CHM 1220 Chapter 20 - Nuclear Chemistry

Review

Protons: Atomic number (Z) is the number of protons in the nucleus of an element

The atomic number defines the identity of an element. No two elements have the same atomic number

6	← Atomic number
C	← symbol
12.001	← Atomic mass

Elements are ordered in the periodic table according to their atomic number.

Neutrons: Atoms of a given element must all have the same number of protons, but different atoms of the same element can have different numbers of neutrons. They are called isotopes

The difference in the number of neutrons makes usually only a slightly difference to the chemical reactivity of a given element, but often a large difference to its nuclear chemistry.

Mass number (A) = # of protons + # of neutrons

Complete the table

Isotopic symbol $^{A}_{Z}X$	Z	A	# of protons	# of neutrons
$^{13}_{6}C$				
		75	34	
	17			20

20.1-20.2 Nuclear reactions and Radioactivity

Nuclear chemistry is the study of the properties and changes of nuclei (with the except, we will completely ignore the electrons in atoms). Unstable nuclei are spontaneously "radioactive" and "decay through "radiation of one form to another.

*Nuclear reaction involves a change in an atom's nucleus, usually producing a different element. Chemical reaction involves only a change in the way that different atoms are combined. No new elements are produced because atoms are rearranged to form new compound.

*Different isotopes of an element have essentially the same behavior in chemical reaction but often have different behavior in nuclear reactions

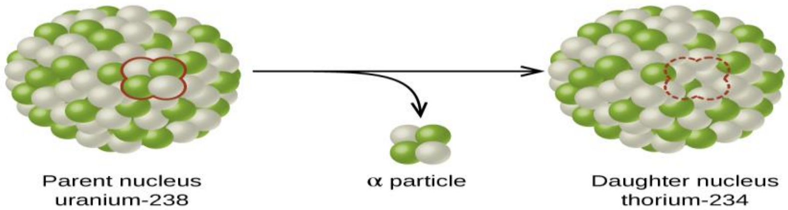
*The energy change produced from a nuclear reaction is far greater than a chemical reaction. The nuclear transformation of 1.0g Uranium-235 releases more than one million times as much energy as the chemical combustion of methane

Other properties

- Radioactive rays can ionize matter.
  - ✓ Cause uncharged matter to become charged
- Radioactive rays have high energy.
- Radioactive rays can penetrate matter.
- Radioactive rays cause phosphorescent chemicals to glow.

Radiation: the spontaneous disintegration of a nucleus resulting in emission of particles or electromagnetic radiation

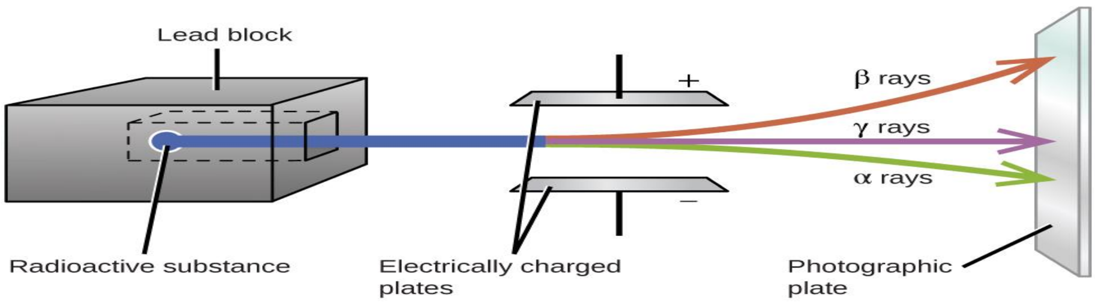
- The spontaneous change of an unstable nuclide into another is **radioactive decay**. The unstable nuclide is called the **parent nuclide**; the nuclide that results from the decay is known as the **daughter nuclide**. The daughter nuclide may be stable, or it may decay itself
- All isotopes of elements with $Z > 83$ are radioactive (unstable)



Types of Radiation

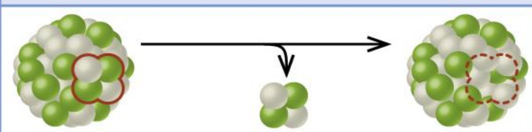
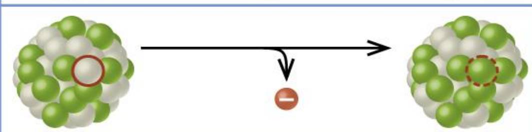
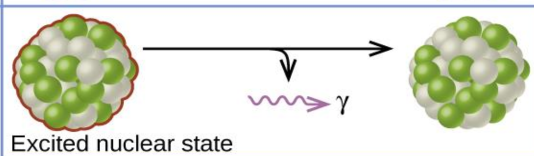
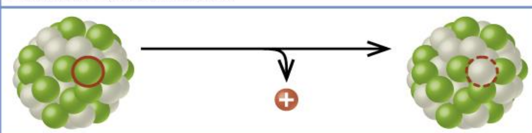
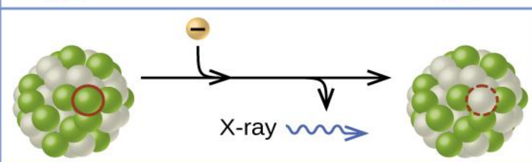
Ernest Rutherford's experiments involving the interaction of radiation with a magnetic or electric field helped him determine that one type of radiation consisted of positively charged and relatively massive α particles; a second type was made up of negatively charged and much less massive β particles; and a third was uncharged electromagnetic waves, γ rays

Alpha particles are positively charged and thus bend towards the negative plate. Beta particles are negatively charged and thus bend towards the positive plate. Gamma particles have no charge and are thus not affected by the charged plates. The degree of "bending" is related to the mass of the particle. Alpha particles are heavier than beta particles and thus are less easily moved in space.



- You must memorize them
- Do not use the alternative symbols (e.g α for alpha or β- for beta emission) included in your text
- Gamma emission and electron capture involve emission of electromagnetic radiation while all other types involve emission of particles

- **Electron Capture** is a process in which the nucleus captures an inner-shell electron, thereby converting a proton into a neutron.
    - Electron capture has the effect of positron emission, converting a proton to a neutron.

Type	Nuclear equation	Representation	Change in mass/atomic numbers
Alpha decay	${}^{A}_{Z}X \rightarrow {}^{4}_{2}He + {}^{A-4}_{Z-2}Y$		A: decrease by 4 Z: decrease by 2
Beta decay	${}^{A}_{Z}X \rightarrow {}^{0}_{-1}e + {}^{A}_{Z+1}Y$		A: unchanged Z: increase by 1
Gamma decay	${}^{A}_{Z}X \rightarrow {}^{0}_{0}\gamma + {}^{A}_{Z}Y$		A: unchanged Z: unchanged
Positron emission	${}^{A}_{Z}X \rightarrow {}^{0}_{+1}e + {}^{A}_{Z-1}Y$		A: unchanged Z: decrease by 1
Electron capture	${}^{A}_{Z}X + {}^{0}_{-1}e \rightarrow {}^{A}_{Z-1}Y$		A: unchanged Z: decrease by 1

Radiation protection requires

- paper and clothing for alpha particles.
- a lab coat or gloves for beta particles.
- a lead shield or a thick concrete wall for gamma rays.

- limiting the amount of time spent near a radioactive source.
- increasing the distance from the source.

TABLE 19.3 Some Properties of Radiation

Type of Radiation	Energy Range	Penetrating Distance in Water[a]
α	3–9 MeV	0.02–0.04 mm
β	0–3 MeV	0–4 mm
X	100 eV–10 keV	0.01–1 cm
γ	10 keV–10 MeV	1–20 cm

[a]Distances at which one-half of the radiation has been stopped.

Nuclear Equations: A way to describe the radioactive processes and transformations that nuclei undergo. Like chemical equation, they must be balanced.

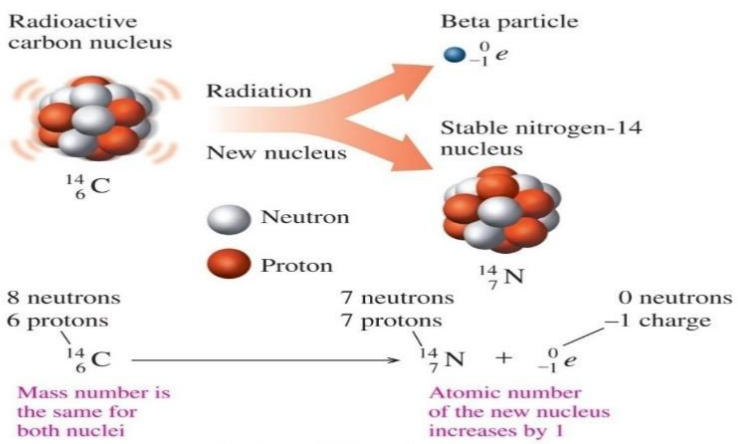
The algebraic sum of the subscripts must be the same on both side of the equation, and the algebraic sum of the superscripts must be the same on both side of the equation.

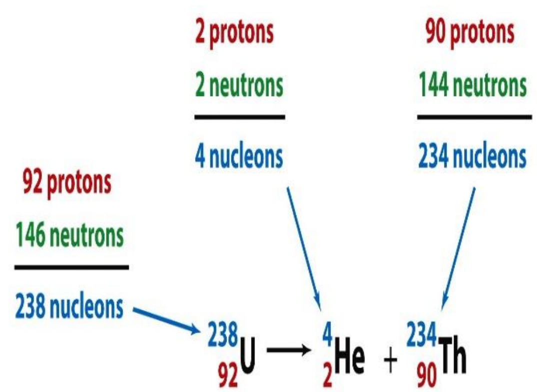
General equation

Unstable nucleus → new stable nucleus + emitted radioactive particles + energy

Except

Unstable nucleus + absorbed electron → new stable nucleus (e-capture)



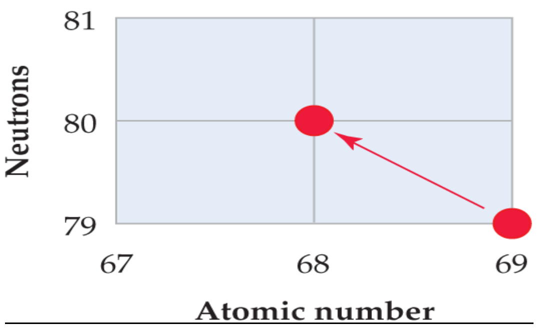


Examples

- Write a balanced nuclear equation for alpha decay of $^{208}_{84}Po$

- The isotope Fe-59 decays by a beta emission

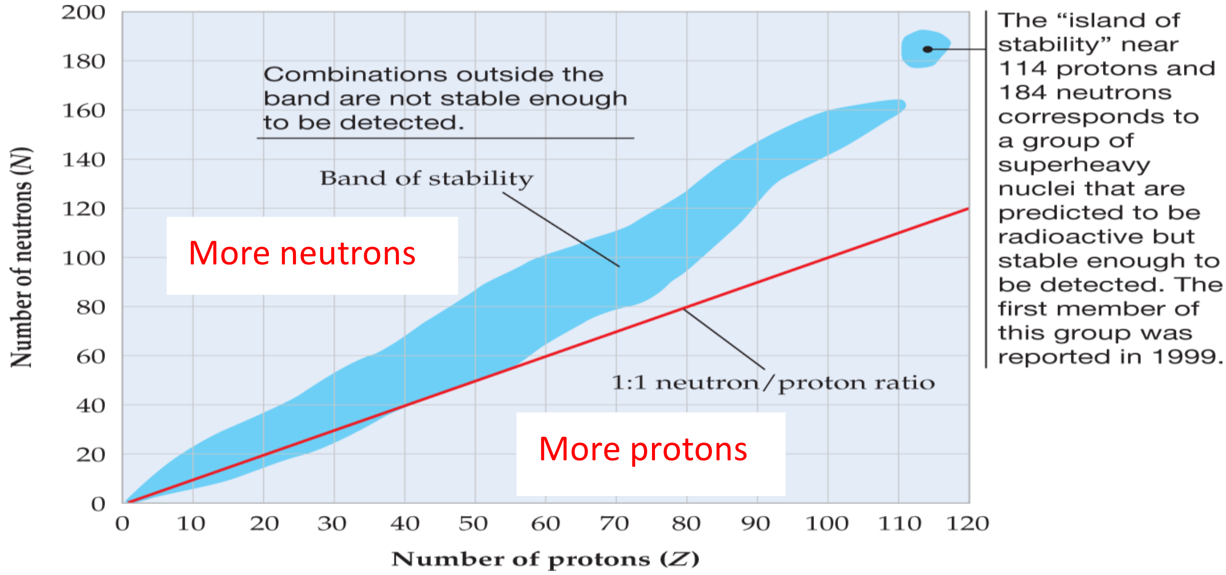
Write a balanced nuclear reaction for the reaction below



20.3 Nuclear Stability and predicting the type of Radioactivity

We want to know why there is radioactivity. What makes the nucleus a stable one? There are no concrete theories to explain this, but there are only general observations based on the available stable isotopes. It depends on a variety of factors and no single rule allows us to predict whether a nucleus is radioactive and might decay unless we observe it. There are some observations that have been made to help us make predictions

- Every element in the periodic table has at least one radioactive isotope.
- Hydrogen is the only element whose most abundant stable isotope, hydrogen-1, contains more protons (1) than neutrons (0).
- The ratio of neutrons to protons gradually increases for elements heavier than calcium.
- All isotopes heavier than bismuth-209 are radioactive, even though they may occur naturally.
- The "band of stability" allows predictions of the type of radioactivity that unstable nuclei will undergo. There are 3600 nuclides in the "band", but only 264 are indefinitely stable (non-radioactive).
- Nuclei above the belt of stability have neutron rich isotopes. These can lower their ratio and move to the belt of stability by emitting a beta particle. This increases number of protons and decreases neutrons
- Nuclei below the belt of stability are protons rich. Nuclei increase their neutrons and decrease protons by positron emission (more common in lighter nuclei) and electron capture (more common in heavier nuclei)
- Nuclei with atomic numbers greater than 84 tend to undergo alpha emission. This emission decreases the number of neutrons and protons by 2 moving the nucleus diagonally toward the belt of stability
- Smaller nuclei ($Z < 20$) with a 1:1 ratio of neutrons to protons are often stable but as the nucleus size increases the ratio of neutron to proton increases for stable nuclei (minimizing proton-proton repulsion).



Neutron to proton ratio
It appears that neutron to proton (n/p) ratio is the dominant factor in nuclear stability.

Strong nuclear force exists between nucleons. The more protons packed together the more neutrons are needed to bind the nucleus together.

Atomic nucleus with an atomic number up to twenty has almost equal number of protons and neutrons. Nuclei with higher atomic numbers have more neutrons to protons. The number of neutrons needed to create a stable nucleus increase more than the number of protons

Nuclei that lie above the band of stability need to decrease the $\frac{\text{neutron}}{\text{proton}}$ ratio

1) Neutron emission: $^{137}_{53}I \longrightarrow {}^{136}_{53}I + {}^{1}_{0}n$ (Loses neutron only. This rarely occur)

2) Beta emission: $^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{-1}e$ (converting neutron within nucleus into a proton, thereby increasing the atomic number of nucleus by one)

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

Nuclei that lie below the band of stability need to increase the $\frac{\text{neutron}}{\text{proton}}$ ratio:

1) Positron emission $^{14}_{6}C \longrightarrow {}^{14}_{5}B + {}^{0}_{1}e$ (loses protons.)

2) Electron capture $^{7}_{4}Be + {}^{0}_{-1}e \longrightarrow {}^{7}_{3}Li$ (loses proton)

Alpha emission $^{226}_{88}Ra \longrightarrow\ ^{222}_{86}Rn + ^{4}_{2}He$ (Common for heavy nuclides. Loses two protons and two neutrons, but actually increases neutron/proton ratio for heavy nuclei with n>p

3)

Even-Odd rules

- Elements with an even atomic number have larger numbers of nonradioactive isotopes than do elements with odd atomic numbers

Z-proton	N-neutron	# of stable isotopes
Even	Even	163 (most stable)
Even	Odd	53
Odd	Even	50
Odd	Odd	4 (least stable)

Magic numbers

There is a special stability associate with having a "magic" number for each (in particular, the number of electrons associated with the noble gases). Magic numbers are natural occurrences isotopes and are stable.

Proton: 2, 8, 20, 28, 50, 82, 114

Neutron: 2, 8, 20. 50, 82, 126, 184

In general, nuclear stability is greater for nuclei containing even numbers of protons and neutrons or both.

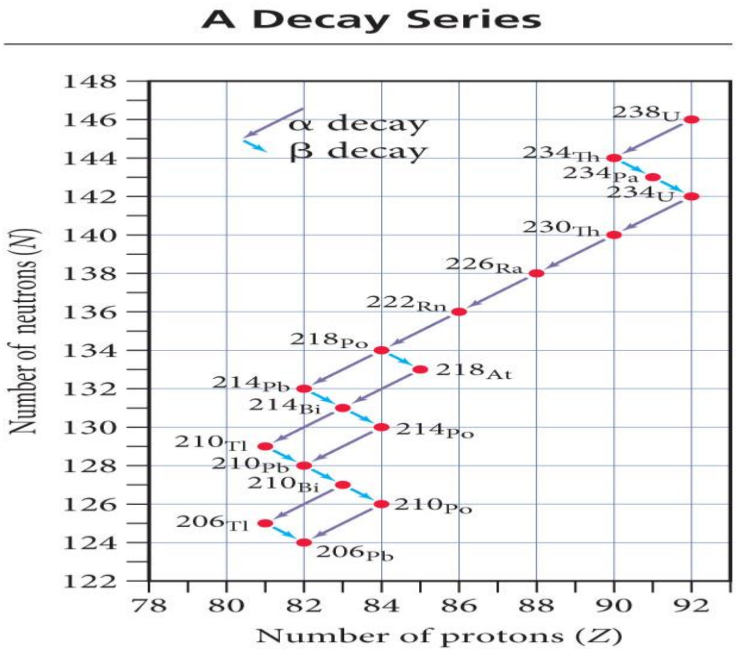
E.g What make these nuclei different from each other?

Sn (Z = 50) has 10 stable isotopes while In (Z = 49) and Sb (Z =51), each has two isotopes.

Example: Using the above chart stable to determine whether the following is stable or unstable. If not then which radioactive particle will be emitted.

a) $^{40}_{20}Ca$ b) $^{210}_{84}Po$ c) Mg-22

Radioactive series:



Atoms with Z > 83 are radioactive and decay in one or more steps involving the primarily alpha and beta decay (with some gamma decay to carry away excess energy).

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

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

Protacium-234 is also radioactive, decaying via alpha and beta until a stable nuclei, Pb-206, is reached.

20.6 Energy Changes During Nuclear Reactions:

As a simple example of the energy associated with the strong nuclear force, consider the helium atom composed of two protons, two neutrons, and two electrons. The total mass of these six subatomic particles may be calculated as:

$$(2p\times 1.0073 \text{ amu}) + (2n\times 1.0087 \text{ amu}) + (2e- \times 0.00055 \text{ amu}) = 4.0331 \text{ amu}$$

However, mass spectrometric measurements reveal that the mass of an e atom is 4.0026 amu, less than the combined masses of its six constituent subatomic particles.

Mass Defect: The loss in mass that occurs when protons and neutrons combine to form a nucleus. The loss in mass is converted into energy that is released during the nuclear reaction and is thus a direct measure of the **binding energy** holding the nucleons together

Δm = mass of nucleons– actual mass

Find the mass defect of a copper-63 nucleus if the actual mass of a copper-63 nucleus is 62.914 amu.

Copper has 29 protons and copper-63 also has (63 - 29) 34 neutrons.

The mass of a proton is 1.00728 amu and a neutron is 1.00867 amu.

- The combined mass or mass of nucleons is calculated:

- Calculate the mass defect.

Δm =

Nuclear Binding energy of a nucleus is the energy needed to break a nucleus into its individual protons and neutrons. The larger binding energy, the more stable the nucleus

Conversion of Mass Defect into Energy

To convert the mass defect into energy:

- Convert the mass defect into kilograms (1 amu = 1.6606×10^{-27} kg)
- Convert the mass defect into its energy equivalent using Einstein's equation, $\Delta E = \Delta mc^2$.

Example: Determine the binding energy of the copper-63 atom.

- Convert the mass defect (calculated in the previous example) into kg.

- Convert this mass into energy using $\Delta E = \Delta mc^2$, where c = 3.00×10^8 m/s.

The energy calculated in the previous example is the nuclear binding energy. However, nuclear binding energy is often expressed as kJ/mol of nuclei or as MeV/nucleon.

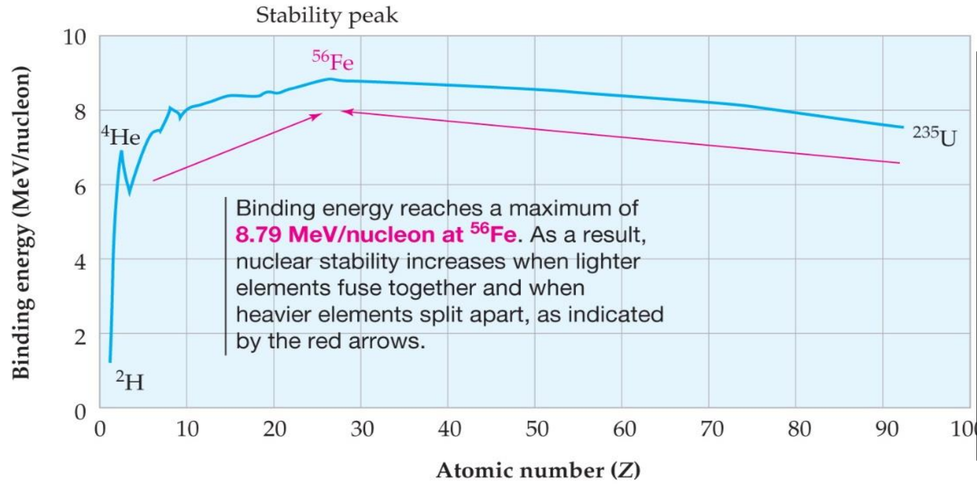
- **To convert the energy to kJ/mol of nuclei** we will simply employ the conversion factors for converting joules into kilojoules (1 kJ = 1000 J) and for converting individual particles into moles of particles (Avogadro's Number).

To convert the binding energy to MeV (megaelectron volts) per nucleon we will employ the conversion factor for converting joules into MeV (1 MeV = 1.602×10^{-13} J) and the number of nucleons (protons and neutrons) which make up the nucleus.

Example: Determine the binding energy per nucleon for Osmium-190.

mass of ^{190}Os = 189.95863 amu mass of neutron = 1.008664 amu mass of proton = 1.007276 amu

The binding energy per nucleon for most nuclei falls around 1.3 x 10^{-12} J (8 MeV). A plot of binding energy per nucleon vs. mass number shows that the most stable nuclei occur around A = 50 to 60. **Fe-56 is the most stable nucleus in the universe.** The existence of a maximum in this curve indicates that energy is released in either a fission or fusion process in which more stable nuclei (i.e., closer to ^{56}Fe) are produced.



Because of the interconversion of mass and energy, the law of conservation of mass and the law of conservation of energy must be *combined*. The combination of mass and energy must be conserved

Binding energy per nucleon generally increases from small atoms to atoms with mass number around 56. Thus fusing atoms to form medium-size atoms (**nuclear fusion**) release energy.

Binding energy per nucleon generally decreases from atoms with mass number around 56 and larger. Thus splitting larger atoms to form medium-sized atoms (**nuclear fission**) also releases energy.

20.4 to 20.5 Radioactive decay rates - Dating

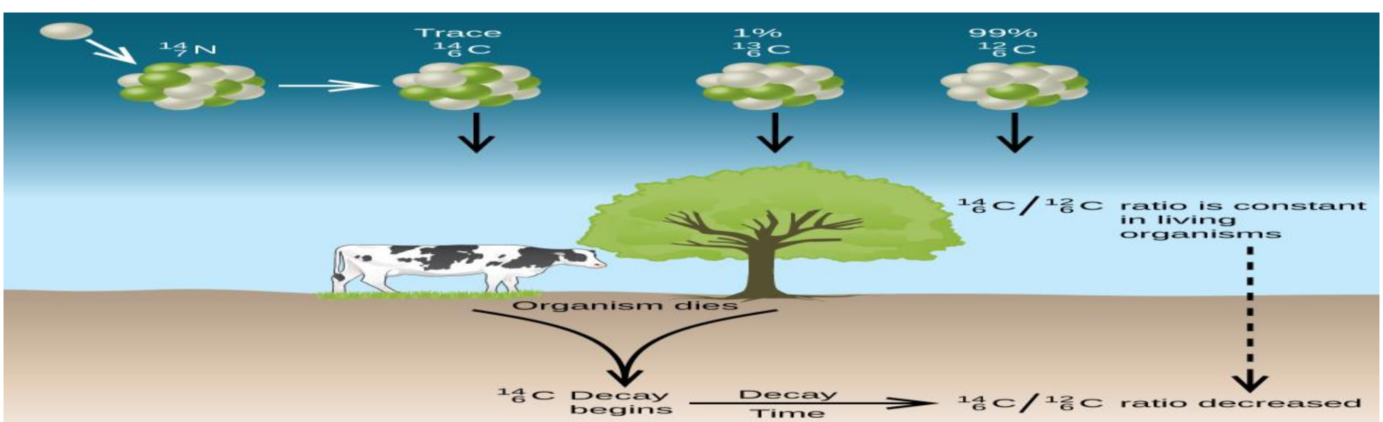
Naturally occurring carbon consists of three isotopes: C-12, which constitutes about 99% of the carbon on earth; C-13 is about 1% of the total; and trace amounts of C-14. Carbon-14 forms in the upper atmosphere by the reaction of nitrogen atoms with neutrons from cosmic rays in space:

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

Carbon-14 eventually enters the food chain via the formation of carbon dioxide and its uptake by plants via photosynthesis. Eating these plants distributes carbon-14 throughout all living organisms

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

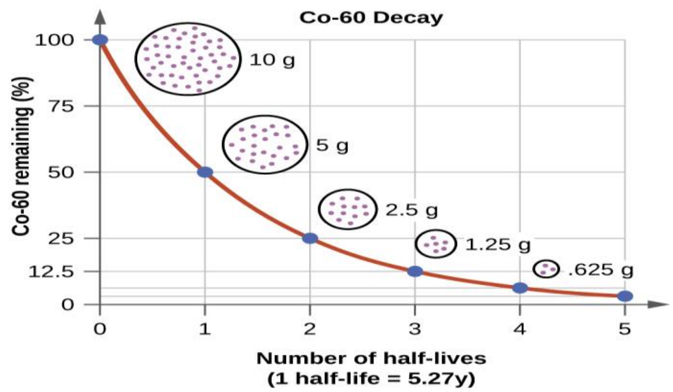
The half-life of carbon-14 is 5730 years:



Because the half-life of any nuclide is constant, the amount of substance remaining in an artifact can serve as a nuclear clock to determine ages of objects. When an animal is alive it maintains a caron-14 to carbon-12 ratio identical to that in the atmosphere. When it dies, this ratio decreases. By measuring this ratio and contrast it to the atmosphere, we can get an approximate age.

Radioactive decay is a first order process, we can adapt the mathematical relationship used in first-order chemical reaction where the concentration is substituted with N, the numbers of nuclei.

$$\ln\left(\frac{N_t}{N_0}\right) = -kt \qquad \text{and} \qquad t_{1/2} = \frac{\ln 2}{k} \qquad \text{then} \qquad k = \frac{\ln 2}{t_{1/2}}$$



Example: A wood artifact is found to have a C-14 activity of 9.1 disintegrations per second. Given that the C-14 activity of an equal mass of fresh-cut wood has a constant value of 15.2 disintegrations per second, determine the age of the artifact. The half-life of C-14 is 5730 years

Example: A rock is found to contain 5.51 mg of U-238 and 1.63 mg of Pb-206. Determine the age of the rock ($t_{1/2}$ of U-238 is 4.51×10^9 yr)