Objective

After completing this section, you should be able to describe the basic structure of the atom.

Key Terms

Make certain that you can define, and use in context, the key terms below.

- atomic number
- atomic weight
- electron
- mass number
- neutron
- proton

The Nuclear Atom

The precise physical nature of atoms finally emerged from a series of elegant experiments carried out between 1895 and 1915. The most notable of these achievements was Ernest Rutherford's famous 1911 alpha-ray scattering experiment, which established that

- Almost all of the mass of an atom is contained within a tiny (and therefore extremely dense) nucleus which carries a positive electric charge whose value identifies each element and is known as the atomic number of the element.

- Almost all of the volume of an atom consists of empty space in which electrons, the fundamental carriers of negative electric charge, reside. The extremely small mass of the electron (1/1840th the mass of the hydrogen nucleus) causes it to behave as a quantum particle, which means that its location at any moment cannot be specified; the best we can do is describe its behavior in terms of the probability of its manifesting itself at any point in space. It is common (but somewhat misleading) to describe the volume of space in which the electrons of an atom have a significant probability of being found as the electron cloud. The latter has no definite outer boundary, so neither does the atom. The radius of an atom must be defined arbitrarily, such as the boundary in which the electron can be found with 95% probability. Atomic radii are typically 30-300 pm.
Protons and Neutrons

The nucleus is itself composed of two kinds of particles. Protons are the carriers of positive electric charge in the nucleus; the proton charge is exactly the same as the electron charge, but of opposite sign. This means that in any [electrically neutral] atom, the number of protons in the nucleus (often referred to as the nuclear charge) is balanced by the same number of electrons outside the nucleus.

Because the electrons of an atom are in contact with the outside world, it is possible for one or more electrons to be lost, or some new ones to be added. The resulting electrically-charged atom is called an ion.

The other nuclear particle is the neutron. As its name implies, this particle carries no electrical charge. Its mass is almost the same as that of the proton. Most nuclei contain roughly equal numbers of neutrons and protons, so we can say that these two particles together account for almost all the mass of the atom.

Atomic Number (Z)

What single parameter uniquely characterizes the atom of a given element? It is not the atom’s relative mass, as we will see in the section on isotopes below. It is, rather, the number of protons in the nucleus, which we call the atomic number and denote by the symbol Z. Each proton carries an electric charge of +1, so the atomic number also specifies the electric charge of the nucleus. In the neutral atom, the Z protons within the nucleus are balanced by Z electrons outside it.

Atomic numbers were first worked out in 1913 by Henry Moseley, a young member of Rutherford's research group in Manchester.

Moseley searched for a measurable property of each element that increases linearly with atomic number. He found this in a class of X-rays emitted by an element when it is bombarded with electrons. The frequencies of these X-rays are unique to each element, and they increase uniformly in successive elements. Moseley found that the square roots of these frequencies give a straight line when plotted against Z; this enabled him to sort the elements in order of increasing atomic number.
You can think of the atomic number as a kind of serial number of an element, commencing at 1 for hydrogen and increasing by one for each successive element. The chemical name of the element and its symbol are uniquely tied to the atomic number; thus the symbol "Sr" stands for strontium, whose atoms all have \( Z = 38 \).

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**Mass Number (A)**

The *mass number* equals the sum of the numbers of protons and the number of neutrons in the nucleus. It is sometimes represented by the symbol \( A \), so

\[ A = Z + N \]

in which \( Z \) is the atomic number and \( N \) is the *neutron number*.

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**Nuclides and their Symbols**

The term *nuclide* simply refers to any particular kind of nucleus. For example, a nucleus of atomic number 7 is a nuclide of nitrogen. Any nuclide is characterized by the pair of numbers \( (Z, A) \). The element symbol depends on \( Z \) alone, so the symbol \(^{26}\text{Mg}\) is used to specify the mass-26 nuclide of magnesium, whose name implies \( Z=12 \). A more explicit way of denoting a particular kind of nucleus is to add the atomic number as a subscript. Of course, this is somewhat redundant, since the symbol Mg always implies \( Z=12 \), but it is sometimes a convenience when discussing several nuclides.

Two nuclides having the same atomic number but different mass numbers are known as *isotopes*. Most elements occur in nature as mixtures of isotopes, but twenty-three of them (including beryllium and fluorine, shown in the table) are monoisotopic. For example, there are three natural isotopes of magnesium: \(^{24}\text{Mg} \) (79% of all Mg atoms), \(^{25}\text{Mg}\) (10%), and \(^{26}\text{Mg}\) (11%); all three are present in all compounds of magnesium in about these same proportions.

<table>
<thead>
<tr>
<th>Z</th>
<th>mass numbers</th>
</tr>
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<tbody>
<tr>
<td>H</td>
<td>1 2 3</td>
</tr>
<tr>
<td>He</td>
<td>2 3 4</td>
</tr>
<tr>
<td>Li</td>
<td>3 6 7</td>
</tr>
<tr>
<td>Be</td>
<td>4 9</td>
</tr>
<tr>
<td>B</td>
<td>5 10 11</td>
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<td>C</td>
<td>6 12 13 14</td>
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<td>N</td>
<td>7 14 15</td>
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<tr>
<td>O</td>
<td>8 16 17 18</td>
</tr>
<tr>
<td>F</td>
<td>9 19</td>
</tr>
<tr>
<td>Ne</td>
<td>10 20 21 22</td>
</tr>
</tbody>
</table>

Approximately 290 isotopes occur in nature. The two heavy isotopes of hydrogen are especially important— so much so that they have names and symbols of their own:
Deuterium accounts for only about 15 out of every one million atoms of hydrogen. Tritium, which is radioactive, is even less abundant. All the tritium on the earth is a by-product of the decay of other radioactive elements.

**Atomic Weights**

Atoms are of course far too small to be weighed directly; weight measurements can only be made on the massive (but unknown) numbers of atoms that are observed in chemical reactions. The early combining-weight experiments of Dalton and others established that hydrogen is the lightest of the atoms, but the crude nature of the measurements and uncertainties about the formulas of many compounds made it difficult to develop a reliable scale of the relative weights of atoms. Even the most exacting weight measurements we can make today are subject to experimental uncertainties that limit the precision to four significant figures at best.

**The Periodic Table**

The elements are arranged in a periodic table, which is probably the single most important learning aid in chemistry. It summarizes huge amounts of information about the elements in a way that facilitates the prediction of many of their properties and chemical reactions. The elements are arranged in seven horizontal rows, in order of increasing atomic number from left to right and top to bottom. The rows are called periods, and they are numbered from 1 to 7. The elements are stacked in such a way that elements with similar chemical properties form vertical columns, called groups, numbered from 1 to 18 (older periodic tables use a system based on roman numerals). Groups 1, 2, and 13–18 are the main group elements, listed as A in older tables. Groups 3–12 are in the middle of the periodic table and are the transition elements, listed as B in older tables. The two rows of 14 elements at the bottom of the periodic table are the lanthanides and the actinides, whose positions in the periodic table are indicated in group 3.

**Contributors**

- Dr. Dietmar Kennepohl FCIC (Professor of Chemistry, Athabasca University)
- Prof. Steven Farmer (Sonoma State University)