxed{\mathbf{n} \rightarrow \mathbf{p} + \mathbf{e}^-}$$

This carbon-14 isotope, which makes up less than one-millionth of 1 percent of the carbon in the atmosphere, is radioactive and has eight neutrons. (The most common isotope, carbon-12, has six neutrons and is not radioactive.) Because both carbon-12 and carbon-14 are forms of carbon, they have the same chemical properties. Both of these isotopes, for example, chemically react with oxygen to form carbon dioxide, which is consumed by plants through the process of photosynthesis. This means that all plants contain a tiny quantity of radioactive carbon-14.

TABLE 2.2 Some Isotopes of Carbon[a]

Symbol		Number of Protons	Number of Electrons	Number of Neutrons
^{11}C		6	6	5
^{12}C	98.9% stable	6	6	6
^{13}C	1.11% stable	6	6	7
^{14}C		6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Carbon has 15 known isotopes of which $\mathbf{^{12}C}$ and $\mathbf{^{13}C}$ are stable.

Chapter 2

Atomic Mass Unit (u) and Atomic Weight

- Atoms have extremely small masses.
- The heaviest known atoms have a mass of approximately 4×10^{-22} g.
- A **mass scale on the atomic level** is used, where an atomic mass unit (u) is the base unit.
  - **$1\ u = 1.66054 \times 10^{-24}$ g**
- Because in the real world we use large amounts of atoms and molecules, we use **average masses** in calculations.
- An average mass is found using all isotopes of an element weighted by their relative abundances. This is the element's **atomic weight**.
- **Atomic Weight = Σ [(isotope mass) × (fractional natural abundance)] for ALL isotopes.**
- The masses of any atom is compared to C-12

Average Atomic Mass of Carbon

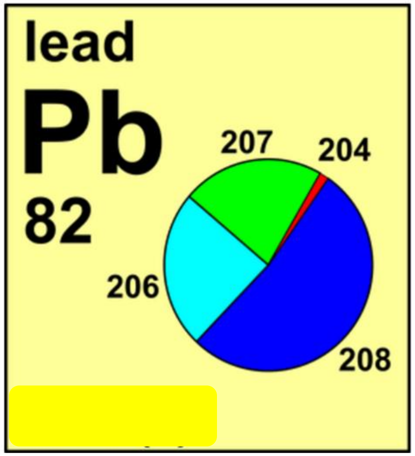
- Natural carbon is a mixture of **^{12}C, ^{13}C, and ^{14}C**
  - Atomic mass of carbon is an average value of these three isotopes
- Composition of natural carbon:
  - ^{12}C atoms (mass = 12 u) - 98.89%
  - ^{13}C atoms (mass = 13.003355 u) - 1.11%

$$\boxed{\begin{aligned} &98.89\% \text{ of } 12\text{ u} + 1.11\% \text{ of } 13.0034\text{ u} \\ &= (0.9889)(12\text{ u}) + (0.0111)(13.0034\text{ u}) \\ &= 12.01\text{ u} \end{aligned}}$$

Exercise

- The element Pb consists of:
  - 1.40% of an isotope with mass 203.973 u
  - 24.10% of an isotope with mass 205.9745 u
  - 22.10% of an isotope with mass 206.9759 u
  - 52.40% of an isotope with mass 207.9766 u

Calculate the average atomic mass for Pb.



Mass Spectrometer

Determines the isotopic composition of natural elements
Determine accurate mass values for individual atoms/molecules

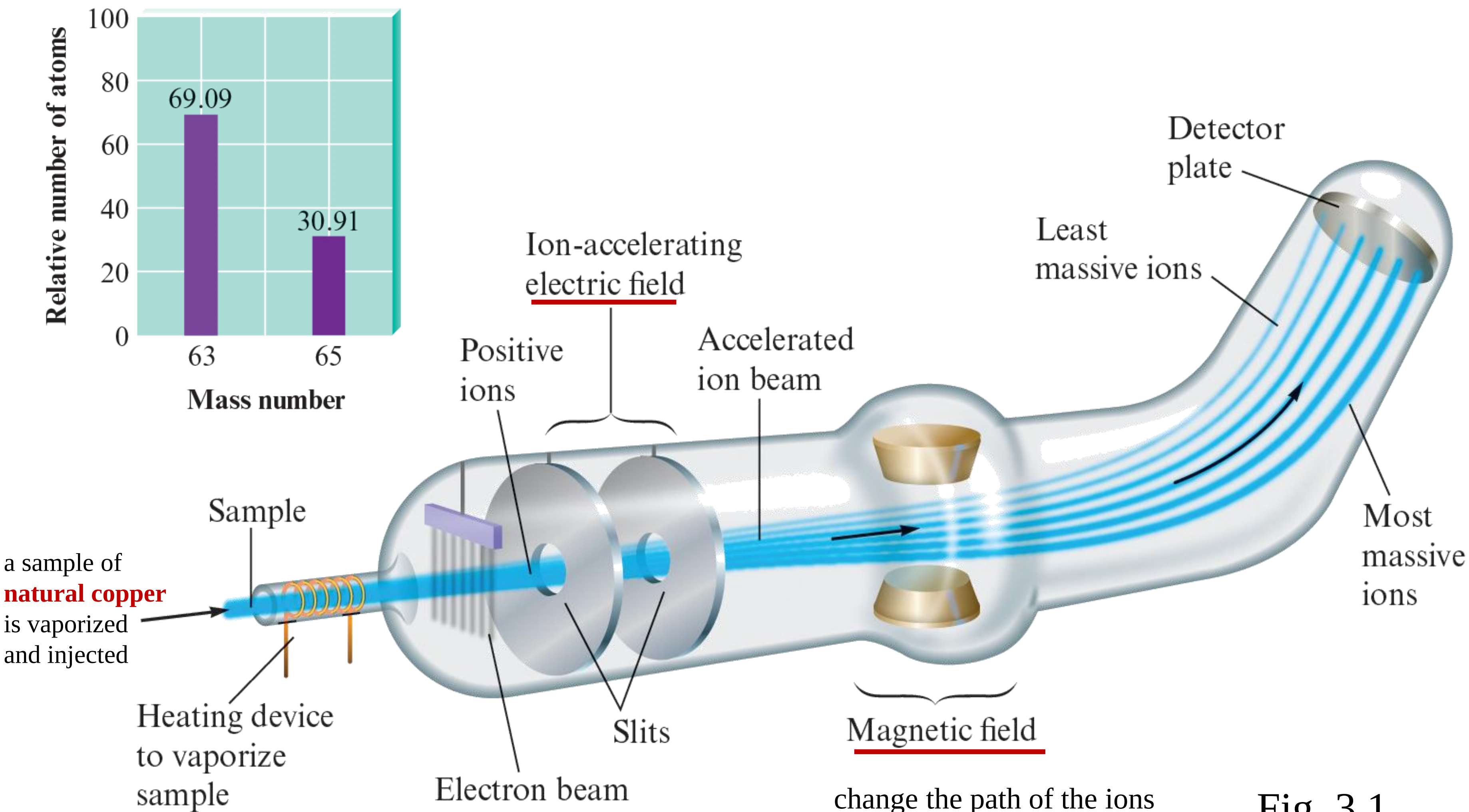
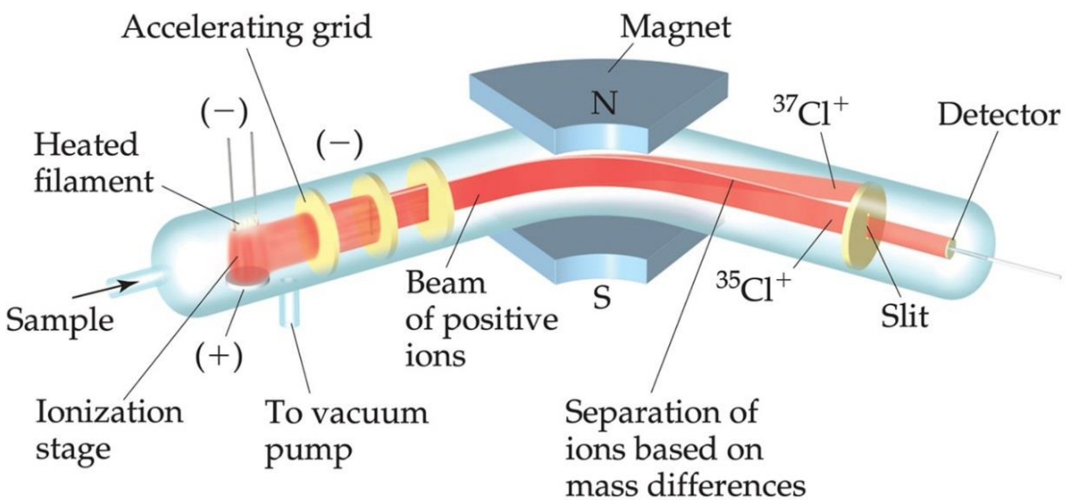


Fig. 3.1



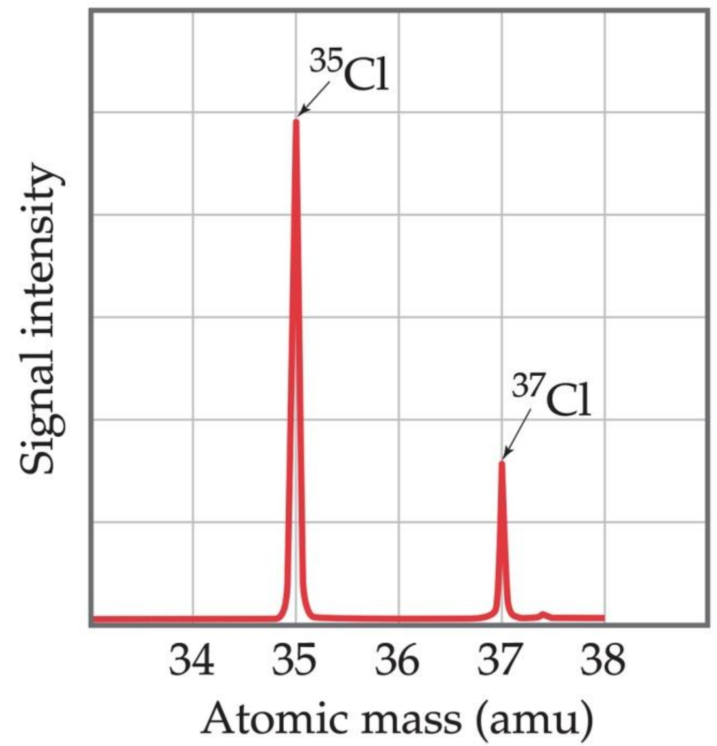
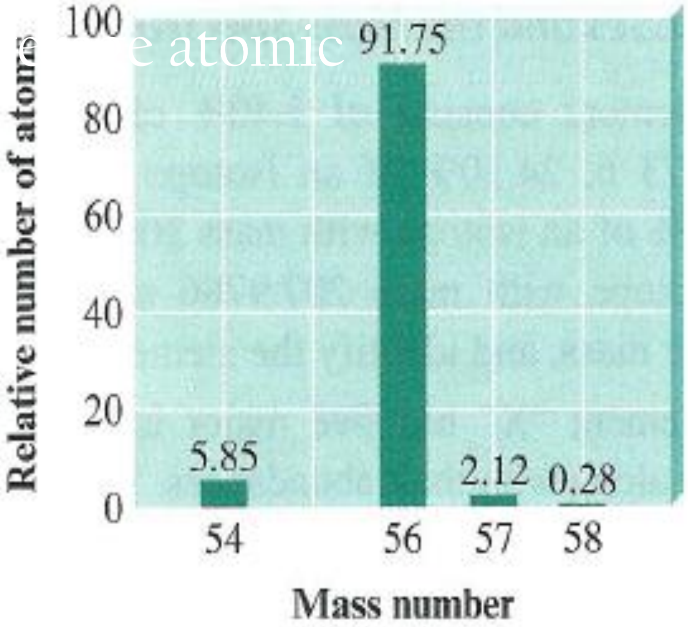


Figure 2.11 A mass spectrometer. Cl atoms are first ionized to form Cl^+ ions, accelerated with an electric field, and finally their path is directed by a magnetic field. **The paths of the ions of the two Cl isotopes diverge as they pass through the field.**

Figure 2.12 Mass spectrum of atomic chlorine. The **fractional abundances of the isotopes ^{35}Cl and ^{37}Cl are indicated by the relative signal intensities** of the beams reaching the detector of the mass spectrometer.

Exercise

The stable isotopes of iron are ^{54}Fe, ^{56}Fe, ^{57}Fe, and ^{58}Fe. The mass spectrum of iron looks like the following:



Radiometric Dating

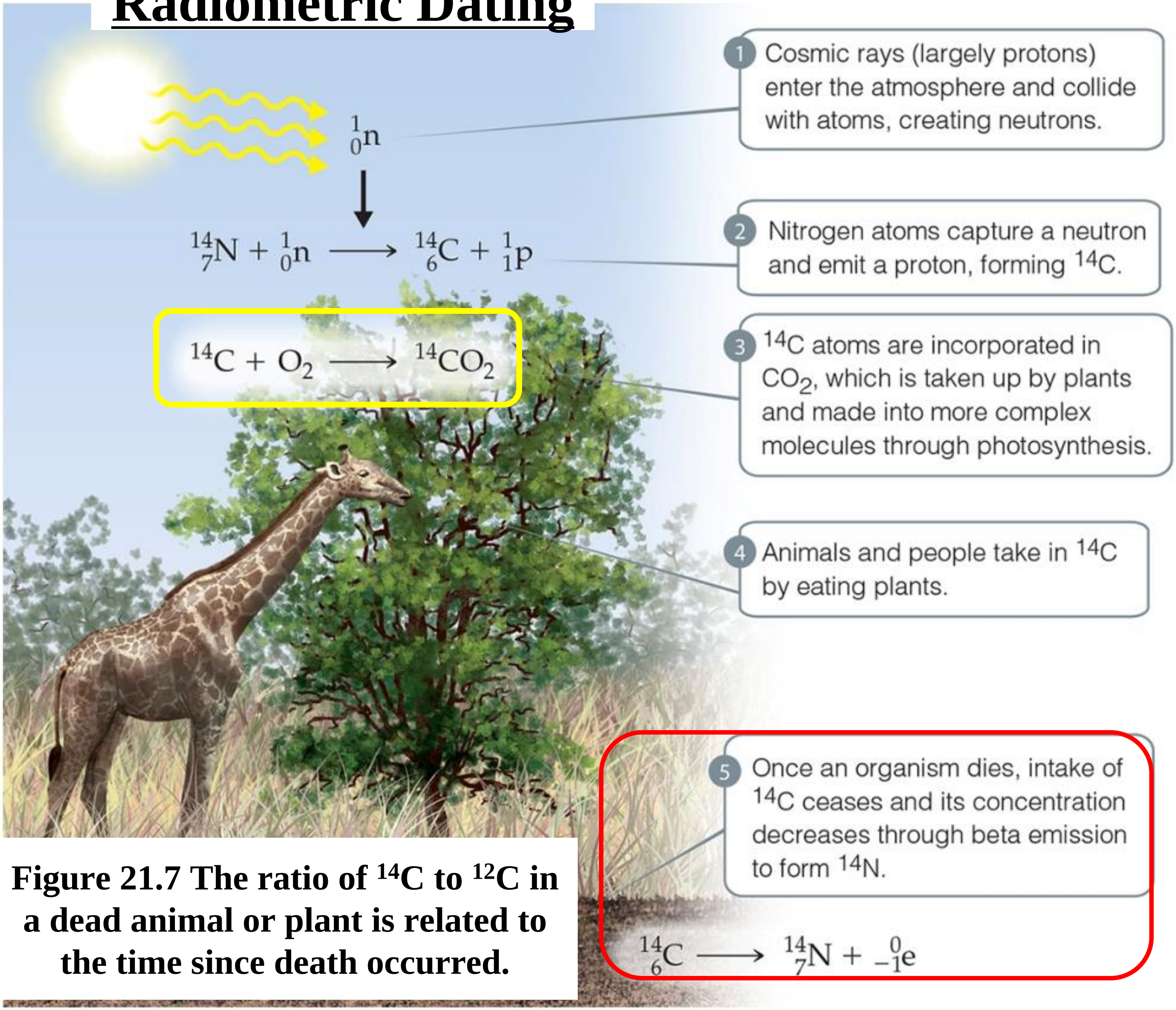


Figure 21.7 The ratio of ^{14}C to ^{12}C in a dead animal or plant is related to the time since death occurred.

Carbon-14 Dating

Origin: C-14 is continuously produced in the atmosphere when high-energy neutrons from space collide with N-14

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

The resulting radiocarbon combines with atmospheric oxygen to form radioactive $^{14}CO_2$, which is incorporated into plants by photosynthesis; animals then acquire ^{14}C by eating the plants. **Carbon-14 nuclide (with a half life of 5730 years) decays by β-particle production.**

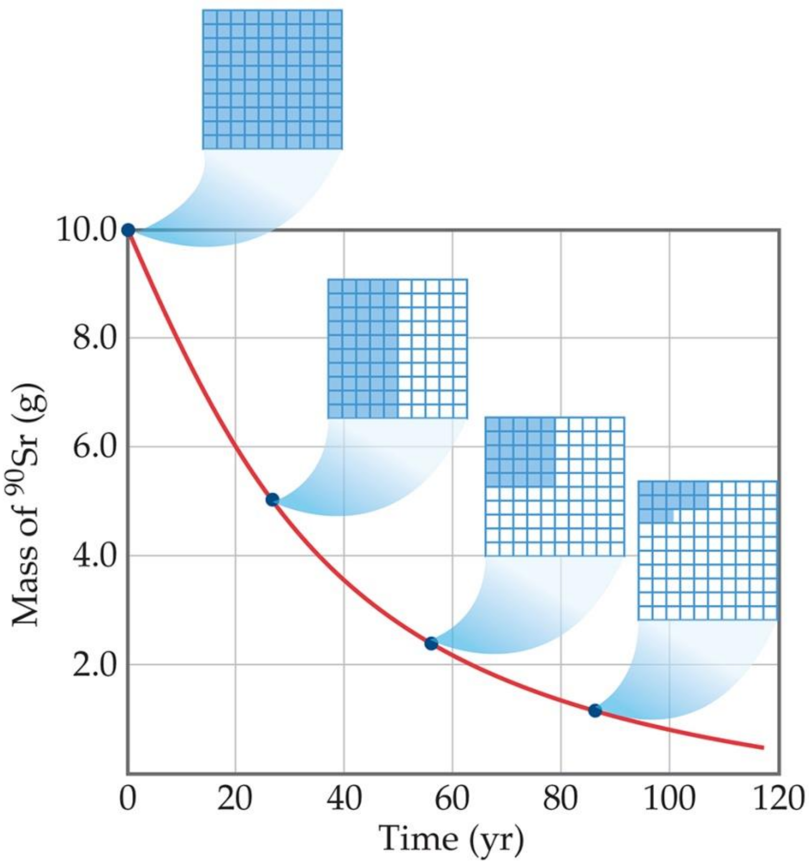
$$^{14}_{6}C \rightarrow ^{0}_{-1}e + ^{14}_{7}N$$

When the animal or plant dies, it stops exchanging carbon with its environment, and from that point onwards the amount of ^{14}C it contains begins to decrease as the ^{14}C undergoes radioactive decay.

- **While still living, the ratio of $^{14}C/^{12}C$ is constant** because the organism replenishes its supply of carbon: CO_2 in air is the ultimate source of all C in an organism.
- Once the organism dies **the ratio of $^{14}C/^{12}C$ decreases**.
- By measuring the **$^{14}C/^{12}C$ ratio** in a once living artifact and comparing it to the **$^{14}C/^{12}C$ ratio** in a living organism, we can tell how long ago the organism was alive.
- The limit for this technique is **50,000 years** old: About 9 half-lives, after which radioactivity from **^{14}C** will be below the background radiation.

Kinetics of Radioactive Decay

- **Radioactive decay is a first-order process.**
- **The kinetics of such a process obey this equation:**



$$\ln \frac{N_t}{N_0} = -kt$$

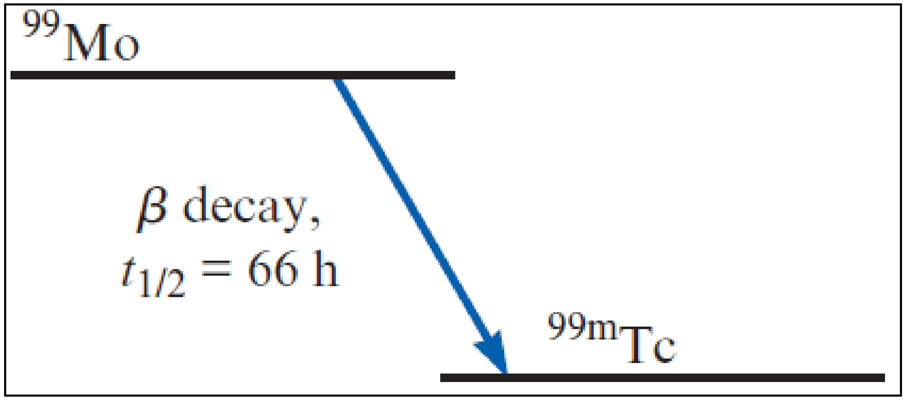
t **-** Time

N_0 **-** Original number of nuclides (at $t = 0$)

N **-** Number of nuclides remaining at time ***t***

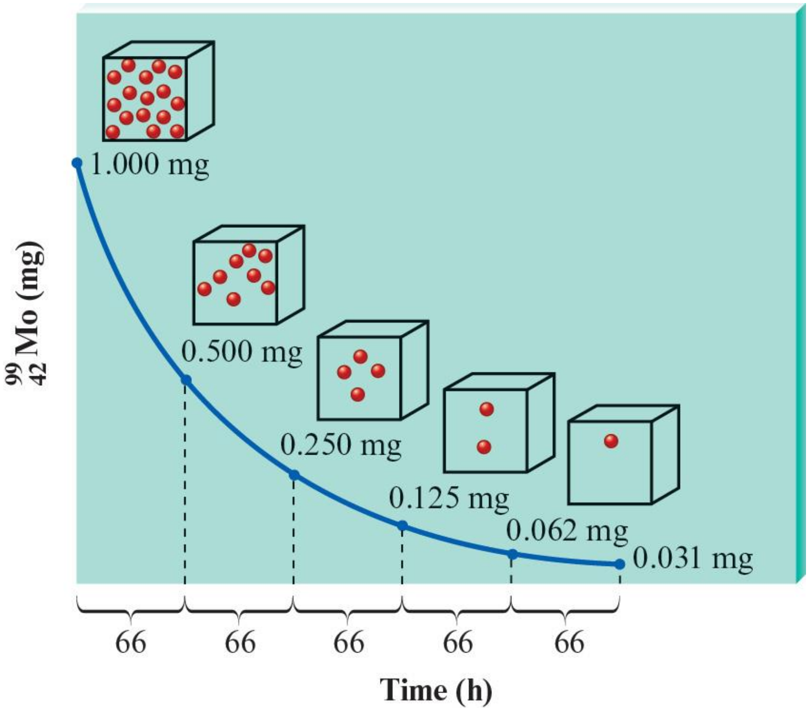
The **half-life** ($t_{1/2}$) of a radioactive sample is defined as the time required for the number of nuclides to reach half of the original value ($N_0/2$).

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



Half-life of molybdenum-99 is 66.0 h

five half-lives for Mo-99



How to determine the age of rocks?

Artifacts derived from nonliving materials can also be dated based on the radioactive minerals they contain. The naturally occurring mineral isotopes uranium-238 and uranium-235, for example, decay very slowly and ultimately become lead—but not the common isotope lead-208. Instead, as was shown in Figure 19.2 uranium-238 decays to lead-206. Uranium-235, on the other hand, decays to lead-207. Thus, the lead-206 and lead-207 that now exist in a uranium-bearing rock were at one time uranium. Older rocks contain higher percentages of these trace isotopes.

If you know the half-lives of uranium isotopes and the percentage of lead isotopes in some uranium-bearing rock, you can calculate the date of the rock's formation. Rocks dated in this manner have been found to be as much as 3.7 *billion* years old. Samples from the Moon have been dated at 4.2 billion years, which is close to the estimated age of our solar system: 4.6 billion years.

***The* half-life of C-14 is 5700 years, *carbon dating* is limited to objects only up to about 50,000 years old; after this time there is too little radioactivity to measure. Other isotopes can be used (U-238:Pb-206 in rock).**

A 1 g sample of carbon from recently living matter contains about 50 trillion billion (5 x 10^{22}) carbon atoms. Of these carbon atoms, about 65 billion (6.5 x 10^{10}) are the radioactive C-14 isotope. This gives the carbon a beta disintegration rate of about eleven decays per minute.

Because of fluctuations in cosmic ray bombardment rates over the centuries, carbon dating has an uncertainty of about 15 percent.

This means, for example, that the straw of an old adobe brick dated to be 500 years old may really be only 425 years old on the low side or 575 years old on the high side. For many purposes, this is an acceptable level of uncertainty.

Question: Suppose that an archaeologist extracts 1 gram of carbon from an ancient ax handle and finds that it is one-fourth as radioactive as 1 gram of carbon extracted from a freshly cut tree branch. About how old is the ax handle?

Energy Changes in Nuclear Reactions

Why are the energies associated with nuclear reactions so large, in many cases orders of magnitude larger than those associated with nonnuclear chemical reactions?

- There is a tremendous amount of energy stored in nuclei.
- Einstein's celebrated equation, $\boldsymbol{E = mc^2}$, relates directly to the calculation of this energy.

***m* must be converted to kilograms, the SI unit of mass, to obtain *E* in joules, the SI unit of energy.**

The **mass changes in chemical reactions** are too small to detect

The mass changes and the associated energy changes in nuclear reactions are much greater than those in chemical reactions.

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th} + 2\,{}^{0}_{0}\gamma$$

When U-238 transmutes (also known as α-particle decay) to two elements, energy is released, partly in the form of the kinetic energy of the alpha particle, partly in the kinetic energy of the thorium atom, and partly in the form of gamma radiation, usually millions of electron-volts (**MeV**)**.**

Where does this energy come from?

The mass change, *m*, is the total mass of the products minus the total mass of the reactants.

233.9942 g + 4.0015 g - 238.0003 g = -0.0046 g

The masses of the nuclei are
U-238, 238.0003 u;
Th-90, 233.9942 u;
and He-4: 4.0015 u.

$$E = \Delta m c^{2} = -4.1 \times 10^{11}\ \mathrm{J}$$

410 billion kJ

The mass change accompanying the radioactive decay of 1 mol of uranium-238, for example, is 50,000 times greater than that for the combustion of 1 mol of CH_4.

Exercise 21.8

How much energy is lost or gained when 1 mol of cobalt-60 undergoes beta decay

The mass of a Co-60 atom is 59.933819 u, and that of a Ni-60 atom is 59.930788 u.

Cobalt-60 nuclide undergoes β^{-} decays

$$^{60}\mathrm{Co} \longrightarrow {}^{60}\mathrm{Ni} + \beta^{-}$$

$$\mathrm{n} \rightarrow \mathrm{p} + \beta^{-} \cdot$$

High **γ-energies release (**amounts to ~20 watts per gram, nearly 30 times larger than that of ^{238}Pu) during the decay process, this result from a significant mass difference between ^{60}Ni and ^{60}Co of 0.003 u.

Mass Defect

An alpha particle has a mass of 4.00150 u.

$^{4}_{2}\mathrm{He}$

$$\text{Mass of two protons} = 2(1.00728\,\mathrm{u}) = 2.01456\,\mathrm{u}$$
$$\text{Mass of two neutrons} = 2(1.00866\,\mathrm{u}) = \underline{2.01732\,\mathrm{u}}$$
$$\text{Total mass} = 4.03188\,\mathrm{u}$$
$$\boxed{\text{Mass difference } \Delta m = 0.03038\,\mathrm{u}}$$

mass defect

The masses of nuclei are always less than the masses of the individual nucleons of which they are composed.

The mass of a nucleon depends upon where it is. In general, a nucleon's mass is greatest when it is free by itself outside of the nucleus. The nucleon's mass is smallest when it is tightly bound within the nucleus.

Nucleus	Mass of Nucleus (u)	Mass of Nucleons (u)	Mass Defect (u)
$^{4}_{2}He$	4.00150	4.03188	0.03038
$^{56}_{26}Fe$	55.92068	56.44914	0.52846
$^{238}_{92}U$	238.00031	239.93451	1.93420

1 u of mass defect = 931.5 MeV

the mass of a nucleon depends upon where it is. In general, a nucleon's mass is greatest when it is free by itself outside of the nucleus. The nucleon's mass is smallest when it is tightly bound within the nucleus.

Nuclear Binding Energy

$$\text{Energy} + {}^{4}_{2}\text{He} \longrightarrow 2\,{}^{1}_{1}\text{H} + 2\,{}^{1}_{0}\text{n}$$

The energy needed to separate a nucleus into its nucleons is called the **nuclear binding energy**.

TABLE 21.7 Mass Defects and Binding Energies for Three Nuclei

Nucleus	Binding Energy (J)	Binding Energy per Nucleon (J)
${}^{4}_{2}\text{He}$	4.53×10^{-12}	1.13×10^{-12}
${}^{56}_{26}\text{Fe}$	7.90×10^{-11}	1.41×10^{-12}
${}^{238}_{92}\text{U}$	2.89×10^{-10}	1.21×10^{-12}

1 MeV = 1.602×10^{-13} J

Values of binding energies per nucleon can be used to compare the stabilities of different combinations of nucleons. **The greater the binding energy per nucleon, the more stable the nucleus.**

Exercise

1. Calculate the binding energy per nucleon (in MeV) for the He-4 nucleus.

2. Calculate the change in energy if 1 mole of oxygen-16 nuclei was formed from neutrons and protons

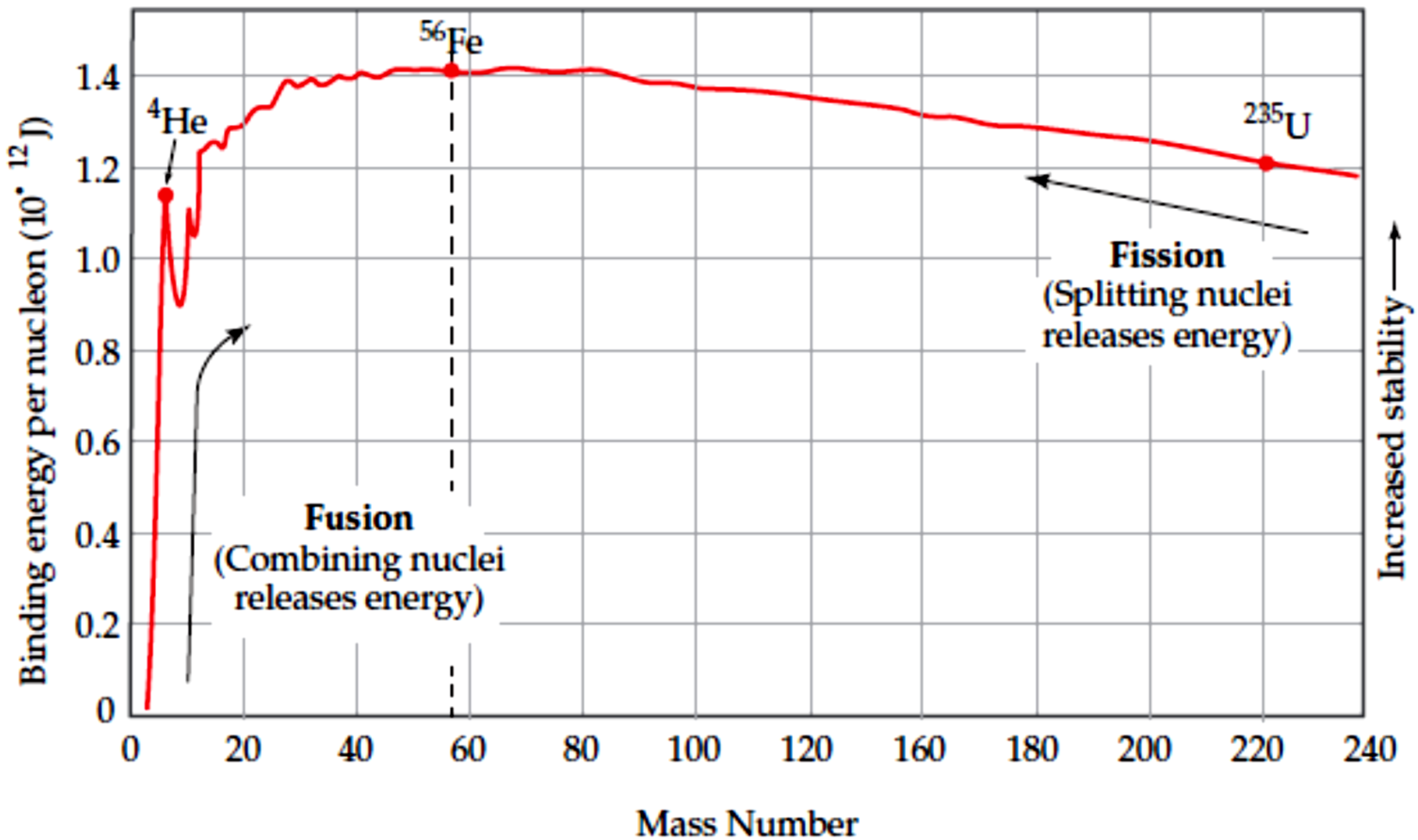
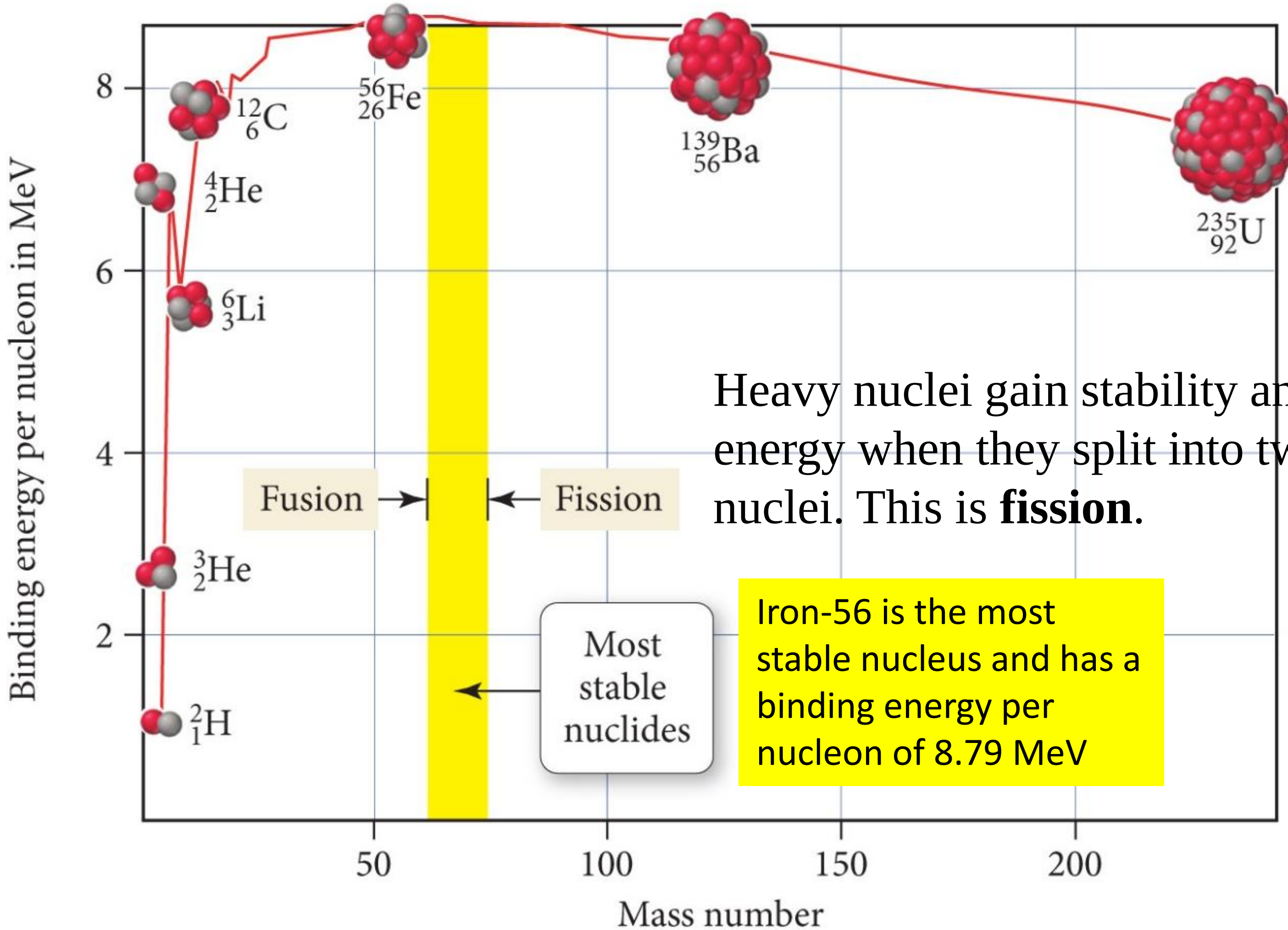


Figure 21.12 Nuclear binding energies. The average binding energy per nucleon increases initially as the mass number increases and then decreases slowly. Because of these trends, fusion of light nuclei and fission of heavy nuclei are exothermic processes.

Effect of Binding Energy on Nuclear Process



Heavy nuclei gain stability and give off energy when they split into two smaller nuclei. This is **fission**.

Iron-56 is the most stable nucleus and has a binding energy per nucleon of 8.79 MeV

Lighter nuclei emit great amounts of energy by being combined in **fusion**.

All nuclear power today is produced by nuclear fission. A more promising long-range source of energy is to be found on the left side of the energy valley, in a process known as nuclear fusion.

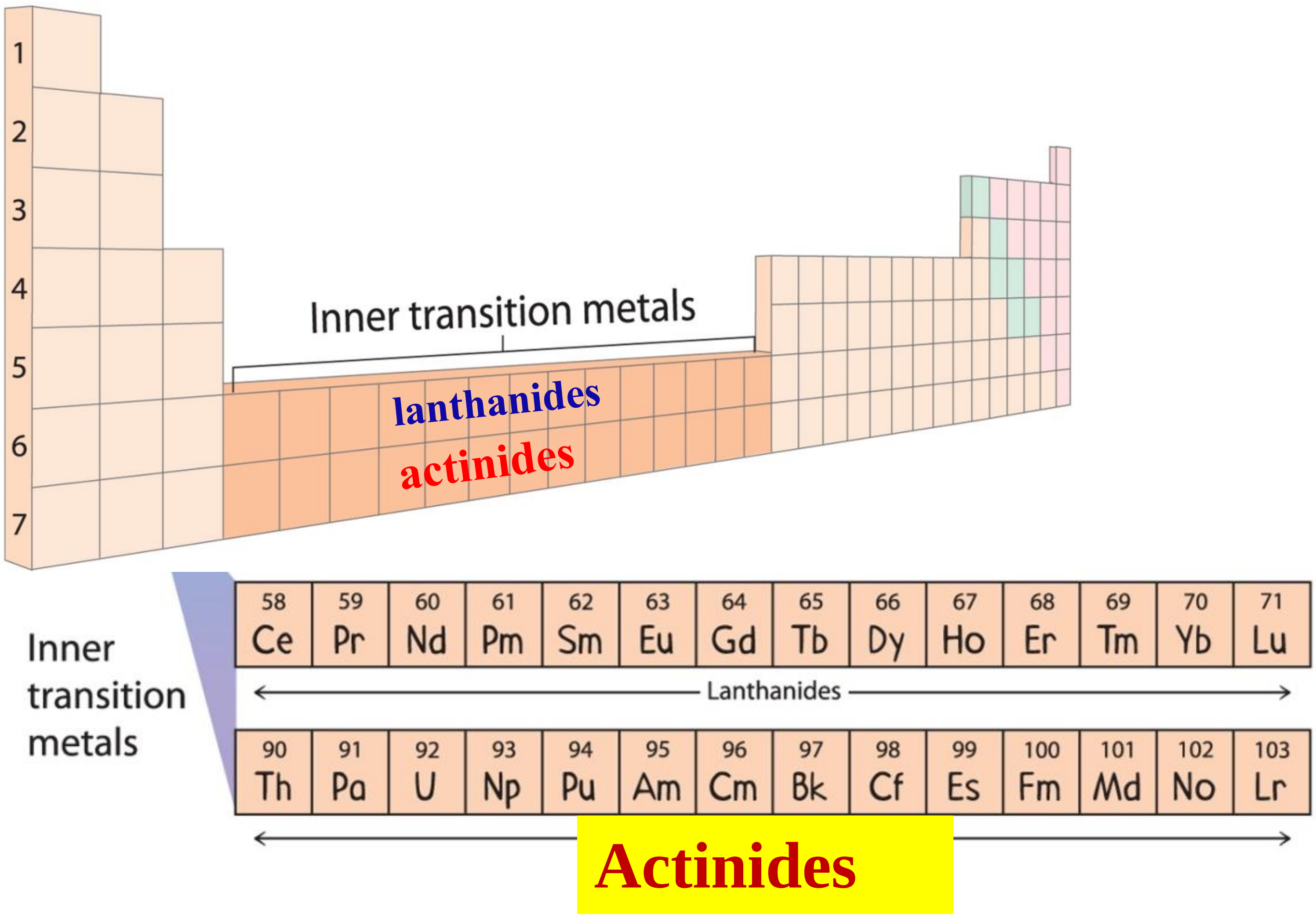
Nuclear Fusion

- **When small atoms are combined, much energy is released. This occurs on the Sun. The reactions are often called thermonuclear reactions.**
- **If it were possible to easily produce energy by this method, it would be a preferred source of energy.**
- **However, extremely high temperatures and pressures are needed to cause nuclei to fuse.**
- **This was achieved using an atomic bomb to initiate fusion in a hydrogen bomb. Obviously, this is not an acceptable approach to producing energy.**

$$^{2}_{1}H + ^{3}_{1}H \longrightarrow ^{4}_{2}He + ^{1}_{0}n$$

13% of worldwide energy comes from nuclear energy, not fusion, but fission.

Actinides


1
2
3
4
5
6
7
Inner transition metals
lanthanides
actinides
Inner transition metals
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
71 Lu
Lanthanides
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No
103 Lr
Actinides

Nuclear Fission

- **Splitting a large nucleus into smaller halves**
- **Releases enormous amounts of energy**

Heavy nuclei can split in many ways. The equations below show two ways U-235 can split after bombardment with a neutron.

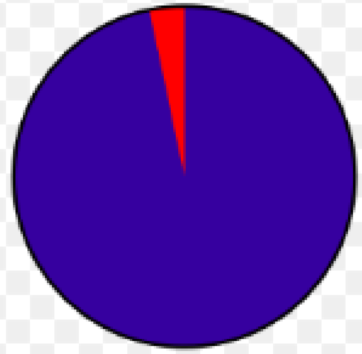
$$^{1}_{0}n + ^{235}_{92}U \rightarrow ^{137}_{52}Te + ^{97}_{40}Zr + 2\,^{1}_{0}n$$

$$^{1}_{0}n + ^{235}_{92}U \rightarrow ^{142}_{56}Ba + ^{91}_{36}Kr + 3\,^{1}_{0}n$$

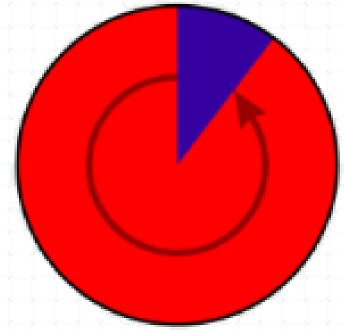
Only 1 part in 140 of naturally occurringuranium is U-235.

Nuclear fission involves the delicate balance between two forces within the nucleus. One force is the *strong nuclear force*, which is a force that holds all the nucleons together. The second force is the repulsive *electric force* occurring among all the like-sign protons. In most nuclei the nuclear strong force dominates. In uranium, however, this domination is weak. If the uranium nucleus is stretched into an elongated shape , the electric forces may push it into an even more elongated shape. If the elongation passes a critical point, the electric forces overwhelm the strong nuclear forces, and the nucleus splits. This

Natural uranium
> 99.2% U-238
0.72% U-235



Low-enriched uranium
(reactor grade)
3-4% U-235



Highly enriched uranium
(weapons grade)
90% U-235

half-life: 703.8 million years

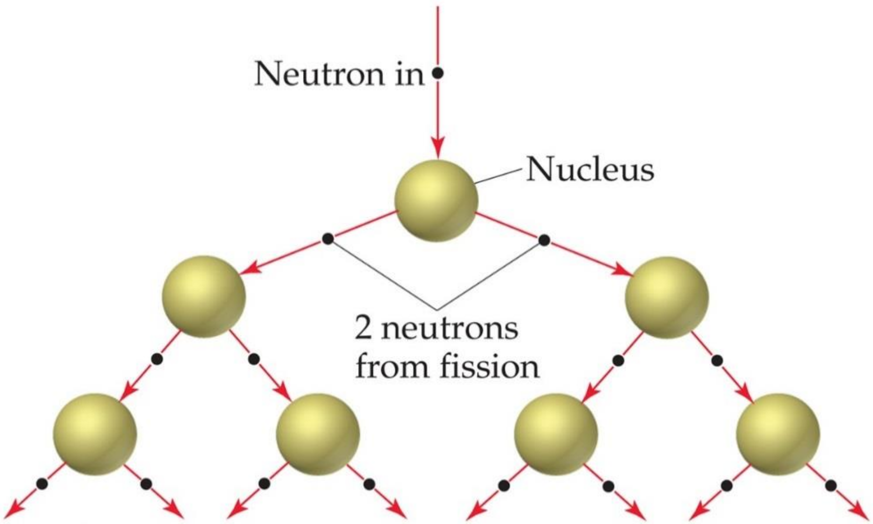
U-238 half life: 4.468 billion years

Constructing a fission bomb is a formidable task. The difficulty is in separating enough uranium-235 from the more abundant uranium-238. Separation is difficult because, except for their slightly different masses, the two isotopes have the same physical and chemical properties. Scientists took more than 2 years to extract enough of the 235 isotope from uranium ore to make the bomb that was detonated at Hiroshima in 1945. To this day, uranium isotope separation, also known as uranium enrichment, remains a difficult process.

Fissionable isotopes include U-235, Pu-239, and Pu-240.

Chain Reaction

- When a product of one reaction begins the next reaction
- ***Neutrons propagate the chain reaction for U-235***



Why do chain reactions not happen in naturally occurring uranium ore deposits?

When the most abundant isotope **U-238 absorbs neutrons** created by U-235, the U-238 typically **does not undergo fission**. So any chain reaction is snuffed out by the neutron absorbing U-238, as well as by the rock in which the ore is imbedded.

$$
{}^{235}_{92}\mathrm{U} \xrightarrow[7.038\times10^{8}\ \mathrm{y}]{\alpha} {}^{231}_{90}\mathrm{Th} \xrightarrow[25.52\ \mathrm{h}]{\beta^-} {}^{231}_{91}\mathrm{Pa} \xrightarrow[3.276\times10^{4}\ \mathrm{y}]{\alpha} {}^{227}_{89}\mathrm{Ac} \left\{ \begin{array}{c} \xrightarrow[21.773\ \mathrm{y}]{98.62\%\beta^-} {}^{227}_{90}\mathrm{Th} \xrightarrow[18.718\ \mathrm{d}]{\alpha} \\ \xrightarrow[21.773\ \mathrm{y}]{1.38\%\alpha} {}^{223}_{87}\mathrm{Fr} \xrightarrow[21.8\ \mathrm{min}]{\beta^-} \end{array} \right\} {}^{223}_{88}\mathrm{Ra} \xrightarrow[11.434\ \mathrm{d}]{\alpha} {}^{219}_{86}\mathrm{Rn}
$$

$$
{}^{219}_{86}\mathrm{Rn} \xrightarrow[3.96\ \mathrm{s}]{\alpha} {}^{215}_{84}\mathrm{Po} \xrightarrow[1.778\ \mathrm{ms}]{\alpha} {}^{211}_{82}\mathrm{Pb} \xrightarrow[36.1\ \mathrm{min}]{\beta^-} {}^{211}_{83}\mathrm{Bi} \left\{ \begin{array}{c} \xrightarrow[2.13\ \mathrm{min}]{99.73\%\alpha} {}^{207}_{81}\mathrm{Tl} \xrightarrow[4.77\ \mathrm{min}]{\beta^-} \\ \xrightarrow[2.13\ \mathrm{min}]{0.27\%\beta^-} {}^{211}_{84}\mathrm{Po} \xrightarrow[0.516\ \mathrm{s}]{\alpha} \end{array} \right\} {}^{207}_{82}\mathrm{Pb}_{\text{(stable)}}
$$

The fission of one atom of uranium-235 generates 202.5 MeV,
~ 2.5 million times more than the energy released from burning coal.

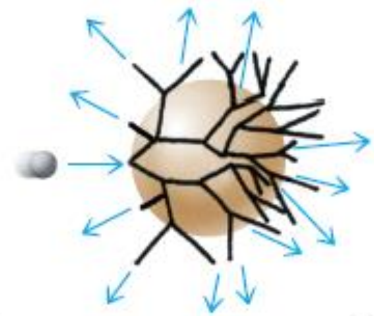
$$
{}^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th}
$$

Thorium-based nuclear power ---better fuel but…

Since 2008, **thorium** in place of uranium has been used to generate nuclear energy for nuclear power.

The deeper you go below the Earth's surface, the hotter it gets. At a depth of merely 30 kilometers the temperature is over 500°C. At greater depths it is so hot that rock melts into magma, which can rise to the Earth's surface to escape as lava. Superheated subterranean water can escape violently to form geysers or more gently to form a soothing natural hot spring. The main reason it gets hotter down below is that the Earth contains an abundance of radioactive isotopes and is heated as it absorbs the radiation from these isotopes. So volcanoes, geysers, and hot springs are all powered by radioactivity. Even the drifting of continents (plate tectonics) is a consequence of the Earth's internal radioactivity.

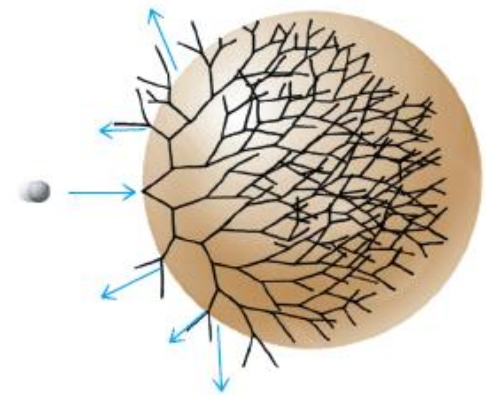
Chain reactions are more effective in large chunks of uranium than in smaller chunks. In smaller chunks, neutrons easily find the surface and escape. As the neutrons escape, the chain reaction no longer builds up.



Neutrons escape small lump of uranium-235.

If less than one neutron causes another fission event (when the sample is too small, too many neutrons escape), the process dies out; the reaction is said to be **subcritical**.

If more than one neutron from each fission event causes another fission event, the process rapidly escalates and the heat buildup causes a **violent explosion**. This situation is described as **supercritical**.

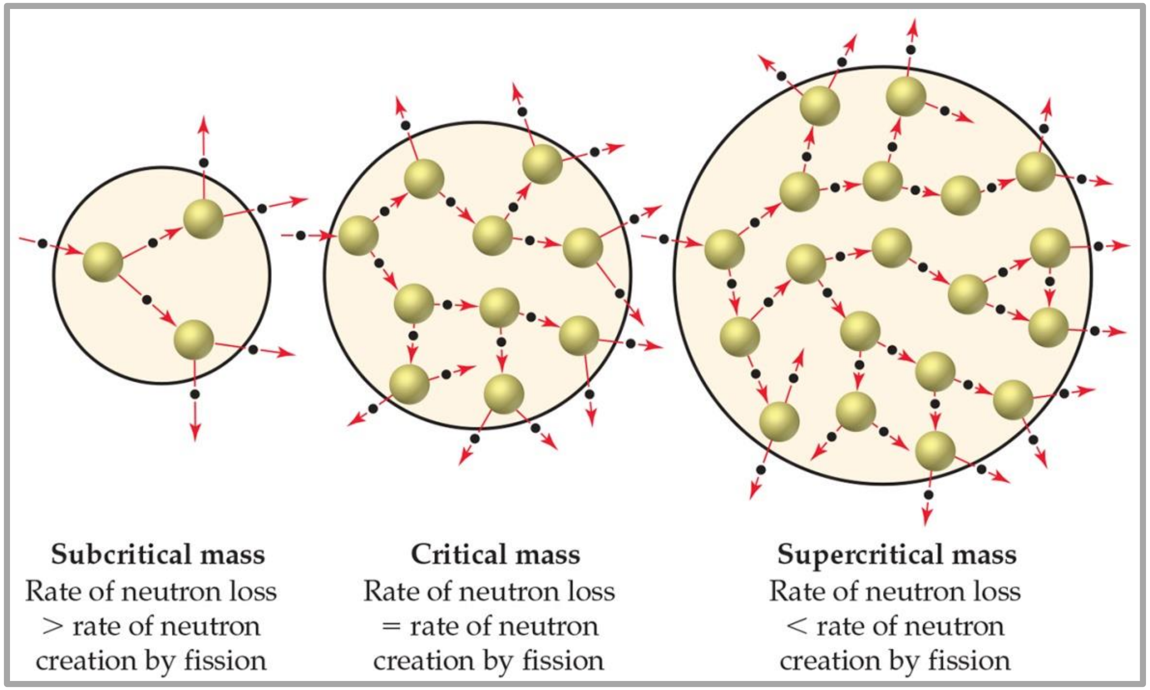


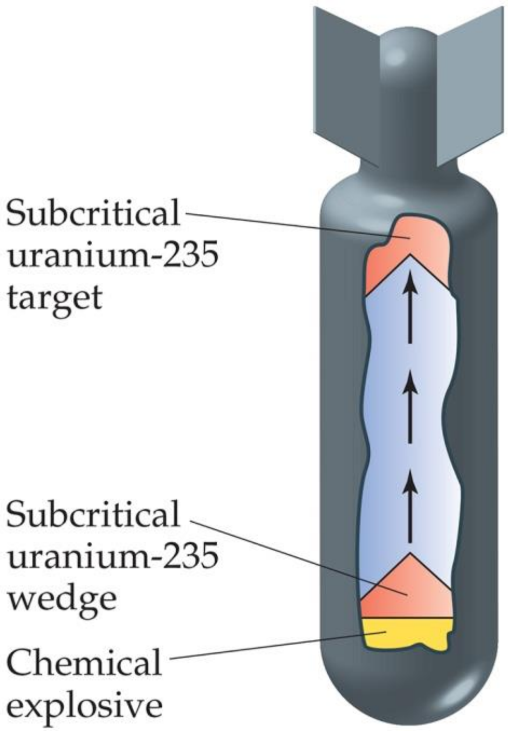
Neutrons trigger more reactions within large lump of uranium-235.

If two small chunks of uranium are suddenly pushed together, they make a larger chunk. Within this larger chunk, neutrons are no longer able to escape as easily.

Critical Mass for Chain Reaction

– The minimum mass that must be present for a chain reaction to be sustained is called the **critical mass**

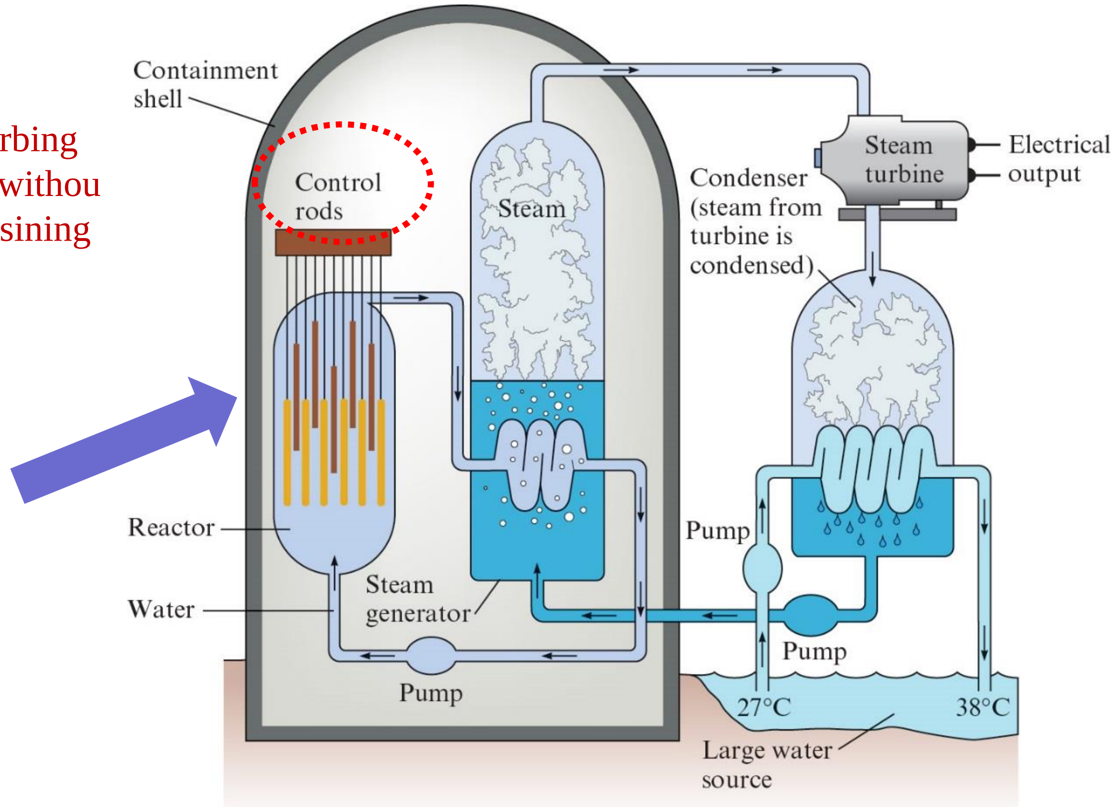




If more than critical mass is present (**supercritical mass**), an explosion will occur. Weapons were created by causing smaller amounts to be forced together to create this mass.

If **exactly one neutron from each fission event** causes another fission event, the process sustains itself at the same level and is said to be **critical**. The critical state requires a <u>certain mass of fissionable material</u>, called the **critical mass.**

capable of absorbing many neutrons without themselves fissining



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

A significant disadvantage of fission power is the generation of radioactive waste products.

Nuclear Power Plants versus Coal-Burning Power Plants

Use about 50 kg of fuel to generate enough electricity for 1 million people

No air pollution

Creat Nuclear waste

Use about 2 million kg of fuel to generate enough electricity for 1 million people

Produce NO_2 and SO_x that add to acid rain

Produce CO_2 that adds to the greenhouse effect

Where Does the Energy from Fission Come From?

1. During nuclear fission, some of the mass of the separate nucleons is converted into energy: $E = m c^2$
2. The difference in mass between the separate nucleons and the combined nucleus is called the **mass defect** (**Δm**).

When a system gains or loses energy, it also gains or loses a quantity of mass:

$$\Delta E = \Delta m\, c^2$$

(Mass of a nucleus is less than that of its nucleons since the process is exothermic)

$${}^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{140}_{56}\mathrm{Ba} + {}^{93}_{36}\mathrm{Kr} + 3\,{}^{1}_{0}\mathrm{n}$$

Mass Reactants		Mass Products	
${}^{235}_{92}\mathrm{U}$	235.04392 amu	${}^{140}_{56}\mathrm{Ba}$	139.910581 amu
${}^{1}_{0}\mathrm{n}$	1.00866 amu	${}^{93}_{36}\mathrm{Kr}$	92.931130 amu
		$3\,{}^{1}_{0}\mathrm{n}$	3(1.00866) amu
Total	**236.05258 amu**		**235.86769 amu**

Mass lost (**Δm**) = 236.05258 amu – 235.86769 amu

Each mole of U-235 that fissions produces about 1.7×10^{13} J of energy

(A very exothermic chemical reaction produces 10^6 J per mole).

Practice

21.9 Indicate the number of protons and neutrons in the following nuclei: (**a**) $^{239}_{94}Pu$, (**b**) $^{142}_{56}Ba$, (**c**) potassium-41.

21.10 Indicate the number of protons and neutrons in the following nuclei: (**a**) $^{214}_{83}Bi$, (**b**) $^{210}_{82}Pb$, (**c**) uranium-235.

21.11 What do these symbols stand for? (**a**) $^{1}_{1}p$, (**b**) $^{0}_{-1}e$, (**c**) $^{0}_{+1}e$

21.12 What do these symbols stand for? (**a**) $^{0}_{0}\gamma$, (**b**) $^{4}_{2}He$, (**c**) $^{1}_{0}n$.

21.13 Write balanced nuclear equations for the following processes: (**a**) radon-198 undergoes alpha emission; (**b**) thorium-234 undergoes beta emission; (**c**) copper-61 undergoes positron emission; (**d**) silver-106 undergoes electron capture.

21.14 Write balanced nuclear equations for the following transformations: (**a**) polonium-210 emits alpha particle; (**b**) neptunium-235 undergoes electron capture; (**c**) fluorine-18 emits beta particle; (**d**) carbon-14 decays by beta emission.

21.17 The naturally occurring radioactive decay series that begins with ${}^{235}_{92}U$ stops with formation of the stable ${}^{207}_{82}Pb$ nucleus. The decays proceed through a series of alpha-particle and beta-particle emissions. How many of each type of emission are involved in this series?

21.18 A radioactive decay series that begins with ${}^{232}_{90}Th$ ends with formation of the stable nuclide ${}^{208}_{82}Pb$. How many alpha-particle emissions and how many beta-particle emissions are involved in the sequence of radioactive decays?

21.29 Complete and balance the following nuclear equations by supplying the missing particle:

(a) ${}^{239}_{94}Pu + {}^{1}_{0}n \longrightarrow {}^{0}_{-1}e + ?$

(b) ${}^{238}_{92}U + {}^{4}_{2}He \longrightarrow 3\,{}^{1}_{0}n + ?$

(c) ${}^{218}_{85}At \longrightarrow {}^{0}_{-1}e + ?$

(d) ${}^{146}_{62}Sm \longrightarrow {}^{142}_{60}Nd + ?$

(e) ${}^{118}_{53}I + {}^{0}_{-1}e \longrightarrow ?$

21.30 Complete and balance the following nuclear equations by supplying the missing particle:

(a) ${}^{106}_{47}Ag + {}^{0}_{-1}e \longrightarrow ?$

(b) ${}^{263}_{106}Sg \longrightarrow {}^{4}_{2}He + ?$

(c) ${}^{216}_{84}Po \longrightarrow {}^{212}_{82}Pb + ?$

(d) ${}^{10}_{5}B + ? \longrightarrow {}^{7}_{3}Li + {}^{4}_{2}He$

Questions

1. Define and make clear distinctions between the terms neutron, nucleon, nucleus, and nuclide

2. What are isotopes? Why do different isotopes of the same element have similar chemistries?

3. Carbon-11 is an unstable isotope. Which kind of radioactive decay would be expected for this isotope?

4. Electron capture transforms K-40 into what nuclide?

5. When the U-235 nuclide is struck by a neutron, the products are two neutrons, Ba-139, and
a) Br-96 (b) Kr-96 (c) Rb-94 (d) Kr-94 (e) Sr-90

6. How does radioactivity allow archeologists to measure the ages of ancient artifacts?