Fundamentals of Light Water Nuclear Reactor Atomic and Nuclear Physics

An Online Continuing Education Course for Engineers

Course Number: R-3019

Credit: 3 Hours / 3 PDH / 3 CPD

Fundamentals of Light Water Nuclear Reactor Atomic and Nuclear Physics

Ralph K. Schwartzbeck, M.S., P.E.

Introduction

In 1905, Albert Einstein proposed the theory of relativity, which included a simple but elegant equation showing that mass and energy are proportional.

$E = mc^2$

where:

E = energy,

m = mass, and

$\boldsymbol{c}$ = speed of light.

This equation is perhaps the most famous equation in the world and is even familiar to school children. However, this relationship between mass and energy alludes to fundamental properties of the universe itself.

The physics that underpin this equation are somewhat complex and not well comprehended by those not well versed in higher mathematics. But for most, the significance of this formula boils down to the fact there is a large amount of energy bound inside the matter surrounding us.

Since Einstein's famous equation was introduced, scientists and engineers have used this equation to provide the basis for a peaceful energy source: the production of electrical energy using the energy that exists within the nucleus of the atom. For example, the nuclear fission process is one example of atomic mass being converted into useful energy in nuclear reactors. Since all nuclear power electrical generating stations rely on nuclear fission, this course will focus on describing the basic concepts of atomic and nuclear physics without the complex mathematics.

Production of electricity by nuclear reactors is considered by many a "renewable" method of energy production. Also, current electricity production by nuclear energy provides the largest "baseload" of electrical generation capacity in the world. Thus, a thorough understanding of how energy is produced via the atom is paramount.

This course will provide an introduction to how energy is produced using Einstein's equation. It will begin with the basics of the atomic structure and conclude with a description of the nuclear fission process. To ensure a working knowledge of the physics involved with the nuclear fission process (without the complex mathematics), this course will also include the topics of mass defect, binding energy, radioactivity, radioactive half-life, and neutron interactions with matter, including other atomic elements.

Atomic Structure

To understand the atomic structure, one must start with the definition of matter. In general, matter is any substance that has mass and occupies space. Our senses can easily distinguish matter because it has physical and chemical properties.

Investigation of the chemical and physical properties of matter indicates that all naturally occurring matter is made up of over 100 basic chemical substances known as ***elements***. Each element has different chemical and physical properties. Elements are familiar to all of us as either 1) metals (such as iron, aluminum, zinc, lead, copper, tin, and mercury) or 2) nonmetals (such as carbon, sulfur, phosphorous, iodine, oxygen, hydrogen, and nitrogen). The smallest particle of an element that retains the properties of that element is called the ***atom***.

Structure of Matter

In 1661, the English chemist Robert Boyle published the modern criterion for an element. He defined an element to be a basic substance that cannot be broken down into any simpler substance after it is isolated from a compound, but can be combined with other elements to form compounds. To date, 105 different elements have been confirmed to exist, and researchers claim to have discovered three additional elements. Of the 105 confirmed elements, 90 exist in nature, and 15 are man-made.

Another basic concept of matter that the Greeks debated was whether matter was continuous or discrete. That is, whether matter could be continuously divided and subdivided into ever smaller particles or whether, eventually, an indivisible particle would be encountered. Democritus, in about 450 B.C., argued that substances were ultimately composed of small, indivisible particles that he labeled atoms. He further suggested that different substances were composed of different atoms or combinations of atoms and that one substance could be converted into another by rearranging the atoms. It was impossible to conclusively prove or disprove this proposal for more than 2000 years.

The modern proof for the atomic nature of matter was first proposed by the English chemist John Dalton in 1803. Dalton stated that each chemical element possesses a particular kind of atom, and any quantity of the element is made up of identical atoms of this kind. What distinguishes one element from another element is the kind of atom of which it consists, and the basic physical difference between kinds of atoms is their weight.

For almost 100 years after Dalton established the atomic nature of atoms, it was considered impossible to divide the atom into even smaller parts. All of the results of chemical experiments during this time indicated that the atom was indivisible. Eventually, experimentation into electricity and radioactivity indicated that particles of matter smaller than the atom did indeed exist.

In 1906, J. J. Thompson won the Nobel Prize in physics for establishing the existence of electrons. ***Electrons*** are negatively charged particles that have 1/1835 the mass of the hydrogen atom. Soon after the discovery of electrons, protons were discovered. ***Protons*** are relatively large particles that have almost the same mass as a hydrogen atom and a positive charge equal in magnitude (but opposite in sign) to that of the electron. The third subatomic particle to be discovered, the neutron, was not found until 1932. The ***neutron*** has almost the same mass as the proton but is electrically neutral. Figure 1 is an illustration of a simple carbon-12 atom showing the basic subatomic particles.

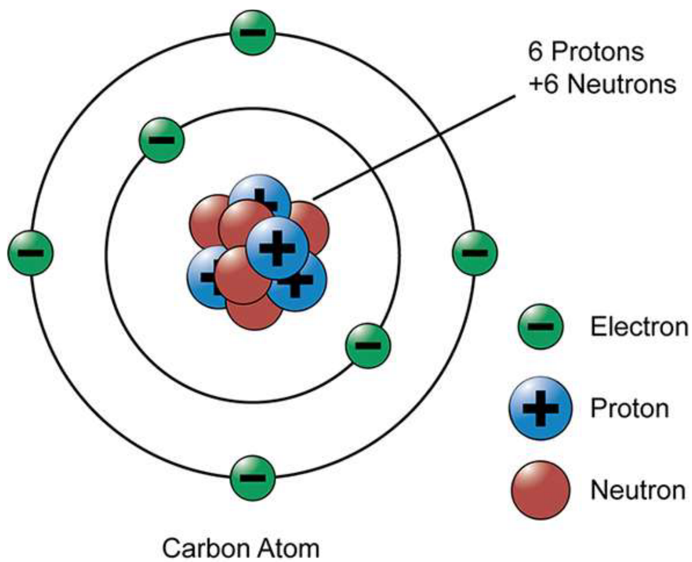

Figure 1: Simple Carbon-12 Atomic Model

Bohr's Model of the Atom

The British physicist Ernest Rutherford postulated that the positive charge in an atom is concentrated in a small region called a ***nucleus*** at the center of the atom, with electrons existing in orbits around it. Niels Bohr, coupling Rutherford's postulation with the quantum theory introduced by Max Planck, proposed that the atom consists of a dense nucleus of protons surrounded by electrons traveling in discrete orbits at fixed distances from the nucleus. An electron in one of these orbits or shells has a specific or discrete quantity of energy (quantum). When an electron moves from one allowed orbit to another allowed orbit, the energy difference between the two states is emitted or absorbed in the form of a single quantum of radiant energy called a photon.

Figure 2 is Bohr's model of the hydrogen atom, showing an electron as having just dropped from the third shell to the first shell with the emission of a photon that has an energy equal to hv (h = Planck's constant = 6.63×10^{-34} J-s and v = frequency of the photon). Bohr's theory was the first to successfully account for the discrete energy levels of this radiation as measured in the laboratory. Although Bohr's atomic model is designed specifically to explain the hydrogen atom, his theories apply generally to the structure of all atoms.

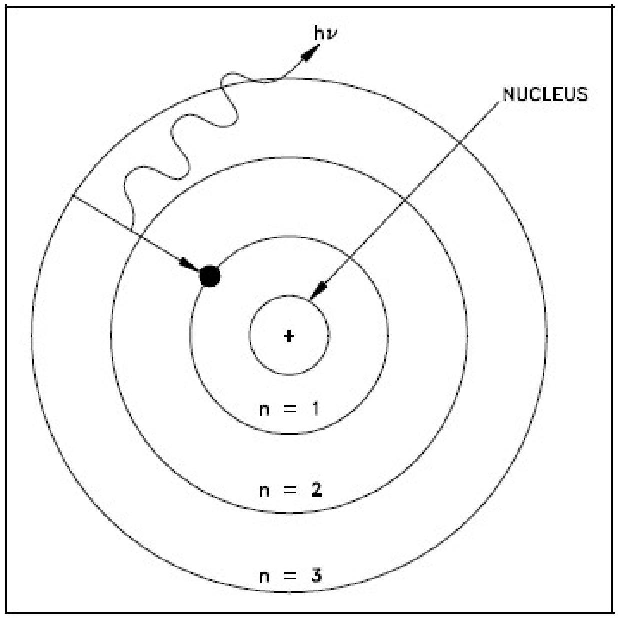

Figure 2: Bohr's Model of the Hydrogen Atom

The most important properties of the three particles that make up an atom are indicated in Table 1. Note that the masses of the proton and neutron are approximately equal and are much greater than the electron.

Table 1: Properties of Subatomic Particles

Particle	**Location**	**Charge**	**Mass**
Neutron	Nucleus	None	1.008665 amu
Proton	Nucleus	+1	1.007277 amu
Electron	Shells around the Nucleus	-1	0.0005486 amu

Measuring Units on the Atomic Scale

The size and mass of atoms are so small that the use of normal measuring units, while possible, is often inconvenient. Therefore, units of measure have been defined for mass and energy on the atomic scale to make measurements more convenient to express. The unit of measure for atomic mass is the ***atomic mass unit (amu)***. One atomic mass unit is equal to 1.66×10^{-24} grams. Note from Table 1 that the mass of a neutron and a proton are approximately equal; both are about 1 amu. The unit for energy is the ***electron volt (eV).*** The electron volt is the amount of energy acquired by a single electron when it falls through a potential difference of one volt.

$1 \text{ eV} = 1.602 \times 10^{-19}$ joules

$= 1.18 \times 10^{-19}$ foot-pounds

$= 1.52 \times 10^{-22}$ BTU

$= 4.45 \times 10^{-26}$ kW-hrs

Each of the elements has a characteristic mass. This is called the ***atomic weight*** of the element. The atomic weight is the mass of the average atom of the element found in nature and is normally given in atomic mass units (amu).

Nuclides

The basic building blocks of the atom's ***nucleus*** are protons and neutrons. Since these particles are found in the atom's nucleus, they are also termed ***nucleons.*** Each type of atom that contains a unique combination of protons and neutrons is called a ***nuclide***. Not all combinations of numbers of protons and neutrons are possible, but about 2500 specific nuclides with unique combinations of neutrons and protons have been identified. Each nuclide is denoted by the chemical symbol of the element, with the atomic number written as a subscript and the mass number written as a superscript, as shown below.

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

where:

X is the symbol of the chemical element

A is the Mass Number

Z is the Atomic Number

Because each element has a unique name, chemical symbol, and atomic number, only one of the three is necessary to identify the element. For this reason, nuclides can also be identified by either the chemical name or the chemical symbol followed by the mass number (for example, U-235 or uranium-235). Another common format is to use the abbreviation of the chemical element with the mass number superscripted (for example, U^{235}). In this course, the format used in the text will usually be the element's name followed by the mass number (i.e., U-235).

Some nuclides occur naturally, and some are produced artificially; some are stable, and some are not. Non-stable nuclides are generally referred to as ***radionuclides***. Nuclides have different nuclear properties depending on the number and type of nucleons in the nucleus. Larger nuclides tend to have more neutrons than protons in their nuclei.

The total number of protons in the nucleus of an atom is called the ***atomic number*** of the atom and is given the symbol **Z**. Each element has a distinct number of protons in its nucleus. For example, all oxygen atoms have exactly eight protons in their nucle… other element has only eight protons in its nucleus. Another example is uranium, which has e…

The number of elec… f protons in the nucleus. The numb… is given the symbol **N**. The ***mass numb…*** number of protons AND neutrons in the… nd by the equation Z + N = A. Each of th… ms of different elements contain a di… ifies the particular element.

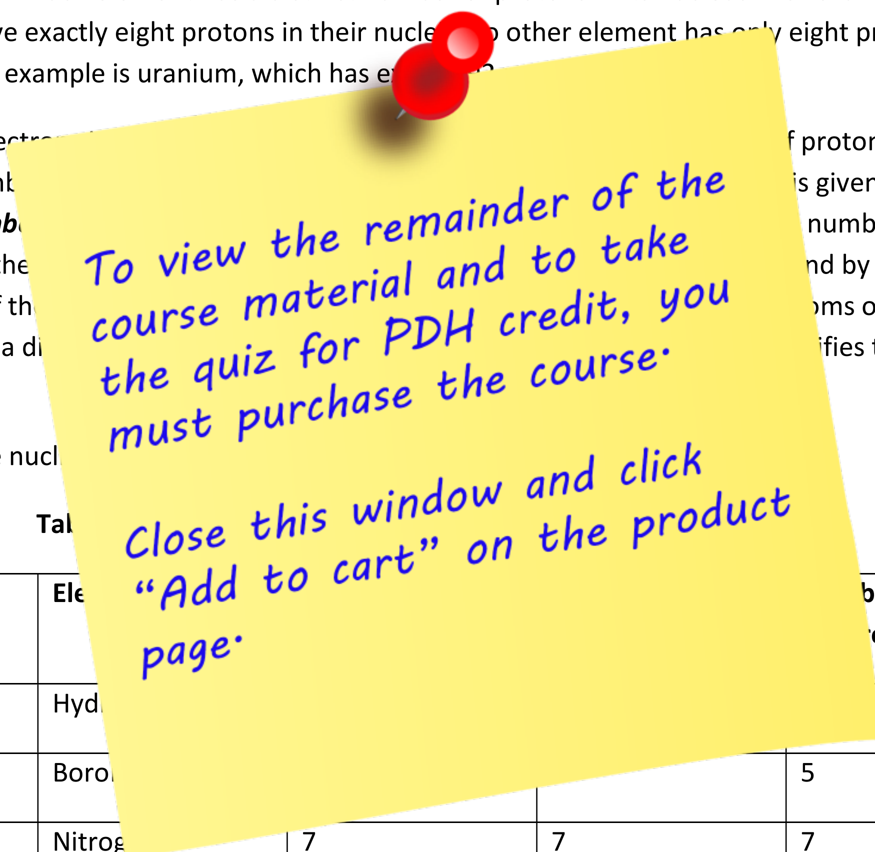

Table 2 shows the nucl…

Tal…

Nuclide	Ele…			…ber of …ons
$_1H^1$	Hyd…			
$_5B^{10}$	Boro…			5
$_7N^{14}$	Nitrog…	7	7	7
$_{48}Cd^{114}$	Cadmium	48	48	66
$_{92}U^{235}$	Uranium	92	92	143
$_{94}Pu^{239}$	Plutonium	94	94	145

Isotopes

Isotopes are nuclides that have the same atomic number (Z) and are, therefore, the same element but differ in the number of neutrons. This results in some atoms of the same element with different atomic