> *"No new creation or destruction of matter is within the reach of chemical agency. We might as well attempt to introduce a new planet into the solar system, or to annihilate one already in existence, as to create or destroy a particle of hydrogen."*
>
> *John Dalton*

Subatomic Structure Review

The understanding that matter is composed of different elements evolved over thousands of years. Substances such as gold and silver were known in ancient times, but they were not understood to be elements. The alchemists, who did not understand elements or atoms, tried to change substances into gold because it was such a valuable substance. They never succeeded because, as we know now, matter is made up of atoms that cannot be changed by ordinary chemical methods. By the end of the nineteenth century chemists had formulated an atomic theory of matter which stated that the atom was the smallest particle of an element. This atomic theory also recognized that atoms were composed of smaller, subatomic particles. Today we recognize more than 110 different elements.

The Chemists' Collection of Subatomic Particles

Particle beam technology became available by the turn of the century, and it enabled scientists to discover the three subatomic particles listed below.

Particle	Mass (g)	Charge
Proton	1.672×10^{-24}	+1
Neutron	1.675×10^{-24}	0
Electron	9.109×10^{-28}	–1

The nature of the electron was determined by experiments on the interaction of beams of electrons (cathode rays) and electrical fields. As a result of Rutherford's alpha-particle experiments in which helium nuclei were projected through metal foils, a new picture of the atom emerged. The atom is viewed as mostly empty space in which the mass is concentrated in a dense, compact core of protons and neutrons known as the nucleus, and this nucleus is surrounded by a diffuse region containing electrons.

Symbols for the Elements

The periodic table of the elements shows the complete collection of known elements with each elemental name represented by a one- or two-letter symbol. These symbols are sometimes written with a superscript and/or a subscript preceding them. In this case, the symbols are called *nuclear symbols*, and they are used to distinguish among the isotopes of an element. All atoms of an element have the same number of protons, but the number of neutrons in an atom can vary. The superscript in a nuclear symbol represents the total number of particles in the nucleus: the sum of the number of protons and neutrons. The subscript indicates the number of protons. Since an atom is electrically neutral, the number of electrons can also be determined from the nuclear symbol because it is equal to the number of protons.

The nuclear symbol of any element is given by

$$^{A}_{Z}\mathrm{Sy}$$

where
A = mass number = number of protons + number of neutrons
Z = atomic number = number of protons
Sy = elemental symbol

For example, the symbol of carbon-12 is $^{12}_{6}\mathrm{C}$. The symbol tells us that an atom of carbon-12 has 6 protons (subscript) and 12 protons + neutrons (superscript). The difference, 12 – 6 = 6, is the number of neutrons. We also know that since the atom is electrically neutral, there are 6 electrons.

In ordinary chemical reactions, atoms never lose or gain protons or neutrons. However, they do lose or gain electrons. In fact, the chemical properties of an element are largely determined by the arrangement of the outer electrons of its atoms.

Mass Units for Elements

Chemists use a variety of units to express both the absolute and the relative mass of each of the elements. These units and the conventions governing their use are the subject of this section.

Atomic Mass Units (amu) Chemists have chosen carbon-12 as the standard for atomic masses. The carbon-12 atom is assigned a mass of exactly 12 atomic mass units. This means that

$$1 \text{ amu} = \frac{1}{12} \text{ the mass of one carbon-12 atom}$$

Also, since 6.02×10^{23} carbon-12 atoms have a mass of 12 grams,

$$1 \text{ amu} = \frac{1}{6.02 \infty 10^{23}} \text{ grams}$$

Isotopes of an element necessarily have different masses because they have different numbers of neutrons. For example, carbon-13, $^{13}_{6}\mathrm{C}$, which has 6 protons, 6 electrons, and 7 neutrons, has a mass of 13.00335 amu.

Carbon-12 and carbon-13 atoms are both present in any sample of carbon. The fractional abundance of carbon-12 is 0.9890, and that of carbon-13 is 0.0110. The fractional abundances for these isotopes must add up to 1. Fractional abundance can also be expressed as a percentage, and the numbers are 98.90% and 1.10% for carbon-12 and carbon-13, respectively. These must add to 100%.

Atomic Mass (Weight) The atomic mass (weight) of an element is defined as the average mass of the naturally-occurring isotopes of the elements, weighted to account for their percentage abundance. The unit of atomic mass is the amu.