Preface
A personal note from the authors to the student

Most biology textbooks are written as if an understanding of their subject were not affected by the existence or nonexistence of God. We hope that your study of *Designs in the Living World* will demonstrate that God's presence or absence has numerous ramifications in life science. For example, if he does not exist, then this ought to become rather obvious when we evaluate the chances that life would arise and perpetuate itself without him. But if God is, and if we proceed to study biology without referring to him, we would come away with the wrong perspective of the living world. We would also participate in a subtle form of plagiarism by failing to give God credit for his creative work.

Although most of biology deals with how living things *are* today rather than how they *were* many years ago, the present can only be fully appreciated in light of the past. It is the usual practice of authors to write from their own viewpoint of both the past and present, and we will do likewise. Believing that God is responsible for both the origin and the history of life, we will consistently acknowledge him throughout the book. That is not meant to imply that we will ignore other positions. Although there are a number of controversies in biology, none has more widespread implications than that of origins. So it is important for you to understand a variety of viewpoints, no matter which you personally hold. Many authors of biology textbooks describe the origin of life and its many forms as if evolution were the only scientific scenario available to thinking people. They treat all other origins concepts by ignoring them, dealing with them only in caricature, or by ridiculing them. Believing that such bias truncates the reader's understanding of a broad and important subject, we organize origins data here in the context of several contrasting systems of thought, including the view which we personally consider correct: young-earth creation.

You cannot understand the subject of origins without examining religious and philosophical roots, as is done here. In this book we also periodically probe the amazing zone of overlap between science and the Bible. *Designs in the Living World* is a science textbook, however, not a treatise on theology or Biblical apologetics.

Another implication of accepting the Bible as God's message to humanity is the responsibility it places on us as the caretakers of the living creation. Speaking of his new human creation, God said, "…let them rule over the fish of the sea and the birds of the air, over the livestock, over all the earth, and over all the creatures that move along the ground." (Genesis 1:26) Good rulers accept the obligation to take care of their dominions, and to do this we need to understand how our activities affect the rest of the creation. Regretfully, environmental concerns, as is the case with origins, become entangled with other beliefs, and it is often hard to distinguish truth from propaganda. We will try to provide you with information that will help you be a better "ruler."

Most biology texts are too large. In order to avoid excessive length, we have limited coverage here to a thorough exposition of those themes which are foundational to all life science. From here you can, if you wish, move into more specialized works on botany, zoology, microbiology, genetics, and the history of life.

Most textbooks in biology are also profusely illustrated in full color. This makes for a very attractive publication, but unfortunately it elevates the price to a level that strains the finances of those for whom the book is intended. Another problem with the practice of using large amounts of illustrative material is that it caters to a pictorial emphasis in education which concomitantly decreases the reader's opportunity to sharpen verbal skills. In order to limit costs and to emphasize the printed word, we have kept the number of illustrations to a minimum. You may use this book together with an inexpensive atlas of photographs such as Van De Graaff & Crawley's *A Photographic Atlas for the Biology Laboratory*, Third Edition (Morton Publishing Company). Page number references to this book are provided where appropriate in the text.

Where we had to cover chemical and other seemingly difficult subjects, we have supplied substantial background explanation to assist uninitiated readers. We have attempted to cover all subjects with enough rigor to satisfy biological enthusiasts. Yet we have also striven to enlighten and to captivate the attention of those who possess only a marginal interest in the field of life science.

Aware of at least some of our shortcomings both as humans and as biologists, we would appreciate learning your suggestions for improvement of this book, no matter whether you are a student or instructor. We are already aware of some of its weaknesses, and we would appreciate further suggestions which can be sent to the authors at 691 Martin-Griffeth Road, Hull, Georgia 30646.

Each author wishes to thank his wife, family, and friends for encouragement. Brenda Lindley-Anderson created some of the illustrations, and Dr. Wayne Frair provided valuable suggestions for improving the first edition.

We hope that you will find this book enjoyable and useful regardless of your position on origins and theology. Without wanting to offend those who disagree with us, and yet without apology to anyone, we pay ultimate tribute and thanks to the One whose constant providence makes all scientific phenomena available.

Lane P. Lester
Dennis L. Englin
George F. Howe

August, 1997

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Raw Materials for Life

Living things are made from the very same raw materials as everything else in creation: **atoms**. And living things are held together in the same way as everything else in creation: **chemical bonds**. That's not to say there's nothing special about life; there's a lot that's very special. But before we get to the special stuff, we're going to study that which is common to all of creation.

Biologists have not always been concerned or even aware of the chemical foundation of life. In the late 1700s Carolus Linnaeus developed a scheme for classifying plants and animals. This framework, still in use today, was based on Linnaeus' conviction that organisms alive today are the descendants of the first living things created. In his studies, chemistry was of very little concern. On the other hand, the biologists of today have come to rely heavily upon the sciences of mathematics, chemistry, and physics. Some of the most significant contributions to the technological advances of the last century have been due to the pooling of resources and knowledge between these various sciences. Vast areas of knowledge toward the understanding of the smaller details of living organisms have been made available by individuals trained in chemistry and physics.

In the middle 1800s Gregor Mendel crossed different varieties of peas and then used mathematics to analyze his data. With this he became the first person to recognize the basic patterns of inheritance that make it possible for organisms to show variety, but within specific limits. He had no information about events at the cellular and chemical levels that produced the results he obtained. Today detailed analyses of the behavior and chemical makeup of chromosomes show both how traits are inherited and the extent to which variations can occur.

Unifying chemical principles

Life involves a great many changes in both matter and energy, and a basic understanding of chemistry will lead us to a study of the biological processes whereby energy from food is acquired, stored, and utilized to allow us to grow and to maintain our bodily functions.

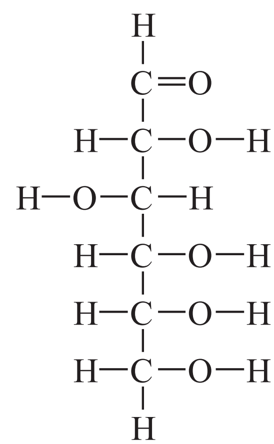
Atoms that have the same chemical behavior are considered to be the same **element**. For example, all iron atoms behave the same chemically. There are 92 naturally-occurring elements and 11 others that have been synthetically produced. Human bodies are composed of about 18.5% carbon, 65% oxygen, 9.5% hydrogen, and 3.2% nitrogen by weight. There are many other elements that are vital to life processes but which occur in extremely small concentrations.

One of the most helpful approaches to understanding science is to look for patterns. When patterns are found for the behavior of matter, reasonable predictions can be made. This predictability is one of the most remarkable evidences of design in the universe. Such patterns are even seen in the basic building blocks of all atoms—the **protons**, **neutrons**, and **electrons**. All iron atoms, for example, have 26 protons. You can be sure that any atom with 26 protons will behave like an iron atom.

Each element may be identified by its name, its symbol, and its **atomic number**. The atomic number is the number of protons possessed by all atoms of that particular element. The chemical symbols may be obvious, as in the case of H for hydrogen, or less so, as in Ag for silver. These less obvious abbreviations come from older Latin names for the elements. Chemical symbols and atomic numbers for some of the more common elements are seen here:

Hydrogen	H	1	Chlorine	Cl	17
Helium	He	2	Potassium	K	19
Carbon	C	6	Calcium	Ca	20
Nitrogen	N	7	Zinc	Zn	30
Oxygen	O	8	Silver	Ag	47
Sodium	Na	11	Gold	Au	79
Sulfur	S	16	Mercury	Hg	80

The chemical symbols used are shorthand notations for the atoms of given elements. When placed together as H_2O, the symbols form a chemical formula that represents a **molecule** which is a combination of two or more atoms. The formula H_2O stands for a molecule of water which consists of two hydrogen atoms and one oxygen atom bonded together. $C_6H_{12}O_6$ represents the glucose sugar molecule. It has 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms held together by chemical bonds. The formula tells how many of each atom are in a molecule; but it does not tell how they are arranged. The arrangement of the atoms in a molecule is given by a **structural formula**, such as the one shown here for glucose:



A structural formula is a two-dimensional picture of the arrangement of the atoms in a molecule. Add up the number of each of the atoms in the structural formula and they will match the numbers for those atoms in the chemical formula.

The shuffling of atoms between molecules in a chemical reaction is shown by a **chemical equation**. Photosynthesis is an example of a series of chemical reactions whereby green plants combine water (H_2O) and carbon dioxide (CO_2) into sugar ($C_6H_{12}O_6$) and oxygen (O_2) molecules, using the energy of sunlight. A chemical equation like the one below for photosynthesis shows how the atoms are rearranged to form new molecules

$$12H_2O + 6CO_2 \longrightarrow C_6H_{12}O_6 + 6O_2 + 6H_2O$$

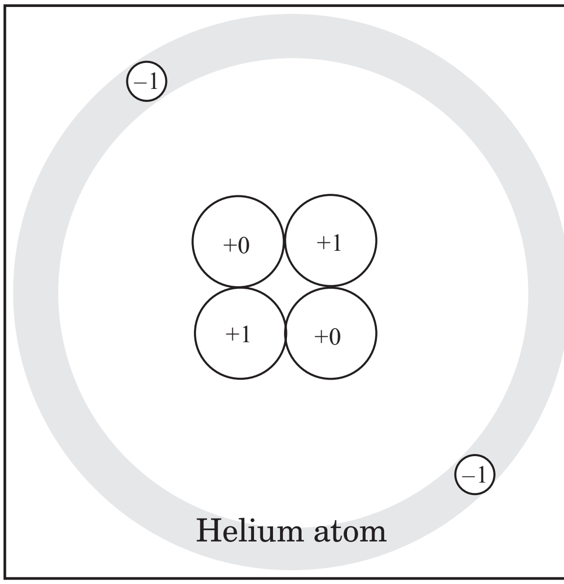
Chemical equation for photosynthesis

during a chemical reaction. Notice that the number of atoms of each element is the same on both sides of the balanced equation. Twelve molecules of water combine with six molecules of oxygen, to yield one sugar, six oxygens, and six new molecules of water. Adding up the numbers of atoms on each side shows that the numbers of atoms of each element are the same on both sides of the equation. The individual atoms have kept their original number and identity but have just been rearranged.

The principle that water molecules have two hydrogen atoms for each oxygen atom and glucose always has two hydrogen atoms and one oxygen atom for every carbon atom is called the **Law of Definite Proportions**. This is a remarkable consistency found wherever people have studied matter in the universe. The Law of Definite Proportions states that a given molecule will always have the same proportion of different atoms.

Law: A description of a relationship between conditions or events which is always true under specific conditions.

As already mentioned, atoms are composed of protons, neutrons and electrons. The protons and neutrons are found together in the **nucleus** of the atom, whereas electrons are outside the nucleus. A fundamental difference between a proton and a neutron involves their charge. Without understanding the nature of such charges, humans have long been aware of them. Thales of Miletus found in about 600 B.C. that



Helium atom

amber which had been rubbed would attract bits of straw. When a rubber comb is run through hair it will act like a magnet, picking up popcorn and bits of paper.

Benjamin Franklin observed that a rubber rod rubbed with fur attracts a glass rod rubbed with silk. He said that the rubbed rubber rod had a greater amount of **negative** charge and that the rubbed glass rod had a greater amount of **positive** charge. He used the (+) and (–) designations because they indicated opposites. This same symbolism is still used today. Opposite charges *attract* each other and like charges *repel* each other. This can be confirmed by observing that a glass rod rubbed with silk repels another glass rod rubbed with silk.

Since Benjamin Franklin's day it has been discovered that the explanation for this behavior of bulk matter lies in the charges of the protons and the electrons. Using Franklin's symbolism, a proton has a charge of +1 and an electron a charge of –1. A neutron has neither a +1 nor a –1 charge (it is neutral), but even though the neutron has zero net charge, it has essentially the same mass as a proton. An electron, on the other hand, has a mass that is negligible when compared to the mass of a proton or a neutron. Therefore, the mass of an atom is reflected in the sum of its protons and neutrons, which is called the **atomic mass**. The chemical identity of an atom, however, depends upon the number of protons in its nucleus and not its number of neutrons. The mass of an atom depends upon the total number of its protons plus neutrons.

The atomic mass and the atomic number can both be written as part of the chemical symbol. The atomic mass is put in front of and slightly above the main symbol and the atomic number is put after and slightly below the main symbol. Thus, the symbol for carbon, which has an atomic number of 6 and usually has an atomic mass of 12 can be written $^{12}C_6$. Remember that the atomic mass is the sum of the number of protons and neutrons, while the atomic number equals the number of protons.

Ordinarily this symbol for carbon is shortened to merely C, understanding that the 12 and the 6 are understood. Thus in the formula for glucose ($C_6H_{12}O_6$) the number 6 slightly below and after the C stands for the fact that there are six atoms of carbon in one glucose molecule and it does not refer to the fact that a carbon atom has six protons.

Here are some questions for you: The symbol for nitrogen can be written $^{14}N_7$. How many protons does a nitrogen atom have? How many neutrons?

The atom $^{14}C_6$ has six protons and eight neutrons (14 – 6 = 8). The atoms $^{12}C_6$ and $^{14}C_6$ have different atomic mass numbers but are nevertheless both still carbon atoms because they have the same atomic number (number of protons) and will behave the same in chemical reactions. The difference between them is that they have different numbers of neutrons (hence, different masses). Furthermore ^{14}C gives off radiation while ^{12}C is nonradioactive. Atoms such as these are **isotopes** of each other because they have the same number of protons even though their neutron numbers differ. The term “isotope” describes a relationship, not an isolated object. An atom cannot be an isotope by itself. It is an isotope in relation to another atom. It is like the fact that you are not a brother or sister by yourself but you must be a brother or a sister to someone else. Not all isotopes are radioactive; some are merely heavier or lighter than the most common isotope. Remember that when the symbol C is used by itself, $^{12}C_6$ is understood; but if we wish to refer to the heavier, radioactive isotope, we use the symbol ^{14}C for short.

Radioactivity: The spontaneous emission of subatomic particles and gamma rays from unstable atomic nuclei.

There are three isotopes of hydrogen:
$^{1}H_{1}$ (1 proton + 0 neutron): hydrogen
$^{2}H_{1}$ (1 proton + 1 neutron): deuterium
$^{3}H_{1}$ (1 proton + 2 neutrons): tritium

The three atoms above are all isotopes of hydrogen because they each have one proton but differing numbers of neutrons. Their chemical behaviors (how they react or do not react with other atoms and molecules) are similar but their nuclear behaviors are different. Hydrogen ($^{1}H_{1}$) is not radioactive. Neither is deuterium, which has twice the mass of hydrogen. Of the three, only tritium is radioactive as an emitter of beta particles, which are highly energetic electrons unattached to atoms.

Consider these three atoms:
$^{14}N_{7}$ (7 protons + 7 neutrons): nitrogen
$^{14}C_{6}$ (6 protons + 8 neutrons): carbon
$^{14}O_{8}$ (8 protons + 6 neutrons): oxygen

Although each of the above atoms has the same atomic mass number (14), they are not isotopes because they have different numbers of protons (atomic numbers).

The living world is filled with order at every possible level of observation. Our bodies are composed of organ systems with different organs serving a common purpose. Each organ consists of various tissues arranged to enable that organ to serve its role. The tissues are built from well-defined patterns of cells, the basic units of life. Individual cells are composed of reproducible systems of membranes, subcellular structures, and associated molecules. This obvious order seems clearly to be the result of a Designer's conscious planning rather than the lucky result of the collisions of molecules.

The chemical elements which make up the molecules of living cells fall into orderly relationships with each other. Elements can be divided into natural groups by the types of bonds they form and the reactions they undergo. This grouping is shown in the familiar (at least to those who've had some chemistry) periodic table. The positions of any two elements on the periodic table can be used to foresee the type of chemical union that will result between them.

To understand something about chemical bonding, it is necessary first to examine the manner in which electrons are arranged in atoms. In 1913 Niels Bohr presented an idea, or **model**, of atomic structure in which the electrons were viewed as being in discreet orbits around the nucleus of the atom like satellites in orbit around the Earth.

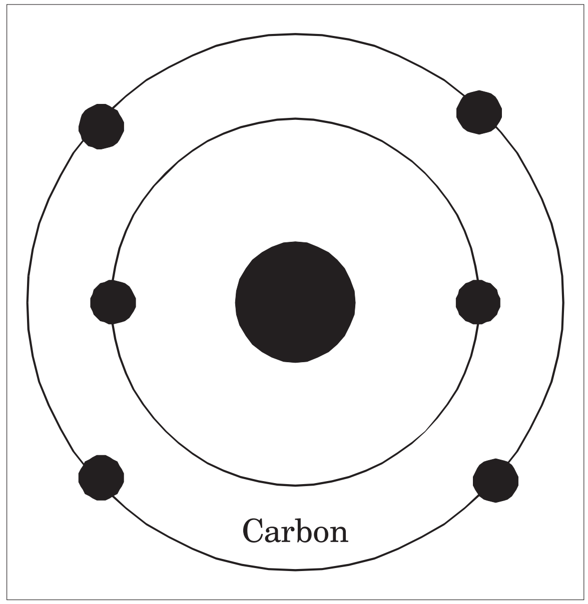
> **Model**: A description of an object or process that cannot be observed directly. A good model accounts for the available information, but it may or may not be an accurate description of the actual object or process.

Later in 1926 Ervin Schrodinger described the electrons as occupying more generalized regions around the nucleus where they had a given probability of being present, such as an 80% probability of finding a particular electron in a specific region. This later view of Schrodinger formed the basis for the current model of the atom in which the regions where electrons are expected to occur are called **orbitals**, not "orbits" as in the Bohr model. Electrons with lower energy levels are located in orbitals closer to the nucleus and those with more energy are found farther from the nucleus.

The tendency of an atom to join with another in a chemical bond depends on the number of electrons in its outermost energy level. The orbital at the first, or lowest, energy level is complete with either **two** electrons or none. All of the orbitals at a higher energy level are complete when they contain a total of **eight** electrons or zero. Atoms that don't already have a complete number of electrons have three ways of achieving a stable number of electrons: lose 'em, gain 'em, or share 'em. The chemical nature of different elements will cause them to tend to use one technique more than another.

A hydrogen atom ($^{1}H_{1}$) has only one proton and usually only one electron at the lowest energy level. As just stated, that energy level has greater stability if it has either zero or two electrons. Such hydrogen atoms will therefore join with other atoms in ways that either provide it with one more electron or get rid of the one it has.

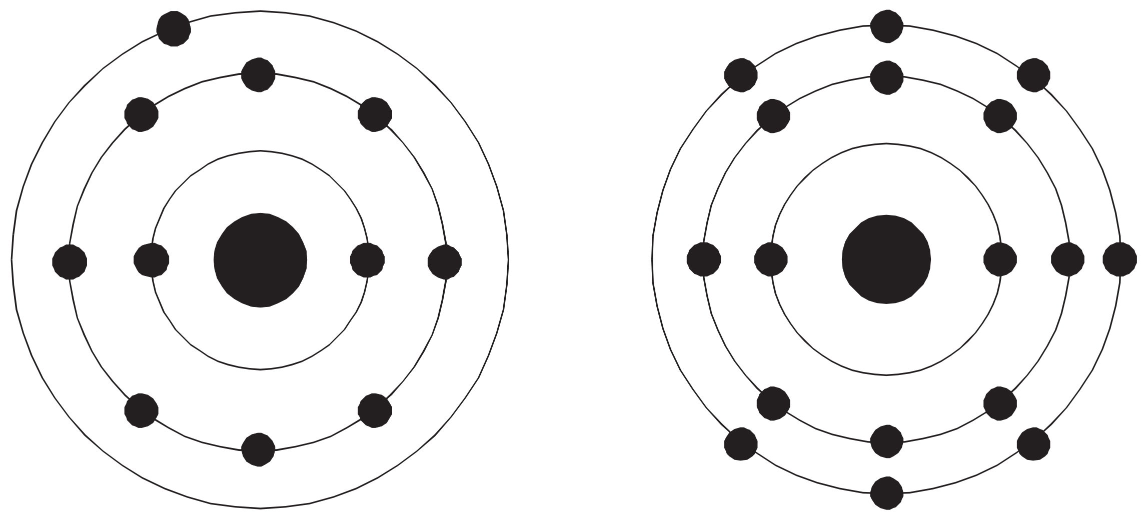
Oxygen ($^{16}O_{8}$) has eight electrons, two of which reside at the first energy level, leaving six more to go into the orbitals of the second energy level. As mentioned before, the orbitals of the second energy level can be stable with a total of eight or zero electrons. Oxygen normally will



form chemical bonds in ways that will add two more electrons, since it is "easier" to gain two than it is to lose six. This is likewise true of nitrogen ($^{14}N_{7}$) which normally gains three electrons to complete the orbitals in its second energy level at a total of eight, although sometimes nitrogen loses five electrons to bring its second level to zero which also is stable.

Carbon ($^{12}C_{6}$) has six protons and hence six electrons. Two of these normally enter and fill the lowest energy orbital which leaves four more electrons to move into orbitals in the second energy level. Since its second level has four electrons, halfway between zero and eight, carbon is a very special atom. It tends neither to gain nor lose electrons but to share four electrons with other atoms so that in effect its second energy level finally controls eight electrons.

There are three situations in which atoms "stick together," and we refer to them as the three types of chemical



bonds. Life is only possible because of the ways these three differ from each other, so it is worthwhile for us to learn something about them.

Ionic bonds

Recall that electrons have a negative electrical charge, and protons a positive charge. When an atom has more or fewer electrons than protons, the overall atom has a charge equal to the difference in the two sets of particles. Such a charged atom is referred to as an **ion**. For example, a calcium atom has 20 protons and often has only 18 electrons. Such a calcium ion has, therefore, two extra positive charges and is symbolized as Ca^{+2}.

Recall also that oppositely-charged objects are attracted to each other. When such atoms are held together this way, it is by ionic bonding. The most abundant ions on earth are Na^{+} and Cl^{-}, and in a dry state are bound to each other by ionic bonds, forming crystals of table salt. Most of these ions on earth, however, are separated in solution in the ocean. The sodium ion has a positive charge by having 10 electrons to its 11 protons. Two of the electrons are in the first orbital, leaving the complete number of eight for the orbital of its second energy level. Similarly, the chlorine ion has 18 electrons to its 17 protons, with the electrons distributed to its energy levels: 2, 8, and 8, again achieving the complete number of eight in the outermost energy level.

Ionic compounds perform critical roles in the body fluids of organisms including producing electrical potentials across membranes, generating nerve impulses, controlling fluid retention by cells, contracting muscle fibers, and initiating heart beats. Ions in solution in body fluids are called electrolytes by physiologists.

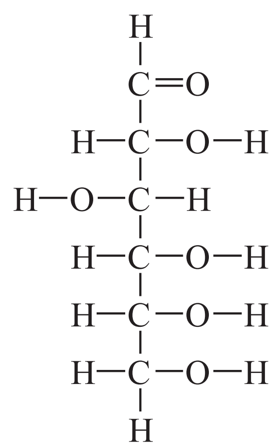
Covalent bonds

The atoms in the molecules of life are primarily held together by covalent bonds, the strongest of the three types of bonds. Here the attraction is not differences in electrical charges, but the sharing of electrons. The electrons being shared spend part of their time around one nucleus and part around the other. In this way, both atoms are able to achieve the stable numbers of zero, two, or eight electrons. The atoms of some elements are found combined into molecules held together by covalent bonds. Hydrogen gas is the simplest example, and this smallest of the atoms joins with another hydrogen atom to share a single pair of electrons. Thus, each nucleus alternates in having either zero electrons or two.

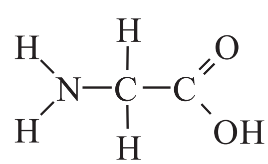
The most abundant gases in our atmosphere, nitrogen and oxygen, are likewise found as paired atoms covalently bonded. A single oxygen atom with eight protons and eight electrons needs two more to be complete (2 in the first energy level, 8 in the second). Two oxygen atoms share two pairs of electrons in order to reach the complete number for both. Nitrogen atoms, with only seven electrons, have to share three pairs of electrons in their covalent bonds in order to be complete.

The carbon atom is the foundation for all of the molecules of life, and its design provides the basis for the hundreds of thousands of different molecules of which living things are composed. Its six electrons are arranged with two at the first energy level and four at the second, causing the atom to need to share electrons at four positions in order to be complete. This flexibility allows the atom to combine in a rich variety of covalent bonds with other atoms. Interestingly, the most common arrangement is for carbon atoms to join together in chains.

Consider again the structural formula of glucose we saw early in this chapter. In a structural formula, the lines represent shared pairs of electrons, and you'll see that each carbon is sharing four pairs of electrons in order to be complete. The carbon atom at the top shares two pairs with the oxygen atom, explaining why there are only three other atoms, instead of four, around that particular atom of carbon.



The ability to interpret structural formulas will be very important in the next chapter, "The molecules of life." Let's consider a molecule that will be encountered early in the next chapter, the amino acid glycine:



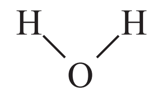
Can you analyze the relationships between the atoms in this molecule? How is each atom made complete in this molecule? One possible bit of confusion is the pair of atoms symbolized as "−OH." That's actually an abbreviated way of expressing "−O−H" in which an oxygen atom shares one pair of electrons with a hydrogen atom and another pair of electrons with some other atom. That particular pair of atoms is so common that the abbreviation is used to minimize crowding in structural formulas. Now take a moment to see if you can tell how each atom's "need" for electrons is satisfied in this molecule.

Nitrogen needs three more electrons to be complete, and you can see how that is accomplished here by sharing three pairs of electrons. The carbon at the right is getting the four electrons it needs by sharing single pairs with the central carbon and a hydrogen, plus two pair of electrons with oxygen. Notice how the "outlying" atoms are also having their "needs" met.

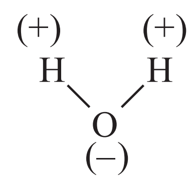
Hydrogen bonds

We move now from studying the strongest type of bond to a consideration of the weakest of the three types. Although it's true that almost all of the bonds in the molecules of life are covalent, there are some extremely important ones that are hydrogen bonds. In fact, it is the weakness of these bonds that makes them important to life, as we'll see in future chapters.

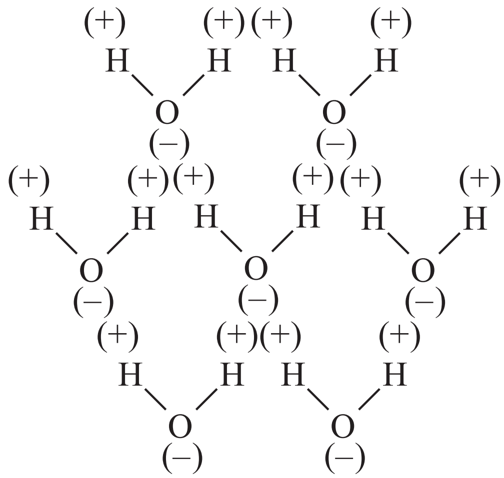
Hydrogen bonds are made possible by the fact that in covalent bonds the sharing of electrons is not always "fair." That is to say, the members of a pair of electrons may not spend an equal amount of time orbiting both atomic nuclei. The simplest example of this kind of behavior is the water molecule:



Its structural formula shows how the oxygen is made complete by sharing a pair of electrons with each of two hydrogens. The oxygen, with eight protons in its nucleus exerts a much stronger pull on the shared electrons than do the hydrogens with only one proton each. So the pairs of shared electrons spend more of the time around the oxygen nucleus than the hydrogen nuclei. This means that, part of the time, the hydrogen nuclei (single protons) are "out there" by themselves, and those positive protons cause that part of the molecule to be slightly positive. Likewise, the oxygen end of the molecule, with the negative electrons spending an extra amount of time there, is slightly negative. We could modify the formula in this way:



The charge signs are placed in parentheses to indicate these are not whole charges as are found in the ions of ionic bonds. These slightly positive and slightly negative ends of the molecules are slightly attracted to each other, and this attraction produces hydrogen bonds. The water molecules tend to arrange themselves, positive to negative, as seen here:



This arrangement is responsible for some of the properties of water, which rivals carbon in its importance to the existence of life on earth. As the U.S. space program has progressed and scientists have speculated on the possibility of life elsewhere in the universe, the majority opinion is that, if life exists anywhere else, it will also have to be based on the chemistry of carbon atoms and water molecules. That is not to say that all a planet needs to support life is to have carbon and water. The earth has many characteristics without which life could not exist here: its distance from its star (sun) and resulting temperature, its size and resulting gravity, its atmosphere with the right mixture of gases, its protective ozone layer and magnetic field which shield us from harmful radiation; the list is long.