Chapter 8: Mass and Energy

Goals of Period 8

Section 8.1: To describe the relationship between energy and mass
Section 8.2: To define nuclear processes
Section 8.3: To discuss the nuclear force
Section 8.4: To determine the stability of nuclei

8.1 The Relationship between Energy and Mass

Einstein's Equation

We have seen that conservation of energy requires that all of the energy put into a process must be accounted for – energy cannot be created or destroyed. Likewise, conservation of mass says that matter cannot be created or destroyed. However, at the beginning of the 20th century Albert Einstein showed that energy can be converted into mass and mass can be converted into energy. Einstein reasoned that that mass is a measure of energy.

If energy is added to a system, the mass of the system increases. Conversely, if energy is removed from a system, the system's mass decreases. For example, winding the spring of a music box stores strain potential energy in the spring, and the mass of the spring has increased. Einstein stated this relationship between mass and energy in the famous equation,

(Equation 8.1)

$$E = Mc^2$$

where

E = energy (joules)
M = mass (kilograms)
c = the speed of light = 3×10^8 m/s

Einstein's Equation Applied to Physical Changes

When we wind up a music box, we stretch the box's spring, and the molecules of the spring's metal are pulled slightly further apart. This is a physical change because winding the spring does not change the chemical composition of the molecules making up the spring. As you wind the spring, you do work against the electromagnetic force holding the molecules together. This work adds energy to the spring in the form of stored strain potential energy. According to Equation 8.1, increasing the energy of the box also increases its mass. However, experience tells us that winding the spring of a music box does not cause any *measurable* difference in the mass of the box.

(Example 8.1)

You do 10 joules of work to wind the spring of a music box. What is the increase in mass of the box as a result of winding the spring?

$$\boldsymbol{M} = \frac{\boldsymbol{E}}{\boldsymbol{c}^2} = \frac{10\ \text{J}}{(3 \times 10^8\ \text{m/s})^2} = \frac{10\ \text{J}}{9 \times 10^{16}\ \text{m}^2/\text{s}^2} = 1.1 \times 10^{-16}\ \text{kg}$$

(Example 8.2)

A music box has a mass of 0.5 kg before its spring is wound. What percent of the mass of the box is the increase in mass as a result of doing 10 joules of work to wind its spring?

$$\frac{1.1 \times 10^{-16}\ \text{kg}}{5 \times 10^{-1}\ \text{kg}} = 2.2 \times 10^{-16} = 2.2 \times 10^{-14}\ \%$$

The increase in mass from a physical change, such as winding a spring, is extremely small. Such a change is not measurable. Next we consider the change in mass as a result of a chemical change.

Einstein's Equation Applied to Chemical Changes

In a chemical change, atoms of elements are rearranged to form new molecules or ions. To do so, the chemical bonds holding atoms together into molecules are broken and new bonds are formed. These chemical bonds result from the attractive electromagnetic forces between atoms. To break a chemical bond, binding energy must be added to overcome these attractive forces. In most cases, binding energy is given off when atoms combine to form chemical compounds. When hydrogen and oxygen molecules combine to form water, energy is given off as shown in Equation 8.2.

(Equation 8.2)

$$2\,H_2 + O_2 \rightarrow 2\,H_2O + \text{energy}$$

According to Einstein's equation, if energy is released when hydrogen and oxygen molecules combine to form water, the mass of the water molecules must be less than the mass of the hydrogen and oxygen molecules. In Example 8.3 we calculate the change in mass when two hydrogen atoms and one oxygen atom combine to form one molecule of water.

(Example 8.3)

Scientists have measured the binding energy of water as 4.75×10^{-19} joules per molecule. That is, 4.75×10^{-19} joules of energy are given off when two hydrogen atoms combine with one oxygen atom to form a molecule of water. What is the decrease in mass of the hydrogen and oxygen atoms when they combine into a water molecule?

$$\boldsymbol{M} = \frac{\boldsymbol{E}}{\boldsymbol{c}^2} = \frac{4.75 \times 10^{-19}\ \text{J}}{(3 \times 10^8\ \text{m/s})^2} = \frac{4.75 \times 10^{-19}\ \text{J}}{9 \times 10^{16}\ \text{m}^2/\text{s}^2} = 5.28 \times 10^{-36}\ \text{kg}$$

(Example 8.4)

The mass of one hydrogen atom is 0.168×10^{-26} kg and the mass of one oxygen atom is 2.66×10^{-26} kg. Find the total mass of these atoms before they combine into water. Then use your answer from Example 8.3 to find the mass of the water molecule formed when the two hydrogen atoms and one oxygen atom combine.

Add the mass of two hydrogen atoms and one oxygen atom to find the total mass of the molecule.

mass of 2 hydrogen atoms =	$2(0.168 \times 10^{-26}$ kg) =	0.336×10^{-26} kg
mass of 1 oxygen atom =	+	$\underline{2.66 \times 10^{-26}\ \text{kg}}$
total mass =		2.996×10^{-26} kg

Subtracting the change in mass (the mass converted into energy) found in Example 8.3 from the total mass of the atoms gives the mass of the water molecule. When numbers written in scientific notation are added or subtracted, all numbers must have the same power of ten.

total mass H and O atoms = 2.996×10^{-26} kg =	$29{,}960{,}000{,}000.00 \times 10^{-36}$ kg
less mass converted into energy	$\underline{-\qquad 5.28 \times 10^{-36}\ \text{kg}}$
mass of one H_2O molecule	$29{,}959{,}999{,}994.72 \times 10^{-36}$ kg

or 2.996×10^{-26} kg

(Example 8.5)

What percent of the mass of the hydrogen and oxygen atoms is the decrease in mass as the result of combining them into a water molecule?

$$\frac{5.28 \times 10^{-36}\ \text{kg}}{2.99 \times 10^{-26}\ \text{kg}} = 1.76 \times 10^{-10} = 1.76 \times 10^{-8}\ \%$$

Concept Check 8.1

One carbon atom combines with two oxygen atoms to form one molecule of carbon dioxide, CO_2. The binding energy of carbon dioxide is 6.54×10^{-19} joules/molecule.

a) What is the decrease in mass of the carbon and oxygen atoms when they combine into one CO_2 molecule?

b) One carbon atom has a mass of 1.99×10^{-26} kg and one oxygen atom has a mass of 2.66×10^{-26} kg. What is the total mass of one carbon atom and two oxygen atoms before they combine?

c) What percent of the mass of the carbon and oxygen atoms is the decrease in mass as the result of combining them into a molecule of CO_2?

From the examples above, we see that the increase or decrease in mass as the result of a chemical change is also extremely small. The amounts of energy given off when molecules break apart or atoms and molecules combine into new compounds is small compared to the mass of the reactants involved. Next we consider the change in mass of when the nuclei of atoms break apart and recombine. Changes to the atomic nucleus result in energies large enough to be measurable.

8.2 Nuclear Processes

The Discovery of Nuclear Processes

Atomic processes such as phase changes and chemical reactions involve rearrangements of electrons or of electron bonds between atoms. These bonds result from the electromagnetic forces between electrons and nuclei and between one electron and another. In such atomic processes nuclei are not disturbed, and atoms of chemical elements such as hydrogen, carbon and oxygen retain their identity. Thus in a chemical reaction such as

$$2H_2 + O_2 \quad \rightarrow \quad 2H_2O$$

the number of H atoms and of O atoms does not change.

We have seen that atomic processes result in changes of mass too small to measure. The binding energy released when changes are made to the electromagnetic force that binds molecules to one another in physical changes and that binds atoms into molecules in chemical changes is not large enough to result in a measurable change in mass. To release useful amounts of energy requires changing the bonds of a stronger force. The strong nuclear force that binds protons and neutrons together into a nucleus releases much larger amounts of binding energy.

By the late 1800's scientists had observed processes that gave off much more energy than expected from any atomic process. For example, uranium atoms were found to give off energetic particles with approximately 3×10^{-12} joules of energy per particle. These particles were named alpha particles (α) and were identified as the nuclei of helium atoms – two protons and two neutrons. We can conclude that the protons and neutrons emitted by an alpha particle must come from the nucleus of a uranium atom. In contrast to atomic processes, nuclear processes involve rearrangement of protons and neutrons in a nucleus, rearrangements of bonds between protons and neutrons to form different nuclei, or even changing a neutron into a proton or a proton into a neutron. Since the number of protons in a nucleus determines the identity of the atomic element, nuclear processes can change atoms of one element to atoms of a different element.

Atomic and Nuclear Processes

Much of the same language we used to describe chemical processes involving atoms is also useful to describe nuclear processes involving nuclei. Terms such as exothermic, endothermic, activation energy, and photon emission have the same meaning for nuclear processes as they do for atomic ones. For example, the chemical reaction that breaks up water molecules into hydrogen and oxygen atoms is endothermic because we have put energy into the process. In this case (electrolysis) it is electrical energy which is supplied. The explosive chemical reaction in which hydrogen atoms combine with oxygen atoms to produce water molecules is exothermic because energy is liberated.

Along with the energy released, this hydrogen-burning reaction makes a chemical product, water, which is harmless. However, you know that such chemical products are not always harmless, especially if made in large quantities: an example is the burning of carbon to produce CO_2. Similarly, as we will see, exothermic nuclear reactions make nuclear products along with the huge amounts of energy they release; these nuclear products can also be undesirable.

The Electron Volt as a Unit of Energy

In the exothermic chemical reaction $2(H_2) + O_2 \rightarrow 2(H_2O)$, 1.59 x 10^4 joules of energy are released for every gram of liquid water formed. These 1.59 x 10^4 joules are produced by the formation of a very large number of water molecules, about 3.34 x 10^{22}.

(Example 8.6)

How much chemical energy is released when one single water molecule is formed?

$$\frac{1.59 \times 10^4 \text{ J}}{3.34 \times 10^{22} \text{ molecules}} = 4.76 \times 10^{-19} \text{ J/molecule}$$

Since this is a very small amount of energy, it is convenient to use a suitably small energy unit, the ***electron volt (eV),*** for such situations.

$$1 \text{ electron volt (eV)} = 1.6 \times 10^{-19} \text{ joules}$$

Expressed in this tiny energy unit, the energy liberated for every molecule of water formed in the hydrogen-burning reaction is 3 eV. This 3 eV represents the difference between the activation energy you must supply (for example, with a match) to break up the H_2 and O_2 molecules into H and O atoms, and the larger amount of energy you get back when the H and O atoms combine to form water.

One eV is the electric potential energy of a single electron when the electron is raised to a voltage of 1 volt. More important for our purposes, energies per atom or per molecule of roughly one eV are typical of processes involving atomic electrons, such as the chemical reaction in this example. Nuclear processes, in contrast, have energies which are typically a million times larger, very roughly 10^6 eV (one million eV or 1 MeV) per nucleus.

Binding Energy

The concept of ***binding energy*** is useful in both atomic and nuclear physics. The binding energy of a molecule is the energy you must supply to break the molecule up into its constituent atoms. Similarly, the binding energy of a nucleus is the energy you must supply to break the nucleus up into free protons and neutrons. For example, the hydrogen molecule H_2 consists of two H atoms bound together by electric forces. To break this H_2 molecule apart into two H atoms requires 4.5 electron volts of energy.

Therefore, the binding energy of H_2 is 4.5 eV. When two H atoms combine to form H_2, an exothermic reaction, the same 4.5 eV of energy is released.

The nucleus of deuterium, or heavy hydrogen, consists of a proton and a neutron bound together by the nuclear force (discussed below). You must supply a very large amount of energy, 2.2 MeV (2.2 million eV) to break this nucleus up into a proton and a neutron. The binding energy of the deuterium nucleus is thus 2.2 MeV. When a proton and a neutron combine to form a deuterium nucleus, an exothermic reaction, 2.2 MeV of energy is released.

Unstable Molecules and Unstable Nuclei

Some chemical compounds are just barely bound together, and their molecules require very small activation energies in the form of heat or mechanical jostling to break apart and liberate several eV of energy. The chemical explosive nitroglycerine is an example of one such unstable compound. At high temperatures, nitroglycerine molecules can break apart spontaneously into smaller molecules over the course of time. These smaller molecules have a total binding energy that is larger than that of nitroglycerine, so energy is released whenever a nitroglycerine molecule breaks up.

Similarly, some nuclei are just barely bound together, and the nuclei can break apart (decay) spontaneously, liberating energies of roughly an MeV per nucleus. These unstable nuclei are said to be radioactive. When a radioactive nucleus decays, it does not fall apart into its constituent protons and neutrons. Rather, it breaks up into two or more pieces that have a total binding energy larger than that of the original unstable nucleus. We will discuss modes of radioactive decay in more detail next period.

Photon Emission

Some chemical reactions, like those that occur when a firefly flashes, result in the emission of light. As we learned in Period 3, an atomic electron can absorb energy and be raised to a higher energy level. Electrons can be raised to a higher energy level than the energy level they normally have by absorbing energy from a photon or from a chemical reaction. This higher-energy arrangement is called an ***excited state*** of the atom. A photon is emitted by each excited atom when it falls back to its normal energy level, called the ***ground state.*** In many cases, this electromagnetic energy is in the visible region of the spectrum and can be seen by your eye. Each such photon has a few electron volts of energy, again typical of atomic processes.

The same thing often happens in nuclear processes. As a result of a nuclear reaction, the protons and neutrons in the nuclei of a reaction product are left in an excited state. When the nucleus falls back to the nuclear ground state arrangement, a photon of radiant energy is emitted by the excited nucleus. However, this photon has an energy typical of nuclear energies, roughly a million eV (1 MeV). These very high energy photons, called gamma rays, are not visible to your eyes, but they are detectable by other means such as Geiger counters. You will do an experiment with a Geiger counter and a gamma ray source to see how much material it takes to absorb such gamma rays. This source, ${}^{60}_{27}Co$, emits gamma rays with an energy of about 1.2 MeV.

Some Comments on Nuclear Safety

We will discuss the biological effects of radiation in a future period. This discussion can be summed up by saying that radiation exposures, which are very small compared to the unavoidable natural radiation background, have immeasurably small effects on people. However, the best rule of thumb is still to keep each person's radiation exposure as low as possible. The added radiation exposure from the sources used in our experiments is negligible compared to what you are already getting from natural sources.

The sources we use in this course are very weak, and the amount of radiation per second emitted by them is small. But we also reduce the exposure by limiting the number of seconds that people are close to the sources. This makes sense -- it's what you do to avoid sunburn, for example. What counts in that case is both how bright the sun is, and how long you are exposed without intervening material like clothes or sunblock. We will reduce radiation exposure in the classroom by keeping all sources in a suitable shielded container in a far corner of the room when they are not in use. The radiation exposure you get from a source decreases rapidly with distance from a source, much as the light striking your body from a light bulb decreases as you move away from the bulb.

8.3 Nuclear Force

In this section, we address the question of why the energies involved in nuclear processes are so large and why nuclei are so small. The answer is that nuclei are held together by a completely new force, which acts over only very small distances, but which is very strong when nucleons are close together.

Protons and neutrons, both of which are found in nuclei, can be thought of as close cousins, with the family name ***nucleons***. Protons and neutrons have nearly the same mass, which is about 2000 times as large as the mass of an electron. The only significant difference between neutrons and protons is that neutrons have zero electric charge, while protons have a positive charge.

Electrons are held in an atom by the attractive electrical force from the protons in the nucleus. But the electrical force between two protons is repulsive, not attractive. So some other force, strong enough to overcome this electrical force, must be operating to hold protons together in a nucleus. The protons and neutrons in a nucleus are held together by a force that is neither electrical nor gravitational – it is an entirely different force, called the nuclear force. You never experience the nuclear force directly in everyday life. The reason for this is that the nuclear force depends on distance in a manner very different from electrical or gravitational forces. Remember that electrical and gravitational forces get weaker with distance in a rather gradual way: they decrease with the inverse square of separation distance D.

The nuclear force, by contrast, is almost zero for distances larger than the size of a small nucleus, but is very strong for smaller distances of separation. We say that the nuclear force has a definite range, i.e. a definite distance of separation at less than which the force abruptly becomes very strong. This range of the nuclear force is about 10^{-15} meters, about the same as the size of an individual proton or neutron, or the radius of a helium nucleus. We may characterize the nuclear force approximately by saying

$\boldsymbol{F_{nuc}}$ = 0 for distances greater than the range of about 10^{-15} m

$\boldsymbol{F_{nuc}}$ = a constant for distances less than about 10^{-15} m

To illustrate the nuclear force with an analogy, we can represent nucleons with ping pong balls. The nuclear force between the balls is as if the balls were covered with a strong glue. Once separated, the balls would experience no force between one another. But if the balls are close enough to touch one another, they stick together very strongly. We never experience the nuclear force directly in everyday life because we are always dealing with separation distances between nuclei that are large compared to the range of the nuclear force. Your body and the objects you touch are made of atoms, which are about 100,000 times larger than the nuclei inside them.

Since the nuclear force is not directly observable in everyday situations, it has required many difficult experiments by physicists to learn about its properties. It turns out to be surprisingly complicated. For instance, at distances very much smaller than the range discussed above, the nuclear force actually becomes repulsive rather than attractive. The result of this is to make the nucleons in a nucleus behave in many ways like a sack of ping pong balls, with nearly constant distance of separation between neighboring nucleons. A simplifying feature of the nuclear force is that it is the same for all nucleons (protons and neutrons): two protons, two neutrons, or a proton and a neutron, attract each other equally strongly. However, electrons do not feel the nuclear force and cannot be held inside a nucleus.

Within its short range, the nuclear force is very strong. This large strength of the nuclear force is the origin of the very large energies involved in nuclear processes. We can get an idea of this strength from the following thought experiment. Imagine that you could somehow bring two protons together, decreasing their distance of separation. The protons repel each other electrically, so you have to push them together, as if you were compressing a spring. You are storing up electric potential energy. When the protons are separated by a distance about the size of an atom, it is not hard to show that this potential energy is about 10 eV, an energy typical of atomic processes. As you continue to squeeze the protons together, the repulsive force increases, and so does the stored potential energy. When the separation gets below nuclear size, to the range of the nuclear force, this electrical potential energy reaches roughly a million eV. At this separation distance, the attractive nuclear force between the protons suddenly becomes very strong and nearly succeeds in holding the protons together despite the electrical repulsion. (In actual fact, two protons don't quite stick together, but a proton and a neutron do.) Thus the energies we are dealing with when the nuclear force is involved are roughly comparable to the electrical potential energy of two protons separated by a distance equal to the range of nuclear forces, or about one MeV.

8.4 Isotopes and Nuclear Stability

As discussed previously, the chemical properties of an element are determined by the number of protons in its nucleus, since this determines how many electrons the neutral atom will have. The number of protons in the nucleus is called the ***atomic number,*** usually denoted by ***Z***. There is a one-to-one correspondence between the atomic number and the element name. A few of the familiar elements and their atomic numbers are shown in Table 8.1

Table 8.1 Common Elements

Atomic Number Z	Element Name	Element Symbol
1	Hydrogen	H
2	Helium	He
5	Carbon	C
8	Oxygen	O
26	Iron	Fe
82	Lead	Pb

The mass of an atom is determined mainly by the total number of nucleons, both protons and neutrons, in its nucleus. This total number of nucleons is called ***the atomic mass number***, denoted by ***A***. The ***neutron number***, denoted by ***N***, is

(Equation 8.3)

$$N = A - Z$$

where

N = the number of neutrons

A = the number of nucleons (neutrons and protons)

Z = the number of protons

(Example 8.7)

How many neutrons are contained in one molecule of Cobalt-60, ${}^{60}_{27}Co$?

$$N = A - Z = 60 \text{ nucleons} - 27 \text{ protons} = 33 \text{ neutrons}$$

In general, the number of neutrons N in the nuclei of a given element Z is not unique. For example, every nucleus of the element carbon has 6 protons (Z = 6). About 99% of the carbon nuclei in the world have 6 neutrons, so for these

$$A = Z + N = 6 \text{ protons} + 6 \text{ neutrons} = 12 \text{ nucleons}$$

Most of the remaining 1% of carbon nuclei in the world contain 7 neutrons, so that for these, A = 6 protons + 7 neutrons = 13 nucleons. It is also possible to produce (via nuclear reactions) carbon nuclei with 4, 5, 8 or 9 neutrons. However, these subspecies of carbon are not bound tightly enough to stick together permanently. These subspecies of carbon are unstable and undergo radioactive decay to other nuclei with larger binding energies.

Nuclei of the same element (same number of protons, Z) that have different numbers of neutrons (different N) are called ***isotopes*** of the element. A standard notation for specifying an isotope is to give its chemical symbol with Z and A as sub- and superscripts. Thus for example, ${}^{12}_{6}C$ and ${}^{13}_{6}C$ are stable isotopes of carbon, but ${}^{14}_{6}C$ is one of the unstable isotopes of carbon. An alternative notation is the element name or symbol followed by the A value, for example, Carbon-14 or C-14.

What determines whether an isotope is stable or unstable? Another way of asking the same question is, what determines whether the binding energy of an isotope is large or small? This is generally a complex topic, but some rules of thumb are useful.

- ***Nuclei with very large Z are unstable.*** The heaviest nucleus which is completely stable has Z = 83. The reason that nuclei with an arbitrarily large number of protons can't hold together comes from the electrical repulsion of each proton by every other proton, and from the short range of the nuclear force. For a very large nucleus, this range is too short to allow nucleons on one side of a nucleus to attract the ones on the opposite side -- they are out of range. So for a sufficiently large nucleus, the repulsive electrical force eventually wins over the attractive, short range nuclear force. Some isotopes with Z larger than 83 are nearly stable, so that the nuclei take billions of years to decay. Two examples are ${}^{238}_{92}U$ (Uranium-238) and ${}^{232}_{90}Th$ (Thorium-232).

- ***Stable nuclei tend to have the same number of neutrons and protons,*** N = Z. One might guess that it would be very easy to stick lots of neutrons together. The nuclear force between neutrons is attractive and, since they are electrically neutral, there are no repulsive electrical forces. However, nature has another surprise for us. It turns out that there is a binding energy penalty for having too many nucleons of the same kind in a nucleus; if too many nucleons of the same kind are present, the binding energy is reduced. The nuclei that are stuck together most tightly (are most tightly bound) tend to have equal numbers of protons and neutrons. This is true of small, light nuclei like carbon or oxygen. Very large, heavy nuclei that contain many protons (like gold, lead and uranium) contain more neutrons than protons in spite of this energy penalty. This is to compensate for the large repulsive electrical forces from the many protons in the nucleus.

Concept Check 8.2

Which of these isotopes are likely to be stable? Which are unstable? Why?

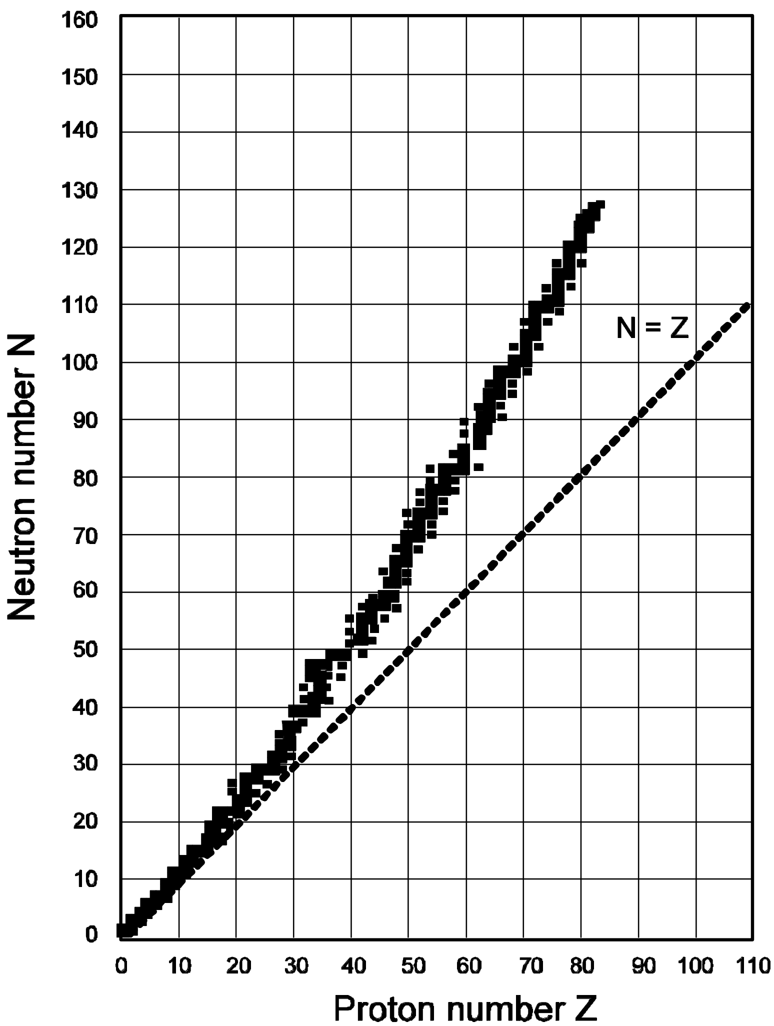
a) ${}^{251}_{98}Cf$ Californium-251

b) ${}^{14}_{7}N$ Nitrogen-14

c) ${}^{24}_{10}Ne$ Neon-24

Figure 8.1 illustrates the stable nuclei as a graph of the number of neutrons versus the number of protons. The dotted line, in comparison, represents nuclei with equal numbers of neutrons and protons. Notice that for nuclei with few protons, stable nuclei have equal numbers of protons and neutrons. Stable nuclei with many protons have more neutrons than protons.

Figure 8.1 Graph of neutron number N vs proton number Z for stable nuclei



Period 8 Summary

8.1: Atoms consist of a nucleus of nucleons – positively charged protons and neutrons with no charge. Nucleons make up 99.9% of the atom's mass.

Surrounding the nucleus is a cloud of electrons. Electrons have negative charge and very little mass. The diameter of the electron cloud is 100,000 times the diameter of the nucleus.

The number of protons (Z) determines which element an atom is.

8.2: Chemical changes involve making and breaking chemical bonds to create new molecules. Nuclear changes involve changes to the nuclei of atoms that result in the formation of atoms of new elements.

The binding energy of a molecule is the energy you must supply to break the molecule up into its constituent atoms. The binding energy of a nucleus is the energy you must supply to break the nucleus up into free protons and neutrons.

8.3: The nuclear force binds nucleons into a nucleus by overcoming the electromagnetic repulsion of the positively charged protons. The nuclear force operates over very short distances of approximately the diameter of an atom.

8.4: Isotopes are nuclei of the same element (same number of protons, Z) that have different numbers of neutrons (N). Some isotopes are unstable and decay into more stable nuclei.

Nuclei with more than 83 protons are unstable because the range of the nuclear force is too short to allow nucleons on one side of a large nucleus to attract nucleons on the other side of the nucleus.

Stable nuclei of smaller atoms tend to have the same number of neutrons and protons. Stable nuclei of large atoms that contain many protons (like gold, lead and uranium) have more neutrons than protons to compensate for the large repulsive electrical forces from the many protons in the nucleus.

Period 8 Exercises

E.1 Protons are held in an atomic nucleus by the strong nuclear force. The attraction of the nuclear force is opposed by which of the following?

a) the repulsive nuclear force between neutrons
b) the attractive nuclear force between electrons
c) the mutual repulsion between like positive charges
d) random molecular collisions
e) the gravitational force

E.2 How much energy would be released if 3 kilograms of mass were completely changed into energy?

a) 3×10^{8} J
b) 1.33×10^{16} J
c) 9×10^{16} J
d) 2.7×10^{17} J
e) 27×10^{24} J

E.3 The binding energy of an H_2 molecule is 4.5 eV. Two grams of H_2 contain 6.0×10^{23} molecules. How many joules of energy are released when 2 grams of H atoms combine to form molecules of H_2? (Hint: 1 eV = 1.6×10^{-19} joules.)

a) 7.2×10^{-19} J
b) 2.2×10^{5} J
c) 4.3×10^{5} J
d) 1.4×10^{24} J
e) 2.7×10^{24} J

E.4 The binding energy of a deuterium nucleus, which consists of a proton and a neutron bound together by the nuclear force, is 2.2 MeV. Two grams of deuterium contain 6.0×10^{23} molecules. How many joules of energy are released in the formation of 2 grams of deuterium from free protons and neutrons?

a) 3.5×10^{-13} J
b) 1.05×10^{11} J
c) 2.1×10^{11} J
d) 1.3×10^{30} J
e) 2.6×10^{30} J

E.5 Which one of the following isotopes would you expect to be stable? Why might each of the others be unstable?

a) $^{232}_{90}Th$ (Thorium-232)

b) $^{40}_{19}K$ (Potassium-40)

c) $^{12}_{6}C$ (Carbon-12)

d) $^{14}_{8}O$ (Oxygen-14)

e) $^{14}_{7}N$ (Nitrogen-14)

Period 8 Review Questions

R.1 Suppose that the size of an atom were represented by the diameter of the OSU oval (about 250 meters). On this scale, how large would the nucleus of the atom be? What sort of object could represent it?

R.2 Why does an H_2 molecule have very nearly the same mass as a deuterium atom?

R.3 What is the difference between a chemical change and a nuclear process? Why is the energy released when H atoms combine to form H_2 molecules so much less than the energy released when protons and neutrons combine to form deuterium molecules?

R.4 What is an isotope of an element? Which of the following isotopes of carbon – ${}^{12}_{6}C$ or ${}^{14}_{6}C$ – is stable? Why?

R.5 Why do stable nuclei of large atoms that contain many protons have more neutrons than protons?