Manahan, Stanley E. "ATOMS AND ELEMENTS"
Fundamentals of Environmental Chemistry
Boca Raton: CRC Press LLC,2001

3 ATOMS AND ELEMENTS

3.1 ATOMS AND ELEMENTS

At the very beginning of this book, chemistry was defined as the science of matter. Chapter 2 examined the nature of matter, largely at a macroscopic level and primarily through its physical properties. Today, the learning of chemistry is simplified by taking advantage of what is known about matter at its most microscopic and fundamental level—atoms and molecules. Although atoms have been viewed so far in this book as simple, indivisible particles, they are in fact complicated bodies. Subtle differences in the arrangements and energies of electrons that make up most of the volume of atoms determine the chemical characteristics of the atoms. In particular, the electronic structures of atoms cause the periodic behavior of the elements, as described briefly in Section 1.3 and summarized in the periodic table in Figure 1.3. This chapter addresses atomic theory and atomic structure in more detail to provide a basis for the understanding of chemistry.

3.2 THE ATOMIC THEORY

The nature of atoms in relation to chemical behavior is summarized in the **atomic theory**. This theory in essentially its modern form was advanced in 1808 by John Dalton, an English schoolteacher, taking advantage of a substantial body of chemical knowledge and the contributions of others. It has done more than any other concept to place chemistry on a sound, systematic theoretical foundation. Those parts of Dalton's atomic theory that are consistent with current understanding of atoms are summarized in Figure 3.1.

Laws that are Explained by Dalton's Atomic Theory

The atomic theory outlined in Figure 3.1 explains the following three laws that are of fundamental importance in chemistry:

1. **Law of Conservation of Mass**: *There is no detectable change in mass in an ordinary chemical reaction.* (This law was first stated in 1798 by "the father of chemistry," the Frenchman Antoine Lavoisier. Since, as shown in Item 5 of Figure 3.1, no atoms are lost, gained, or changed in chemical reactions, mass is conserved.)

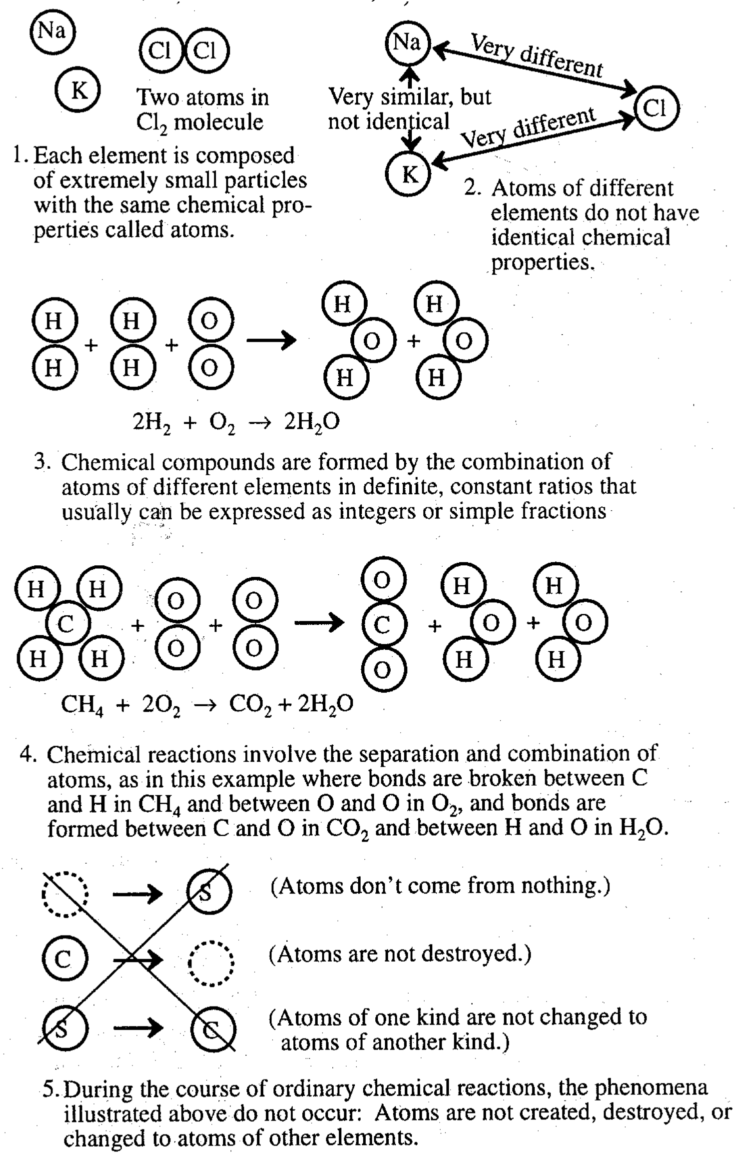


Figure 3.1 Illustration of Dalton's atomic theory.

2. **Law of Constant Composition**: *A specific chemical compound always contains the same elements in the same proportions by mass.* (If atoms always combine in definite, constant ratios to form a particular chemical compound, as implied in Item 3 of Figure 3.1, the elemental composition of the compound by mass always remains the same.)

3. **Law of Multiple Proportions**: *When two elements combine to form two or more compounds, the masses of one combining with a fixed mass of the other are in ratios of small whole numbers* (as illustrated for two compounds composed only of carbon and hydrogen below).

For CH_4: Relative mass of hydrogen = 4 × 1.0 = 4.0 (because there are 4 atoms of H, atomic mass 1.0)

Relative mass of carbon = 1 × 12.0 = 12.0 (because there is 1 atom of C, atomic mass 12.0)

$$\text{C/H ratio for } CH_4 = \frac{\text{Mass of C}}{\text{Mass of H}} = \frac{12.0}{4.0} = 3.0$$

For C_2H_6: Relative mass of hydrogen = 6 × 1.0 = 6.0 (because there are 6 atoms of H, atomic mass 1.0)

Relative mass of carbon = 2 × 12.0 = 24.0 (because there are 2 atoms of C, atomic mass 12.0)

$$\text{C/H ratio for } C_2H_6 = \frac{\text{Mass of C}}{\text{Mass of H}} = \frac{24.0}{6.0} = 4.0$$

Comparing C_2H_6 and CH_4: $$\frac{\text{C/H ratio for } C_2H_6}{\text{C/H ratio for } CH_4} = \frac{4.0}{3.0} = 4/3$$

Note that this is a ratio of small whole numbers.

Small Size of Atoms

It is difficult to imagine just how small an individual atom is. An especially small unit of mass, the *atomic mass unit, u*, is used to express the masses of atoms. This unit was defined in Section 1.3 as a mass equal to exactly 1/12 that of the carbon-12 isotope. An atomic mass unit is only 1.66×10^{-24} g. An average atom of hydrogen, the lightest element, has a mass of only 1.0079 u. The average mass of an atom of uranium, the heaviest naturally occurring element, is 238.03 u. To place these values in perspective, consider that a signature written by ballpoint pen on a piece of paper typically has a mass of 0.1 mg (1×10^{-4} g). This mass is equal to 6 x 10^{19} u. It would take almost 60 000 000 000 000 000 000 hydrogen atoms to equal the mass of ink in such a signature!

Atoms visualized as spheres have diameters of around 1-3 × 10^{-10} meters (100-300 picometers, pm). By way of comparison, a small marble has a diameter of around 1 cm (1×10^{-2} m), which is about 100 000 000 times that of a typical atom.

Atomic Mass

As defined in Section 1.3, the *atomic mass* of an element is the average mass of all atoms of the element relative to carbon-12 taken as exactly 12. Since atomic masses are *relative* quantities, they can be expressed without units. Or atomic masses may be given in atomic mass units. For example, an atomic mass of 14.0067 for nitrogen means that the average mass of all nitrogen atoms is 14.0067/12 as great as the mass of the carbon-12 isotope and is also 14.0067 u.

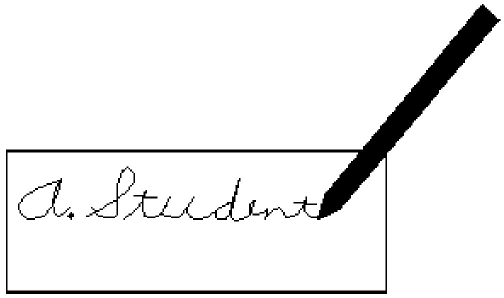


Figure 3.2 The ballpoint pen ink in a typical signature might have a mass of 6 x 10^{19} u.

3.3 SUBATOMIC PARTICLES

Small as atoms are, they in turn consist of even smaller entities called **subatomic particles**. Although physicists have found several dozen of these, chemists need consider only three—*protons*, *neutrons*, and *electrons* (these were introduced briefly in Section 1.3). These subatomic particles differ in mass and charge. Like the atom, their masses are expressed in atomic mass units.

The proton, *p*, has a mass of 1.007277 u and a unit charge of +1. This charge is equal to 1.6022×10^{-19} coulombs, where a coulomb is the amount of electrical charge involved in a flow of electrical current of 1 ampere for 1 second.

The neutron, *n*, has no electrical charge and a mass of 1.009665 u. The proton and neutron each have a mass of essentially 1 u and are said to have a *mass number* of 1. (Mass number is a useful concept expressing the total number of protons and neutrons, as well as the approximate mass, of a nucleus or subatomic particle.)

The electron, *e*, has an electrical charge of -1. It is very light, however, with a mass of only 0.00054859 u, about 1/1840 that of the proton or neutron. Its mass number is 0. The properties of protons, neutrons, and electrons are summarized in Table 3.1.

Table 3.1 Properties of Protons, Neutrons, and Electrons

Subatomic particle	Symbol[1]	Unit charge	Mass number	Mass in μ	Mass in grams
Proton	*p*	+1	1	1.007277	1.6726×10^{-24}
Neutron	*n*	0	1	1.008665	1.6749×10^{-24}
Electron	*e*	-1	0	0.000549	9.1096×10^{-28}

1 The mass number and charge of each of these kinds of particles can be indicated by a superscript and subscript, respectively in the symbols ${}^{1}_{1}p$, ${}^{1}_{0}n$, ${}^{0}_{-1}e$.

Although it is convenient to think of the proton and neutron as having the same mass, and each is assigned a mass number of 1, it is seen in Table 3.1 that their exact masses differ slightly from each other. Furthermore, the mass of an atom differs slightly from the sum of the masses of subatomic particles composing the atom. This is because of the energy relationships involved in holding the subatomic particles together in atoms so that the masses of the atom's constituent subatomic particles do not add up to exactly the mass of the atom.

3.4 THE BASIC STRUCTURE OF THE ATOM

As mentioned in Section 1.3, protons and neutrons are located in the *nucleus* of an atom; the remainder of the atom consists of a cloud of rapidly moving electrons. Since protons and neutrons have much higher masses than electrons, essentially all the mass of an atom is in its nucleus. However, the electron cloud makes up virtually all of the volume of the atom, and the nucleus is very small.

Atomic Number, Isotopes, and Mass Number of Isotopes

All atoms of the same element have the same number of protons and electrons equal to the *atomic number* of the element. (When reference is made to atoms here it is understood that they are electrically neutral atoms and not ions consisting of atoms that have lost or gained 1 or more electrons.) Thus, all helium atoms have 2 protons in their nuclei, and all nitrogen atoms have 7 protons.

In Section 1.3, *isotopes* were defined as atoms of the same element that differ in the number of neutrons in their nuclei. It was noted in Section 3.3 that both the proton and neutron have a *mass number* of exactly l. Mass number is commonly used to denote isotopes. The **mass number of an isotope** *is the sum of the number of protons and neutrons in the nucleus of the isotope*. Since atoms with the same number of protons—that is, atoms of the same element—may have different numbers of neutrons, several isotopes may exist of a particular element. The naturally occurring forms of some elements consist of only one isotope, but isotopes of these elements can be made artificially, usually by exposing the elements to neutrons produced by a nuclear reactor.

It is convenient to have a symbol that clearly designates an isotope of an element. Borrowing from nuclear science, a superscript in front of the symbol for the element is used to show the mass number and a subscript before the symbol designates the number of protons in the nucleus (atomic number). Using this notation denotes carbon-12 as ${}^{12}_{6}C$.

Exercise: Fill in the blanks designated with letters in the table below:

Element	Atomic number	Mass number of isotope	Number of neutrons in nucleus	Isotope symbol
Nitrogen	7	14	7	${}^{14}_{7}N$
Chlorine	17	35	(a)___________	${}^{35}_{17}N$
Chlorine	(b)__________	37	(c)___________	(d)___________
(e)_______	6	(f)__________	7	(g)___________
(h)_______	(i)___________	(j)__________	(k)___________	${}^{11}_{5}B$

Answers: (a) 18, (b) 17, (c) 20, (d) ${}^{37}_{17}Cl$, (e) carbon, (f) 13, (g) ${}^{13}_{6}C$, (h) boron, (i) 5, (j) 11, (k) 6

Electrons in Atoms

In Section 1.3, electrons around the nucleus of an atom were depicted as forming a cloud of negative charge. The energy levels, orientations in space, and behavior of electrons vary with the number of them contained in an atom. In a general sense, the arrangements of electrons in atoms are described by *electron configuration*, a term discussed in some detail later in this chapter.

Attraction between Electrons and the Nucleus

The electrons in an atom are held around the nucleus by the attraction between their negative charges and the positive charges of the protons in the nucleus. Opposite electrical charges attract, and like charges repel. The forces of attraction and repulsion are expressed quantitatively by **Coulomb's law**,

$$F = \frac{Q_1Q_2}{d} \tag{3.4.1}$$

where F is the force of interaction, Q_1 and Q_2 are the electrical charges on the bodies involved, and d is the distance between the bodies.

Coulomb's law explains the attraction between negatively charged electrons and the positively charged nucleus in an atom. However, the law does not explain why electrons move around the nucleus, rather than coming to rest on it. The behavior of electrons in atoms—as well as the numbers, types, and strengths of chemical bonds formed between atoms—is explained by *quantum theory*, a concept that is discussed in some detail in Sections 3.11-3.16.

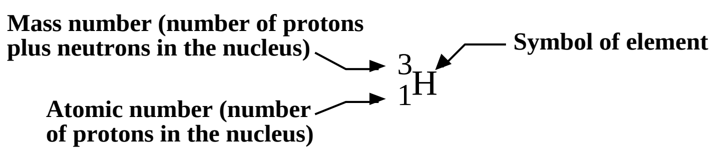
3.5 DEVELOPMENT OF THE PERIODIC TABLE

The *periodic table*, which was mentioned briefly in connection with the elements in Section 1.3, was first described by the Russian chemist Dmitri Mendeleev in 1867. It was based upon Mendeleev's observations of periodicity in chemical behavior without any knowledge of atomic structure. The periodic table is discussed in more detail in this chapter along with the development of the concepts of atomic structure. Elements are listed in the periodic table in an ordered, systematic way that correlates with their electron structures. The elements are placed in rows or **periods** of the periodic table in order of increasing atomic number such that there is a periodic repetition of elemental properties across the periods. The rows are arranged such that elements in vertical columns called **groups** have similar chemical properties reflecting similar arrangements of the outermost electrons in their atoms.

With some knowledge of the ways that electrons behave in atoms it is much easier to develop the concept of the periodic table. This is done for the first 20 elements in the following sections. After these elements are discussed, they can be placed in an abbreviated 20-element version of the periodic table. Such a table is shown in Figure 3.9.

3.6 HYDROGEN, THE SIMPLEST ATOM

The simplest atom is that of **hydrogen**, as it has only one positively charged proton in its nucleus, that is surrounded by a cloud of negative charge formed by only one electron. By far the most abundant kind of hydrogen atom has no neutrons in its nucleus. Having only one proton with a mass number (see Section 3.3) of 1, and 1 electron with a mass number of zero, the mass number of this form of hydrogen is 1 + 0 = 1. There are, however, two other forms of hydrogen atoms having, in addition to the proton, 1 and 2 neutrons, respectively, in their nuclei. The three different forms of elemental hydrogen are *isotopes* of hydrogen that all have the same number of protons, but different numbers of neutrons in their nuclei. Only about 1 of 7000 hydrogen atoms is ${}^{2}_{1}H$, deuterium, mass number 2 (from 1 proton + 1 neutron). The mass number of *tritium*, which has 2 neutrons, is 1 (from 1 proton + 2 neutrons) = 3. These three forms of hydrogen atoms can be designated as ${}^{1}_{1}H$, ${}^{2}_{1}H$, and ${}^{3}_{1}H$. The meaning of this notation is reviewed below:



Designation of Hydrogen in the Periodic Table

The atomic mass of hydrogen is 1.0079. This means that the average atom of hydrogen has a mass of 1.0079 u (atomic mass units); hydrogen's atomic mass is simply 1.0079 relative to the carbon-12 isotope taken as exactly 12. With this information it is possible to place hydrogen in the periodic table with the designation shown in Figure 3.3.

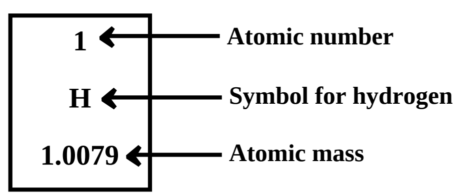


Figure 3.3 Designation of hydrogen in the periodic table.

Showing Electrons in Hydrogen Atoms and Molecules

It is useful to have some simple way of showing the hydrogen atom's single electron in chemical symbols and formulas. This is accomplished with **electron-dot symbols** or **Lewis symbols** (after G. N. Lewis), which use dots around the symbol of an element to show outer electrons (those that may become involved in chemical bonds). The Lewis symbol for hydrogen is

$$\mathbf{H\cdot}$$

As mentioned in Section 1.3, elemental hydrogen consists of molecules made up of 2 H atoms held together by a chemical bond consisting of two shared electrons. Just as it is useful to show electrons in atoms with a Lewis symbol, it is helpful to visualize molecules and the electrons in them with **electron-dot formulas** or **Lewis formulas** as shown for the H_2 molecule in Figure 3.4.

$H\cdot \rightarrow \leftarrow \cdot H$	H_2	H:H
Elemental hydrogen does not exist as individual H atoms.	**Instead, it exists as molecules made up of 2H atoms with the chemical formula H_2.**	**The covalent bond holding the two H atoms together consists of 2 shared electrons shown in the Lewis formula of H_2 above.**

Figure 3.4. Lewis formula for hydrogen molecules.

Properties of Elemental Hydrogen

Pure hydrogen is colorless and odorless. It is a gas at all but very low temperatures; liquid hydrogen boils at -253˚C and solidifies at -259˚C. Hydrogen gas has the lowest density of any pure substance. It reacts chemically with a large number of elements, and hydrogen-filled balloons that float readily in the atmosphere also explode convincingly when touched with a flame because of the following very rapid reaction with oxygen in the air:

$$2H_2 + O_2 \rightarrow 2H_2O \qquad (3.5.1)$$

Production and Uses of Elemental Hydrogen

Elemental hydrogen is one of the more widely produced industrial chemicals. For use in chemical synthesis and other industrial applications, it is commonly made by **steam reforming** of methane (natural gas, CH_4) under high-temperature, high-pressure conditions:

$$CH_4 + H_2O \xrightarrow[\text{30 atm P}]{\text{800˚C T}} CO + 3H_2 \qquad (3.5.2)$$

Hydrogen is used to manufacture a number of chemicals. One of the most important of these is ammonia, NH_3. Methanol (methyl alcohol, CH_3OH) is a widely used industrial chemical and solvent synthesized by the following reaction between carbon monoxide and hydrogen:

$$CO + 2H_2 \rightarrow CH_3OH \qquad (3.5.3)$$

Methanol made by this process can be blended with gasoline to yield a fuel that produces relatively less pollutant carbon monoxide; such “oxygenated gasoline additives” are now required for use in some cities urban areas. Gasoline is upgraded by the chemical addition of hydrogen to some petroleum fractions. Synthetic petroleum can be made by the addition of hydrogen to coal at high temperatures and pressures. A widely used process in the food industry is the addition of hydrogen to unsaturated vegetable oils in the synthesis of margarine and other hydrogenated fats and oils.

3.7 HELIUM, THE FIRST ATOM WITH A FILLED ELECTRON SHELL

In Section 3.4 it was mentioned that the *electron configurations* of atoms determine their chemical behavior. Electrons in atoms occupy distinct *energy levels*. At this point, it is useful to introduce the concept of the **electron shell** to help explain electron energy levels and their influence on chemical behavior. Each electron shell can hold a maximum number of electrons. An atom with a **filled electron shell** is especially content in a chemical sense, with little or no tendency to lose, gain, or share electrons. Elements with these characteristics exist as gas-phase atoms and are called *noble gases*. The high stability of the noble gas electron configuration is explained in more detail in Section 3.9.

Examined in order of increasing atomic number, the first element consisting of atoms with filled electron shells is helium, He, atomic number 2. All helium atoms contain 2 protons and 2 electrons. Virtually all helium atoms are ${}^{4}_{2}He$ that contain 2 neutrons in their nuclei; the ${}^{3}_{2}He$ isotope containing only 1 neutron in its nucleus occurs to a very small extent. The atomic mass of helium is 4.00260.

The two electrons in the helium atom are shown by the Lewis symbol illustrated in Figure 3.5. These electrons constitute a *filled electron shell*, so that helium is a *noble gas* composed of individual helium atoms that have no tendency to form chemical bonds with other atoms. Helium gas has a very low density of only 0.164 g/L at 25˚C and 1 atm pressure.

A helium atom has a filled electron shell containing 2 electrons.

It can be represented by the Lewis symbol above.

Figure 3.5. Two representations of the helium atom having a filled electron shell

Occurrence and Uses of Helium

Helium is extracted from some natural gas sources that contain up to 10% helium by volume. It has many uses that depend upon its unique properties. Because of its very low density compared to air, helium is used to fill weather balloons and airships. Helium is non-toxic, odorless, tasteless, and colorless. Because of these properties and its low solubility in blood, helium is mixed with oxygen for breathing by deep sea divers and persons with some respiratory ailments. Use of helium by divers avoids the very painful condition called “the bends” caused by bubbles of nitrogen forming from nitrogen gas dissolved in blood.

Liquid helium, which boils at a temperature of only 4.2 K above absolute zero is especially useful in the growing science of **cryogenics**, which deals with very low temperatures. Some metals are superconductors at such temperatures so that helium is used to cool electromagnets that develop very powerful magnetic fields for a relatively small magnet. Such magnets are components of the very useful chemical tool known as nuclear magnetic resonance (NMR). The same kind of instrument modified for clinical applications and called MRI is used as a medical diagnostic tool.

3.8 LITHIUM, THE FIRST ATOM WITH BOTH INNER AND OUTER ELECTRONS

The third element in the periodic table is lithium (Li), atomic number 3, atomic mass 6.941. The most abundant lithium isotope has 4 neutrons in its nucleus, so it has a mass number of 7 and is designated ${}^{7}_{3}Li$. A less common isotope, ${}^{6}_{3}Li$ has only 3 neutrons.

Lithium is the first element in the periodic table that is a *metal*. Mentioned in Section 2.2, metals tend to have the following properties:

- Characteristic **luster** (like freshly-polished silverware or a new penny)
- **Malleable** (can be pounded or pressed into various shapes without breaking)
- **Conduct electricity**
- Chemically, tend to **lose electrons** and form cations (see Section 1.4) with charges of +1 to +3.

Lithium is the lightest metal with a density of only 0.531 g/cm^3.

Uses of Lithium

Lithium compounds have a number of important uses in industry and medicine. Lithium carbonate, Li_2CO_3, is one of the most important lithium compounds and is used as the starting material for the manufacture of many other lithium compounds. It is an ingredient of specialty glasses, enamels, and specialty ceramic ware having low thermal expansion coefficients (minimum expansion when heated). Lithium carbonate is widely prescribed as a drug to treat acute mania in manic-depressive and schizo-affective mental disorders. Lithium hydroxide, LiOH, is an ingredient in the manufacture of lubricant greases and in some long-life alkaline storage batteries.

The lithium atom has three electrons. As shown in Figure 3.6, lithium has both **inner electrons**—in this case 2 contained in an **inner shell**—as in the immediately preceding noble gas helium, and an **outer electron** that is farther from, and less strongly attracted to, the nucleus. The outer electron is said to be in the atom's **outer shell**. The inner electrons are, on the average, closer to the nucleus than is the outer electron, are very difficult to remove from the atom, and do not become involved in chemical bonds. Lithium's outer electron is relatively easy to remove from the atom, which is what happens when ionic bonds involving Li^+ ion are formed (see Section

1.4). The distinction between inner and outer electrons is developed to a greater extent later in this chapter.

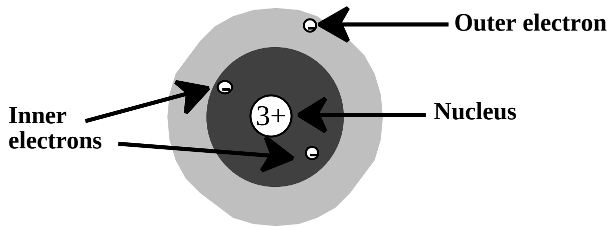
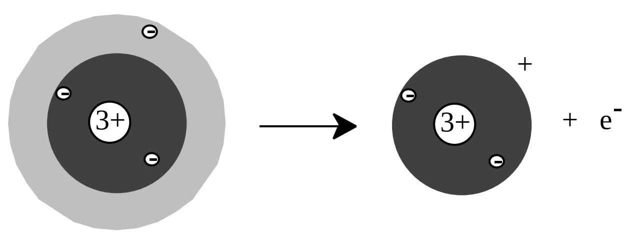


Figure 3.6 An atom of lithium, Li, has 2 inner electrons and 1 outer electron. The latter can be lost to another atom to produce the Li^+ ion, which is present in ionic compounds (see Section 1.4).

In atoms such as lithium that have both outer and inner electrons, the Lewis symbol shows only the outer electrons. Therefore, the Lewis symbol of lithium is

$$\mathrm{Li}\cdot$$

A lithium atom's loss of its single outer electron is shown by the **half-reaction** (one in which there is a net number of electrons on either the reactant or product side) in Figure 3.7. The Li^+ product of this reaction has the very stable helium core of 2 electrons. The Li^+ ion is a constituent of ionic lithium compounds in which it is held by attraction for negatively charged anions (such as Cl^-) in the crystalline lattice of the ionic compound. The tendency to lose its outer electron and to be stabilized in ionic compounds determines lithium's chemical behavior.



$$\mathrm{Li}\cdot \longrightarrow Li^+ + e^-$$

A neutral lithium atom loses an electron **to form a positively charged lithium ion** **and an electron that is acquired by an atom of some other element.**

Figure 3.7. Half-reaction showing the formation of Li^+ from an Li atom. The Li^+ ion has the especially stable helium core of just 2 electrons. The atom to which the electron is lost is not shown, so this is a half-reaction.

3.9 THE SECOND PERIOD, ELEMENTS 4-10

In this section, elements 4-10 will be discussed and placed in the periodic table to complete a period in the table.

Beryllium, Atomic Number 4

Each atom of **beryllium**—atomic number 4, atomic mass 9.01218—contains 4 protons and 5 neutrons in its nucleus. The beryllium atom has two inner electrons and two outer electrons, the latter designated by the two dots in the Lewis symbol below:

$$\mathrm{Be}:$$

Beryllium can react chemically by losing 2 electrons from the beryllium atom. This occurs according to the half reaction

$$\mathrm{Be}: \rightarrow \mathrm{Be}^{2+} + 2\mathrm{e}^{-} \text{ (lost to another atom)} \tag{3.9.1}$$

in which the beryllium atom, Lewis symbol Be:, loses two e^- to form a beryllium ion with a charge of +2. The loss of these two outer electrons gives the beryllium atom the same stable helium core as that of the Li^+ ion discussed in the preceding section.

Beryllium is melted together with certain other metals to give homogeneous mixtures of metals called **alloys**. The most important beryllium alloys are hard, corrosion-resistant, non-sparking, and good conductors of electricity. They are used to make such things as springs, switches, and small electrical contacts. A very high melting temperature of about 1290˚C combined with good heat absorption and conduction properties has led to the use of beryllium metal in aircraft brake components.

Beryllium is an environmentally and toxicologically important element because it causes **berylliosis**, a disease marked by lung deterioration. Inhalation of Be is particularly hazardous, and atmospheric standards have been set at very low levels.

Boron, Atomic Number 5

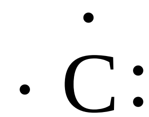
Boron, B, has an atomic number of 5 and an atomic mass of 10.81. Most boron atoms have 6 neutrons in addition to 5 protons in their nuclei; a less common isotope has 5 protons. Two of boron's 5 electrons are in a helium core and 3 are outer electrons as shown by the Lewis symbol

$$\dot{\mathrm{B}}:$$

Boron—along with silicon, germanium, arsenic, antimony, and tellurium—is one of a few elements, called **metalloids**, with properties intermediate between those of metals and nonmetals. Although they have a luster like metals, metalloids do not form positively charged ions (cations). The melting temperature of boron is very high, 2190˚C. Boron is added to copper, aluminum, and steel to improve their properties. It is used in control rods of nuclear reactors because of the good neutronabsorbing properties of the $^{10}_{5}B$ isotope. Some chemical compounds of boron, especially boron nitride, BN, are noted for their hardness. Boric acid, H_3BO_3, is used as a flame retardant in cellulose insulation in houses. The oxide of boron, B_2O_3, is an ingredient of fiberglass, used in textiles and insulation.

Carbon, Atomic Number 6

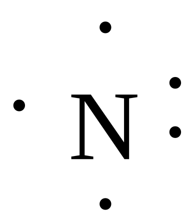
Atoms of **carbon**, C, have 2 inner and 4 outer electrons, the latter shown by the Lewis symbol



The carbon-12 isotope with 6 protons and 6 neutrons in its nucleus, $^{12}_{6}C$ constitutes 98.9% of all naturally occurring carbon. The $^{13}_{6}C$ isotope makes up 1.1% of all carbon atoms. As discussed in Chapter 25, radioactive carbon-14, $^{14}_{6}C$, is present in some carbon sources.

Carbon is an extremely important element with unique chemical properties without which life could not exist. All of organic chemistry (Chapter 10) is based upon compounds of carbon, and it is an essential element in life molecules (studied as part of biochemistry, Chapter 11). Carbon atoms are able to bond to each other to form long straight chains, branched chains, rings, and three-dimensional structures. As a result of its self-bonding abilities, carbon exists in several elemental forms. These include powdery carbon black; very hard, clear diamonds; and graphite so soft that it is used as a lubricant. Activated carbon prepared by treating carbon with air, carbon dioxide, or steam at high temperatures is widely used to absorb undesirable pollutant substances from air and water. Carbon fiber has been developed as a structural material in the form of composites consisting of strong strands of carbon bonded together with special plastics and epoxy resins.

Nitrogen, N, composes 78% by volume of air in the form of diatomic N_2 molecules. The atomic mass of nitrogen is 14.0067, and the nuclei of N atoms contain 7 protons and 7 neutrons. Nitrogen has 5 outer electrons, so its Lewis symbol is

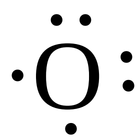


Like carbon, nitrogen is a nonmetal. Pure N_2 is prepared by distilling liquified air, and it has a number of uses. Since nitrogen gas is not very chemically reactive, it is used as an inert atmosphere in some industrial applications, particularly where fire or chemical reactivity may be a hazard. People have been killed by accidentally entering chambers filled with nitrogen gas, which acts as a simple asphyxiant with no odor to warn of its presence. Liquid nitrogen boils at a very cold -190˚C. It is widely used to maintain very low temperatures in the laboratory, for quick-freezing foods, and in freeze-drying processes. Freeze-drying is used to isolate fragile biochemical compounds from water solution, for the concentration of environmental samples to be analyzed for pollutants, and for the preparation of instant coffee and other dehydrated foods. It has potential applications in the concentration and isolation of hazardous waste substances.

Like carbon, nitrogen is an essential element for life processes. Nitrogen is an ingredient of all of the amino acids found in proteins. Nitrogen compounds are fertilizers essential for the growth of plants. The **nitrogen cycle**, which involves incorporation of N_2 from the atmosphere into living matter and chemically bound nitrogen in soil and water, then back into the atmosphere again, is one of nature's fundamental cycles. Nitrogen compounds, particularly ammonia (NH_3), and nitric acid (HNO_3), are widely used industrial chemicals.

Oxygen, Atomic Number 8

Like carbon and nitrogen, **oxygen**, atomic number 8, is a major component of living organisms. Oxygen is a nonmetal existing as molecules of O_2 in the elemental gas state, and air is 21% oxygen by volume. Like all animals, humans require oxygen to breathe and to maintain their life processes. The nuclei of oxygen atoms contain 8 protons and 8 neutrons, and the atomic mass of oxygen is 15.9994. The oxygen atom has 6 outer electrons, as shown by its Lewis symbol below:



In addition to O_2, there are two other important elemental oxygen species in the atmosphere. These are atomic oxygen, O, and ozone, O_3. These species are normal constituents of the stratosphere, a region of the atmosphere that extends from about 11 kilometers to about 50 km in altitude. Oxygen atoms are formed when high-energy ultraviolet radiation strikes oxygen molecules high in the stratosphere:

$$O_2 \xrightarrow[\text{radiation}]{\text{Ultraviolet}} O + O \qquad (3.9.3)$$

The oxygen atoms formed by the above reaction combine with O_2 molecules,

$$O + O_2 \rightarrow O_3 \qquad (3.9.4)$$

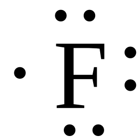
to form ozone molecules. These molecules make up the **ozone layer** in the stratosphere and effectively absorb additional high-energy ultraviolet radiation. If it weren't for this phenomenon, the ultraviolet radiation would reach the Earth's surface and cause painful sunburn and skin cancer in exposed people. However, ozone produced in photochemical smog at ground level is toxic to animals and plants.

The most notable chemical characteristic of elemental oxygen is its tendency to combine with other elements in energy-yielding reactions. Such reactions provide the energy that propels automobiles, heats buildings, and keeps body processes going. One of the most widely used chemical reactions of oxygen is that with hydrocarbons, particularly those from petroleum and natural gas. For example, butane (C_4H_{10}, a liquifiable gaseous hydrocarbon fuel) burns in oxygen from the atmosphere, a reaction that provides heat in home furnaces, water heaters, and other applications:

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O \tag{3.9.5}$$

Fluorine, Atomic Number 9

Fluorine, F, has 7 outer electrons, so its Lewis symbol is



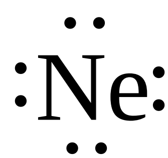
Under ordinary conditions, elemental fluorine is a greenish-yellow gas consisting of F_2 molecules.

Fluorine compounds have many uses. One of the most notable of these is the manufacture of chlorofluorocarbon compounds known by the trade name Freon. These are chemical combinations of chlorine, fluorine, and carbon, an example of which is dichlorodifluoromethane, Cl_2CF_2. These compounds used to be widely employed as refrigerant fluids and blowing agents to make foam plastics; they were once widely used as propellants in aerosol spray cans. Uses of chlorofluorocarbons have now been largely phased out because of their role in destroying stratospheric ozone (discussed with oxygen, above).

Neon, Atomic Number 10

The last element in the period of the periodic table under discussion is **neon**. Air is about 2 parts per thousand neon by volume, and neon is obtained by the distillation of liquid air. Neon is especially noted for its use in illuminated signs that consist of glass tubes containing neon, through which an electrical current is passed, causing the neon gas to emit a characteristic glow.

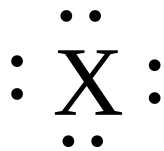
In addition to 10 protons, most neon atoms have 10 neutrons in their nuclei, although some have 12, and a very small percentage have 11. As shown by its Lewis symbol,



the neon atom has 8 outer electrons, which constitute a *filled electron shell*, just as the 2 electrons in helium give it a filled electron shell. Because of this “satisfied” outer shell, the neon atom has no tendency to acquire, give away, or share electrons. Therefore, neon is a *noble gas*, like helium, and consists of individual neon atoms.

Stability of the Neon Noble Gas Electron Octet

In going through the rest of the periodic table it can be seen that all other atoms with 8 outer electrons, like neon, are also noted for a high degree of chemical stability. In addition to neon, these noble gases are argon (atomic number 18), krypton (atomic number 36), xenon (atomic number 54), and radon (atomic number 86). Each of these may be represented by the Lewis symbol



where X is the chemical symbol of the noble gas. It is seen that these atoms each have 8 outer electrons, a group known as an **octet** of electrons. In many cases, atoms that do not have an octet of outer electrons acquire one by losing, gaining, or sharing electrons in chemical combination with other atoms; that is, they acquire a **noble gas outer electron configuration**. For all noble gases except helium, which has only 2 electrons, the noble gas outer electron configuration consists of eight electrons. The tendency of elements to acquire an 8-electron outer electron configuration, which is very useful in predicting the nature of chemical bonding and the formulas of compounds that result, is called the **octet rule**. Although the use of the octet rule to explain and predict bonding is discussed in some detail in Chapter 4, at this point it is useful to show how it explains bonding between hydrogen and carbon in methane, as illustrated in Figure 3.8.

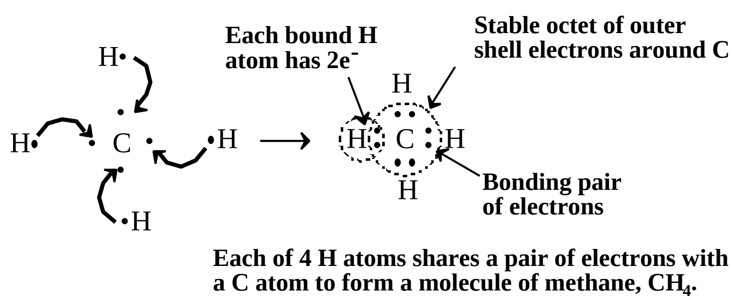


Figure 3.8 Illustration of the octet rule in methane.

3.10 ELEMENTS 11-20, AND BEYOND

The abbreviated version of the periodic table will be finished with elements 11 through 20. The names, symbols, electron configurations, and other pertinent information about these elements are given in Table 3.2. An abbreviated periodic table with these elements in place is shown in Figure 3.9. This table shows the Lewis symbols of all of the elements to emphasize their orderly variation across periods and similarity in groups of the periodic table.

The first 20 elements in the periodic table are very important. They include the three most abundant elements on the earth's surface (oxygen, silicon, aluminum); all

Table 3.2 Elements 11-20

Atomic number	Name and Lewis symbol	Atomic mass	Number of outer e^-	Major properties and uses
11	Sodium, Na·	22.9898	1	Soft, chemically very reactive metal. Nuclei contain 11 *p* and 12 *n*.
12	Magnesium, Mg:	24.312	2	Lightweight metal used in aircraft components, extension ladders, portable tools. Chemically very reactive. Three isotopes with 12, 13, 14 *n*.
13	Aluminum, ·Al:	26.9815	3	Lightweight metal used in aircraft, automobiles, electrical transmission line. Chemically reactive, but forms self-protective coating.
14	Silicon, ·Si:	28.086	4	Nonmetal, 2nd most abundant metal in Earth's crust. Rock constituent. Used in semiconductors.
15	Phosphorus, ·P:	30.9738	5	Chemically very reactive nonmetal. Highly toxic as elemental white phosphorus. Component of bones and teeth, genetic material (DNA), fertilizers, insecticides.
16	Sulfur, ·S:	32.064	6	Brittle, generally yellow nonmetal. Essential nutrient for plants and animals, occurring in amino acids. Used to manufacture sulfuric acid. Present in pollutant sulfur dioxide, SO_2.
17	Chlorine, ·Cl:	35.453	7	Greenish-yellow toxic gas composed of molecules of Cl_2. Manufactured in large quantities to disinfect water and to manufacture plastics and solvents.
18	Argon, :Ar:	39.948	8	Noble gas used to fill light bulbs and as a plasma medium in inductively coupled plasma atomic emission analysis of elemental pollutants.
19	Potassium, K·	39.098	1	Chemically reactive alkali metal very similar to sodium in chemical and physical properties. Essential fertilizer for plant growth as K^+ ion.
20	Calcium, Ca:	40.078	2	Chemically reactive alkaline earth metal with properties similar to those of magnesium.

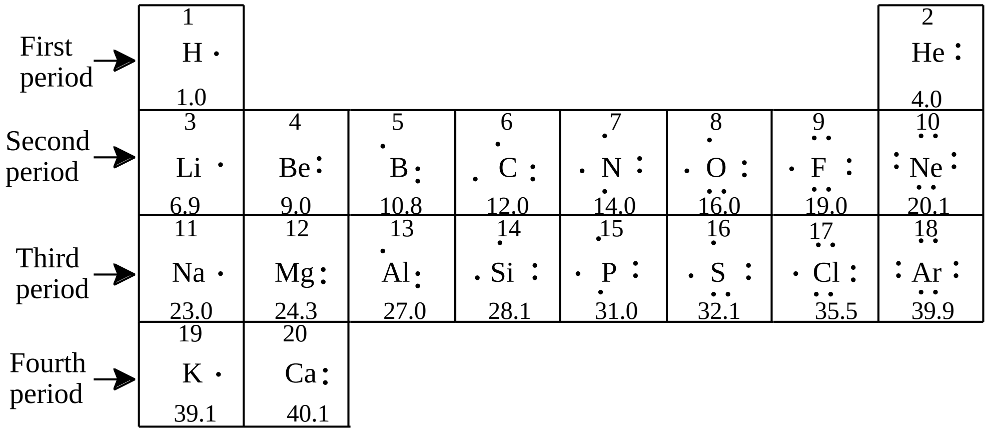


Figure 3.9 Abbreviated 20-element version of the periodic table showing Lewis symbols of the elements

elements of any appreciable significance in the atmosphere (hydrogen in H_2O vapor, N_2, O_2, carbon in CO_2, argon, and neon); the elements making up most of living plant and animal matter (hydrogen, oxygen, carbon, nitrogen, phosphorus, and sulfur); and elements such as sodium, magnesium, potassium, calcium, and chlorine that are essential for life processes. The chemistry of these elements is relatively straightforward and easy to relate to their atomic structures. Therefore, emphasis is placed on them in the earlier chapters of this book. It is helpful to remember their names, symbols, atomic numbers, atomic masses, and Lewis symbols.

As mentioned in Section 1.3, the vertical columns of the table contain *groups* of elements that have similar chemical structures. Hydrogen, H, is an exception and is not regarded as belonging to any particular group because of its unique chemical properties. All elements other than hydrogen in the first column of the abbreviated table are **alkali metals**—lithium, sodium, and potassium. These are generally soft silvery-white metals of low density that react violently with water to produce hydroxides (LiOH, NaOH, KOH) and with chlorine to produce chlorides (LiCl, NaCl, KCl). The **alkaline earth** metals—beryllium, magnesium, calcium—are in the second column of the table. When freshly cut, these metals have a grayish-white luster. They are chemically reactive and have a strong tendency to form doubly charged cations (Be^{2+}, Mg^{2+}, Ca^{2+}) by losing two electrons from each atom. Another group notable for the very close similarities of the elements in it consists of the **noble gases** in the far right column of the table. Each of these—helium, neon, argon—is a monatomic gas that does not react chemically.

The Elements beyond Calcium

The electron structures of elements beyond atomic number 20 are more complicated than those of the lighter elements. The complete periodic table in Figure 1.3 shows, among the heavier elements, the transition metals, including chromium, manganese, iron, cobalt, nickel, and copper; the lanthanides; and the actinides, including thorium, uranium, and plutonium. The transition metals include a number of metals that are important in industry and in life processes. The actinides contain elements familiar to those concerned with nuclear energy, nuclear warfare, and related issues. A list of the known elements through atomic number 109 is given on page 120 at the end of this chapter.

3.11 A MORE DETAILED LOOK AT ATOMIC STRUCTURE

So far, this chapter has covered some important aspects of atoms. These include the facts that an atom is made of three major subatomic particles, and consists of a very small, very dense, positively charged nucleus surrounded by a cloud of negatively charged electrons in constant, rapid motion. The first 20 elements have been discussed in some detail and placed in an abbreviated version of the periodic table. Important concepts introduced so far in this chapter include:

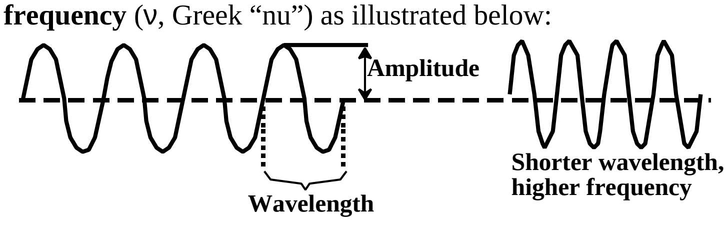
- Dalton's atomic theory
- Electron shells
- Inner shell electrons
- Octet rule
- Lewis symbols to represent outer e^-.
- Significance of filled electron shells
- Outer shell electrons
- Abbreviated periodic table

The information presented about atoms so far in this chapter is adequate to meet the needs of many readers. These readers may choose to forgo the details of atomic structure presented in the rest of this chapter without major harm to their understanding of chemistry. However, for those who wish to go into more detail, or who do not have a choice, the remainder of this chapter discusses in more detail the electronic structures of atoms as related to their chemical behavior and introduces the quantum theory of electrons in atoms.

Electromagnetic Radiation

The **quantum theory** explains the unique behavior of charged particles that are as small and move as rapidly as electrons. Because of its close relationship to electromagnetic radiation, an appreciation of quantum theory requires an understanding of the following important points related to electromagnetic radiation:

- Energy can be carried through space at the speed of light, 3.00×10^9 meters per second (m/s) in a vacuum, by **electromagnetic radiation**, which includes visible light, ultraviolet radiation, infrared radiation, microwaves, and radio waves.
- Electromagnetic radiation has a **wave character**. The waves move at the speed of light, c, and have characteristics of **wavelength** (λ), amplitude, and **frequency** (ν, Greek "nu") as illustrated below:



- The wavelength is the distance required for one complete cycle and the frequency is the number of cycles per unit time. They are related by the following equation:

$$\nu\lambda = c$$

  where ν is in units of cycles per second (s^{-1}, a unit called the **hertz**, Hz) and λ is in meters (m).

- In addition to behaving as a wave, electromagnetic radiation also has characteristics of particles.
- The dual wave/particle nature of electromagnetic radiation is the basis of the **quantum theory** of electromagnetic radiation, which states that radiant energy can be absorbed or emitted only in discrete packets called **quanta** or **photons**. The energy, E, of each photon is given by

$$E = h\nu$$

  where h is Planck's constant, 6.63×10^{-34} J-s (joule x second).

- From the preceding, it is seen that *the energy of a photon is higher when the frequency of the associated wave is higher* (and the wavelength shorter).

3.12 QUANTUM AND WAVE MECHANICAL MODELS OF ELECTRONS IN ATOMS

The *quantum theory* introduced in the preceding section provided the key concepts needed to explain the energies and behavior of electrons in atoms. One of the best clues to this behavior, and one that ties the nature of electrons in atoms to the properties of electromagnetic radiation, is the emission of light by energized atoms. This is easiest to explain for the simplest atom of all, that of hydrogen, which consists of only one electron moving around a nucleus with a single positive charge. Energy added to hydrogen atoms, such as by an electrical discharge through hydrogen gas, is re-emitted in the form of light at very specific wavelengths (656, 486, 434, 410 nm in the visible region). The highly energized atoms that can emit this light are said to be "excited" by the excess energy originally put into them and to be in an **excited state**. The reason for this is that the electrons in the excited atoms are forced farther from the nuclei of the atoms and, when they return to a lower energy state, energy is emitted in the form of light. The fact that very specific wavelengths of light are emitted in this process means that electrons can be present only in specified states at highly specific energy levels. Therefore, the transition from one energy state to a lower one involves the emission of a specific energy of electromagnetic radiation (light). Consider the equation

$$E = h\nu \qquad (3.12.1)$$

that relates energy to frequency, ν, of electromagnetic radiation. If a transition of an electron from one excited state to a lower one involves a specific amount of energy, E, a corresponding value of ν is observed. This is reflected by a specific wavelength of light according to the following relationship:

$$\lambda = \frac{c}{\nu} \qquad (3.12.2)$$

The first accepted explanation of the behavior outlined above was the **Bohr theory** advanced by the Danish physicist Neils Bohr in 1913. Although this theory has been shown to be too simplistic, it had some features that are still pertinent to atomic structure. The Bohr theory visualized an electron orbiting the nucleus (a proton) of the hydrogen atom in orbits. Only specific orbits called **quantum states** were allowed. When energy was added to a hydrogen atom, its electron could jump to a higher orbit. When the atom lost its energy as the electron returned to a lower orbit, the energy lost was emitted in the form of electromagnetic radiation as shown in Figure 3.10. Because the two energy levels are of a definite magnitude according to quantum theory, the energy lost by the electron must also be of a definite energy, E = $h\nu$. Therefore, the electromagnetic radiation (light) emitted is of a specific frequency and wavelength.

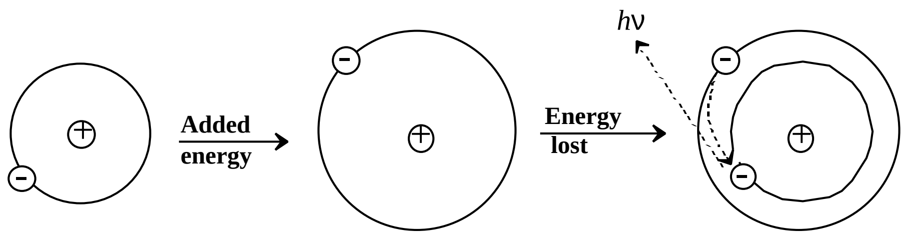


Figure 3.10 According to the Bohr model, adding energy to the hydrogen atom promotes an electron to a higher energy level. When the electron falls back to a lower energy level, excess energy is emitted in the form of electromagnetic radiation of a specific energy, E = $h\nu$.

The Wave Mechanical Model of Atomic Structure

Though shown to have some serious flaws and long since abandoned, the Bohr model laid the groundwork for the more sophisticated theories of atomic structure that are accepted today and introduced the all-important concept that *only specific energy states are allowed for an electron in an atom.* Like electromagetic radiation, electrons in atoms are now visualized as having a dual wave/particle nature. They are treated theoretically by the **wave mechanical model** as **standing waves** around the nucleus of an atom. The idea of a standing wave can be visualized for the string of a musical instrument as represented in Figure 3.11. Such a wave does not move along the length of a string because both ends are anchored, which is why it is called a *standing wave.* Each wave has **nodes**, which are points of zero displacement. Because there must be a node on each end where the string is anchored, the standing waves can only exist as multiples of *half-wavelengths*.

According to the *wave mechanical* or *quantum mechanical* model of electrons in atoms, the known *quantization* of electron energy in atoms occurs because only specific multiples of the standing wave associated with an electron's movement are allowed. Such a phenomenon is treated mathematically with the **Schrödinger equa-**

tion, resulting from the work of Erwin Schrödinger, first published in 1926. Even for the one-electron hydrogen atom, the mathematics is quite complicated, and no attempt will be made to go into it here. The Schrödinger equation is represented as

$$H\psi = E\psi \qquad (3.12.3)$$

where ψ is the Greek psi. Specifically, ψ is the **wave function**, a function of the electron's energy and the coordinates in space where it may be found. The term H is an **operator** consisting of a set of mathematical instructions, and E is the sum of the kinetic energy due to the motion of the electron and the potential energy of the mutual attraction between the electron and the nucleus.

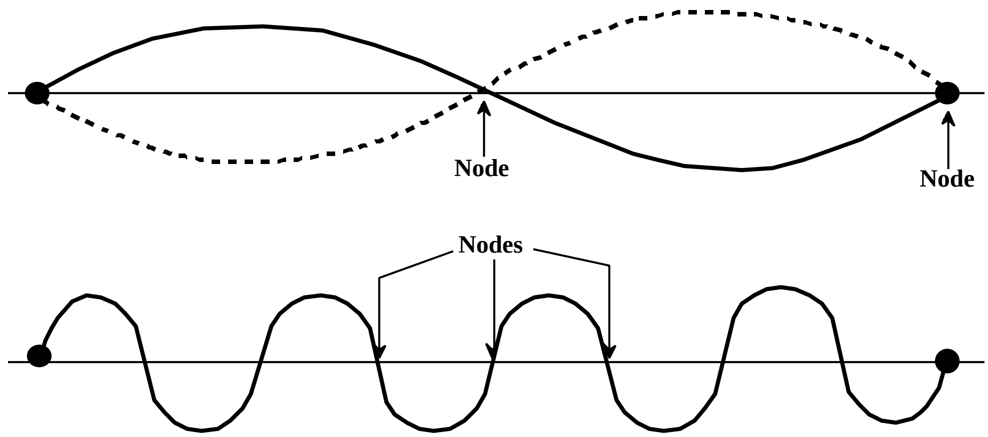


Figure 3.11 A string anchored at both ends (such as on a stringed musical instrument) can only vibrate at multiples of waves or half-waves. The illustration above shows one wave (top) and four waves (bottom). Each point at which a wave intersects the horizontal lines is called a node.

Solutions to the Schrödinger Equation, Electrons in Orbitals

With its movement governed by the laws of quantum mechanics, it is not possible to know where an electron is relative to an atom's nucleus at any specific instant. However, the Schrödinger equation permits calculation of the probability of an electron being in a specified region. Solution of the equation gives numerous wave functions, each of which corresponds to a definite energy level and to the probability of finding an electron at various locations and distances relative to the nucleus (Figure 3.12). The wave function describes **orbitals**, *each of which has a characteristic energy and region around the nucleus where the electron has certain probabilities of being found*. The term orbital is used because quantum mechanical calculations do not give specific orbits consisting of defined paths for electrons, like planets going around the sun.

The square of the wave function, ψ^2, calculated and integrated for a small segment of volume around a nucleus, is proportional to the probability of finding an electron in that small volume, which can be regarded as part of an *electron cloud.* With this view, the electron is more likely to be found in a region where the cloud is relatively more dense. The cloud has no definite outer limits, but fades away with increasing distance from the nucleus to regions in which there is essentially no probability of finding the electron.

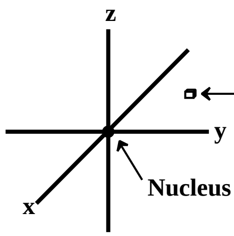


Figure 3.12. The xyz coordinate system used to describe the orientations in space of orbitals around the nucleus of an atom.

Multielectron Atoms and Quantum Numbers

Wave mechanical calculations describe various **energy levels** for electrons in an atom. Each energy level has at least one orbital. *A single orbital may contain a maximum of only two electrons*. Although the Schrödinger approach was originally applied to the simplest atom, hydrogen, it has been extended to atoms with many electrons (multielectron atoms). Electrons in atoms are described by four **quantum numbers**, which are defined and described briefly below. With these quantum numbers and a knowledge of the rules governing their use, it is possible to specify the orbitals allowed in a particular atom. *An electron in an atom has its own unique set of quantum numbers; no two electrons can have exactly identical quantum numbers.*

The Principal Quantum Number, n

Main energy levels corresponding to *electron shells* discussed earlier in this chapter are designated by a **principal quantum number**, **n**. Both the size of orbitals and the magnitude of the average energy of electrons contained therein increase with increasing n. Permitted values of n are 1, 2, 3, . . . , extending through 7 for the known elements.

The Azimuthal Quantum Number, l

Within each main energy level (shell) represented by a principal quantum number, there are **sublevels** (subshells). Each sublevel is denoted by an **azimuthal quantum number**, ***l***. For any shell with a principal quantum number of n, the possible values of l are 0, 1, 2, 3, . . . ,(n - 1). This gives the following:

- For n = 1, there is only 1 possible subshell, l = 0.
- For n = 2, there are 2 possible subshells, l = 0, 1.
- For n = 3, there are 3 possible subshells, l = 0, 1, 2.
- For n = 4, there are 4 possible subshells, l = 0, 1, 2, 3.

From the above it is seen that the maximum number of sublevels within a principal energy level designated by n is equal to n. Within a main energy level, sublevels denoted by different values of l have slightly different energies. Furthermore, l designates *different shapes of orbitals.*

The italicized letters *s*, *p*, *d*, and *f*, corresponding to l values of 0, 1, 2, and 3, respectively, are frequently used to designate sublevels. The number of electrons present in a given sublevel is limited to 2, 6, 10, and 14 for *s*, *p*, *d*, and *f*, sublevels, respectively. It follows that, since each orbital can be occupied by a maximum of 2 electrons, there is only 1 orbital in the *s* sublevel, 3 orbitals in the *p* sublevel, 5 in the *d* and 7 in the *f* . The value of the principal quantum number and the letter designating the azimuthal quantum number are written in sequence to designate both the shell and subshell. For example, 4*d* represents the *d* subshell of the fourth shell.

The Magnetic Quantum Number, m_l

The **magnetic quantum number**, m_l, is also known as the **orientational quantum number**. It designates the orientation of orbitals in space relative to each other and distinguishes orbitals within a subshell from each other. It is called the magnetic quantum number because the presence of a magnetic field can result in the appearance of additional lines among those emitted by electronically excited atoms (that is, in atomic emission spectra). The possible values of m_l in a subshell with azimuthal quantum number l are given by $m_l = +l, +(l - 1), \ldots, 0, \ldots, -(l - 1), -l$. As examples, for $l = 0$, the only possible value of m_l is 0, and for $l = 3$, m_l, may have values of 3, 2, 1, 0, -1, -2, -3.

Spin Quantum Number , m_s

The fourth and final quantum number to be considered for electrons in atoms is the **spin quantum number** m_s, which can have values of only $+1/2$ or $-1/2$. It results from the fact that an electron spins in either of two directions and generates a tiny magnetic field with an associated magnetic moment. *Two electrons can occupy the same orbital only if they have opposite spins so that their magnetic moments cancel each other.*

Quantum Numbers Summarized

The information given by each quantum number is summarized below:

- The value of the principal quantum number, n, gives the shell and main energy level of the electron.
- The value of the azimuthal quantum number, l, specifies sublevels with somewhat different energies within the main energy levels and describes the shapes of orbitals.
- The magnetic quantum number, m_l, distinguishes the orientations in space of orbitals in a subshell and may provide additional distinctions of orbital shapes.

- The spin quantum number, m_s, with possible values of only $+1/2$ and $-1/2$, accounts for the fact that each orbital can be occupied by a maximum number of only 2 electrons with opposing spins.

Specification of the values of 4 quantum numbers describes each electron in an atom, as shown in Table 3.3.

Table 3.3 Quantum Numbers for Electrons in Atoms

n[1]	l[2]	m_l[3]	m_s[4]	
1	0 (1*s*)	0	$(m_s = +1/2 \text{ or } -1/2)$	1 orbital
2	0 (2*s*)	0	$(m_s = +1/2 \text{ or } -1/2)$	1 orbital
	1 (2*p*)	-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	3 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
3	0 (3*s*)	0	$(m_s = +1/2 \text{ or } -1/2)$	1 orbital
	1 (3*p*)	-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	3 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
	2 (3*d*)	-2	$(m_s = +1/2 \text{ or } -1/2)$	
		-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	5 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
		+2	$(m_s = +1/2 \text{ or } -1/2)$	
4	0 (4*s*)	0	$(m_s = +1/2 \text{ or } -1/2)$	1 orbital
	1 (4*p*)	-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	3 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
	2 (4*d*)	-2	$(m_s = +1/2 \text{ or } -1/2)$	
		-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	5 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
		+2	$(m_s = +1/2 \text{ or } -1/2)$	
	3 (4*f*)	-3	$(m_s = +1/2 \text{ or } -1/2)$	3 orbitals
		-2	$(m_s = +1/2 \text{ or } -1/2)$	
		-1	$(m_s = +1/2 \text{ or } -1/2)$	
		0	$(m_s = +1/2 \text{ or } -1/2)$	7 orbitals
		+1	$(m_s = +1/2 \text{ or } -1/2)$	
		+2	$(m_s = +1/2 \text{ or } -1/2)$	
		+3	$(m_s = +1/2 \text{ or } -1/2)$	

[1] n, principal quantum number denoting an energy level, or shell
[2] l, azimuthal quantum number of a sublevel or subshell
[3] m_l, magnetic quantum number designating an orbital; each entry in this column designates an orbital capable of holding 2 electrons with opposing spins.
[4] m_s, spin quantum number, which may have a value of $+1/2$ or $-1/2$.

For each electron in an atom, the orbital that it occupies and the direction of its spin are specified by values of n, *l*, m_l, and m_s, assigned to it. These values are unique for each electron; *no two electrons in the same atom can have identical values of all four quantum numbers*. This rule is known as the **Pauli exclusion principle**.

3.13 ENERGY LEVELS OF ATOMIC ORBITALS

Electrons in atoms occupy the lowest energy levels available to them. Figure 3.13 is an energy level diagram that shows the relative energy levels of electrons in various atomic orbitals. Each dash, —, in the figure represents *one* orbital that is a potential slot for *two* electrons with opposing spins.

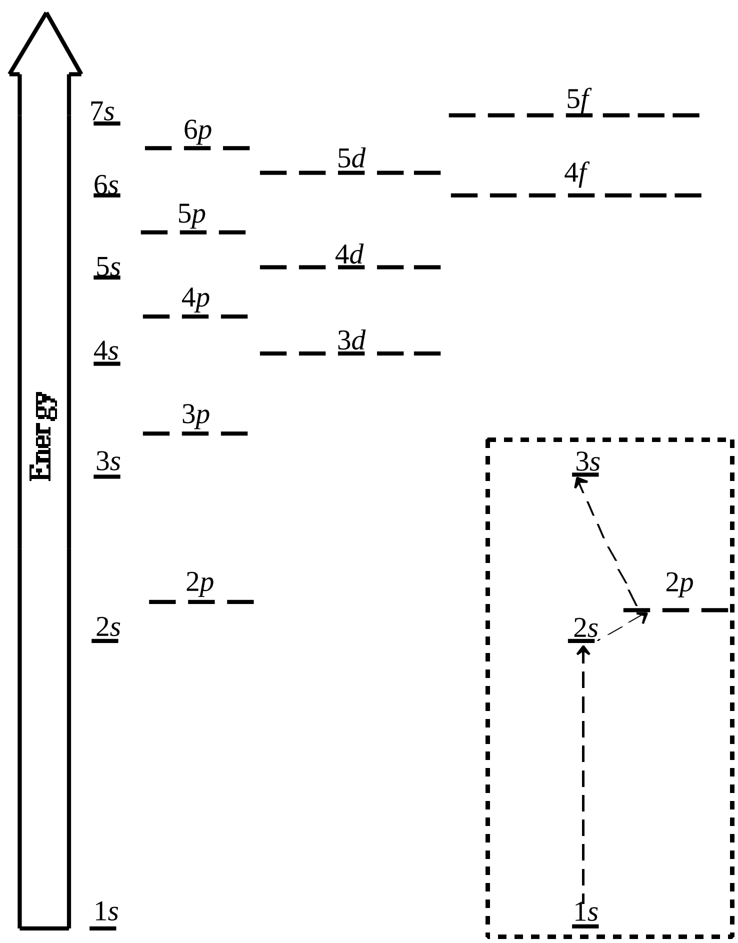


Figure 3.13 Energy levels of atomic orbitals. Each orbital capable of containing 2 electrons is shown with a dash, —. The order of placing electrons in orbitals is from the lowest-lying orbitals up, as shown in the inset.

The electron energies represented in Figure 3.13 can be visualized as those required to remove an electron from a particular orbital to a location completely away from the atom. As indicated by its lowest postion in the diagram, the most energy would be needed to remove an electron in the 1*s* orbital; comparatively little energy is required to remove electrons from orbitals having higher n values, such as 5, 6, or 7. Furthermore, the diagram shows decreasing separation of the energy levels (values of n) with increasing n. It also shows that some sublevels with a particular value of n have lower energies than sublevels for which the principal quantum number is n - 1. This is first seen from the placement of the 4*s* sublevel below that of the 3*d* sublevel. As a consequence, with increasing atomic number, the 4*s* orbital acquires 2 electrons before any electrons are placed in the 3*d* orbitals.

The value of the principal quantum number, n, of an orbital is a measure of the relative distance of the maximum electron density from the nucleus. An electron with a lower value of n, being on the average closer to the nucleus, is more strongly attracted to the nucleus than an electron with a higher n value that is at a relatively greater distance from the nucleus.

Hund's Rule of Maximum Multiplicity

Orbitals that are in the same sublevel, but that have different values of m_l, have the same energy. This is first seen in the 2*p* sublevel, where the three orbitals with m_l of -1, 0, +1 (shown as three dashes on the same level in Figure 3.13) all have the same energies. The order in which electrons go into such a sublevel follows **Hund's rule of maximum multiplicity**, which states that *electrons in a sublevel are distributed to give the maximum number of unpaired electrons* (that is, those with parallel spins having the same sign of m_s). Therefore, the first three electrons to be placed in the 2*p* sublevel would occupy the three separate available orbitals and would have the same spins. Not until the fourth electron out of a maximum number of 6 is added to this sublevel are two electrons placed in the same orbital.

3.14 SHAPES OF ATOMIC ORBITALS

In trying to represent different shapes of orbitals, it is important to keep in mind that they do not contain electrons within finite volumes, because there is no outer boundary of an orbital at which the probability of finding an electron drops to exactly 0. However, an orbital can be drawn as a figure ("fuzzy cloud") around the atom nucleus within which there is a relatively high probability (typically 90%) of finding an electron within the orbital. These figures are called **contour representations of orbitals** and are as close as one can come to visualizing the shapes of orbitals. Each point on the surface of a contour representation of an orbital has the same value of ψ^2 (square of the wave function of the Schrödinger equation, Section 3.13), and the entire surface encloses the volume within which an electron spends 90% of its time. The two most important aspects of an orbital are its size and shape, both of which are rather well illustrated by a contour representation.

Figure 3.14 shows the contour representations of the first three *s* orbitals. These are spherically shaped and increase markedly in size with increasing principal quantum number, reflecting increased average distance of the electron from the nucleus.

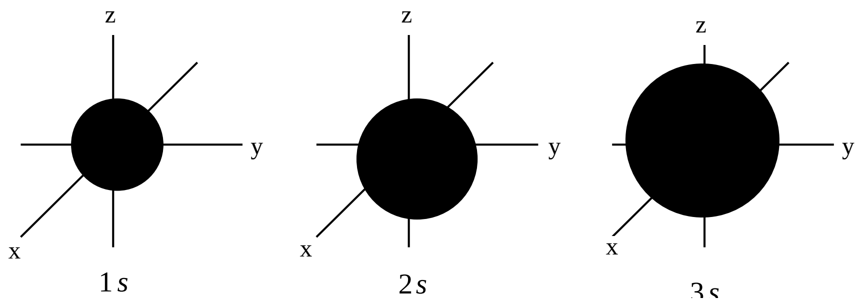


Figure 3.14 Contour representations of s orbitals for the first three main energy levels.

Figure 3.15 shows contour representations of the three 2*p* orbitals. These are seen to have different orientations in space; they have directional properties. In fact, *s* orbitals are the only ones that are spherically symmetrical. The shapes of orbitals are involved in molecular geometry and help to determine the shapes of molecules. The shapes of *d* and *f* orbitals are more complex and variable than those of *p* orbitals and are not discussed in this book.

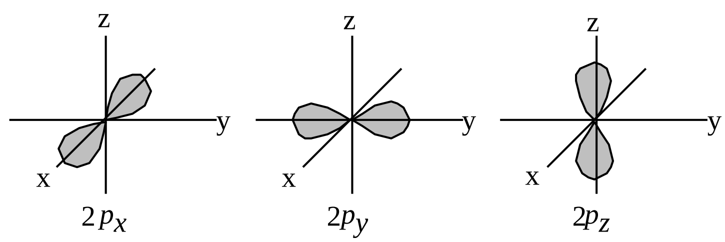


Figure 3.15 Contour representations of 2*p* orbitals.

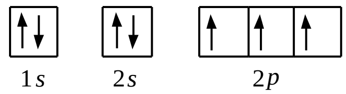
3.15 ELECTRON CONFIGURATION

Electron configuration is a means of stating which kinds of orbitals contain electrons and the numbers of electrons in each kind of orbital of an atom. It is expressed by the number and letter representing each kind of orbital and superscript numbers telling how many electrons are in each sublevel. The hydrogen atom's one electron in the 1*s* orbital is shown by the following notation:

Sublevel notation
↓
$1s^1$
↑
Number of electrons in sublevel

Nitrogen, atomic number 7, has 7 electrons, of which 2 are paired in the 1*s* orbital, 2 are paired in the 2*s* orbital and 3 occupy singly each of the 3 available 2*p* orbitals. This electron configuration is designated as $1s^2 2s^2 2p^3$.

The **orbital diagram** is an alternative way of expressing electron configurations in which each separate orbital is represented by a box. Individual electrons in the orbitals are shown as arrows pointing up or down to represent opposing spins (m_s = $+1/2$ or $-1/2$). The orbital diagram for nitrogen is the following:



This gives all the information contained in the notation $1s^22s^22p^3$, but emphasizes that the three electrons in the three available 2*p* orbitals each occupy separate orbitals, a condition that is a consequence of Hund's rule of maximum multiplicity (Section 3.13).

Most of the remainder of this chapter is devoted to a discussion of the placement of electrons in atoms with increasing atomic number, its effects upon the chemical behavior of atoms, and how it leads to a systematic organization of the elements in the periodic table. This placement of electrons is in the order shown in increasing energy levels from bottom to top in Figure 3.13 as illustrated for nitrogen in Figure 3.16.

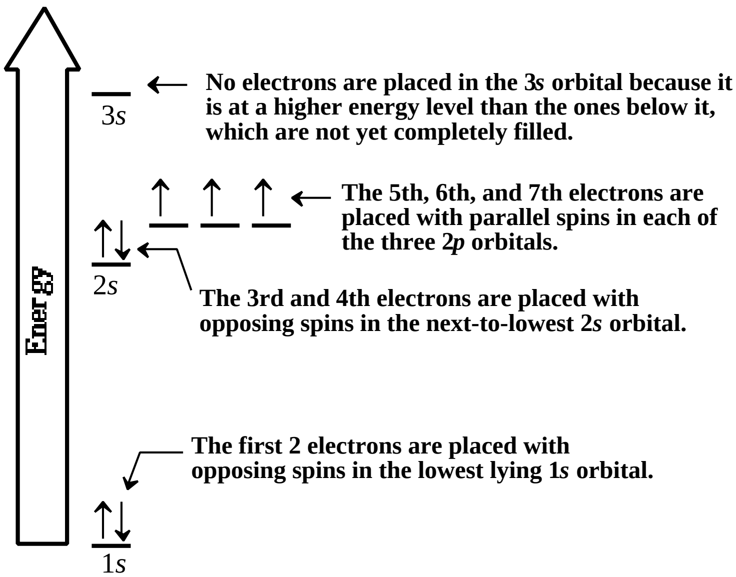


Figure 3.16 Placement of electrons in orbitals for the 7-electron nitrogen atom according to the energy level diagram.

3.16 ELECTRONS IN THE FIRST 20 ELEMENTS

The most meaningful way to place electrons in the orbitals of atoms is on the basis of the periodic table. This enables relating electron configurations to chemical properties and the properties of elements in groups and periods of the periodic table. In this section, electron configurations are deduced for the first 20 elements and given in an abbreviated version of the periodic table.

Electron Configuration of Hydrogen

The 1 electron in the hydrogen atom goes into its lowest-lying 1*s* orbital. Figure 3.17 summarizes all the information available about this electron and its electron configuration.

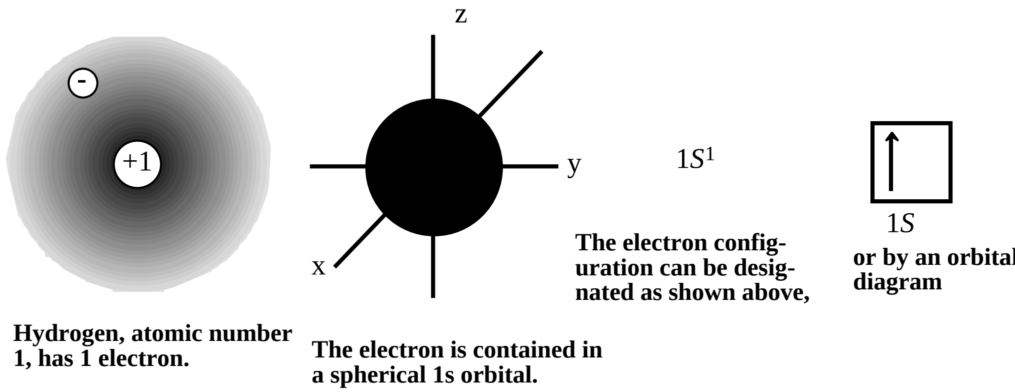


Figure 3.17 Representation of the electron on the hydrogen atom.

Electron Configuration of Helium

An atom of helium, atomic number 2, has 2 electrons, both contained in the 1*s* orbital and having quantum numbers n = 1, l = 0, m_l = 0, and $m_s = +1/2$ and $-1/2$. Both electrons have the same set of quantum numbers except for m_s. The electron configuration of helium can be represented as $1s^2$, showing that there are 2 electrons in the 1*s* orbital. Two is the maximum number of electrons that can be contained in the first principal energy level; additional electrons in atoms with atomic number greater than 2 must go into principal energy levels with n greater than 1. There are several other noble gas elements in the periodic table, of which neon and argon have already been mentioned, but helium is the only one with a filled shell of 2 electrons—the rest have stable outer shells of 8 electrons.

Electron Configurations of Elements 2–20

The electron configurations of elements with atomic number through 20 are very straightforward. Electrons are placed in order of orbitals with increasing energy as shown in Figure 3.13. This order is

$$1s^22s^22p^63s^23p^64s^2$$

Note that in this configuration the electrons go into the 4*s* orbital before the 3*d* orbital, which lies at a slightly higher energy level. In filling the *p* orbitals, it should also be kept in mind that 1 electron goes into each of three *p* orbitals before pairing occurs. In order to follow the discussion of electron configurations for elements through 20, it is useful to refer to the abbreviated periodic table in Figure 3.20.

Lithium

For lithium, atomic number 3, two electrons are placed in the 1*s* orbital, leaving the third electron for the 2*s* orbital. This gives an electron configuration of $1s^22s^1$. The two 1*s* electrons in lithium are in the stable noble gas electron configuration of helium and are very difficult to remove. These are lithium's *inner electrons* and, along with the nucleus, constitute the **core** of the lithium atom. The electron configuration *of the core* of the lithium atom is $1s^2$, the same as that of helium. Therefore, the lithium atom is said to have a *helium core*. *The core of any atom consists of its nucleus plus its inner electrons, those with the same electron configuration as the noble gas immediately preceding the element in the periodic table*.

Valence Electrons

Lithium's lone $1s$ electron is an outer electron contained in the *outer shell* of the atom. Outer shell electrons are also called **valence electrons**, and are the electrons that can be shared in covalent bonding or lost to form cations in ionic compounds. Lithium's valence electron is illustrated in Figure 3.18.

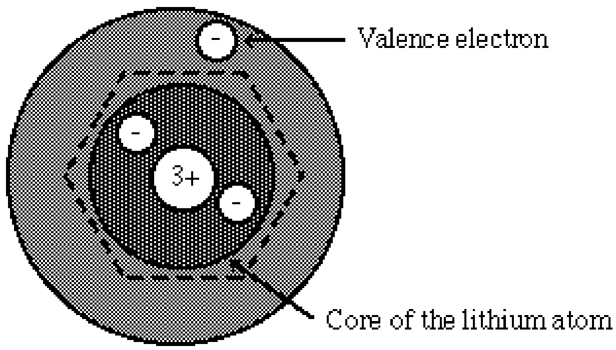


Figure 3.18 Core (kernel) and valence electrons of the lithium atom.

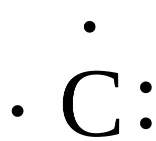
Beryllium

Beryllium, atomic number 4, has 2 inner electrons in the $1s$ orbital and 2 outer electrons in the $2s$ orbital. Therefore, beryllium has a *helium core*, plus 2 *valence electrons*. Its electron configuration is $1s^22s^2$.

Filling the 2*p* orbitals

Boron, atomic number 5, is the first element containing an electron in a $2p$ orbital. Its electron configuration is $1s^22s^22p^1$. Two of the five electrons in the boron atom are contained within the spherical orbital closest to the nucleus, and 2 more are in the larger spherical $2s$ orbital. The lone $2p$ electron is in an approximately dumbell-shaped orbital in which the average distance of the electron from the nucleus is about the same as that of the $2s$ electrons. This electron could have any one of the three orientations in space shown for p orbitals in Figure 3.15.

The electron configuration of carbon, atomic number 6, is $1s^22s^22p^2$. The four outer (valence) electrons are shown by the Lewis symbol



Two of the four valence electrons are represented by a pair of dots, :, to indicate that these electrons are paired in the same orbital. These are the two 2*s* electrons. The other two are shown as individual dots to represent two unpaired 2*p* electron in separate orbitals.

The 7 electrons in nitrogen, N, are in a $1s^22s^22p^3$ electron configuration. Nitrogen is the first element to have at least one electron in each of 3 possible *p* orbitals (see Figure 3.19).

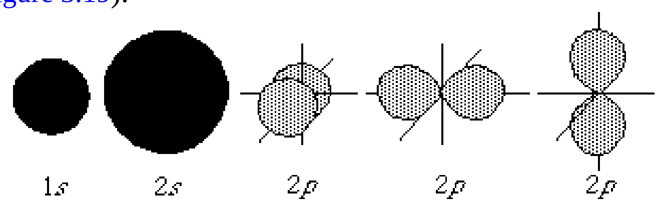
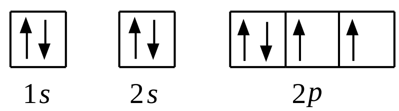
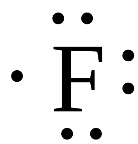


Figure 3.19. The N atom contains 2 electrons in the 1*s* orbital, 2 in the 2*s* orbital, and 1 in each of three separate 2*p* orbitals oriented in different directions in space.

The next element to be considered is oxygen, which has an atomic number of 8. Its electron configuration is $1s^22s^22p^4$. It is the first element in which it is necessary for 2 electrons to occupy the same *p* orbital, as shown by the following orbital diagrams:

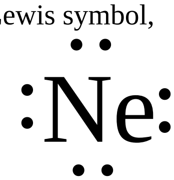


The electron configuration of fluorine, atomic number 9, is $1s^22s^22p^5$. Its Lewis symbol,



shows that the fluorine atom has only one unpaired electron of the 7 electrons in its valence shell. In its chemical reactions, fluorine seeks to obtain another electron to give a stable *octet.*

Neon, atomic number 10, is at the end of the second period of the periodic table and is a noble gas as shown by its Lewis symbol,



denoting a filled octet of electrons. Its electron configuration is $1s^22s^22p^6$. In this configuration the $2s^22p^6$ portion stands for the outer electrons. Like neon, all other elements with the outer electron configuration $\mathbf{ns^2np^6}$ are noble gases located in the far right of the periodic table (for example, argon, with the electron configuration $1s^22s^22p^6\mathbf{3s^23p^6}$).

Filling the 3*s*, 3*p*, and 4*s* orbitals

The 3*s* and 3*p* orbitals are filled in going across the third period of the periodic table from sodium through argon. Atoms of these elements, like all atoms beyond neon, have their 10 innermost electrons in the neon electron configuration of $1s^22s^22p^6$. Therefore, these atoms have a *neon core*, which may be designated {Ne}.

{Ne} stands for $1s^22s^22p^6$

With this notation, the electron configuration of element number 11, sodium may be shown as $\{Ne\}3s^1$, which is an abbreviation for $1s^22s^22p^63s^1$. The former notation has some advantage in simplicity, while showing the outer electrons specifically. In the example just cited it is easy to see that sodium has 1 outer shell 3*s* electron, which it can lose to form the Na^+ ion with its stable noble gas neon electron configuration.

At the end of the third period is located the noble gas argon, atomic number 18, with the electron configuration $1s^22s^22p^63s^23p^6$. For elements beyond argon, that portion of the electron configuration identical to argon's may be represented simply as {Ar}. Therefore, the electron configuration of potassium, atomic number 19, is $1s^22s^22p^63s^23p^64s^1$, abbreviated $\{Ar\}4\ s^1$, and that of calcium, atomic number 20, is $1s^22s^22p^63s^23p^64s^2$, abbreviated $\{Ar\}4\ s^2$. This completes the periodic table through element number 20 as shown in Figure 3.20.

1 H $1s^1$							2 He $1s^2$
3 Li $\{He\}2s^1$	4 Be $\{He\}2s^2$	5 B $\{He\}2s^22p^1$	6 C $\{He\}2s^22p^2$	7 N $\{He\}2s^22p^3$	8 O $\{He\}2s^22p^4$	9 F $\{He\}2s^22p^5$	10 Ne $\{He\}2s^22p^6$
11 Na $\{Ne\}3s^1$	12 Mg $\{Ne\}3s^2$	13 Al $\{Ne\}3s^23p^1$	14 Si $\{Ne\}3s^23p^2$	15 P $\{Ne\}3s^23p^3$	16 S $\{Ne\}3s^23p^4$	17 Cl $\{Ne\}3s^23p^5$	18 Ar $\{Ne\}3s^23p^6$
19 K $\{Ar\}4s^1$	20 Ca $\{Ar\}4s^2$						

Figure 3.20 Abbreviated periodic table showing the electron configurations of the first 20 elements.

3.17 ELECTRON CONFIGURATIONS AND THE PERIODIC TABLE

There are several learning devices to assist expression of the order in which atomic orbitals are filled (electron configuration for each element). However, it is of little significance to express electron configurations without an understanding of the meaning of the configurations. By far the most meaningful way to understand this important aspect of chemistry is within the context of the periodic table, as outlined in Figure 3.21. To avoid clutter, only atomic numbers of key elements are shown in this table. The double-pointed arrows drawn horizontally across the periods are labeled with the kind of orbital being filled in that period.

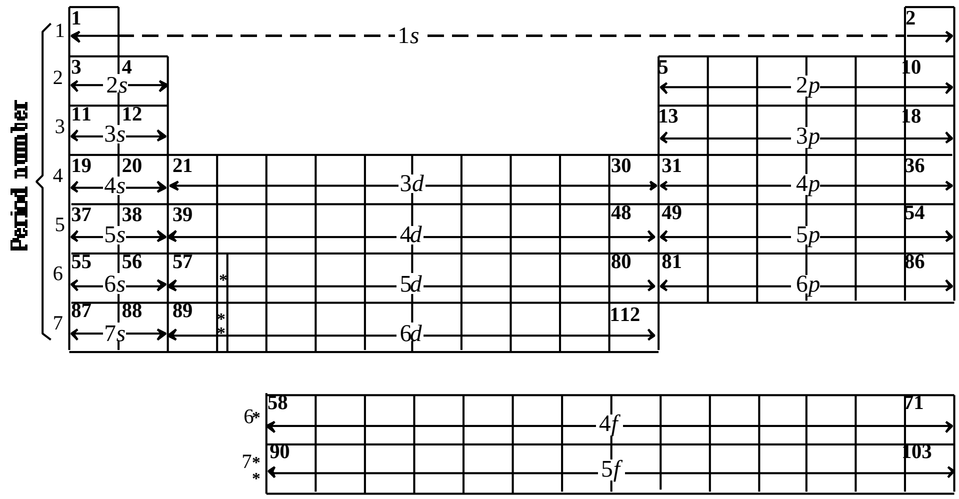


Figure 3.21 Outline of the periodic table labeled to show the filling of atomic orbitals. The type of element being filled with increasing atomic number is shown by the labeled horizontal lines across the periods. The left point of each arrow, ←, marks an element in which the filling of a new kind of orbital starts, and the right point of each arrow, →, marks an element for which the filling of a new kind of orbital is completed. The atomic numbers are given for each element at which the filling of a kind of orbital begins or is completed.

The first step in using the periodic table to figure out electron configurations is to note that the periods are numbered 1 through 7 from top to bottom along the left side of the table. These numbers correspond to the principal quantum numbers (n values) for the orbitals that become filled across the period for both *s* and *p* orbitals. Therefore, in the first group of elements—those with atomic numbers 1, 3, 11, 19, 37, 55, and 87—the last electron added is in the n*s* orbital, for example, 5*s* for Rb, atomic number 37. The last electron added to each of the second group of elements —those with atomic numbers 4, 12, 20, 38, 56, and 88—is the second electron going into the n*s* orbital. For example, in the third period, Mg, atomic number 12, has a filled 3*s* orbital containing 2 electrons. For the group of elements in which the *p* orbitals start to be filled—those in the column with atomic numbers 5, 13, 31, 49, 81—the last electron added to each atom is the first one to enter an n*p* orbital. For example for element 31, Ga, which is contained in the 4th period, the outermost electron is in the 4*s* orbital. In going to the right across each period in this part of the periodic table, the outermost electrons are, successively, np^2, np^3, np^4, np^5, and np^6. Therefore, for elements with atomic numbers 14, 15, 16, 17, and 18 in the 3rd period, the outermost electrons are 3p^2, 3p^3, 3p^4, 3p^5, and 3p^6. Each of the noble gases beyond helium has filled n*p* orbitals with a total of 6 outermost *p* electrons in the three n*p* orbitals.

Each set of five *d* orbitals becomes filled for the *transition metals* in the three horizontal periods beginning with atomic numbers 21, 39, and 57 and ending with, successively, atomic numbers 30, 48, and 80. *For each of these orbitals the value of* *n is **1 less** than the period number in which the orbitals become filled*. The first *d* orbitals to become filled are the lowest-lying ones possible, the 3*d* orbitals, which become filled in the fourth period. Across the 5th period, where the 5*s* and 5*p* orbitals become filled for elements 37-38 and 49-54, respectively, the 4*d* orbitals (n 1 less than the period number) become filled for the transition metals, atomic numbers 39-48.

Below the main body of the periodic table are the **inner transition elements** consisting of two rows of elements that are actually parts of the 6th and 7th periods, respectively. The first *f* orbitals begin to fill with the first of the *lanthanides*, element number 58 (Figure 3.26). The principal quantum numbers of their *f* orbitals are *2 less than* their period numbers. These are 4*f* orbitals, of which there are 7 that are filled completely with element number 71. The 4*f* orbitals are filled in the *6th* period. The 5*f* orbitals are filled with the actinide elements, atomic numbers 90-103.

With Figure 3.26 in mind, it is possible to write the expected electron configurations of any of the elements; these are given for all elements in Table 3.4. Consider the following examples:

- Atomic No. 16, S: 3rd period, 16 electrons, Ne core, outermost electrons 3*p*, electron configuration {Ne}$3s^23p^4$
- Atomic No. 23, V: 4th period, 23 electrons, Ar core, outermost electrons 3*d*, electron configuration {Ar}$4s^23d^3$
- Atomic No. 35, Br: 4th period, 35 electrons, Ar core, all 3*d* orbitals filled, outermost electrons 4*p*, electron configuration {Ne}$4s^23d^{10}4p^5$
- Atomic No. 38, Sr: 5th period, 38 electrons, Kr core, outermost electrons 5*s*, electron configuration {Kr}$5s^2$
- Atomic No. 46, Pd: 5th period, 46 electrons, Kr core, outermost electrons 4*d*, electron configuration {Kr}$5s^24d^8$
- Atomic No. 77, Ir: 6th period, 77 electrons, Xe core, all 4*f* orbitals filled in the 6th period, outermost electrons 5*d*, electron configuration {Xe}$6s^24f^{14}5d^7$

Considering the above, it is possible to write the expected electron configurations of any of the elements. The actual electron configurations are given in Table 3.4. In some cases, these vary slightly from those calculated according to the rules. outlined above. These exceptions occur because of the relatively higher stabilities of two half-filled sets of outermost orbitals, or orbitals in which one is half-filled and one entirely filled. The examples below illustrate this point:

- Cr, atomic number 24. Rules predict {Ar}$4s^23d^4$. However the actual electron configuration is {Ar}$4s^13d^5$ because this gives the slightly more stable electron configuration with *half-filled* 4*s and* 3*d* orbitals.
- Cu, atomic number 29. Rules predict {Ar}$4s^23d^9$. However the actual electron configuration is {Ar}$4s^13d^{10}$ because this gives *half-filled* 4*s* and *filled* 3*d* orbitals.

Table 3.4 Electron Configurations of the Elements

Atomic number	Symbol	Electron configuration	Atomic number	Symbol	Electron configuration
1	H	$1s^{1}$	55	Cs	{Xe}$6s^{1}$
2	He	$1s^{2}$	56	Ba	{Xe}$6s^{2}$
3	Li	{He}$2s^{1}$	57	La	{Xe}$6s^{2}5d^{1}$
4	Be	{He}$2s^{2}$	58	Ce	{Xe}$6s^{2}4f^{1}5d^{1}$
5	B	{He}$2s^{2}2p^{1}$	59	Pr	{Xe}$6s^{2}4f^{3}$
6	C	{He}$2s^{2}2p^{2}$	60	Nd	{Xe}$6s^{2}4f^{4}$
7	N	{He}$2s^{2}2p^{3}$	61	Pm	{Xe}$6s^{2}4f^{5}$
8	O	{He}$2s^{2}2p^{4}$	62	Sm	{Xe}$6s^{2}4f^{6}$
9	F	{He}$2s^{2}2p^{5}$	63	Eu	{Xe}$6s^{2}4f^{7}$
10	Ne	{He}$2s^{2}2p^{6}$	64	Gd	{Xe}$6s^{2}4f^{7}5d^{1}$
11	Na	{Ne}$3s^{1}$	65	Tb	{Xe}$6s^{2}4f^{9}$
12	Mg	{Ne}$3s^{2}$	66	Dy	{Xe}$6s^{2}4f^{10}$
13	Al	{Ne}$3s^{2}3p^{1}$	67	Ho	{Xe}$6s^{2}4f^{11}$
14	Si	{Ne}$3s^{2}3p^{2}$	68	Er	{Xe}$6s^{2}4f^{12}$
15	P	{Ne}$3s^{2}3p^{3}$	69	Tm	{Xe}$6s^{2}4f^{13}$
16	S	{Ne}$3s^{2}3p^{4}$	70	Yb	{Xe}$6s^{2}4f^{14}$
17	Cl	{Ne}$3s^{2}3p^{5}$	71	Lu	{Xe}$6s^{2}4f^{14}5d^{1}$
18	Ar	{Ne}$3s^{2}3p^{6}$	72	Hf	{Xe}$6s^{2}4f^{14}5d^{2}$
19	K	{Ar}$4s^{1}$	73	Ta	{Xe}$6s^{2}4f^{14}5d^{3}$
20	Ca	{Ar}$4s^{2}$	74	W	{Xe}$6s^{2}4f^{14}5d^{4}$
21	Sc	{Ar}$4s^{2}3d^{1}$	75	Re	{Xe}$6s^{2}4f^{14}5d^{5}$
22	Ti	{Ar}$4s^{2}3d^{2}$	76	Os	{Xe}$6s^{2}4f^{14}5d^{6}$
23	V	{Ar}$4s^{2}3d^{3}$	77	Ir	{Xe}$6s^{2}4f^{14}5d^{7}$
24	Cr	{Ar}$4s^{1}3d^{5}$	78	Pt	{Xe}$6s^{1}4f^{14}5d^{9}$
25	Mn	{Ar}$4s^{2}3d^{5}$	79	Au	{Xe}$6s^{1}4f^{14}5d^{10}$
26	Fe	{Ar}$4s^{2}3d^{6}$	80	Hg	{Xe}$6s^{2}4f^{14}5d^{10}$
27	Co	{Ar}$4s^{2}3d^{7}$	81	Tl	{Xe}$6s^{2}4f^{14}5d^{10}6p^{1}$
28	Ni	{Ar}$4s^{2}3d^{8}$	82	Pb	{Xe}$6s^{2}4f^{14}5d^{10}6p^{2}$
29	Cu	{Ar}$4s^{1}3d^{10}$	83	Bi	{Xe}$6s^{2}4f^{14}5d^{10}6p^{3}$
30	Zn	{Ar}$4s^{2}3d^{10}$	84	Po	{Xe}$6s^{2}4f^{14}5d^{10}6p^{4}$
31	Ga	{Ar}$4s^{2}3d^{10}4p^{1}$	85	At	{Xe}$6s^{2}4f^{14}5d^{10}6p^{5}$
32	Ge	{Ar}$4s^{2}3d^{10}4p^{2}$	86	Rn	{Xe}$6s^{2}4f^{14}5d^{10}6p^{6}$
33	As	{Ar}$4s^{2}3d^{10}4p^{3}$	87	Fr	{Rn}$7s^{1}$
34	Se	{Ar}$4s^{2}3d^{10}4p^{4}$	88	Ra	{Rn}$7s^{2}$
35	Br	{Ar}$4s^{2}3d^{10}4p^{5}$	89	Ac	{Rn}$7s^{2}6d^{1}$
36	Kr	{Ar}$4s^{2}3d^{10}4p^{6}$	90	Th	{Rn}$7s^{2}6d^{2}$
37	Rb	{Kr}$5s^{1}$	91	Pa	{Rn}$7s^{2}5p^{2}6d^{1}$
38	Sr	{Kr}$5s^{2}$	92	U	{Rn}$7s^{2}5p^{3}6d^{1}$
39	Y	{Kr}$5s^{2}4d^{1}$	93	Np	{Rn}$7s^{2}5f^{4}6d^{1}$
40	Zr	{Kr}$5s^{2}4d^{2}$	94	Pu	{Rn}$7s^{2}5f^{6}$
41	Nb	{Kr}$5s^{1}4d^{4}$	95	Am	{Rn}$7s^{2}5f^{7}$
42	Mo	{Kr}$5s^{1}4d^{5}$	96	Cm	{Rn}$7s^{2}5f^{7}6d^{1}$
43	Tc	{Kr}$5s^{2}4d^{5}$	97	Bk	{Rn}$7s^{2}5f^{9}$
44	Ru	{Kr}$5s^{1}4d^{7}$	98	Cf	{Rn}$7s^{2}5f^{10}$
45	Rh	{Kr}$5s^{1}4d^{8}$	99	Es	{Rn}$7s^{2}5f^{11}$
46	Pd	{Kr}$4d^{10}$	100	Fm	{Rn}$7s^{2}5f^{12}$
47	Ag	{Kr}$5s^{1}4d^{10}$	101	Md	{Rn}$7s^{2}5f^{13}$
48	Cd	{Kr}$5s^{2}4d^{10}$	102	No	{Rn}$7s^{2}5f^{14}$
49	In	{Kr}$5s^{2}4d^{10}5p^{1}$	103	Lr	{Rn}$7s^{2}5f^{14}6d^{1}$
50	Sn	{Kr}$5s^{2}4d^{10}5p^{2}$	104	Rf	{Rn}$7s^{2}5f^{14}6d^{2}$
51	Sb	{Kr}$5s24d^{10}5p^{3}$	105	Ha	{Rn}$7s^{2}5f^{14}6d^{3}$
52	Te	{Kr}$5s24d^{10}5p^{4}$	106	Unh	{Rn}$7s^{2}5f^{14}6d^{4}$
53	I	{Kr}$5s24d^{10}5p^{5}$	107	Uns	{Rn}$7s^{2}5f^{14}6d^{5}$
54	Xe	{Kr}$5s24d^{10}5p^{6}$	108	---	{Rn}$7s^{2}5f^{14}6d^{6}$
			109	Une	{Rn}$7s^{2}5f^{14}6d^{7}$

CHAPTER SUMMARY

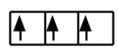
The chapter summary below is presented in a programmed format to review the main points covered in this chapter. It is used most effectively by filling in the blanks, referring back to the chapter as necessary. The correct answers are given at the end of the summary.

Briefly, the basic parts of the atomic theory are [1]____,
[2]____,
[3]____,
[4]____,
and [5]____.
Three fundamental laws that are explained by the atomic theory are [6]____________________, [7]____________________, and [8]____________________ The atomic mass unit is used to [9]____________________ and is defined as [10]____________________. Most of the volume of an atom is composed of [11]____________________. The three subatomic particles of concern to chemists, their charges, and mass numbers are[12]____________________. The atomic number and mass number of $^{14}_{7}N$ are [13]____________________. A systematic arrangement of elements that places those with similar chemical properties and electron configurations in the same groups is the [14] ____________________. Most hydrogen atoms have a nucleus consisting of [15]__________. The notation Ca: is an example of [16]____________________ and H:H is an example of [17]____________________. Helium is the first element with a [18]____________________. Lithim is the first element having both [19]__________ and [20]__________ electrons. Four general characteristics of metals are [21]____.
The Lewis symbols of elements 3–10 are [22]____________________.The octet rule is [23]____.
The names of the elements in the third period of the periodic table are [24]____. The noble gases in the first 20 elements are [25]____________________.
The unique behavior of charged particles that are as small and move as rapidly as electrons is explained by [26]____________________. Electromagnetic radiation has a characteristic [27]__________and [28]__________ related by the equation [29]__________. According to the quantum theory, radiant energy may be absorbed or emitted only in discrete packets called [30]__________, the energy of which is given by the equation [31]__________. The Bohr model introduced the all-important concept that [32]____.

The wave mechanical model of electrons treats them as [33]____ around the nucleus of an atom. The wave mechanical model of electrons in atoms is treated by the [34]________________________________ equation expressed mathematically as [35]__________________. In this equation ψ is the [36]____________________________ and is a function of the electron's [37]____. According to the wave mechanical model electrons occupy [38]____________ each of which has [39]____.

A single orbital can contain a maximum of [40]___ electrons. An electron in an atom is described by four [41]_________________ which may not be [42]____________________ for any two electrons in an atom. The symbol n represents the [43]_________________ which may have values of [44]_________________________. The symbol *l* represents the [45]________________ quantum number, which can have values of [46]_________________________. The symbol m_l represents [47]________________________________ with possible values of [48]____. The symbol m_s is the [49]____________________________ which may have values of only [50]________________________________. Using standard notation for electron configuration (starting $1s^2 2s^2 2p^6$) the order of filling of orbitals and the maximum number of electrons in each is [51]____.

The orbital diagram for the *p* electrons in nitrogen,



illustrates the rule that [52]____.

The entire surface of a contour representation of an orbital encloses [53]____.

The contour representation of an *s* orbital is that of [54]__________________________ whereas that of a *p* orbital is shaped like a [55]________________________________.

[56]________________________________ is a means of stating which kinds of orbitals contain electrons and the numbers of electrons in each kind of orbital of an atom. It is expressed by the number and letter representing [57]___________________________ and superscript numbers telling [58]____.

The electron configurations of phosphorus (P), potassium (K), and arsenic (As) are, respectively, [59]____.

The part of an atom consisting of its nucleus and the electrons in it equivalent to those of the noble gas immediately preceding the element in the periodic table is the [60]_____________ of the atom. The electronic configuration *of the core* of the lithium atom is $1s^2$, the same as that of helium. Therefore, the lithium atom is said to have a [61]___________________. Electrons that can be shared in covalent bonding or lost to form cations in ionic compounds are called [62]_________________________. All elements with the outer electron configuration ns^2np^6 are [63]__________________. In respect to period number in the periodic table, the principal quantum number of *s* electrons is [64]________________________________, that of *p* electrons is [65]____, that of *d* electrons is [66]____, and that of *f* electrons is [67]________________________________. The types of elements in which *d* orbitals become filled are [68]________________. The electron configuration of Cr, atomic number 24, is {Ar}$4s^1 3d^5$,

which appears to deviate slightly from the rules because it gives a [69]______________
____.

Answers to Chapter Summary

1. Elements are composed of small objects call atoms.
2. Atoms of different elements do not have identical chemical properties.
3. Chemical compounds are formed by combination of atoms of different elements in definite ratios.
4. Chemical reactions involve the separation and combination of atoms.
5. During the course of ordinary chemical reactions, atoms are not created, destroyed, or changed to atoms of other elements.
6. law of conservation of mass
7. law of multiple proportions
8. law of constant composition
9. express masses of atoms
10. exactly 1/12 mass of carbon-12 isotope
11. electrons around the nucleus
12. proton (+1, 1), electron (-1, 0), neutron (0, 1)
13. 7 and 14
14. periodic table
15. 1 proton
16. an electron-dot symbol or Lewis symbol
17. an electron-dot formula or Lewis formula
18. filled electron shell
19. inner
20. outer
21. luster, malleable, conduct electricity, tend to lose electrons to form cations
22. 3 Li·, 4 Be:, 5 ·B:, 6 ·C:, 7 ·N:, 8 ·O:, 9 ·F:, 10 :Ne:
23. the tendency of elements to acquire an 8-electron outer electron configuration in chemical compounds
24. sodium, magnesium, aluminum, silicon, phosphorus, sulfur, chlorine, and argon
25. helium, neon, argon
26. quantum theory
27. wavelength
28. frequency
29. $\nu\lambda = c$
30. quanta or photons
31. $E = h\nu$
32. only specific energy states are allowed for an electron in an atom
33. standing waves
34. Schrödinger equation
35. $H\psi = E\psi$
36. wave function
37. energy and the coordinates in space where it may be found
38. orbitals

39. characteristic energy and region around the nucleus where the electron has certain probabilities of being found
40. two
41. quantum numbers
42. identical
43. principal
44. 1, 2, 3, 4, 5, 6, 7. . .
45. azimuthal
46. 0, 1, 2, 3, . . . ,(n - 1)
47. magnetic quantum number
48. $+l$, $+(l - 1)$, . . . ,0, . . . , $-(l - 1)$, $-l$
49. spin quantum number
50. $+1/2$ or $-1/2$
51. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2$
52. electrons in a sublevel are distributed to give the maximum number of unpaired electrons
53. the volume within which an electron spends 90% of its time
54. a sphere
55. dumbell, or two "pointed" spheres touching at the nucleus
56. Electron configuration
57. each kind of orbital
58. how many electrons are in each sublevel
59. $1s^2 2s^2 2p^6 3s^2 3p^3$, $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$, $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$
60. core
61. helium core
62. valence electrons
63. noble gases located in the far right of the periodic table
64. the same as the period number
65. the same as the period number
66. one less than the period number
67. two less than the period number
68. transition metals
69. slightly more-stable electron configuration with half-filled 4*s* and 3*d* orbitals

QUESTIONS AND PROBLEMS

Section 3.2. The Atomic Theory

1. Match the law or observation denoted by letters below with the portion of Dalton's atomic theory that explains it denoted by numbers:

   A. Law of Conservation of Mass B. Law of Constant Composition

   C. Law of Multiple Proportions D. The reaction of C with O_2 does not produce SO_2.

   1. Chemical compounds are formed by the combination of atoms of different elements in definite constant ratios that usually can be expressed as integers or simple fractions.

2. During the course of ordinary chemical reactions, atoms are not created or destroyed

3. During the course of ordinary chemical reactions, atoms are not changed to atoms of other elements.

4. Illustrated by groups of compounds such as $CHCl_3$, CH_2Cl_2, or CH_3Cl.

2. Particles of pollutant fly ash may be very small. Estimate the number of atoms in such a small particle assumed to have the shape of a cube that is 1 micrometer (μm) to the side. Assume also that an atom is shaped like a cube 100 picometers (pm) on a side.

3. Explain why it is incorrect to say that atomic mass is the mass of any atom of an element. How is atomic mass defined?

Section 3.3. Subatomic Particles

4. The $^{12}_{6}C$ isotope has a mass of exactly 12 u. Compare this with the sum of the masses of the subatomic particles that compose this isotope. Is it correct to say that the mass of an isotope is exactly equal to the sum of the masses of its subatomic particles? Is it close to the sum?

5. What is the distinction between the mass of a subatomic particle and its mass number?

6. Add up the masses of all the subatomic particles in an isotope of carbon-12. Is the sum exactly 12? Should it be?

7. Fill in the blanks in the table below

Subatomic particle	Symbol	Unit Charge	Mass number	Mass in u	Mass in grams
Proton	(a)_____	(b)_____	(c)_____	(d)_______	(e)_________
(f)______	*n*	0	1	(g)_______	(h)_________
Electron	*e*	(i)______	(j)______	(k)_______	9.1096×10^{-28}

Section 3.4. The Basic Structure of the Atom

8. Define what is meant by *x*, *y*, and A in the notation $^{y}_{x}A$.

9. Describe what happens to the magnitude and direction of the forces between charged particles (electrons, protons, nuclei) of (a) like-charge and (b) unlike charge with distance and magnitude of charge.

Section 3.6. Hydrogen, the Simplest Atom

10. What is the Lewis symbol of hydrogen and what does it show? What is the Lewis formula of H_2 and what does it show?

11. In many respects the properties of elemental hydrogen are unique. List some of these properties and some of the major uses of H_2.

Section 3.7. Helium, the First Atom with a Filled Electron Shell

12. Give the Lewis symbol of helium and explain what it has to do with (a) electron shell, (b) filled electron shell, (c) and noble gases.

13. Where is helium found, and for what purpose is it used?

Section 3.8. Lithium, The First Atom with Both Inner and Outer Electrons

14. Using dots to show <u>all</u> of its electrons, give the Lewis symbol of Li. Explain how this symbol shows (a) inner and outer electrons and electron shells, (b) valence electrons, (c) and how Li^+ ion is formed.

15. Discuss the chemical and physical properties of lithium that indicate that it is a metal.

Section 3.9. The Second Period, Elements 4-10

16. What is a particular health concern with beryllium?

17. Based upon its electronic structure, suggest why boron behaves like a metalloid, showing properties of both metals and nonmetals.

18. Carbon has two isotopes that are of particular importance. What are they and why are they important?

19. Why might carbon be classified as a “life element”?

20. What two species other than O_2 are possible for elemental oxygen, particularly in the stratosphere:

21. What do particular kinds of fluorine compounds have to do with atmospheric ozone?

22. In reference to neon define and explain the significance of (a) noble gas, (b) octet of outer shell electrons, (c) noble gas outer electron configuration, (d) octet rule.

Section 3.10. Elements 11-20, and Beyond

23. To which class of elements do lithium, sodium, and potassium belong? What are their elemental properties?

Section 3.11. A More Detailed Look at Atomic Structure

24. Define and explain (a) electromagnetic radiation, (b) wave character of electromagnetic radiation, (c) quanta (photons).

Section 3.12. Quantum and Wave Mechanical Models of Electrons in Atoms

25. What is the significance of the fact that very specific wavelengths of light are emitted when atoms in an excited state revert back to a lower energy state (ground state)?

26. What are the major accomplishments and shortcomings of the Bohr theory?

27. How are electrons visualized in the wave mechanical model of the atom?

28. What is the fundamental equation for the the wave mechanical or quantum mechanical model of electrons in atoms? What is the significance of ψ in this equation?

29. Why are orbitals not simply called “orbits” in the quantum mechanical model of atoms?

30. Name and define the four quantum numbers used to describe electrons in atoms.

31. What does the Pauli exclusion principle say about electrons in atoms?

Section 3.13. Energy Levels of Atomic Orbitals

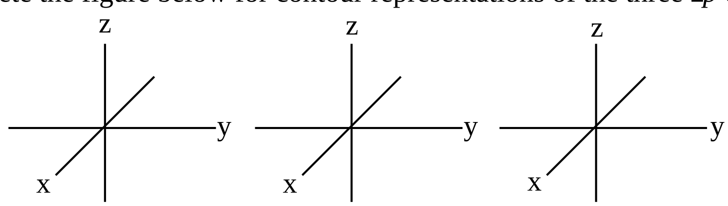
32. Use the notation employed for electron configurations to denote the order in which electrons are placed in orbitals through the 4*f* orbitals.

33. What is a statement and significance of Hund’s Rule of Maximum Multiplicity?

Section 3.14. Shapes of Atomic Orbitals

34. Discuss shapes and sizes of *s* orbitals with increasing principal quantum number.

35. Complete the figure below for contour representations of the three 2*p* orbitals.



$2p_x$ $2p_y$ $2p_z$

3.15. Electron Configuration

36. An atom of nitrogen has an electron configuration of $1s^22s^22p^3$. Give the electron con-figuration of an atom with 5 more electrons and also give the orbital diagram of such an atom.

3.16. Electrons in the First 20 Elements

37. Give the electron configurations of carbon, neon, aluminum, phosphorus and calcium.

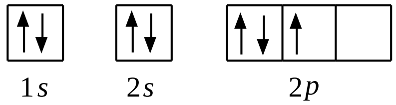
38. Give the symbols of the elements with the following electron configurations:

(a) $1s^22s^3$
(b) $1s^22s^22p^63s^1$
(c) $1s^22s^22p^63s^23p^4$
(d) $1s^22s^22p^63s^23p^64s^1$.

39. What is the core of an atom? Specifically, what is the neon core?

40. What are valence electrons? What are the electron configurations of the valence electrons in aluminum?

41. What is wrong with the orbital diagram below?



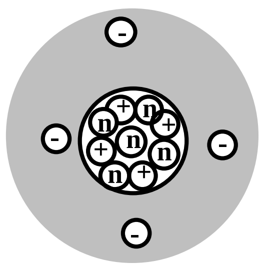
Section 3.17. Electron Configurations and the Periodic Table

42. State the rule that gives which orbitals are filled for transition elements in relation to the period number of the elements involved. Do the same for the inner transition elements.

43. Explain why the electron configuration of Cr is {Ar}$4s^1 3d^5$ and why that of Cu, is {Ar}$4s^1 3d^{10}$.

General Questions

44. The isotope,



could be designated as

A $^{7}_{3}$Li, B. $^{9}_{4}$Be, C. $^{11}_{5}$B, D. $^{13}_{6}$C, E. $^{20}_{10}$Ne

45. Of the following, the statement that is part of Dalton's atomic theory is

A. Atoms of different elements are likely to have identical chemical properties.

B. Two or more different elements combine in definite, constant ratios to form chemical compounds.

C. Atoms of one element may be converted to atoms of other elements during chemical reactions.

D. The total mass of the products of chemical reactions may be different from the total mass of the reactants.

E. Each element is composed of diatomic molecules, such as O_2.

46. Four of the following elements share a common characteristic insofar as they related to the periodic table. The one that does not share this characteristic is

A. Neon B. Carbon C. Calcium D. Oxygen E. Nitrogen

47. Of the following, the statement that is **untrue** pertaining to atoms is

A. About half of the volume is contained in the nucleus.
B. Essentially all of the mass is in the nucleus.
C. Essentially all the volume is composed of a cloud of electrons.
D. An atom of $^{1}_{1}H$ has no neutrons.
E. A neutral atom has equal numbers of protons and electrons.

List of the Elements[1]

Atomic number	Name	Symbol	Atomic mass	Atomic number	Name	Symbol	Atomic mass
1	Hydrogen	H	1.00794	56	Barium	Ba	137.327
2	Helium	He	4.0026	57	Lanthanum	La	138.9055
3	Lithium	Li	6.941	58	Cerium	Ce	140.115
4	Beryllium	Be	9.01218	59	Praseodymium	Pr	140.9077
5	Boron	B	10.811	60	Neodymium	Nd	144.24
6	Carbon	C	12.011	61	Promethium	Pm	145
7	Nitrogen	N	14.0067	62	Samarium	Sm	150.36
8	Oxygen	O	15.9994	63	Europium	Eu	151.965
9	Fluorine	F	18.9984	64	Gadolinium	Gd	157.25
10	Neon	Ne	20.1797	65	Terbium	Tb	158.925
11	Sodium	Na	22.9898	66	Dysprosium	Dy	162.50
12	Magnesium	Mg	24.305	67	Holmium	Ho	164.9303
13	Aluminum	Al	26.98154	68	Erbium	Er	167.26
14	Silicon	Si	28.0855	69	Thulium	Tm	168.9342
15	Phosphorus	P	30.9738	70	Ytterbium	Yb	173.04
16	Sulfur	S	32.066	71	Lutetium	Lu	174.967
17	Chlorine	Cl	35.4527	72	Hafnium	Hf	178.49
18	Argon	Ar	39.948	73	Tantalum	Ta	180.9497
19	Potassium	K	39.0983	74	Tungsten	W	183.85
20	Calcium	Ca	40.078	75	Rhenium	Re	186.207
21	Scandium	Sc	44.9559	76	Osmium	Os	190.2
22	Titanium	Ti	47.88	77	Iridium	Ir	192.22
23	Vanadium	V	50.9415	78	Platinum	Pt	195.08
24	Chromium	Cr	51.9961	79	Gold	Au	196.9665
25	Manganese	Mn	54.938	80	Mercury	Hg	200.59
26	Iron	Fe	55.847	81	Thallium	Tl	204.383
27	Cobalt	Co	58.9332	82	Lead	Pb	207.2
28	Nickel	Ni	58.6934	83	Bismuth	Bi	208.98
29	Copper	Cu	63.546	84	Polonium	Po	209
30	Zinc	Zn	65.39	85	Astatine	At	210
31	Gallium	Ga	69.723	86	Radon	Rn	222
32	Germanium	Ge	72.61	87	Francium	Fr	223
33	Arsenic	As	74.9216	88	Radium	Ra	226.0254
34	Selenium	Se	78.96	89	Actinium	Ac	227.0278
35	Bromine	Br	79.904	90	Thorium	Th	232.038
36	Krypton	Kr	83.8	91	Protactinium	Pa	231.0359
37	Rubidium	Rb	85.4678	92	Uranium	U	238.0289
38	Strontium	Sr	87.62	93	Neptunium	Np	237.048
39	Yttrum	Y	88.9056	94	Plutonium	Pu	244
40	Zirconium	Zr	91.224	95	Americium	Am	243
41	Niobium	Nb	92.9064	96	Curium	Cm	247
42	Molybdenum	Mo	95.94	97	Berkelium	Bk	247
43	Technetium	Tc	98	98	Californium	Cf	251
44	Ruthenium	Ru	101.07	99	Einsteinium	Es	252
45	Rhodium	Rh	102.9055	100	Fermium	Fm	257.1
46	Palladium	Pd	106.42	101	Mendelevium	Md	258.1
47	Silver	Ag	107.8682	102	Nobelium	No	255
48	Cadmium	Cd	112.411	103	Lawrencium	Lr	260
49	Indium	In	114.82	104	Ruthefordium	Rf	261.11
50	Tin	Sn	118.710	105	Dubnium	Db	262.11
51	Antimony	Sb	121.757	106	Seaborgium	Sg	263.12
52	Tellurium	Te	127.60	107	Bohrium	Bh	262.12
53	Iodine	I	126.9045	108	Hassium	Hs	265
54	Xenon	Xe	131.29	109	Meitnerium	Mt	266
55	Cesium	Cs	132.9054				

[1] Elements above atomic number 92 have been made artificially.