ATOMIC NATURE OF MATTER

All matter is composed of atoms. The atom is the smallest amount of matter that retains the properties of an element. Atoms themselves are composed of smaller particles, but these smaller particles no longer have the same properties as the overall element.

EO 1.1 **STATE the characteristics of the following atomic particles, including mass, charge, and location within the atom:**

a. Proton
b. Neutron
c. Electron

EO 1.2 **DESCRIBE the Bohr model of an atom.**

EO 1.3 **DEFINE the following terms:**

a. Nuclide **c. Atomic number**
b. Isotope **d. Mass number**

EO 1.4 **Given the standard $^{A}_{Z}X$ notation for a particular nuclide, DETERMINE the following:**

a. Number of protons
b. Number of neutrons
c. Number of electrons

EO 1.5 **DESCRIBE the three forces that act on particles within the nucleus and affect the stability of the nucleus.**

Structure of Matter

Early Greek philosophers speculated that the earth was made up of different combinations of basic substances, or elements. They considered these basic elements to be earth, air, water, and fire. Modern science shows that the early Greeks held the correct concept that matter consists of a combination of basic elements, but they incorrectly identified the elements.

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.

Subatomic Particles

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 it is electrically neutral.

Bohr 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 1 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 = $h\nu$. (h = Planck's constant = 6.63×10^{-34} J-s and ν = 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. Additional information on electron shell theory can be found in the Chemistry Fundamentals Handbook.

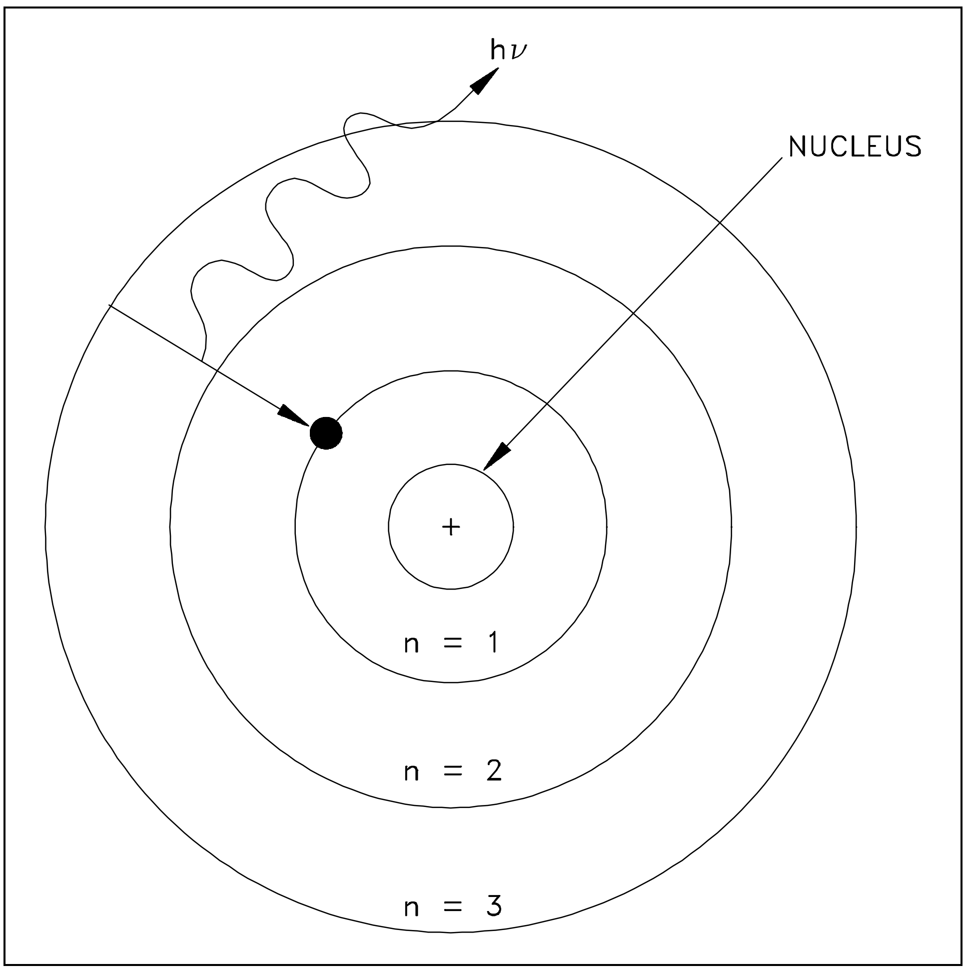

Figure 1 Bohr's Model of the Hydrogen Atom

Properties of the three subatomic particles are listed in Table 1.

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 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. 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 mass is the atomic mass unit (amu). One atomic mass unit is equal to 1.66 x 10^{-24} grams. The reason for this particular value for the atomic mass unit will be discussed in a later chapter. Note from Table 1 that the mass of a neutron and a proton are both 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. One electron volt is equivalent to 1.602 x 10^{-19} joules or 1.18 x 10^{-19} foot-pounds.

Nuclides

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. The number of electrons in an electrically-neutral atom is the same as the number of protons in the nucleus. The number of neutrons in a nucleus is known as the neutron number and is given the symbol N. The *mass number* of the nucleus is the total number of nucleons, that is, protons and neutrons in the nucleus. The mass number is given the symbol A and can be found by the equation Z + N = A.

Each of the chemical elements has a unique atomic number because the atoms of different elements contain a different number of protons. The atomic number of an atom identifies the particular element.

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 in Figure 2. 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, ^{235}U). In this handbook the format used in the text will usually be the element's name followed by the mass number. In equations and tables, the format in Figure 2 will usually be used.

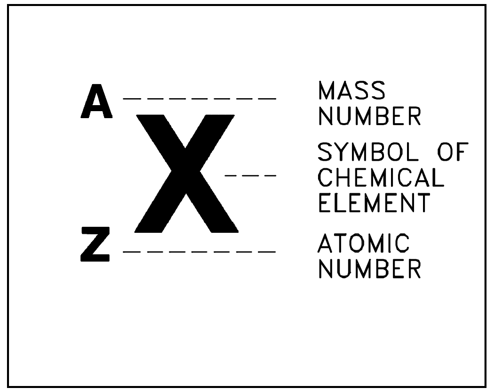

Figure 2 Nomenclature for Identifying Nuclides

Example:

State the name of the element and the number of protons, electrons, and neutrons in the nuclides listed below.

$^{1}_{1}H$

$^{10}_{5}B$

$^{14}_{7}N$

$^{114}_{48}Cd$

$^{239}_{94}Pu$

Solution:

The name of the element can be found from the Periodic Table (refer to Chemistry Fundamentals Handbook) or the Chart of the Nuclides (to be discussed later). The number of protons and electrons are equal to Z. The number of neutrons is equal to Z - A.

Nuclide	Element	Protons	Electrons	Neutrons
$^{1}_{1}H$	hydrogen	1	1	0
$^{10}_{5}B$	boron	5	5	5
$^{14}_{7}N$	nitrogen	7	7	7
$^{114}_{48}Cd$	cadmium	48	48	66
$^{239}_{94}Pu$	plutonium	94	94	145

Isotopes

Isotopes are nuclides that have the same atomic number and are therefore the same element, but differ in the number of neutrons. Most elements have a few stable isotopes and several unstable, radioactive isotopes. For example, oxygen has three stable isotopes that can be found in nature (oxygen-16, oxygen-17, and oxygen-18) and eight radioactive isotopes. Another example is hydrogen, which has two stable isotopes (hydrogen-1 and hydrogen-2) and a single radioactive isotope (hydrogen-3).

The isotopes of hydrogen are unique in that they are each commonly referred to by a unique name instead of the common chemical element name. Hydrogen-1 is almost always referred to as hydrogen, but the term protium is infrequently used also. Hydrogen-2 is commonly called deuterium and symbolized $^{2}_{1}D$. Hydrogen-3 is commonly called tritium and symbolized $^{3}_{1}T$. This text will normally use the symbology $^{2}_{1}H$ and $^{3}_{1}H$ for deuterium and tritium, respectively.

Atomic and Nuclear Radii

The size of an atom is difficult to define exactly due to the fact that the electron cloud, formed by the electrons moving in their various orbitals, does not have a distinct outer edge. A reasonable measure of atomic size is given by the average distance of the outermost electron from the nucleus. Except for a few of the lightest atoms, the average atomic radii are approximately the same for all atoms, about 2×10^{-8} cm.

Like the atom the nucleus does not have a sharp outer boundary. Experiments have shown that the nucleus is shaped like a sphere with a radius that depends on the atomic mass number of the atom. The relationship between the atomic mass number and the radius of the nucleus is shown in the following equation.

$$r = (1.25 \times 10^{-13} \text{ cm}) A^{1/3}$$

where:

r = radius of the nucleus (cm)
A = atomic mass number (dimensionless)

The values of the nuclear radii for some light, intermediate, and heavy nuclides are shown in Table 2.

TABLE 2
Calculated Values for Nuclear Radii

Nuclide	Radius of Nucleus
$^{1}_{1}H$	1.25×10^{-13} cm
$^{10}_{5}B$	2.69×10^{-13} cm
$^{56}_{26}Fe$	4.78×10^{-13} cm
$^{178}_{72}Hf$	7.01×10^{-13} cm
$^{238}_{92}U$	7.74×10^{-13} cm
$^{252}_{98}Cf$	7.89×10^{-13} cm

From the table, it is clear that the radius of a typical atom (e.g. 2×10^{-8} cm) is more than 25,000 times larger than the radius of the largest nucleus.

Nuclear Forces

In the Bohr model of the atom, the nucleus consists of positively-charged protons and electrically-neutral neutrons. Since both protons and neutrons exist in the nucleus, they are both referred to as nucleons. One problem that the Bohr model of the atom presented was accounting for an attractive force to overcome the repulsive force between protons.

Two forces present in the nucleus are (1) electrostatic forces between charged particles and (2) gravitational forces between any two objects that have mass. It is possible to calculate the magnitude of the gravitational force and electrostatic force based upon principles from classical physics.

Newton stated that the *gravitational force* between two bodies is directly proportional to the masses of the two bodies and inversely proportional to the square of the distance between the bodies. This relationship is shown in the equation below.

$$F_g = \frac{G\ m_1\ m_2}{r^2}$$

where:

F_g = gravitational force (newtons)
m_1 = mass of first body (kilograms)
m_2 = mass of second body (kilograms)
G = gravitational constant (6.67×10^{-11} N-m^2/kg^2)
r = distance between particles (meters)

The equation illustrates that the larger the masses of the objects or the smaller the distance between the objects, the greater the gravitational force. So even though the masses of nucleons are very small, the fact that the distance between nucleons is extremely short may make the gravitational force significant. It is necessary to calculate the value for the gravitational force and compare it to the value for other forces to determine the significance of the gravitational force in the nucleus. The gravitational force between two protons that are separated by a distance of 10^{-20} meters is about 10^{-24} newtons.

Coulomb's Law can be used to calculate the force between two protons. The *electrostatic force* is directly proportional to the electrical charges of the two particles and inversely proportional to the square of the distance between the particles. Coulomb's Law is stated as the following equation.

$$F_e = \frac{K\ Q_1\ Q_2}{r^2}$$

where:

F_e = electrostatic force (newtons)
K = electrostatic constant (9.0×10^9 N-m^2/C^2)
Q_1 = charge of first particle (coulombs)
Q_2 = charge of second particle (coulombs)
r = distance between particles (meters)

Using this equation, the electrostatic force between two protons that are separated by a distance of 10^{-20} meters is about 10^{12} newtons. Comparing this result with the calculation of the gravitational force (10^{-24} newtons) shows that the gravitational force is so small that it can be neglected.

If only the electrostatic and gravitational forces existed in the nucleus, then it would be impossible to have stable nuclei composed of protons and neutrons. The gravitational forces are much too small to hold the nucleons together compared to the electrostatic forces repelling the protons. Since stable atoms of neutrons and protons do exist, there must be another attractive force acting within the nucleus. This force is called the nuclear force.

The *nuclear force* is a strong attractive force that is independent of charge. It acts equally only between pairs of neutrons, pairs of protons, or a neutron and a proton. The nuclear force has a very short range; it acts only over distances approximately equal to the diameter of the nucleus (10^{-13} cm). The attractive nuclear force between all nucleons drops off with distance much faster than the repulsive electrostatic force between protons.

TABLE 3
Forces Acting in the Nucleus

Force	Interaction	Range
Gravitational	Very weak attractive force between all nucleons	Relatively long
Electrostatic	Strong repulsive force between like charged particles (protons)	Relatively long
Nuclear Force	Strong attractive force between all nucleons	Extremely short

In stable atoms, the attractive and repulsive forces in the nucleus balance. If the forces do not balance, the atom cannot be stable, and the nucleus will emit radiation in an attempt to achieve a more stable configuration.

Summary

The important information in this chapter is summarized on the following page.

Atomic Nature of Matter Summary

- Atoms consist of three basic subatomic particles. These particles are the proton, the neutron, and the electron.

- Protons are particles that have a positive charge, have about the same mass as a hydrogen atom, and exist in the nucleus of an atom.

- Neutrons are particles that have no electrical charge, have about the same mass as a hydrogen atom, and exist in the nucleus of an atom.

- Electrons are particles that have a negative charge, have a mass about eighteen hundred times smaller than the mass of a hydrogen atom, and exist in orbital shells around the nucleus of an atom.

- The Bohr model of the atom consists of a dense nucleus of protons and neutrons (nucleons) surrounded by electrons traveling in discrete orbits at fixed distances from the nucleus.

- Nuclides are atoms that contain a particular number of protons and neutrons.

- Isotopes are nuclides that have the same atomic number and are therefore the same element, but differ in the number of neutrons.

- The atomic number of an atom is the number of protons in the nucleus.

- The mass number of an atom is the total number of nucleons (protons and neutrons) in the nucleus.

- The notation ${}^{A}_{Z}X$ is used to identify a specific nuclide. "Z" represents the atomic number, which is equal to the number of protons. "A" represents the mass number, which is equal to the number of nucleons. "X" represents the chemical symbol of the element.

  Number of protons = Z
  Number of electrons = Z
  Number of neutrons = A - Z

- The stability of a nucleus is determined by the different forces interacting within it. The electrostatic force is a relatively long-range, strong, repulsive force that acts between the positively charged protons. The nuclear force is a relatively short-range attractive force between all nucleons. The gravitational force the long range, relatively weak attraction between masses, is negligible compared to the other forces.