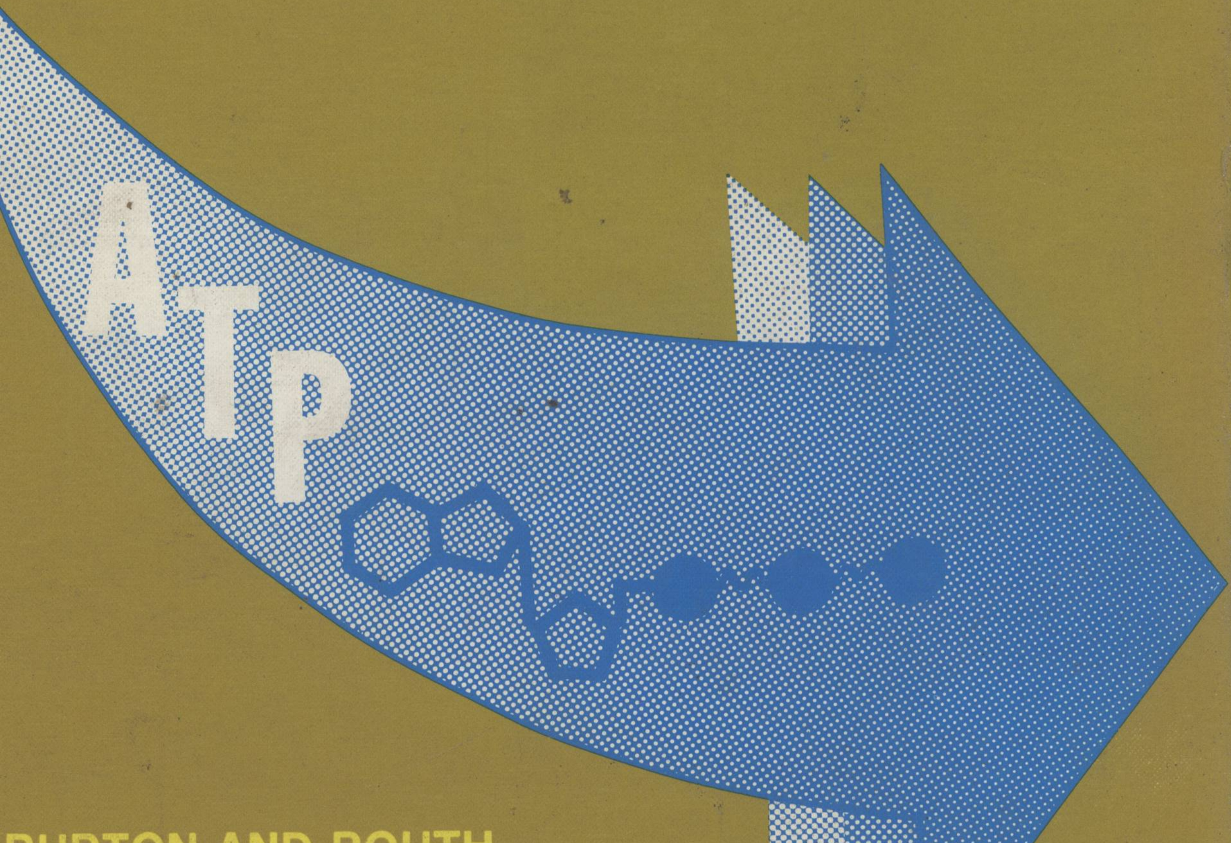
Essentials of Organic and Biochemistry



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PREFACE

Scientific knowledge in the past 10 to 15 years has grown so rapidly that it has been exceedingly difficult to keep science courses abreast of the times. Prior to this period of rapid development, the beginning college course in chemistry, for example, consisted mainly of inorganic and theoretical chemistry with an introduction to qualitative analysis. So many interesting events took place in the fields of organic chemistry and biochemistry that an attempt was made to integrate the principles of inorganic, organic, and biochemistry in beginning college chemistry courses. Several textbooks, including *Essentials of General, Organic and Biochemistry* by the present authors and Darrell Eyman, were written for students not planning to become chemistry majors. Several courses covering material in these three fields were designed for students in nursing and allied health fields, as well as for those interested in a general survey of chemistry as part of a broad background in preparation for other professions.

More recently, several textbooks have appeared that emphasize the chemistry of the environment and the multiple problems of pollution. The development of science and the occurrence of textbooks of this type, coupled with the increased popularity of science articles in the news, has stimulated a quest for scientific knowledge in the present generation of students. The healthy growth of student participation in college and university committees and administration has resulted in the demand for more realistic and relevant chemistry courses in the curriculum.

The trend at present, therefore, is toward courses that meet the needs of students, whether they are in liberal arts, medical technology, nursing, or other health-related fields. A major movement in this direction is a demand for a one semester course in general chemistry followed by a one semester course in organic and biochemistry.

The present text is designed to introduce the principles of organic and biochemistry to students who have had one or more semesters of general chemistry. The authors are convinced that it is no longer sufficient to present only a general description of the fundamentals of organic and biochemistry. They have, therefore, attempted to include an explanation of why chemical reactions follow certain pathways and to explore the mechanisms of important reactions in these two fields of chemistry. In the first section, the chemistry of organic functional groups and the chemical mechanisms of the reactions they undergo are emphasized. The chemical nature and properties of insecticides, pheromones, plastics, natural polymers, and compounds of biochemical interest are introduced in this section. This material is followed by a brief but fairly complete section on biochemistry. In addition to an explanation of the chemistry and metabolism of the major foodstuffs, this section contains a description of the cell, its enzymes and coenzymes, and energy-releasing mechanisms. The chemistry of heredity and DNA and RNA and the biochemistry of drugs and drug action are the subjects of other chapters. A laboratory manual *Experiments in Organic and Biochemistry* has also been prepared to accompany the text.

The authors are indebted to the many students they have had the opportunity to teach in beginning chemistry courses. They would also like to thank their colleagues, especially Darrell Eyman, with whom they shared the teaching of these students, for valuable suggestions and criticisms. Finally, the authors are most grateful to their publisher for thoughtful advice and continued interest in the preparation of the manuscript for publication.

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REVIEW OF FUNDAMENTAL CONCEPTS

CHAPTER 1

The *objectives* of this chapter are to enable the student to:

1. Understand the basic concepts of atomic and molecular orbitals.
2. Write the electronic configurations and Lewis symbols for the representative elements and their ions.
3. Explain the formation of an ionic compound and the nature of ionic bonds.
4. Explain the formation of covalent bonds and how the forces of interaction arise in a covalent bond.
5. Recognize the trends of electronegativity in the periodic table and their usefulness in predicting ionic and polar covalent bonds.
6. Understand atomic orbital hybridization and its relationship to molecular geometry.
7. Draw contributing forms of the resonance hybrid of molecules not adequately described by a single Lewis formula.
8. Define hydrogen bonding and explain how hydrogen bonding imparts some abnormal properties to molecules.
9. Define solute, solvent, solution, and electrolyte.
10. Understand the concept of K_w, K_a, and K_b.
11. Calculate pH, pOH, and pK of an aqueous solution.
12. Describe a titration and understand the significance of a titration curve.
13. Calculate $[H^+]$ and $[OH^-]$ from K_w, pH, or pOH data.
14. Understand the significance of a buffer solution and its role in biochemistry.

BONDING

Two types of bonds are important in chemistry: the ionic bond and the covalent bond. The best way to understand ionic and covalent bonds is in terms of **orbitals.** An **atomic orbital** can be defined as a region in space centered around the nucleus of the atom. The orbital can be thought of as a surface that corresponds to the area in which the electrons in that orbital are most likely to be found. Each atomic orbital may contain zero, one, or two electrons, but not more than two. All orbitals are not identical—they differ in shape, size, and energy.

The shapes of the atomic orbitals of hydrogen are shown in Figure 1–1. Note that

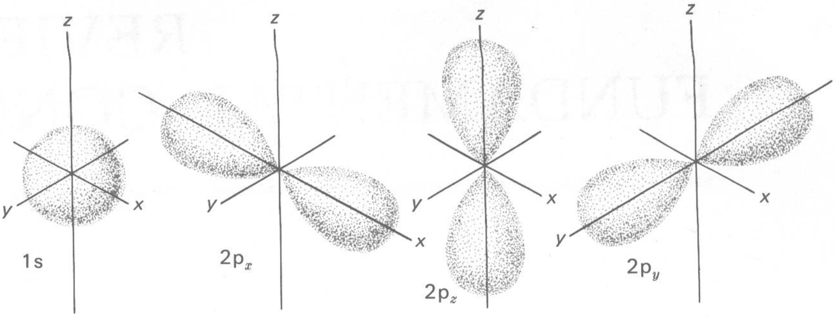


FIGURE 1-1 The shapes of the 1s- and 2p-atomic orbitals of hydrogen.

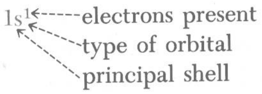
the **s-orbital** is spherical in shape, indicating that the electrons in this orbital may be in any orientation with respect to the nucleus, whereas the **p-orbitals** consist of two lobes of electron density on either side of the nucleus and are directional in shape.

The electrons of an atom are not arbitrarily assigned to any orbital, but are arranged around the nucleus in spherical shells of increasing radius. The lowest energy shell, given the designation 1, is called the K shell. The next highest energy shell, designated 2, is called the L shell; the third shell is M; and the fourth shell is N. Within any given shell (K, L, M, N) the electrons occupy certain orbitals, called s, p, d, and f, which are indicated in order of increasing energy. Therefore, a 2s-orbital is lower in energy than a 2p-orbital, a 3p-orbital is lower in energy than a 3d-orbital, and so on. The energy levels and sublevels and their relative energies are shown in Figure 1–2.

As shown in Figure 1–2, not every shell contains s, p, d, or f orbitals. For example, the K shell may contain only 1s electrons, the L shell only 2s and 2p electrons, the M shell only 3s, 3p, and 3d electrons, and so on. Note also that in the M and N shells the orbitals overlap and the 4s-orbitals are lower in energy than the 3d-orbitals. In addition, since the orbitals become larger the further away they are from the nucleus, the maximum number of electrons that are theoretically possible in any shell increases as depicted in Figure 1–2.

In order to correctly assign the schematic tabulation of the electrons in any given atom (called its electronic configuration), two restrictions must be kept in mind. First, the **Pauli exclusion principle** states that any orbital may contain no more than two electrons and these electrons must have opposite spins (indicated by ↑ and ↓). Secondly, **Hund's Rule** states that a second electron does not occupy an orbital until each orbital of equivalent energy contains at least one electron. Consequently, as a corollary to this rule, we can state that electrons occupy an orbital only if **all** orbitals of lower energy are filled. Therefore, an electron will occupy a 1s-orbital in preference to a 2s-orbital, a 2s-orbital in preference to a 2p-orbital, and so on.

A shorthand method of expressing electronic configuration is by indicating the shell number and the number of electrons in each type of orbital as shown below:



The notation above is for the hydrogen atom which has only one electron in an s-orbital in the K shell. The electronic configuration for helium which has two electrons is $1s^2$. Carbon which has six electrons has the electronic configuration $1s^22s^22p^2$. The electronic configurations for the first ten elements are shown in Table 1–1 Note that in the electronic configuration notation, the sum of the right superscripts always equals the atomic number

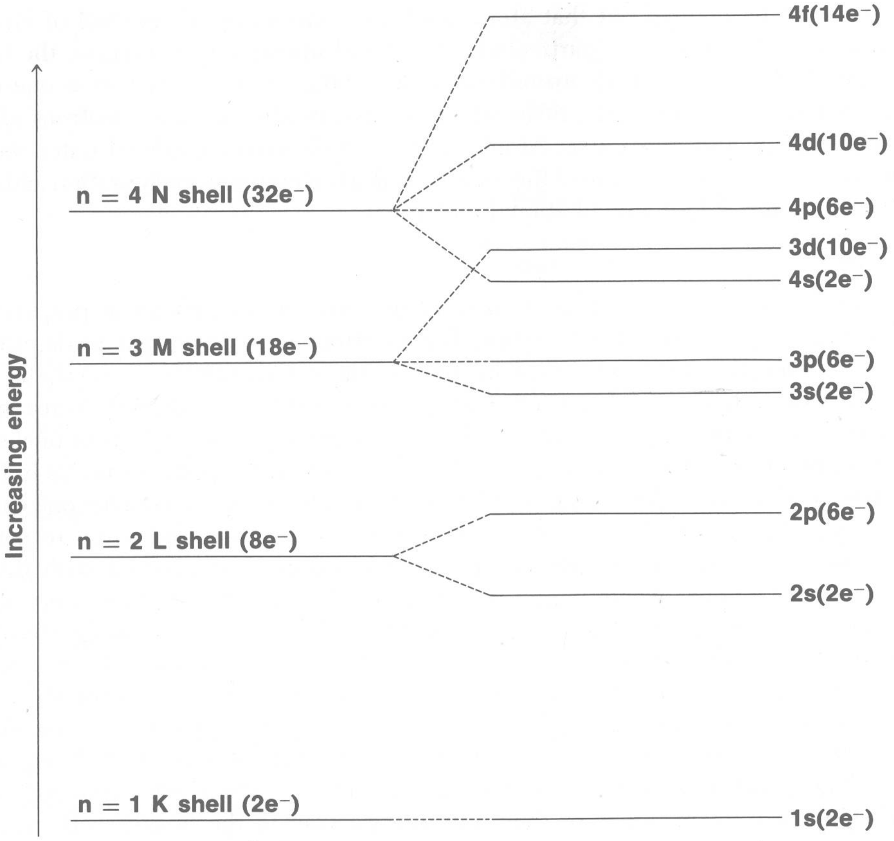


FIGURE 1-2 Energy levels and sublevels in the first four quantum levels or shells.

of the atoms. For ions the electronic configuration is indicated by simply increasing or decreasing the sum of the right superscripts to indicate the number of electrons added or lost as shown below:

$$\mathrm{Li}^{+} = 1s^2$$
$$\mathrm{F}^{-} = 1s^2 2s^2 2p^6$$

TABLE 1-1 ELECTRONIC CONFIGURATIONS OF THE FIRST TEN ELEMENTS

	1s	2s	2p		
H	↑				
He	↑↓				
Li	↑↓	↑			
Be	↑↓	↑↓			
B	↑↓	↑↓	↑		
C	↑↓	↑↓	↑	↑	
N	↑↓	↑↓	↑	↑	↑
O	↑↓	↑↓	↑↓	↑	↑
F	↑↓	↑↓	↑↓	↑↓	↑
Ne	↑↓	↑↓	↑↓	↑↓	↑↓

It should be pointed out that atoms (and also ions) whose outer shell of electrons has reached a maximum are particularly stable and unreactive. Therefore, the helium atom in which the K shell electrons have reached the maximum of two is one of the so-called inert gases. Similarly, neon which has the maximum eight electrons allowed in the L shell is also unreactive. Atoms or ions which have completed outer shells of electrons are said to have attained the inert or rare gas electronic configuration and desire to neither gain nor lose any additional electrons.

Valence Electrons

The electronic configuration of an element determines its chemical properties, as well as many of its physical properties. The electrons most important in determining chemical properties are those found in the outermost quantum level, or shell. These electrons are called the **valence electrons.** The periodically repeated occurrence of elements with similar properties is due to the periodically repeated occurrence of elements with the same number of valence electrons. For example, in all of the elements in Period 2, starting with lithium, the first shell is filled with two electrons. Lithium has one electron in the second shell; beryllium, two; boron, three; carbon, four; and so on to neon, in which the second shell is completed with eight electrons. The element with the next highest atomic number would naturally start a new period under lithium, since its first two shells are completely filled with electrons, and it has one electron in the third shell. Following sodium, magnesium has its first two shells completely filled with electrons and two extra electrons in the third shell; aluminum has three electrons in the third shell; and so on to argon, which has the third shell filled with eight electrons. The element with the next highest atomic weight would then start a new row, or series, being placed under lithium and sodium. This is potassium, with the first two shells completely filled, with eight electrons in the third shell and one electron in the fourth shell. Since the number of electrons necessary to fill a given shell (starting with shell number one) is 2, 8, 18, and 32, it is not surprising that 2, 8, 18, or 32 elements are needed to complete a horizontal row, or period, in the periodic table. The relationships just discussed may be seen in the following tabulation:

Electrons in	*Li*	*Be*	*B*	*C*	*N*	*O*	*F*	*Ne*
First shell	2	2	2	2	2	2	2	2
Second shell	1	2	3	4	5	6	7	8

	Na	*Mg*	*Al*	*Si*	*P*	*S*	*Cl*	*Ar*
First shell	2	2	2	2	2	2	2	2
Second shell	8	8	8	8	8	8	8	8
Third shell	1	2	3	4	5	6	7	8

	K	*etc.*
First shell	2	etc.
Second shell	8	etc.
Third shell	8	etc.
Fourth shell	1	etc.

Another symbolism for indicating the electronic configuration of an atom is very helpful in predicting chemical properties. The **Lewis symbol** is written as the letter(s) denoting the element, surrounded by dots symbolizing the number of valence electrons. The letter(s) of the symbol represent the nucleus and all electrons in inner closed shells, and the dots represent the valence electrons. Examples of Lewis symbols are:

$$\mathrm{Li}\cdot \qquad \mathrm{Be}: \qquad \cdot\mathrm{B}: \qquad :\dot{\mathrm{C}}\cdot \qquad :\underset{\cdot}{\dot{\mathrm{N}}}\cdot \qquad :\underset{\cdot}{\ddot{\mathrm{O}}}\cdot \qquad :\underset{\cdot\cdot}{\ddot{\mathrm{F}}}\cdot \qquad :\underset{\cdot\cdot}{\ddot{\mathrm{Ne}}}:$$

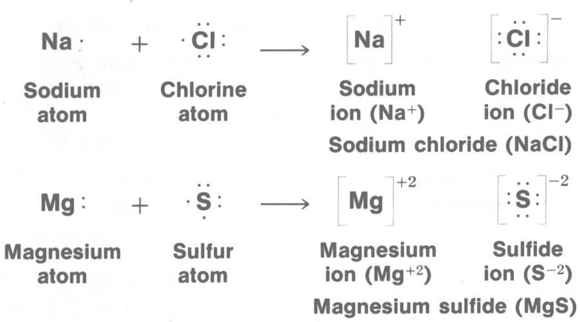


FIGURE 1-3 The process of electron transfer commonly occurs in the formation of inorganic salts.

The Lewis symbols for elements in one vertical column, or group, in the periodic table differ only in letters, since each has the same number of valence electrons. Thus, the Lewis symbols for nitrogen and phosphorus, both in Group VA, are $:\underset{\cdot}{\mathrm{N}}\cdot$ and $:\dot{\underset{\cdot}{\mathrm{P}}}\cdot$.

IONIC BONDS

The nature of ionic bonding may be illustrated by considering the formation of sodium chloride. In the overall reaction, it appears that sodium atoms react with chlorine atoms to form the compound sodium chloride. As seen in Figure 1–3, however, an electron is transferred from the outer shell of the sodium atom to the outer shell of the chlorine atom. This loss of an electron from sodium occurs readily and requires a relatively small amount of energy, which is called the ionization potential. When this electron is presented to the chlorine atom, the chlorine atom readily accepts it, with the release of energy, which is called the electron affinity. The sodium ions and the chloride ions that are formed in the process are much more stable than the atoms, and resemble the inert gases neon and argon, respectively, in their valence electron configuration. The ions differ from the inert gas atoms in that they are no longer neutral but bear positive (Na) and negative (Cl) charges. These opposite charges attract each other and are responsible for the strong ionic bond between sodium ions and chloride ions in the compound sodium chloride. Compounds in which the atoms are held together by ionic bonds are called ionic compounds.

A compound formed by the transfer of two electrons to yield an outer shell of eight valence electrons in each ion is magnesium sulfide. A relatively small amount of energy is required to remove the two outer electrons of magnesium. They are accepted by the outer shell of the sulfur atom with the release of energy. This electron transfer is illustrated in Figure 1–3. The ionic bond between magnesium and sulfide ions results from the strong electrostatic attraction between the oppositely charged ions.

In simple chemical reactions, atoms are commonly represented by the symbol alone. However, ions alone, or in compounds, are represented as the symbol bearing either positive or negative charges. For example, note ions such as Na^+ and Mg^{++} and compounds such as Na^+Cl^- and $Mg^{++}S^{--}$.

The number of electrons transferred in forming an ion corresponds to the combining capacity of an atom. The bonding resulting from electron transfer is called electrovalence, or ionic bonding. It can readily be seen that the combining capacity is equal to the number of electrons gained or lost by an atom when it is converted into an ion.

COVALENT BONDS

About 1916, it was suggested that two atoms may combine by sharing valence electrons. The process of joining atoms to form molecules by the sharing of electrons

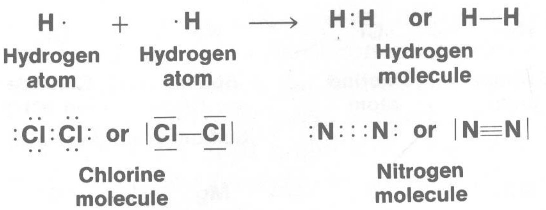


FIGURE 1–4 The sharing of electrons, or covalence, illustrated by hydrogen, chlorine, and nitrogen molecules.

is called covalence. As an example, hydrogen gas consists of molecules that contain two hydrogen atoms held together by a force resulting from the sharing of a pair of electrons, as shown in Figure 1–4. This figure illustrates the use of Lewis symbols for molecules with covalent bonds. In the symbol for the molecule, a pair of dots placed between the atoms represents a bond.

Hydrogen chlorine, nitrogen, and other diatomic gaseous elements show similar behavior in sharing electrons to the extent that their valence shells are filled. In the chlorine molecule, each chlorine atom with seven valence electrons is able to attain a complete shell by sharing one pair of electrons with the other chlorine atom. In the nitrogen molecule, a filled shell is attained only if three pairs of electrons are shared. These examples are represented schematically in Figure 1–4 with Lewis formulas.

It is stated that the hydrogen and chlorine molecules are held together by a single bond, but that the nitrogen molecule has a triple bond. Double bonds are also known to exist in many molecules.

The rationalization of the existence of a force between atoms which share a pair of electrons is aided by the following hypothetical experiment. In this experiment, diagrammatically represented in Figure 1–5 a, two hydrogen atoms are allowed to approach one another until their 1s-atomic orbitals overlap and mutually occupy the space between the nuclei. The result of the overlap is an **electron cloud** associated with the molecule. An electron cloud associated with more than one nucleus is referred to as a molecular orbital. In the molecular orbital formed here by the combination of atomic

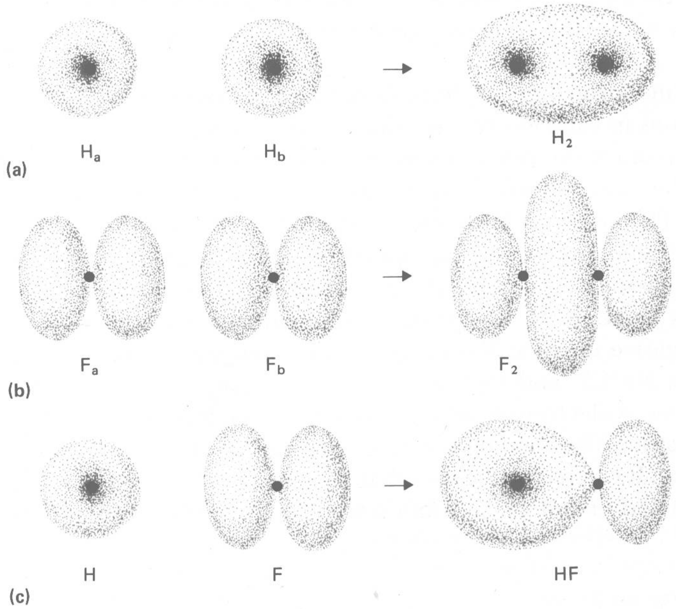


FIGURE 1–5 (*a*) Hypothetical approach of the electron clouds of two hydrogen atoms and the resulting electron cloud of the hydrogen molecule. (*b*) A similar approach of the p-orbitals of two fluorine atoms. (*c*) The approach of a 1s-orbital of a hydrogen and a 2p-orbital of a fluorine.

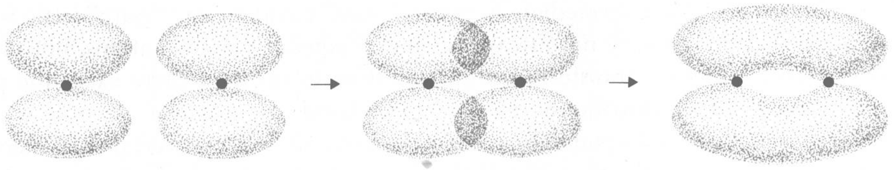


Figure 1-6 Side-by-side overlap of atomic p-orbitals to give a pi-bond.

orbitals, there is a high probability of finding the electrons in the region between the nuclei.

Before the atoms approach one another, the only forces which exist are between the electron and the nucleus of a given atom. When the atoms are brought close together, four new forces arise. Two of these tend to destabilize the molecule relative to the separated atoms. These are the repulsions of the nuclei for one another and the repulsions of the electrons for one another. These repulsions are expected, since it is known that similarly charged objects repel one another. One of the attractive forces arises because the electron originating from atom H_a is attracted to the nucleus of atom H_b. The same force holds for the electron on atom H_b interacting with the nucleus of atom H_a. The second attractive force arises from the interaction of the opposed small magnetic fields associated with the oppositely spinning electrons. The summation of the contributions of these four forces determines the total force of interaction between the two atoms and, consequently, the net stabilization of the molecule relative to the separated atoms.

The results of a similar experiment involving the approach of fluorine atoms is shown in Figure 1–5 b, where it is seen that atomic 2p-orbitals on the approaching atoms overlap to form a region between the nuclei where the probability of finding electrons is quite high. Figure 1–5 c shows a third type of experiment in which a 1s-orbital on hydrogen overlaps a 2p-orbital on fluorine to form a covalent bond.

In each of the three hypothetical experiments shown in Figure 1–5, the overlap of the electron clouds of approaching atoms results in an increased electron density along the internuclear axis. A covalent bond, resulting from increased electron density along the axis connecting the nuclei of adjacent atoms, is called a **sigma-bond** (σ-bond).

A covalent bond between adjacent atoms can also result from the overlap of p-orbitals in a side-by-side fashion, as shown in Figure 1–6. Overlap of electron clouds in this manner results in increased electron density above and below the internuclear axis. A bond of this type is called a **pi-bond** (π-bond).

POLAR BONDS

"Pure" covalent bonds, as are found in homonuclear diatomic molecules, and "pure" ionic bonds, as are found in salts like NaCl, represent the extremes of bonding. In the former, the electron pair is shared equally between atoms (Fig. 1–7), whereas in the latter there is no sharing of electrons, but rather a complete transfer of an electron from one

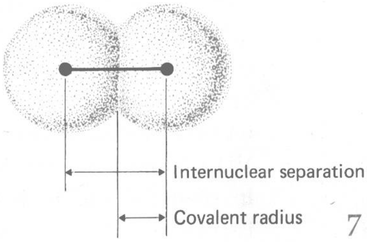
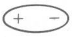


Figure 1-7 Covalent radius of a homonuclear diatomic molecule.

atom to another. Bonding intermediate between "pure" covalent and "pure" ionic is very common. Since electrons may not always be shared equally between atoms, but yet are not necessarily completely transferred from one atom to another, there exists the possibility of unequal charge distribution in a covalent bond.

For example, in the compound hydrogen chloride, HCl, the shared pair of electrons is attracted more by the chlorine end of the molecule than by the hydrogen end. The result is an unequal charge distribution, with the chlorine end more negative and the hydrogen end of the molecule more positive. Hydrogen chloride may be represented as follows:

HCl or $\mathrm{H}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Cl}}}:$ or $\mathrm{H}^+:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{Cl}}}:^-$ or 

Even though the molecule is electrically neutral, the center of the positive charge does not coincide with the center of the negative charge. The molecule is called a dipole and is said to possess a dipole moment. Dipole moments are often designated by the symbol ↦. In this symbol, the pointed end corresponds to the negative end of the dipole; the length is related to the magnitude; and the orientation of the dipole is indicated by the orientation of the symbol. When placed in an electrical field, such molecules will line up with their negative ends facing the positive plate (electrode) and their positive ends facing the negative plate (electrode). The dipole character of these compounds gives rise to the names polar bonds and polar covalent compounds.

Compounds whose atoms equally share a pair of electrons will have the center of their positive charge coinciding with the center of their negative charge. These compounds do not exhibit dipole characteristics and are called nonpolar molecules with nonpolar bonds. Hydrogen molecules (H_2) and chlorine molecules (Cl_2) are examples of nonpolar covalent compounds. In general, if a molecule is composed of two of the same kind of atoms, the bond between them will be nonpolar and the molecule will be nonpolar. If two different atoms make up the molecule, the bond is polar and the molecule is polar.

ELECTRONEGATIVITY

It has already been stated that in the compound hydrogen chloride the chlorine atom has a greater attraction for the electron pair than does the hydrogen atom. The attraction of an atom for shared electrons depends on the amount of energy required in the transfer of electrons from or to its outer electron shell. The attraction for valence electrons varies from element to element and is called its electronegativity. The relative electronegativity values that have been determined for some of the common elements are represented in a partial periodic table as shown in Table 1–2. These values are related to the ability of the atoms to attract shared electrons and thus increase the negative charge of their end of a molecule.

The **electronegativity** of an element is related to its ionization potential and its electron affinity. Elements which have high ionization potentials and high electron affinities exhibit high electronegativities. Inspection of Table 1–2 will reveal that elements with the highest electronegativities are found in the upper right corner of the periodic table. There is also a tendency toward increasing values across a period to the right and toward decreasing values moving down the elements of a group. If two elements with greatly different electronegativities combine, the bond will be highly polar or ionic in nature. Metals in Groups IA and IIA combining with the nonmetals in Groups VIA and VIIA almost always form ionic compounds with ionic bonds. When the combination involves two elements with similar values, the bonds are usually covalent. If the compound is covalent, the atoms with the greatest difference in electronegativity will form the more