Unit 4: The Nucleus

Structure of An Atom

- Nucleus Contains Protons (+) and Neutrons
- Electrons

Some Radioactive Decays

- α decay (Helium-4 Nucleus)
- β⁻ decay (electron)
- γ decay (gamma-ray photon)
- proton decay (spherical nucleus)
- proton decay (deformed nucleus)

Student Name: ____________ **Key** ______________

Class Period: ________
Unit 4 Vocabulary:

1. Artificial transmutation: Changing one element into another by bombarding it with particle bullets in a particle accelerator.
2. Atomic Mass Unit (amu): 1/12 the mass of a C-12 atom; the approximate mass of either a proton or neutron.
3. Atomic Number: The number that identifies an element, equal to an atom’s number of protons.
4. Deflect: Change in direction due to an outside force.
5. Emit: To give off something.
6. Half-life: The time it takes for half the mass of a radioactive isotope to undergo decay. Half-life is also the period of time in which any given nucleus has a 50% chance of undergoing radioactive decay.
7. Isotope: Atoms of the same element that contain different numbers of neutrons and therefore differ in atomic mass as well.
8. Mass Defect: The mass that was lost during a nuclear change that was converted into energy via $E=mc^2$.
9. Mass Number: The sum total of all the protons and neutrons in an atomic nucleus.
10. Natural Radioactivity (Radioactive Decay): The spontaneous breakdown of an unstable nucleus into a more stable nucleus and a product of decay (alpha, beta-negative, beta-positive, or gamma).
11. Neutron: The particle that has no charge and has a mass of 1 amu.
12. Nuclear Charge: The net positive charge of the nucleus, which is equal to the number of protons in the nucleus.
13. Nuclear Fission: The process whereby a large nucleus is *split* by artificial transmutation into smaller nuclei with the release of a large amount of energy derived from the conversion of a tiny bit of mass into energy.
Unit 4 Vocabulary (Cont’d):

14. Nuclear Fusion: The process whereby two small nuclei are *combined* to form one larger nucleus with the release of a huge amount of energy derived from the conversion of a tiny bit of mass into energy.

15. Nucleon: A particle that exists in the nucleus (protons and neutrons).

16. Nucleus: The central core of the atom, having a net positive charge, and consisting of the atom’s protons and neutrons.

17. Particle Accelerator: A device that uses electromagnetic fields to speed up charged particles.

18. Proton: A particle that represents a unit charge of +1 and a mass of 1 amu.

19. Weight-Average Mass: The average mass of a sample of an element that is determined by the mass and abundance of every isotope of that element.
Unit 4 Homework Assignments:

<table>
<thead>
<tr>
<th>Assignment:</th>
<th>Date:</th>
<th>Due:</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Atoms are the smallest pieces an element may be broken down to and still retain the properties of that element.

- The word element comes from the Greek word atomos, meaning “indivisible”, or unbreakable.
- Individual atoms are so tiny they cannot be seen directly. They may be detected through x-ray crystallography or atomic force microscopes, but only indirectly.
- 602,300,000,000,000,000 atoms of hydrogen combined have a mass of 1 gram, or about the same as a single small paper clip.
- That same number of hydrogen atoms in gaseous form would fill 11 empty 2-liter soda bottles and still have the same mass as the one small paper clip.
Atoms are constructed of the following sub atomic (smaller than atom sized) particles.

1. Nucleons (Particles WITHIN the Nucleus):
   i. **Protons**: Protons have a mass of one atomic mass unit (amu), or equivalent to $1.66 \times 10^{-24}$ grams. Protons have a nuclear charge of positive one ($+1$). Protons are found inside the atom’s nucleus, and the number of protons in the atom’s nucleus provides the atomic number of the atom. The atomic number (the number of protons) is what identifies the element and its properties. Note the diagram on pg 8.
   ii. **Neutrons**: Neutrons have a mass of one atomic mass unit (amu) and no charge. Neutrons are found inside the atom’s nucleus, and the number of neutrons added to the number of protons provides the mass number of the atom. The number of neutrons does not change the identity of the element. As an example, oxygen (element symbol O) is found most commonly in a form with mass number 16. Oxygen’s proton number ALWAYS is eight, in O-16 there would be eight neutrons as well ($8 + 8 = 16$). The
number of protons and neutrons does NOT have to be equal. Atoms of an element may have differing numbers of neutrons within the same substance. **Isotopes** have **differing** number **counts** of **neutrons** within the nuclei of the atom. The most common isotope is the isotope with the mass equal to the average atomic mass given on the period table rounded to the nearest whole number. As oxygen has a given mass (period table) of 15.9994, the most common isotope of oxygen is O-16, or oxygen with a mass number of 16. See the diagram below.

---

**Watch Crash Course Chemistry The Nucleus video**

https://www.youtube.com/watch?v=FSyAehMdpyI
2. **Particles outside the Nucleus:**

   i. **Electrons:** Electrons have a mass of $\frac{1}{1836}$ amu ($9.11 \times 10^{-28}$ grams) and a nuclear charge of -1. Electrons are found orbiting the nucleus in energy levels. Atoms gain, lose, or share electrons when they form chemical bonds. If electrons are gained and lost between atoms, the bond is an ionic bond. If electrons are shared between atoms, the bond is a covalent bond. The **NUMBER of ELECTRONS** in the atom **EQUALS** the number of **PROTONS**. Atoms are neutrally charged, so the + charge of the protons and the - charge of the electrons need to be equal to provide an overall neutral charge to the entire atom. In the given example (oxygen) for protons and neutrons, as oxygen has eight protons, oxygen needs to have eight electrons in orbit around the nucleus. **Electrons** play a HUGE role in chemistry, as they are the ‘glue’ that holds EVERYTHING together.
Atom structure and designations:

The diagram below shows an idealized representation of a helium atom. (Helium (He): atomic number = 2, mass number = 4)

1. How many protons does an atom of iron (Fe) contain?
   i. Iron has an atomic number of 26, therefore iron contains 26 protons.
2. What is the nuclear charge of an atom of iron (Fe)?
   i. Iron contains 26 protons, therefore iron has 26 electrons.
3. How many atomic mass units (amu) do the protons in an iron (Fe) atom weigh?
   i. Iron has 26 protons, with each proton having a mass of 1 amu, therefore the mass of protons in an iron atom is 26 amu.
4. How many electrons does an iron (Fe) atom contain?
   i. The number of electrons in an atom must equal the number of protons in that atom. Iron contains 26 protons; therefore iron must contain 26 electrons. Looking at the electron configuration (shown below) for iron (Fe) you see 2-8-14-2, which when added together equals 26, or the total number electrons in iron. (See the key on the Periodic Table of Elements on page 9 of the Reference Table.)

5. What is the most common isotope of iron (Fe)?
   i. The most common isotope of iron is found by taking the average of ALL the atomic masses of the isotopes of iron (55.847). The average is then rounded to the nearest whole number (56). That rounded whole number is assigned as the mass number of iron, or written as Fe-56.
6. How many neutrons are in the nucleus of the most common isotope (Fe-56) of iron?

   i. To find the number of neutrons in an atom’s nucleus, take the mass number (protons and neutrons in nucleus) as given. Then subtract the atomic (element) number. The result is the neutron number count.

   Fe-56 has 26 protons, therefore 56 - 26 protons = 30 neutrons

7. Why does iron (Fe) have an average mass of 55.847 amu?

   i. There are four commonly occurring isotopes of iron (Fe).

<table>
<thead>
<tr>
<th>Mass of Isotope (amu)</th>
<th>Notation 1</th>
<th>Notation 2</th>
<th># protons (atomic #)</th>
<th># neutrons (mass # - atomic #)</th>
<th># electrons (atomic #)</th>
<th>% Abundance in Nature</th>
</tr>
</thead>
<tbody>
<tr>
<td>54</td>
<td>Fe-54</td>
<td>$^{54}\text{Fe}$</td>
<td>26</td>
<td>54 - 26 = 28</td>
<td>26</td>
<td>5.845%</td>
</tr>
<tr>
<td>56</td>
<td>Fe-56</td>
<td>$^{56}\text{Fe}$</td>
<td>26</td>
<td>56 - 26 = 30</td>
<td>26</td>
<td>91.754%</td>
</tr>
<tr>
<td>57</td>
<td>Fe-57</td>
<td>$^{57}\text{Fe}$</td>
<td>26</td>
<td>57 - 26 = 31</td>
<td>26</td>
<td>2.119%</td>
</tr>
<tr>
<td>58</td>
<td>Fe-58</td>
<td>$^{58}\text{Fe}$</td>
<td>26</td>
<td>58 - 26 = 32</td>
<td>26</td>
<td>0.282%</td>
</tr>
</tbody>
</table>

**Writing atomic symbols:**

[Diagram showing atomic symbols]
Topic: **Weight-Average Atomic Mass**

Objective: How do we calculate weight-average atomic mass?

The atomic **masses** given on the periodic table are WEIGHT-AVERAGED masses. This means they are calculated using both the masses of each isotope and their percent abundances as found in nature.

- For simplicity, weight-average masses will be rounded to the thousandths place.

Calculating weight-average mass of an element:

\[
WAM = (\text{mass}_{\text{isotope}_1} \times \%_{/100}) + (\text{mass}_{\text{isotope}_2} \times \%_{/100}) + (\text{mass}_{\text{isotope}_3} \times \%_{/100}) + \ldots
\]

Therefore, for the four isotopes of iron (Fe):

<table>
<thead>
<tr>
<th>Mass of Isotope (amu)</th>
<th>Notation 1</th>
<th>% Abundance in Nature</th>
</tr>
</thead>
<tbody>
<tr>
<td>54</td>
<td>Fe-54</td>
<td>5.845%</td>
</tr>
<tr>
<td>56</td>
<td>Fe-56</td>
<td>91.754%</td>
</tr>
<tr>
<td>57</td>
<td>Fe-57</td>
<td>2.119%</td>
</tr>
<tr>
<td>58</td>
<td>Fe-58</td>
<td>0.282%</td>
</tr>
</tbody>
</table>

The weight-average mass (WAM) of iron (Fe) calculates as:

\[
WAM = (54 \times \frac{5.845}{100}) + (56 \times \frac{91.754}{100}) + (57 \times \frac{2.119}{100}) + (58 \times \frac{0.282}{100})
\]

\[
WAM = (3.156) + (51.382) + (1.208) + (0.164)
\]

\[
WAM = 55.910 \text{ amu}
\]

The above calculation is as such: (5.845% of 54 = 3.156 amu) plus (91.754% of 56 = 51.382 amu) plus (2.1195 of 57 = 1.208 amu) plus (0.282% of 58 = 0.164 amu) equals the WAM of 55.910 amu.
Note that the result on the previous page is still not exactly the same as the listed weight-average mass on the period table. The isotope information used for these problems comes from the National Nuclear Data Center at Brookhaven National Laboratory. The weight-average mass on the Periodic Table may include other isotopes of iron, which are radioactive and make up a very tiny percentage of iron’s mass. The weight-average mass is based on the abundance of the naturally occurring isotopes of that element.

Also, protons and neutrons DO NOT have a mass of EXACTLY 1 amu. Neutrons have a tiny fraction more mass than protons. An atomic mass unit (amu) is calculated as an average mass, found by taking the mass of a $^{12}$C nucleus and dividing it by 12. $1 \text{ amu} = \frac{1}{12}$ of $^{12}$C mass

- Relative atomic masses are based on $^{12}$C = 12.000

**Weight-Average example:**

i. Boron (B) consists of two natural isotopes. $^{10}$B has a mass of 10 amu and is 19.80% of all B atoms. $^{11}$B has a mass of 11 amu and is 80.20% of all B atoms. Therefore, 19.80% of 10 = 1.98 amu plus 80.20% of 11 = 8.822 amu equals 10.802 amu.

$$WAM = (10.0 \times \frac{19.80}{100}) + (11.0 \times \frac{80.20}{100}) = 10.802 \text{ amu}$$
Atomic Structure (The Nucleus) homework

Circle your answers for the multiple choice questions below.

1. One atomic mass unit is defined as weighing
   a) 1/16th the mass of $^{16}\text{O}$
   b) 1/32nd the mass of $^{32}\text{S}$
   c) 1/12th the mass of $^{12}\text{C}$
   d) 1/10th the mass of $^{10}\text{B}$

2. How many electrons does it take to have the same mass as one proton?
   a) 1  
   b) 100  
   c) 957  
   d) 1836

3. $^{12}\text{C}$ contains 6 protons, 6 neutrons, and 6 electrons. Which of those, if removed or changed, would alter the identity of the element?
   a) Proton  
   b) Neutron  
   c) Electron

4. Which of the following represents isotopes of the same element? (Note that there is a subscripts Z and superscript Z.)
   a) $^{8}\text{O}$ and $^{18}\text{O}$
   b) $^{8}\text{O}$ and $^{9}\text{F}$
   c) $^{16}\text{O}$ and $^{18}\text{O}$
   d) $^{9}\text{F}$ and $^{10}\text{Ne}$

5. A nucleus of $^{25}\text{Mg}$ contains how many neutrons?
   a) 12  
   b) 13  
   c) 25  
   d) 37

6. What is the nuclear charge of a nucleus of $^{32}\text{P}$?
   a) +15  
   b) +17  
   c) +31  
   d) +47

7. A nucleus of $^{25}\text{Mg}$ contains how many protons?
   a) 12  
   b) 13  
   c) 25  
   d) 37

8. What is the atomic charge of an atom of $^{32}\text{P}$?
   a) +15  
   b) +17  
   c) +31  
   d) 0

9. How many electrons are found outside the nucleus of $^{13}\text{N}$?
   a) 6  
   b) 7  
   c) 13  
   d) 19

Cont’d next page
Complete the table below. (1 point EACH)

<table>
<thead>
<tr>
<th>Isotope</th>
<th># Protons</th>
<th># Neutrons</th>
<th># Electrons</th>
<th>Nuclear Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{39}_{19}\text{K}$</td>
<td>19</td>
<td>20</td>
<td>19</td>
<td>19</td>
</tr>
<tr>
<td>$^{42}_{20}\text{Ca}$</td>
<td>20</td>
<td>22</td>
<td>20</td>
<td>20</td>
</tr>
<tr>
<td>$^{56}_{26}\text{Fe}$</td>
<td>26</td>
<td>30</td>
<td>26</td>
<td>26</td>
</tr>
<tr>
<td>$^{232}_{92}\text{U}$</td>
<td>92</td>
<td>140</td>
<td>92</td>
<td>92</td>
</tr>
</tbody>
</table>

Find the most common isotope for each of the following elements. (1 point EACH)

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Mass (from Periodic Table)</th>
<th>Most Common Isotope</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca</td>
<td>40.08</td>
<td>Ca-40</td>
</tr>
<tr>
<td>Br</td>
<td>79.904</td>
<td>Br-80</td>
</tr>
<tr>
<td>Zn</td>
<td>65.409</td>
<td>Zn-65</td>
</tr>
<tr>
<td>Hg</td>
<td>200.59</td>
<td>Hg-201</td>
</tr>
</tbody>
</table>

Weight-Average Mass Problems: (Show all work; round to thousandths)

10. What is the average atomic mass for thallium, Tl, if the two isotopes of Tl have the given masses and abundances?

$^{203}\text{Tl}$ has a mass of 203 amu with an abundance of 29.5%

$^{205}\text{Tl}$ has a mass of 205 amu and an abundance of 70.5%

$$W.A.M. = (203 \text{ amu} \times 0.295) + (205 \text{ amu} \times 0.705) = 204.410 \text{ amu} = ^{204}\text{Tl}$$
Nuclear Stability:

The **larger** (more massive) a **nucleus** is, the **harder** it is for that nucleus to remain **intact**.

- When a nucleus is **radioactive**, it gives off **decay** particles and **changes** from one **element** into another. This process is also known as radioactive decay or natural transmutation.
- Atoms with an atomic number of 1 through 83 have at least one stable (nonradioactive) isotope, but **ALL isotopes** starting with atomic 84 or **greater** are radioactive (**unstable**).
- There are Three Types of Natural Decay. Natural decay allows atoms of one element to decay/transmute into atoms of a different element.
The Three Types of Natural Decay:

<table>
<thead>
<tr>
<th>Decay Type</th>
<th>Notation</th>
<th>Mass</th>
<th>Charge</th>
<th>What happens to the atom when it undergoes this type of decay</th>
</tr>
</thead>
</table>
| Alpha (α)   | $^4_2$He | 4 (made of 2 protons and 2 neutrons) | +2 (from the 2 protons) | The nucleus loses 2 protons (atomic mass decreases by 2) and 4 total particles (mass decreases by 4). It turns into a different element. $^{238}_{92}$U $\rightarrow ^4_2$He + $^{234}_{90}$Th
|             |          |      |        | radioactive alpha nucleus particle nucleus                  |
|             |          |      |        | The alpha particles released by uranium in the Earth's crust build up underground in porous rock, where they gain electrons and turn into actual atoms of pure helium. This is where we get the helium that is in balloons! |
| Beta (β−)   | $^0_{−1}$e | 0 (made of an electron) | −1 | A neutron in the nucleus decays to form a proton (atomic number increases by 1, but mass stays the same) and an electron (the beta particle) which leaves the nucleus at high speeds. $^{42}_{19}$K $\rightarrow ^0_{−1}$e + $^{42}_{20}$Ca
|             |          |      |        | radioactive beta nucleus particle nucleus                   |
| Position (β+) | $^0_{−1}$e | 0 (made of an anti-electron (positron)) | +1 | A proton in the nucleus decays to form a neutron (atomic number decreases by 1, but mass stays the same) and a positron (the antimatter form of an electron) which leaves the nucleus at high speeds. $^{53}_{26}$Fe $\rightarrow ^0_{+1}$e + $^{53}_{25}$Mn
|             |          |      |        | radioactive positron nucleus particle nucleus               |

You are able to find out the type of decay a selected radioactive isotope will undergo by looking on Reference Table N.

Gamma Particles:

There is one more type of decay, gamma decay (γ), a form of high-energy light given off as the nucleus stabilizes. Gamma rays do not change the element, has no mass or charge, and is highly energetic. **Gamma rays require** a LOT of concrete, or a thick plate of lead to stop, while alpha and beta particles are stopped by paper or skin.
The diagram below shows the different paths taken by the types of radioactive particles as they pass through an electric field. At the left is a shielded container of a radioactive substance that only allows decay particles to be emitted in a straight line. Downstream are positively and negatively charged electric plates. Positive-charged particles (alpha and positron) are attracted and deflect towards the negatively charged plate. Negative-charged particles (beta) are attracted and deflect towards the positively charged plate. Gamma rays have no charge; therefore gamma rays do not deflect.

Watch Beta decay video
https://www.youtube.com/watch?v=Yln_pmy-mWk

Watch Positron Decay Video
https://www.youtube.com/watch?v=bFxBt_QtKKug
The Proton/Neutron Ratio:

The ratio of neutron:proton (n:p) in a stable atom varies with the size of the atom. Small atoms are stable as a 1:1 ratio. As the atom mass gets larger, more neutrons are needed for stability which may push the stable n:p ratio as high as 1.5:1. This creates a zone of stability, inside which the isotopes are stable. Outside the ‘zone’, nuclei either have too many or too few neutrons to be stable, and therefore decay by emitting α, β, or γ particles to bring the ratio back into the zone of stability.

ALL isotopes of ALL elements above $^{83}$Bi are unstable and undergo radioactive decay.
Topic: **Nuclear Energy**

Objective: How does the lost mass emerge as energy?

**Nuclear Energy:**

Radioactive decay gives off energy. In fact, the energy from radioactive decay has kept the Earth’s core molten for 4.5 billion years. The molten core produces the Earth’s magnetic field, important for life as it shields the planet from most high-energy solar particles.

The energy that is emitted during a nuclear decay comes from the **MASS DEFECT**. This is the tiny bit of mass that is **lost** and **converted** into energy. While the Mass Defect seems to be in violation of the Laws of Conservation of Mass and Energy, those Laws cover physical and chemical changes, not nuclear changes. The equation used to calculate the energy given off during a nuclear transmutation is \( E=mc^2 \), discovered by Einstein, but used by Lise Mitner to make nuclear power possible. \( E \) is energy, in joules. The symbol \( m \) is the mass that was destroyed by the nuclear change, in kg. The symbol \( c^2 \) is the speed of light squared which is a huge number (9.00 x 10\(^{16}\) m\(^2\)/sec\(^2\)). You are not expected to calculate this energy in Regents Chemistry, but you do need to understand that

**A HUGE amount of energy may be created by destroying a TINY amount of nuclear mass.**
How do we detect radioactivity?

1. Radioactivity can expose protected photographic film.
   a. Antoine Henri Becquerel placed a uranium rock on top of photographic film, causing the film to be exposed.

2. Radioactive isotopes cause phosphorescent (glow-in-the-dark) materials to glow.
   a. Marie and Pierre Curie discovered radium and polonium, both more radioactive per gram than uranium. Radium causes phosphorescent materials to glow even if no light was present to charge them. Radium once was used to paint clock and watch dials, but the radium caused cancer in those (including Marie Curie) that worked closely with it.

3. Radioactive decay particles are charged, and they ionize (give charge) to matter they pass through.
   a. The Geiger Counter was developed to detect alpha radiation due to ionization of a gas in a tube. It was later improved by Müeller to detect any type of ionizing radiation. It displays counts per time period.
## Uranium Decay Process

**Objective:** Find the decay steps of $^{238}\text{U}$ as it decays.

### Uranium Decay:

$^{238}\text{U}$ is unstable and **decays** into a more stable nuclei. It takes 14 **discrete decay steps** until a stable, non-radioactive nucleus is reached. The daughter nuclide of one step becomes the parent nuclide of the next step.

<table>
<thead>
<tr>
<th>Parent Nuclide</th>
<th>Daughter Nuclide</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{238}\text{U}$</td>
<td>$^4_2\text{He} + ^{234}\text{Th}$</td>
</tr>
<tr>
<td>$^{234}\text{Th}$</td>
<td>$^0_{-1}\text{e} + ^{234}\text{Pa}$</td>
</tr>
<tr>
<td>$^{234}\text{Pa}$</td>
<td>$^0_{-1}\text{e} + ^{234}\text{U}$</td>
</tr>
<tr>
<td>$^{238}\text{U}$</td>
<td>$^4_2\text{He} + ^{230}\text{Th}$</td>
</tr>
<tr>
<td>$^{230}\text{Th}$</td>
<td>$^4_2\text{He} + ^{226}_{88}\text{Ra}$ Continue until Step 14.</td>
</tr>
<tr>
<td>$^{226}\text{Ra}$</td>
<td>$^4_2\text{He} + ^{222}_{86}\text{Rn}$</td>
</tr>
<tr>
<td>$^{222}\text{Rn}$</td>
<td>$^4_2\text{He} + ^{218}_{84}\text{Po}$</td>
</tr>
<tr>
<td>$^{218}\text{Po}$</td>
<td>$^4_2\text{He} + ^{214}_{82}\text{Pb}$</td>
</tr>
<tr>
<td>$^{214}\text{Pb}$</td>
<td>$^0_{-1}\text{e} + ^{214}_{83}\text{Bi}$</td>
</tr>
<tr>
<td>$^{214}\text{Bi}$</td>
<td>$^0_{-1}\text{e} + ^{214}_{84}\text{Po}$</td>
</tr>
<tr>
<td>$^{214}\text{Po}$</td>
<td>$^4_2\text{He} + ^{210}_{82}\text{Pb}$</td>
</tr>
<tr>
<td>$^{210}\text{Pb}$</td>
<td>$^0_{-1}\text{e} + ^{210}_{83}\text{Bi}$</td>
</tr>
<tr>
<td>$^{210}\text{Bi}$</td>
<td>$^0_{-1}\text{e} + ^{210}_{84}\text{Po}$</td>
</tr>
<tr>
<td>$^{210}\text{Po}$</td>
<td>$^4_2\text{He} + ^{206}_{82}\text{Pb}$</td>
</tr>
</tbody>
</table>
Natural Radioactivity homework

Circle your answer for each multiple choice question.

1. Which particle type is given off by a decaying nucleus of $^{37}$K?
   a) Alpha  
   b) Beta  
   c) Positron  
   d) Gamma

2. Which particle type has the greatest mass?
   a) Alpha  
   b) Beta  
   c) Positron  
   d) Gamma

3. Which product of decay may only be stopped by a thick sheet of lead?
   a) Alpha  
   b) Beta  
   c) Positron  
   d) Gamma

4. Which element has no stable isotopes, only having radioactive isotopes?
   a) Si  
   b) Am  
   c) Cu  
   d) C

5. Which decay particle will deflect towards a positively charged plate?
   a) Alpha  
   b) Beta  
   c) Positron  
   d) Gamma

6. Which particle, when emitted during decay, will cause the atomic number of the decaying nucleus to increase by one?
   a) Alpha  
   b) Beta  
   c) Positron  
   d) Gamma

7. You think you may have discovered a chunk of radioactive rock. Describe one way you could check if this rock is radioactive.

   Place on film; use Geiger-Mueller counter; place near phosphorescent material
Natural Radioactivity homework

Given each of the unstable isotopes below with their symbols, mass number, and decay type, determine the daughter nuclide (decay product) by writing the complete nuclear equation. (Use atomic numbers and symbols as found in the periodic table)

(One point EACH)

<table>
<thead>
<tr>
<th>Equation</th>
<th>Daughter Nuclide</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{87}<em>{36}$Kr $\rightarrow$ $^{0}</em>{-1}$e + $^{87}_{37}$Rb</td>
<td></td>
</tr>
<tr>
<td>$^{212}<em>{84}$Po $\rightarrow$ $^{4}</em>{2}$α + $^{208}_{82}$Pb</td>
<td></td>
</tr>
<tr>
<td>$^{14}<em>{6}$C $\rightarrow$ $^{0}</em>{-1}$e + $^{14}_{7}$N</td>
<td></td>
</tr>
<tr>
<td>$^{6}<em>{2}$He $\rightarrow$ $^{0}</em>{-1}$e + $^{6}_{3}$Li</td>
<td></td>
</tr>
<tr>
<td>$^{19}<em>{10}$Ne $\rightarrow$ $^{0}</em>{+1}$e + $^{19}_{9}$F</td>
<td></td>
</tr>
<tr>
<td>$^{240}<em>{92}$U $\rightarrow$ $^{4}</em>{2}$He + $^{236}_{90}$Th</td>
<td></td>
</tr>
<tr>
<td>$^{20}<em>{9}$F $\rightarrow$ $^{0}</em>{-1}$e + $^{20}_{10}$Ne</td>
<td></td>
</tr>
<tr>
<td>$^{208}<em>{81}$Tl $\rightarrow$ $^{0}</em>{-1}$e + $^{208}_{82}$Pb</td>
<td></td>
</tr>
</tbody>
</table>
Radioactive decay is a random process. It is not possible to predict when a particular nucleus will decay, but we can make fairly accurate predictions regarding how many nuclei in a sample will decay in a given period of time.

i. **Half-life**: The half-life of a radioactive isotope is defined as the period of time that must elapse for one-half of the nuclei in the sample to undergo radioactive decay.

ii. During one half-life period:

   a) **Half** of the number of radioactive nuclei in the sample undergoes decay into a new, more stable isotope.

   - If a sample contains 1000 radioactive nuclei to begin with, 500 nuclei will undergo decay over one half-life.

   b) **Half** the mass of the radioactive isotope is converted into a new, more stable isotope.

   - If a sample contains 4.0 grams of radioactive nuclei to begin with, after one half-life 2.0 grams of radioactive material will remain undecayed, with the other 2.0 grams having transmuted into a new, more stable isotope.
c) A Gieger-Müeller counter’s count-per-time period will be one-half of the count it started with.

- If a Gieger-Müeller counter is showing 400 counts per minute to begin with, after one-half life, the counter will show 200 counts per minute.

iii. After one half-life, half (50%) of the original radioactive material will have undergone decay, and 50% is more stable.

iv. After a second half-life, one-quarter (25%) of the original radioactive material will remain undecayed, and 75% is more stable.

v. After a third half-life, one-eighth (12.5%) of the original radioactive material will remain undecayed, and 87.5% is more stable.

vi. Note that the actual half-life time is constant. If one half-life is 10 days, two half-lives would take 20 days, three half-lives 30 days, and so on.

Reference Table N shows the half-lives of some radioactive isotopes.

Watch Bozeman Science Radiation and Radioactive Decay video
https://www.youtube.com/watch?v=oFdR_yMKOCw
Solving Half-Life Problems:

If you are given how much of the radioactive isotope you have to begin with, you want to find how much of the radioactive isotope will remain after a future amount of time has elapsed.

1. Step 1: Determine how many half-lives will occur. Take how much total time is given and divide it by the length of the half-life.
2. Step 2: Divide by two the amount (mass, percent, fraction, or number of nuclei) you began with as many times as there will be half-lives.

Examples of Half-Life Problems:

1. The half-life of radioactive $^{222}\text{Rn}$ is 3.8 days. If your basement today contains 20.0 grams of $^{222}\text{Rn}$, how much radioactive $^{222}\text{Rn}$ will remain after 19 days if no more $^{222}\text{Rn}$ enters your basement during that time?
   i. # of half-lives = time elapsed/half-life time = 19 d/3.8 d = 5 half-lives
   ii. Then divide by 2 the starting amount of $^{222}\text{Rn}$ five times.

$$\frac{20.0 \ g}{2} \rightarrow \frac{10.0 \ g}{2} \rightarrow \frac{5.00 \ g}{2} \rightarrow \frac{2.50 \ g}{2} \rightarrow \frac{1.25 \ g}{2} \rightarrow 0.625 \ g$$ $^{222}\text{Rn}$ remains
2. A laboratory sample of $^{32}\text{P}$ triggers 400 counts per minute (cpm) in a Gieger-Müeller counter to begin with. How many days will it take for the $^{32}\text{P}$ to decay so that there are only 50 cpm?
   i. You need to find the number of half-lives it takes for the Gieger-Müeller counter to go from 400 to 50 cpm. Divide 400 by 2 until get 50.
   
   \[
   \frac{400 \text{ cpm}}{2} \rightarrow \frac{200 \text{ cpm}}{2} \rightarrow \frac{100 \text{ cpm}}{2} \rightarrow 50 \text{ cpm} \quad (3 \text{ divisions} = 3 \text{ half-lives})
   \]

   ii. Look for the half-life of $^{32}\text{P}$ in Reference Table N and see that it is 14.3 days. Now multiply the number of elapsed half-lives by the length of each half-life (h-l).
   iii. days/h-l $\times$ 3 half-lives = 42.0 days to go from 400 cpm to 50 cpm

3. A cylinder containing 5.0 L of pure radioactive $^{19}\text{Ne}$ is left to sit for 103.2 seconds. What percentage of the original radioactive $^{19}\text{Ne}$ will remain after that time?
   i. Look in Reference Table N for the half-life of $^{19}\text{Ne}$: 17.2 seconds
   ii. # of half-lives = time elapsed/h-l length = 103.2 secs/17.2 secs = 6 h-l
   iii. When we began, 100\% of the $^{19}\text{Ne}$ sample was pure, so divided 100 by 2 a total of 6 times to find the remaining percentage of pure $^{19}\text{Ne}$.
   
   \[
   \frac{100\%}{2} \rightarrow \frac{50\%}{2} \rightarrow \frac{25\%}{2} \rightarrow \frac{12.5\%}{2} \rightarrow \frac{6.25\%}{2} \rightarrow \frac{3.125\%}{2} \rightarrow 1.5625\% \text{ of } ^{19}\text{Ne remains}
   \]
   iv. This means that after 6 half-lives only about 1.6\% of the original $^{19}\text{Ne}$ remains in the cylinder.
Solving Half-Life Problems:

If you are given how much of the decayed radioactive isotope you have currently, you want to find how much of the radioactive isotope in the past you began with.

1. Step 1: Determine how many half-lives have occurred. Take how much total time has elapsed and divide it by the length of the half-life.

2. Step 2: Double the amount (mass, percentage, fraction, or number of nuclei) as many times as there have been half-lives.

Examples of Half-Life Problems:

1. The half-life of the radiological assay $^{99}\text{Tc}$ is 6.0 hours. If 10.0 µg of $^{99}\text{Tc}$ remain after 24 hours, how much $^{99}\text{Tc}$ was originally administered to the patient?
   i. $\# \text{ of half-lives} = \frac{\text{time elapsed}}{\text{half-life length}} = \frac{24 \text{ hrs}}{6.0 \text{ hrs}} = 4 \text{ h-l elapsed}$
   ii. Take the radioactive amount remaining (10.0 µg) and double it four times.
      $10.0 \times 2 (20) \rightarrow 20.0 \times 2 (20) \rightarrow 40.0 \times 2 (80) \rightarrow 80.0 \times 2 = 160 \mu\text{g}$ of $^{99}\text{Tc}$ initially
   iii. This means that there was originally 160 µg of $^{99}\text{Tc}$ initially
2. A laboratory sample of $^{32}\text{P}$ currently triggers 100.0 counts-per-minute (cpm) in a Gieger-Müeller counter. How long ago did the $^{32}\text{P}$ sample trigger 1600.0 cpm?
   
   i. Find out how many half-lives it takes for the Gieger-Müeller counter to go from 1600.0 cpm to 100.0 cpm. Divide 1600.0 by 2 until you get 100.0 as an answer.
   
   $1600/2 \rightarrow 800/2 \rightarrow 400/2 \rightarrow 200/2 = 100$ means four h-l have elapsed
   
   ii. Look for the half-life for $^{32}\text{P}$ in Reference Table N and find it to be 14.3 days.
   
   iii. Multiply the h-l length of $^{32}\text{P}$ (14.3 days) x 4 h-l elapsed = 57.2 days
   
   iv. That means 57.2 days ago the Gieger-Müeller counter would have recorded 1600 clicks per minute.
Topic: **Radioactive Half-Life**

Objective: How can we calculate the age of matter?

**Solving Half-Life Problems:**

If you are given how much of the radioactive isotope decayed over a given amount of time, you want to find half-life length of time of the material.

1. Step 1: Determine how many half-lives have occurred. Take the original amount (mass, percentage, fraction, or number of nuclei) of radioactive material, then divide the original amount by two until you get to the stated remaining amount.

2. Divide the given time elapsed by the number of half-lives you calculated above. That results in the length of time per half-life.

**Examples of Half-Life Problems:**

1. A radioactive sample is placed next to a Gieger-Müeller counter and continuously monitored. In 20.0 hours the counter’s reading changes from 500 counts-per-minute (cpm) to 125 cpm. How long is the half-life of the sample material?
   i. Find out how many half-lives it will take for the counter to go from 500 to 125 cpm.

   \[
   500/2 \rightarrow 250/2 = 125 \text{ means that two half-lives have occurred.}
   \]

   ii. Half-life = time elapsed/ # of h-l = 20.0 hours/2 h-l = 10.0 hr h-l
2. A sample of pure radioactive isotope is left to decay. After 40.0 days the sample is placed in a mass spectrometer instrument, which determines that only 25% of the original radioactive isotope remains. How long is the half-life of this isotope?

i. Find out how many half-lives it will take for the original 100% to decay to 25% remaining.

100%/2 \rightarrow 50%/2 = 25% means that it takes two half-lives for the original 100% radioactive sample to decay to 25% radioactive remaining.

ii. Half-life = time elapsed/ # of half-lives = 40.0 days/2 half-lives = 20.0 days per half-life

---

Half-life is the time for 1/2 of sample to change into another isotope or element.

Radioactive Dating:

Radioactive Dating is used to determine the age of a substance that contains a radioactive isotope of known half-life.

- Step 1: Determine how many times the original amount (mass, percentage, fraction, or number of nuclei) of radioactive matter may be divided by two until you reach the final amount. This number is the number of half-lives that have occurred.
- Step 2: Multiply the number of half-lives by the duration (length) of the element half-life as found in Reference Table N.

Examples of Radioactive Dating Problems:

1. The oldest rocks on Earth have been found to contain 50% $^{238}$U and 50% $^{206}$Pb (the final stable decay product of $^{238}$U). What is the age of these rocks on Earth?
   
   i. Find out how many half-lives have occurred to go from 100% $^{238}$U to 50% $^{238}$U.

   $100%/2 \rightarrow 50\%$ means one half-life has occurred

   ii. Age of Sample = # of half-lives x h-l length (Reference Table N)
iii. Half-life of $^{238}\text{U}$ from Reference Table N = $4.51 \times 10^9$ years

iv. Age of Sample = # of half-lives x h-l duration = $1 \text{ h-l} \times (4.51 \times 10^9 \text{ yrs})$

= $4.51 \times 10^9$ years age of these rocks

2. An ancient paper scroll is discovered and found that only 25% of the original concentration of $^{14}\text{C}$ (a natural C isotope found in all organic material) remains. How old is this scroll?

i. Find out how many half-lives have elapsed to go from 100% $^{14}\text{C}$ to 25% $^{14}\text{C}$.

$100%/2 \rightarrow 50%/2 = 25\%$ means that two half-lives have elapsed

ii. According to Reference Table N the half-life of $^{14}\text{C}$ is 5730 years.

iii. Age of Sample = # of half-lives x h-l duration = $2 \text{ h-l} \times 5730 \text{ yrs}$ = 11,460 years old is the age of the scroll
Topic: **Using Radioactive Isotopes**

Objective: What are the uses of radioactive isotopes?

**Using Radioactive Isotopes:**

Radioactive isotopes have a ‘bad’ connotation, but many are useful, and even helpful when used correctly.

<table>
<thead>
<tr>
<th>Radioactive Isotope</th>
<th>Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-14</td>
<td>Used to determine the age of biological remains (archaeology)</td>
</tr>
<tr>
<td>I-131</td>
<td>Used to detect and cure hyperthyroidism (overactive thyroid)</td>
</tr>
<tr>
<td>Co-60</td>
<td>Used as a source of radiation for radiotherapy of cancer</td>
</tr>
<tr>
<td>Tc-99m</td>
<td>Used to image blood vessels, especially in the brain, to detect tumors</td>
</tr>
<tr>
<td>Pu-239</td>
<td>Used as a highly fissionable fuel source for nuclear power or nuclear weapons</td>
</tr>
<tr>
<td>Am-241</td>
<td>Used in tiny amounts in smoke detectors as a source of ions to make a current</td>
</tr>
<tr>
<td>U-235</td>
<td>Used as fissionable fuel source for nuclear power or nuclear weapons</td>
</tr>
<tr>
<td>U-238</td>
<td>Used to determine the age of uranium-containing rock formations (geology)</td>
</tr>
</tbody>
</table>

**Irradiation of food:** Radiation kills **bacteria**, allowing longer shelf life without being pasteurized. Pasteurization requires heating, which can change food taste and quality. **Irradiation** doesn’t change the flavor. Some milk is sold irradiated, especially in Europe and Asia.

**Medical Radioactive Isotopes:** Radioactive **isotopes** may be used in **medicine** to **treat** cancers or as tools to **detect** (trace) problems. Treatment and tracer isotopes usually have rather short half-lives (hours to days) and can be eliminated quickly from the body by normal metabolic processes. Some isotopes of longer half-lives (weeks to months) are implanted in radioactive ‘seeds’ very close to the cancer tumor being treated. $^{125}$Iodine and $^{103}$Palladium are used in this way.
Notes page:
Half-Life homework

Circle your answer for each multiple choice question.

1. Which isotope listed is used for detecting brain tumors?
   a) $^{99}\text{Tc}$  
   b) $^{131}\text{I}$  
   c) $^{238}\text{U}$  
   d) $^{60}\text{Co}$

2. Which isotope listed is used to determine the age of rocks?
   a) $^{99}\text{Tc}$  
   b) $^{131}\text{I}$  
   c) $^{238}\text{U}$  
   d) $^{60}\text{Co}$

3. Which isotope listed is used to treat hyperthyroidism?
   a) $^{99}\text{Tc}$  
   b) $^{131}\text{I}$  
   c) $^{238}\text{U}$  
   d) $^{60}\text{Co}$

Radioactive Half-life and Radioactive Dating problems.

*Show* all your work for each problem and draw a box around your final answer which must *include* the required units.

4. *What* is the half-life of a radioactive isotope if 25% of the original mass of the isotope remains after 20.0 days?

   
   100% $\rightarrow$ 50% $\rightarrow$ 25%
   0 days 10 d 20 d

   20 days/2 h.l. = 10 day half-life

5. The half-life of cesium-137 is 30.2 years. How much $^{137}\text{Cs}$ was *originally present* if after 120.8 years 6.0 g of $^{137}\text{Cs}$ remained?

   120.8 yr/30.2 yr per h.l. = 4 h.l.

   0 h.l 1 h.l. 2 h.l. 3 h.l. 4 h.l.

   96.0 g $\rightarrow$ 48.0 g $\rightarrow$ 24.0 g $\rightarrow$ 12.0 g $\rightarrow$ 6.0 g

6. The half-life of barium-131 is 12.0 days. How many grams of $^{131}\text{Ba}$ remain after 60.0 days if the initial $^{131}\text{Ba}$ sample had a mass of 10.0 g?

   60.0 d/12.0 d per h.l. = 5 h.l.

   0 h.l 1 h.l. 2 h.l. 3 h.l. 4 h.l. 5 h.l.

   10.0 g $\rightarrow$ 5.00 g $\rightarrow$ 2.50 g $\rightarrow$ 1.25 g $\rightarrow$ 0.625 g $\rightarrow$ 0.313 g

Cont’d next page
7. How much phosphorous-32 was *originally present* if after 71.5 days there were 2.0 grams of $^{32}$P? ($^{32}$P half-life is 14.3 days)

$$71.5 \text{ d}/14.3\text{d per h.l.} = 5 \text{ h.l.}$$

$$\begin{array}{c}
\text{64.0g} & \leftarrow & \text{32.0g} & \leftarrow & \text{16.0g} & \leftarrow & \text{8.00g} & \leftarrow & \text{4.00g} & \leftarrow & \text{2.00g} \\
0 \text{ h.l.} & \quad & 1 \text{ h.l.} & \quad & 2 \text{ h.l.} & \quad & 3 \text{ h.l.} & \quad & 4 \text{ h.l.} & \quad & 5 \text{ h.l.}
\end{array}$$

8. A Gieger-Müller counter detects 300.0 counts per minute when a sample of neon-19 is placed near it. Using Reference Table N, calculate how long it will take for the Gieger-Müller counter to record 75 counts per minute in the $^{19}$Ne sample? $^{19}$Ne = 17.2 sec/h.l.

$$300 \text{ cpm}/75 \text{ cpm} = 25\% \text{ of original}$$

$$17.2 \text{ sec/h.l.} \times 2 \text{ h.l.} = 34.4 \text{ secs}$$

$$300 \text{ cpm} \rightarrow 150 \text{ cpm} \rightarrow 75 \text{ cpm}$$

9. A nuclear weapon test 58.2 years ago created strontium-90, which dispersed into the surrounding environment. A soil test today measures 20.0 $\mu$g of $^{90}$Sr in a 1 kg soil sample. How many $\mu$g of $^{90}$Sr per 1 kg of soil would have been just after the weapons test? (Use Reference Table N)

$$58.2 \text{ yr}/29.1\text{yr per h.l.} = 2 \text{ h.l.}$$

$$\begin{array}{c}
\text{80.0$\mu$g} & \leftarrow & \text{40.0$\mu$g} & \leftarrow & \text{20.$\mu$g} \\
0 \text{ h.l.} & \quad & 1 \text{ h.l.} & \quad & 2 \text{ h.l.}
\end{array}$$

10. You are given a 5000 year old biological sample your first day on a new job. Your new boss, who is not a nuclear chemist, tells you to use the ratio of $^{238}$U to $^{206}$Pb in the sample to determine the age of the sample. How would you politely explain to the boss that dating the sample this way may not be the best procedure?

$^{238}$U has a h.l. of $4.5 \times 10^9$ yrs. This is far too long for a 5,000 year old biological sample. $^{14}$C would be better for this age biological sample, as $^{14}$C h.l. is 5,700 yrs and all organisms take in $^{14}$C during life.
Artificial Transmutation:

Artificial Transmutation is the process of changing one element into another element.

Stars are the fusion ‘furnaces’ that generate natural elements. Stars may contain 92 natural elements, ranging from hydrogen (\(1\)H) to uranium (\(92\)U). Any element larger than Neptunium (\(93\)Np) to Ununoctium (\(118\)Uuo) is called a transuranium element and were created by people on Earth.

Describe what is needed to create a new element:

i. A sample target nuclei, usually a very massive one;

ii. A particle ‘bullet’ that has a charge (almost always a positive charge, i.e. an alpha particle or another nucleus, but electrons may be accelerated as well);

iii. A particle accelerator to make the charged particle ‘bullet’ move fast enough to collide with the target nuclei with enough force to combine both sub-atomic particles into a different element.
A particle accelerator uses **electromagnetic** fields to **accelerate** charged particles. Neutral particles, such as neutrons or gamma rays, cannot be accelerated in a particle accelerator.

Generally the larger the collider, the faster the charged ‘bullets’ may travel, and the heavier and heavier elements may be created.
**Topic:** Artificial Transmutation

**Objective:** How may we determine the products of transmutation?

**Determine the Products of Artificial Transmutation:**

**Step 1:** One of the products will be unknown. Add up the mass numbers on the side that all particles are known. The mass numbers of the other side should add up to the same thing.

**Step 2:** Add up the atomic numbers (charges) on the side where all particles are *known*. The charges of the other side should add up to the same thing.

**Step 3:** Look up the atomic number (if 3 or more) on the Periodic Table and identify the element you have. If the atomic number is 2 or less (including -1), look up the identity of the particle on Reference Table O.

\[
_{7}^{14}\text{N} + _{2}^{4}\text{He} \rightarrow _{8}^{17}\text{O} + \text{X}
\]

- On the top: \(14 + 4 = 17 + x\), so the mass is 1.
- On the bottom: \(7 + 2 = 8 + x\), so the charge is 1.
- This is a PROTON, according to Reference Table O.

\[
_{13}^{27}\text{Al} + _{2}^{4}\text{He} \rightarrow \text{X} + _{8}^{1}\text{H}
\]

- On the top: \(27 + 4 = x + 1\), so the mass is 30.
- On the bottom: \(13 + 2 = x + 0\), so the charge is 15.
- Element 15 is PHOSPHOROUS (P).
Transuranium Elements: (Atomic number beyond 92 Uranium)

$$^{239}_{94}Pu + ^{4}_{2}He \rightarrow ^{242}_{96}Cm + X$$

On the top: 239 + 4 = 242 + x, so the mass is 1.
On the bottom: 94 + 2 = 96 + x, so the charge is 0.
This is a NEUTRON, according to Reference Table O.

$$^{249}_{98}Cf + ^{12}_{6}C \rightarrow X + 4 ^{1}_{0}n$$
X is $^{257}_{104}$Rf

On the top: 249 + 12 = x + (4 × 1), so 281 = x + 4, the mass is 257.
On the bottom: 98 + 6 = x + (4 × 0), so 104 = x + 0, the charge is 104.
Element 104 is RUTHERFORDIUM (Rf).

Natural Decay vs Artificial Transmutation:

<table>
<thead>
<tr>
<th>Unique To Natural Decay</th>
<th>COMMON TO BOTH</th>
<th>Unique to Artificial Transmutation</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{234}<em>{91}$Pa $\rightarrow ^{0}</em>{1}e + ^{234}_{92}$U</td>
<td>Both form new elements from old ones.</td>
<td>Stable nucleus is forced to change into a less stable nucleus of a new element.</td>
</tr>
<tr>
<td>The left side of the equation has only the unstable nucleus, the right side has both the decay particle and the new, more stable nucleus.</td>
<td>In both, the masses on top of each side add up to the same, and the charges on the bottom of each side add up to the same.</td>
<td>The left side of the equation has the target nucleus and the particle bullet, the right side shows the results of that collision.</td>
</tr>
<tr>
<td>Produces energy through the destruction of mass.</td>
<td>Both follow Einstein’s equation $E=mc^2$. A tiny bit of mass (mass defect) is destroyed and energy is created.</td>
<td>Produces energy through the destruction of mass, however much more energy has to go into the process than comes out of it.</td>
</tr>
</tbody>
</table>
Topic: **Nuclear Power - Fission**

Objective: How are radioactive isotopes used for useful energy?

**Nuclear Power Generation:**

**Nuclear Power**-using Einstein’s \( E=mc^2 \) equation to create energy by the destruction of mass.

*Note: The Law of Conservation of Mass (& Energy) is applicable to physical & chemical processes, NOT nuclear processes.)*

**Nuclear Fission:** a few nuclei larger than \(^{56}\text{Fe}\) may be split into smaller nuclei, destroying a tiny amount of mass while releasing vast amounts of energy.

- Nuclear fission generates thousands of times more energy per unit of mass than fossil fuel energy production.
- Nuclear fission currently powers many commercial nuclear electricity reactors plus ships (aircraft carriers and submarines).
- **Nuclear Fission Reactors:**
  - Nuclear fission power generation in nuclear reactors use the following reaction to generate energy (which heats water):

\[
^{235}_{92}\text{U} + ^{1}_{0}\text{n} \rightarrow ^{92}_{36}\text{Kr} + ^{141}_{56}\text{Ba} + 3 ^{1}_{0}\text{n} + \text{energy}
\]
In this reaction, $^{235}\text{U}$ is the fissionable fuel in fuel rods containing uranium oxide pellets (enriched from a natural 0.72% $^{235}\text{U}$ to about 3% $^{235}\text{U}$). When a $^{235}\text{U}$ nucleus is struck by a slow-moving neutron, the $^{235}\text{U}$ nucleus splits into two (usually different sized) nuclei. The split (fission) also release 2 or 3 fast moving neutrons, which if slowed down, maybe used to split more $^{235}\text{U}$ nuclei.

In order to control the speed of the freed neutrons, the fuel rods are placed in a moderator, usually water, in the USA. The water also works to cool the fuel rods, and to carry the energy (as heat) out of the reactor. If the water cannot control the speed of the neutron flow, inert control rods are placed between the fuel rods that absorb the free neutrons, and can completely shut down the reaction when needed.
Topic: **Nuclear Power - Fusion**

Objective: How are radioactive isotopes used for useful energy?

**Nuclear Fusion:**

- Nuclear Fusion combines small nuclei into larger nuclei while losing some mass.
- The difference in mass between the two smaller nuclei and the final larger nuclei is called the mass defect and is converted to massive amounts of energy.
- Nuclear fusion generates millions of times more energy per unit of mass than fossil fuel energy production.
- Nuclear Fusion needs extreme temperatures and pressures to strip atoms of their electrons, forming a different phase of matter known as plasma. Plasma is formed entirely of positively charged atomic nuclei.
- Under intense heat and pressure, the positively charged plasma nuclei (normally repulsive) are forced together and join (fuse). The most common is the fusion of two hydrogen nuclei.

\[
^2_1H + ^2_1H \rightarrow ^4_2He + \text{energy}
\]

\[
^2_1H + ^2_1H \rightarrow ^4_2He + ^1_0n + \text{energy}
\]
Inside the Plasma

Tritium Nucleus

Helium Nucleus (3.5 MeV)

A fusion reaction occurs! 17.6 MeV of energy is released!

Deuterium Nucleus

E = mc²

Free Neutron (14.1 MeV)
**Topic: Nuclear Power - Summary**

Objective: How are radioactive isotopes used for useful energy?

### Nuclear Power Summary:

How are Nuclear Fission and Fusion compared to each other?

<table>
<thead>
<tr>
<th>Unique To Nuclear Fission</th>
<th>COMMON TO BOTH</th>
<th>Unique to Nuclear Fusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{235}<em>{92}$U + $^{1}</em>{0}$n $\rightarrow$ $^{92}<em>{36}$Kr + $^{141}</em>{56}$Ba + 3 $^{1}_{0}$n + energy</td>
<td>Both generate their energy the same way...by converting small amounts of mass (MASS DEFECT) into extraordinary amounts of energy.</td>
<td>$^{2}<em>{1}$H + $^{2}</em>{1}$H $\rightarrow$ $^{4}_{2}$He + energy</td>
</tr>
<tr>
<td>Reaction splits a large nucleus apart to form two smaller ones.</td>
<td></td>
<td>Reaction combines two small nuclei together to form one larger one.</td>
</tr>
<tr>
<td>Reaction is unknown in the natural world, is a form of artificial transmutation</td>
<td>Reaction requires temperatures of millions of degrees and vast pressures</td>
<td>All stars are powered by nuclear fusion</td>
</tr>
<tr>
<td>Reaction can take place at any temperature or pressure</td>
<td>Reaction has not been made energy-efficient enough for use</td>
<td>Hydrogen is the most abundant element in the universe</td>
</tr>
<tr>
<td>Reaction is currently being used to produce electricity for our use</td>
<td>Produces THOUSANDS of times more energy than conventional chemical explosives</td>
<td>Produces MILLIONS of times more energy than conventional chemical explosives</td>
</tr>
<tr>
<td>Requires mining to extract uranium ore</td>
<td>Produces radioactive wastes</td>
<td>Produces essentially no radioactive waste</td>
</tr>
</tbody>
</table>

Watch Crash Course Chemistry Nuclear Energy video

https://www.youtube.com/watch?v=FU6y1XIADdg
Notes page:
Nuclear Power homework

Circle your answer for each multiple choice question.

1. Which of the following isotopes may act as fissionable fuel in a fission reaction?
   a) $^{14}\text{C}$  b) $^{235}\text{U}$  c) $^{238}\text{U}$  d) $^{241}\text{Am}$

2. Which of the following nuclear reactions represents an example of the process of artificial transmutation?
   a) $^{96}_{42}\text{Mo} + ^2_1\text{H} \rightarrow ^1_0\text{n} + ^{97}_{43}\text{Tc}$  
   b) $^{104}_{47}\text{Ag} \rightarrow ^0_0\text{e} + ^{104}_{48}\text{Cd}$  
   c) $^{235}_{92}\text{U} + ^1_0\text{n} \rightarrow ^{92}_{36}\text{Kr} + ^{141}_{56}\text{Ba} + 3^1_0\text{n}$  
   d) $^2_1\text{H} + ^2_1\text{H} \rightarrow ^4_2\text{He}$

3. Which of the following nuclear reactions represents an example of the process of natural radioactive decay?
   a) $^{96}_{42}\text{Mo} + ^2_1\text{H} \rightarrow ^1_0\text{n} + ^{97}_{43}\text{Tc}$  
   b) $^{104}_{47}\text{Ag} \rightarrow ^0_0\text{e} + ^{104}_{48}\text{Cd}$  
   c) $^{235}_{92}\text{U} + ^1_0\text{n} \rightarrow ^{92}_{36}\text{Kr} + ^{141}_{56}\text{Ba} + 3^1_0\text{n}$  
   d) $^2_1\text{H} + ^2_1\text{H} \rightarrow ^4_2\text{He}$

4. Which of the following nuclear reactions represents an example of the process of nuclear fusion?
   a) $^{96}_{42}\text{Mo} + ^2_1\text{H} \rightarrow ^1_0\text{n} + ^{97}_{43}\text{Tc}$  
   b) $^{104}_{47}\text{Ag} \rightarrow ^0_0\text{e} + ^{104}_{48}\text{Cd}$  
   c) $^{235}_{92}\text{U} + ^1_0\text{n} \rightarrow ^{92}_{36}\text{Kr} + ^{141}_{56}\text{Ba} + 3^1_0\text{n}$  
   d) $^2_1\text{H} + ^2_1\text{H} \rightarrow ^4_2\text{He}$

5. Which of the following nuclear reactions represents an example of the process of nuclear fission?
   a) $^{96}_{42}\text{Mo} + ^2_1\text{H} \rightarrow ^1_0\text{n} + ^{97}_{43}\text{Tc}$  
   b) $^{104}_{47}\text{Ag} \rightarrow ^0_0\text{e} + ^{104}_{48}\text{Cd}$  
   c) $^{235}_{92}\text{U} + ^1_0\text{n} \rightarrow ^{92}_{36}\text{Kr} + ^{141}_{56}\text{Ba} + 3^1_0\text{n}$  
   d) $^2_1\text{H} + ^2_1\text{H} \rightarrow ^4_2\text{He}$

Cont’d next page
**Nuclear Power homework**

Complete the following equations by writing the correct atomic number, atomic mass, and the correct symbol for the missing particle ‘X’.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Particle X</th>
<th>Is this an example of natural decay or artificial transmutation?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) $^{40}<em>{20}$Ca + X $\longrightarrow$ $^{40}</em>{10}$K + $^{1}_{1}$H</td>
<td>Neutron or $^{1}_{0}$n</td>
<td>Artificial</td>
</tr>
<tr>
<td>2) $^{56}<em>{42}$Mo + $^{2}</em>{1}$H $\longrightarrow$ $^{1}<em>{0}$n + $^{97}</em>{43}$Tc</td>
<td>$^{97}_{43}$Tc</td>
<td>Artificial</td>
</tr>
<tr>
<td>3) $^{64}<em>{26}$Fe + $^{4}</em>{2}$He $\longrightarrow$ $^{2}<em>{1}$H + $^{66}</em>{26}$Fe</td>
<td></td>
<td>Artificial</td>
</tr>
<tr>
<td>4) $^{246}<em>{96}$Cm + $^{12}</em>{6}$C $\longrightarrow$ $^{4}<em>{0}$n + $^{254}</em>{102}$No</td>
<td>$^{254}_{102}$No</td>
<td>Artificial</td>
</tr>
<tr>
<td>5) $^{82}<em>{35}$Br $\longrightarrow$ $^{82}</em>{36}$Kr + X</td>
<td>$\beta^{-} or 0^{-1}$e</td>
<td>Natural</td>
</tr>
<tr>
<td>6) $^{19}<em>{10}$Ne $\longrightarrow$ $^{0}</em>{1}$e + X</td>
<td>$^{19}_{9}$F</td>
<td>Natural</td>
</tr>
<tr>
<td>7) $^{37}<em>{18}$Ar + $^{0}</em>{1}$e $\longrightarrow$ X</td>
<td>$^{37}_{17}$Cl</td>
<td>Natural process; NOT decay</td>
</tr>
<tr>
<td>8) $^{90}<em>{42}$Mo + $^{0}</em>{1}$n $\longrightarrow$ $^{90}_{43}$Tc + X</td>
<td>$\beta^{-} or 0^{-1}$e</td>
<td>Artificial</td>
</tr>
<tr>
<td>9) $^{40}<em>{18}$Ar + X $\longrightarrow$ $^{43}</em>{19}$K + $^{1}_{1}$H</td>
<td>$^4_2\alpha or ^4_2$He</td>
<td>Artificial</td>
</tr>
<tr>
<td>10) X $\longrightarrow$ $^{23}<em>{11}$Na + $^{0}</em>{1}$e</td>
<td>$^{23}_{12}$Mg</td>
<td>Natural</td>
</tr>
</tbody>
</table>
Notes page: