

Chapter 2: Small Molecules: Structure and Behavior

Introduction

- Ultimately, life is an emergent property. The characteristics and properties of the atoms and then the molecules of which life is composed are fundamental to the nature of life.
- This chapter provides essential information on the characteristics and properties of important atoms and molecules, which will be useful in the study of biology - the study of life. This information will be useful later in this course as well as when studying physiology, genetics, ecology and for many other fields of biology.

I. Atom: The Constituents of Matter

- All matter is composed of atoms.
- Atoms vary in size but all are amazingly small.
- Each atom consists of at least one proton and one electron (as does a hydrogen atom, the simplest atom).
- The proton would be found in the center of the electron's orbit, in the region called the nucleus.
- The nucleus possesses a positive electrical charge relative to the electron, which would be negatively charged. Really, the charges are simply opposite and the + or - was assigned arbitrarily, but are now convention.
- Atoms have mass.
 - Mass is a measure of quantity.
 - For an atom, the mass mostly comes from the proton and a neutrally charged body called a neutron.
- A neutron if present is found in the nucleus, the same region where the proton is found.
- What is the mass in grams of a proton? 1.7×10^{-24} grams.
 - Grams are abbreviated g, and are a widely used unit of measure.
 - An electron has even less mass than 9×10^{-28} g. When adding contributors, electrons can be ignored.
- For an atom, the number of electrons equals the number of protons, and so they are neutrally charged.
- Atoms are extremely small, and are mostly empty space.

A. Elements: A Collection of One Type of Atom

- Hydrogen, the element, consists solely of hydrogen atoms.
- About 100 different elements are found in the universe.
- Information on elements is arranged in a logical order on a table called the Periodic Table. (*See 2.1, The periodic Table*)

- The Periodic Table arranges elements left to right and top to bottom, depending on their atomic number, which is also the number of protons in their atoms.
- The Periodic Table is arranged in columns based on similarities in elements' properties.
- About 98% of the mass of life forms are carbon, hydrogen, nitrogen, oxygen, phosphorus and sulfur.
- Whereas *elements* contain just one kind of atom, substances that are composed of more are called *compounds*.

B. The number of protons identifies the element.

- The atomic number is a unique number for each element.
- The atomic number is the number of protons found in an atom of an element.
- The mass number is the total number of protons plus the number of neutrons. The mass number is given 1 for each proton and 1 for each neutron. (*See Figure (top left of page 18)*)
- Mass number is used as the weight of the atom in units called Daltons.
- Each element has either a one or two character symbol: H is hydrogen, C is carbon, Na is sodium, and Fe is iron.
- Each box of the periodic table has a designation like $^{12}_6\text{C}$. The number 12 is for the mass of carbon; the number 6 is for the atomic number.

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| <ul style="list-style-type: none"> • Question: How many protons does carbon have? Answer: 6 • Question: How many neutrons does carbon commonly have? Answer: 6 (because $12 - 6 = 6$) |
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C. Isotopes differ in the number of neutrons.

- All atoms of an element have the same number of protons, but not necessarily neutrons.
- Some atoms of what would be the same element can have different numbers of neutrons.
- Atoms of the same element that have different atomic weights are called isotopes. (*See Figure 2.2*)
 - The common form of hydrogen is ^1H . One of every 6,500 hydrogen atoms on earth has a neutron, which makes it ^2H . The ^3H form has 2 neutrons and weights 3 times that of ^1H .
 - ^2H is called deuterium, ^3H is called tritium.
 - Carbon exists as ^{12}C , ^{13}C and ^{14}C in nature. Carbon's isotopes are just called carbon. Most isotopes are named according to the number of their protons.
- The mass number for a certain element on the periodic table is often not close to a whole number, because the mass number is calculated based on a natural mixture of the isotopes found in nature for that element. For example, xenon

has a mass number of 131.30, because it is the average of several different isotopes in the ratios commonly found.

- Some isotopes are radioisotopes.
 - Radioisotopes emit energy as alpha, beta and gamma radiation from some of their nuclei.
 - Radioactive decay transforms the original atom into another type, usually another element.
 - This means that a change in the number of protons occurs.
 - Uranium 238 (for its mass number), atomic number 92, becomes thorium 234, atomic number 90. This means it ends up with two fewer protons.
- It is even more interesting that carbon 14, atomic number 6, loses a particle and becomes nitrogen 14, atomic number 7. This means it somehow acquires an additional proton and has one less neutron.

D. Electron behavior determines chemical bonding.

- Electrons move at speeds approaching that of light.
 - The location of electrons cannot be determined, but the area it travels can. This region is called the orbit of the electron.
 - If the nucleus of an atom was expanded to the size of a grain of salt, the electrons would be spinning at speeds approaching that of light about as far away from the grain as the radius of the largest dome in the world. Moreover, the electrons would be far too small to see.
- Two at the most electrons can occupy each orbital.
 - Electron shells are composed of orbitals.
 - Electron shells can be viewed as energy levels.
 - K shell is innermost.
 - K has just one orbital called the *s* orbital.
 - The *s* fills first and has the lowest energy.
 - Hydrogen has one electron, helium two in an *s* orbital.
 - The L shell is next closest to the nucleus and has one *s* and three *p* orbitals. (See Figure 2.4)
 - L can accommodate eight electrons, two per orbit.
 - Usually, L is filled when 10 or more electrons are present.
 - Then come shells M, N, O, P and Q.
 - The outermost shell determines the nature of an atom's reactivity.
 - Mostly, if eight electrons are in the outer shell, which consists of four orbitals, the atom is stable and tends not to react.
 - Inert elements include helium, neon, argon, and krypton. (See Figure 2.5)
 - Other atoms, which don't have 8 in the outermost shell, share, gain or lose electrons to arrive at a stable state.
 - Sharing creates molecules from atoms.

- The tendency to form atomic relationships resulting in 8 electrons in the outermost shell is called the rule of eight or the octet rule.
- Of elements important to life, hydrogen and phosphorus are exceptions to the octet rule: hydrogen is stable with two, phosphorus, commonly with 10.

II. Chemical Bonds: Linking Atoms Together

- Bonds vary in strength; the strongest involve sharing of electrons. (*See Table 2.1*)
- The strongest types of bonds are covalent.
- Next strongest are ionic and hydrogen bonds. They have approximately 1/10 the strength of covalent.
- Covalent bonds involve sharing one or more pairs of electrons. Two hydrogen atoms can share one each to make a pair of shared electrons.

A. Covalent bonds consist of shared pairs of electrons.

- In hydrogen molecules, H_2 , the pair of electrons share a common orbital and spend equal amounts of time around each of the two nuclei. (*See Figure 2.6*)
- The nuclei stay a distance from each other due to mutual repelling positive charges.
- Shared electron orbital(s) bond together and repulsion by the nuclei push away. The balanced distance is 0.1 nm.
- Carbon has an atomic number of 6. Two electrons occupy the innermost shell. Four are found in the outer L shell.
- Carbon needs 4 more to complete its L shell. (*See Table 2.2*)
- Four hydrogen atoms can bond to generate CH_4 - methane.
- CH_4 shared orbitals form a tetrahedron. (*See Figure 2.7*)
- There are some compounds with single covalent bonds, but also some have double and some triple.
- C_2H_4 has a double-bonded pair of carbon atoms.
- N_2 , nitrogen molecules, the form of nitrogen found in the atmosphere, has a triple covalent bond ($N \equiv N$).
- Equal Versus Unequal Sharing of Electrons
 - N_2 , H_2 , O_2 and other bonds between atoms of the same element share electrons equally.
 - This is because their nuclei are the same in terms of the number of protons, the distances of electrons from the nuclei, and their electrons are attached to each atom equally.
 - Equal sharing is also the case when different atoms have about the same attractions for electrons. Carbon and hydrogen have about the same attraction.
 - Even atoms that hold electrons tightly share equally if the other atom in the molecule has the same attraction for electrons.

- When there is a difference in the attraction in a molecule and sharing is unequal, the electron spends most of its time around the atom with the greater attraction for the electron.
- The attraction of electrons is called electronegativity. (*See Table 2.3*)
- Examples of atoms that form covalent bonds that have high electronegativity would be oxygen and nitrogen.
- Oxygen has six electrons in its outermost shell and needs 2 more to fill it.
- When oxygen forms covalent bonds with atoms that have weaker electronegativities like carbon or hydrogen, the electrons are shared but unequally in terms of time.
- Unequal sharing of electrons causes a partial charge, a negative charge, around the most electronegative atom, and a partial positive charge around the less electronegative atom.
- The bond is called covalent because it involves sharing; it is *polar* covalent because the sharing is unequal. (*See Figure 2.8*)
- δ^- , delta negative, is the symbol for a partial negative charged region of a molecule.
- δ^+ is the symbol for a partial positive charged region of a molecule.
- Molecules that have polar covalent bonds are called polar molecules.
- Some molecules are so large that they are said to have polar regions and non-polar regions.
- Non-polar molecules share electrons relatively evenly between atoms, like hydrogen and carbon does. Carbon has four in its outermost shell and needs four to complete it. Hydrogen has one and needs one. Their attraction for electrons is nearly equal. Hydrocarbons are classified as non-polar.

B. Hydrogen bonds may form between molecules.

- The δ^- portion of a water molecule, the area around the oxygen, is hydrogen bonded to the δ^+ portion of another water molecule, the area around the hydrogen. (*See Figure 2.9*)
- The δ^- region around the oxygen can attract two other water molecules. This is because it has two additional electrons orbiting it most of the time, one from each of the two hydrogens.
- The each of the δ^+ regions of the hydrogen atoms in a water molecule attracts the δ^- region of one other water molecule.
- Each of these attractions is called a hydrogen bond. Hydrogen bonds differ from covalent bonds in one very important way: hydrogen bonds are not sharing electrons.
- Hydrogen bonds can occur between molecules besides water.
- For example, hydrogen bonds hold together the DNA molecules found within our cells.

- Although hydrogen bonds are weak, they tend to be additive. They are of profound biological importance.

C. Ions form bonds by electrical attraction.

- Unlike the sharing or even unequal sharing that characterizes covalent bonds, ionic bonds do not involve sharing of electrons.
- Ions are formed when an atom totally loses or gains one or more electrons. (*See Figure 2.10*)
- Ions are always symbolized by + or - superscripts. For example, sodium ion is Na^+ .
- The + symbol is for the electrical charge of the ion. A single + means the ion has one more proton than electron. A single - means the ion has one more electron than proton, for example Cl^- , the chloride ion found in common table salt.
- Positively charged ions are called cations.
- Negatively charged ions are called anions.
- Na^+ is a cation; Cl^- is an anion.
- Some ions are minus or plus more than one electron.
 - Calcium ion is Ca^{2+} and has 2 more protons than electrons.
 - The Ca^{2+} ion is a cation and is said to be divalent.
 - The aluminum ion is Al^{3+} .
 - Iron can be found either as Fe^{2+} (ferrous) or Fe^{3+} (ferric). In fact, in solution Fe^{2+} sometimes changes to Fe^{3+} and vice versa.
 - Cuprous ion is Cu^+ and cupric ion is Cu^{2+} .
- Also, common in biological systems, are complex ions. Examples are NH_4^+ (ammonium ion), SO_4^{2-} (sulfate ion) and PO_4^{2-} (phosphate ion).
- Even large biological molecules like proteins and DNA molecules have regions that have charges.
- Ionic bonds are formed by the attractions of the opposite charges.
- Table salt has chloride and sodium ions which are held together by opposite attraction.
- These attractions are strong, but when introduced into water, the partial charges of the water molecules can easily interfere with the ionic bonds and "dissolve" table salt. (*See Figure 2.11*)

D. Polar and non-polar molecules interact best among themselves.

- Like attracts like.
- Polar molecules tend to be hydrophilic (hydro means water, philic means loving). Substances that are ionic or polar often dissolve in water. This is due to interactions between the charges. To dissolve in water is for the molecules of the substance to interact and intermingle with the water molecules.

- Non-polar tend to aggregate with other non-polar molecules rather than with polar water and are called hydrophobic (phobic means hating). (*See Figure 2.12*)
 - This is an exaggeration when it comes to hydrocarbons, though.
 - Hydrocarbons tend to separate from water and tend to aggregate together.
 - If you mix mineral oil and water and then allow it to clear, some water molecules, but very few, will still be associated with the oil and vice versa.
 - Hydrocarbons are not completely non-polar.
 - Carbon has an electronegativity of 2.5 and hydrogen's is 2.1.
 - If non-polar molecules really failed to dissolve in polar solvents like water, use of oxygen would be impossible. Remember that oxygen molecules, O₂, are non-polar because they share their electrons evenly.
- Non-polar molecules are also attracted to each other via relatively weak attractions called van der Waals forces.
- These are transient interactions caused by brief variations in electron distributions, which cause an attraction to closely adjacent molecules.
- Even though each interaction is weak, the same can have important effects and van der Waal forces are biologically significant.

III. Chemical Reactions: Atoms Change Partners

- Chemical reactions occur when atoms change partners. (*See Figure 2.13*)
 - Methane, when reacted with oxygen, forms water and CO₂.
 - CH₄ + O₂ → CO₂ + H₂O is an unbalanced equation.
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| <ul style="list-style-type: none"> • Because the left side of the equation, the reactants side, has a different number of hydrogen and oxygen atoms than the right, the products side, a numerical correction must be made. • The single rule for balancing chemical equations is to balance the numbers while keeping the adjustment values as small as possible. (This is unless the actual events of the reaction are known and are different from the equation solved using this rule.) • Here, we would adjust the equation: CH₄ + 2O₂ → CO₂ + 2H₂O. • One molecule of methane plus two molecules of oxygen yields one molecule of carbon dioxide and two molecules of water. • The equation must balance because matter is neither created nor destroyed. |
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- Heat is released from the reaction described above.
 - Changes in energy usually accompany chemical reactions.
 - The heat and light released existed in the molecules of the reactants as potential energy.
 - Plants put surplus collected solar energy into molecules for example.
 - We can measure the potential energy of molecules and express it in units of heat called calories.

- A calorie is the amount of heat required to raise the temperature of one gram of pure water from 14.5 degrees C to 15.5 degrees C.
- For years, there has been an effort to change to a new unit system called Joules (J).
- One J = 0.239 cal and 1 cal = 4.184J.
- The calorie unit that is used popularly is actually 1,000 of the calorie units we use in science.

IV. Water: Structure and Properties

- Life now has been found in layers of rock, below the surface of the earth.
- In the absence of sunlight and oxygen, organisms derive energy from chemical sources.
- Life can exist without the need for sunlight or oxygen, yet the indispensable ingredient is water. Some scientists think that wherever liquid water exists for a long enough time, life might be found in our universe.

- Water has some very remarkable properties.
- A water molecule is composed of 1 oxygen and 2 hydrogen atoms (H₂O).
- The shape, polarity and ability to form hydrogen bonds give water unusual properties. (*See Figure on page 26*)
- Water is an excellent solvent; it takes a lot of heat to change its temperature compared to its weight; it has high cohesion; and, it expands when frozen

A. Water has a unique structure and special properties.

- Ice Floats
 - Ice is held in a crystalline structure by the orientation of the water molecules' hydrogen bonds. (*See Figure 2.15*)
 - Hydrogen bonds form between each molecule and four others (2 for the δ^- region and 2 for each of the 2 δ^+ regions).
 - These four hydrogen bonds actually increase the space water molecules take up compared to when the water is a liquid and has fewer, approximately 3.5, hydrogen bonds at any given time.
 - Therefore, ice floats. Few other substances expand when they change from liquid to solid phase. Glass, or silicon, is another that does.
 - The floating helps keep lakes from freezing solid.
- Melting and Freezing
 - Compared to other nonmetallic substances, ice requires a lot of heat to melt.
 - Melting 1 mole of water requires 5.9 kJ or 1410.1 calories of energy.
 - Each mole of water weighs 18 grams, so this would be 78.34 calories per gram!
 - Again, this is because of the hydrogen bonds.
 - Heat is released when ice forms.

- This property helps water moderate earth's temperature.
- Specific Heat
 - It takes a lot of heat energy to change the temperature of liquid water.
 - This is due to hydrogen bonds again. To raise the temperature, bonds must be broken more rapidly.
 - It is difficult to find any substance, metal or nonmetal, that takes more heat per gram than water to raise its temperature.
 - Specific heat is the value equal to the number of calories needed to raise one gram of a substance 1°C. The specific heat of water is 1.
 - Like melting and freezing, this property also moderates the fluctuation of earth's temperature.
- Evaporation and Cooling
 - It takes a lot of heat to change water to a gaseous state.
 - Heat is absorbed to break hydrogen bonds.
 - Evaporation has a cooling effect by absorbing calories.
 - Condensing has the opposite effect, releasing heat.
 - Evaporation and condensation (rain or snow) has the effect of transporting energy.
 - Evaporation is an important means for cooling large multicellular organisms.
 - A person can survive a temperature well in excess of their body, yet maintain normal body temperature by sweating and evaporating the sweat.
 - Evaporation occurs less readily in high humidity, which is why hot, humid areas seem hotter at the same temperature than those places that have low humidity.
- Cohesion and Surface Tension
 - Water has a cohesive strength even though it is a liquid.
 - It is the attraction of water molecules to each other that allow it to be transported from the roots to the treetops.
 - Water has high surface tension. (*See Figure 2.16*)
 - Even though water is liquid, diving into water from a significant height will cause injury due to water's surface tension.

B. Most biological substances are dissolved in water.

- A solution is a substance dissolved in a liquid.
- A solution can also be one liquid dissolved in another.
- To dissolve is to intermingle at the molecular level.
- Many reactions occur in aqueous solutions.
- Molecules that dissolve readily in water are either polar, ionic or both.
- Hydrophilic (water-loving) biological substances are found dissolved in water and many are contained in cells by the hydrophobic (water-hating) molecules that make up cell membranes.

Important Chemical Quantitative Analysis Concepts

- Molecular weight or mass is the sum of the atomic weights of all the atoms in a molecule. (*See Figure 2.17*)
 - The molecular weight of H₂ is 2.
 - The molecular weight of water is 2 (1 for each hydrogen) + 16 (for 1 oxygen) = 18.
- A solvent is the fluid used to dissolve a solute.
- Scientists work with moles and not simply mass when studying chemical reactions.
 - A mole is the amount of an ion or compound's molecular weight in grams.
 - A mole of H₂ would be 2 grams of H₂.
 - A mole of H₂O would be 18 grams of H₂O.
 - One liter of water weighs about 1000 grams; so, each liter has about 1000/18 or 55.56 moles.
 - A mole contains a known number of molecules.
 - From the known weight of a proton, it has been possible to determine how many molecules are in a mole.
 - That number is close to 6.03×10^{23} molecules per mole.
 - Every chemist, physicist and most biologists knows Avogadro's number.
 - *This number is very useful to remember.*
- A molar solution is a certain number of moles of a compound in a liter of water and that compound.
 - One mole of sodium chloride would be the atomic weight of sodium (23) plus the atomic weight of chlorine (35.5) or 58.5 in grams. When 58.5 grams of sodium chloride are dissolved in some water and then additional water is added to create a final volume of a liter, the solution is 1 molar.
 - A molar solution is abbreviated M or is symbolized by placing brackets, [], around the other symbols.
 - There are 1000 ml per liter. There are 1000 μ l per ml.
 - A 10 ml volume of a 1 M solution would contain 10/1000 or 1/100th of a mole of molecules of the solute.
 - 1mM is 1/1000 of a molar solution.
 - 1 μ M is 1/1000 of a mM.
 - One liter (l) is 1000 milliliters (ml); 1 ml is 1000 microliters (μ l); 1 μ l is 1000 nanoliters (nl); 1 nl is 1000 picoliters (pl).
 - Grams follow the same pattern; distances in meters, also.

V. Acids, Bases and the pH Scale

- Some substances that dissolve in water influence the concentration of hydrogen protons (H⁺) in the solution.

- The change in concentration of H^+ can profoundly affect biological molecules.

A. Acids donate H^+ ; bases accept H^+

- At a certain pH, an acid will increase the H^+ concentration of an aqueous solution. It can do this by either contributing to the H^+ concentration or by decreasing the $(OH)^-$ concentration. What is important is that the net H^+ concentration is increased.
 - If HCl (hydrochloric acid) is added to water, it dissolves and ionizes into H^+ and Cl^- ions. Because H^+ concentration increases, such a solution becomes acidic.
 - $-COOH$ is a carboxyl group. It is a common organic acid functional group. The reaction is $-COOH \rightarrow COO^- + H^+$.
- Bases accept H^+ (or increase the concentration of $(OH)^-$). Bases cause a net reduction in the H^+ concentration.
 - An example is the dissolving of NaOH in water: $NaOH \rightarrow Na^+ + OH^-$. The increase in $[(OH)^-]$ causes the same amount of $[H^+]$ to decrease.
 - Another is an amino group, common to amino acids, which make up proteins: $-NH_2 + H^+ \rightarrow -(NH_3)^+$. Again, this reaction lowers the H^+ concentration of water.

B. Water is a weak acid.

- Water can slightly ionize to release H^+ . However, pure water is primarily found in molecular, not ionic form.
- In the absence of any atmospheric CO_2 or impurities of sulfur, nitrogen or other compounds that can influence pH, pure water has a pH of 7, which is neither acid or base. When an H^+ is released, so is an $(OH)^-$, so pure water is equally a weak base. $H_2O \rightarrow H^+ + (OH)^-$.

C. The pH is the measure of hydrogen ion concentration.

- The pH 7 means the concentration of hydrogen ions (or hydronium ions, which are H^+ attached to a water molecule) is $[1 \times 10^{-7}]$. This is 1×10^{-7} moles per liter of water.

- Remember from the above boxed section that each liter of water contains 55.56 moles and of these molecules, 1×10^{-7} moles are ionized water, when the pH is neutral. This would be $1 \times 10^{-7}/55.56$ or 1 H^+ for every 5.6×10^8 water molecules (or 1 H^+ for per 560,000,000 water molecules).
- When water is at pH 7, both the H^+ and $(OH)^-$ are both at a concentration of 10^{-7} molar. (See Figure 2.18)
 - When water is pH 8 it is 10^{-8} molar for H^+ and 10^{-6} molar for $(OH)^-$. The higher the number of the pH the greater the OH^- concentration and the lower the H^+ concentration.
 - When an aqueous solution is at pH 6, the H^+ concentration is at 10^{-6} or is 10X greater than at pH 7, and the OH^- concentration becomes 10^{-8} or 1/10 that of

pH 7.

- See if you can find the pattern and predict the concentration of H^+ and $(OH)^-$ at pH 13.
 - Even strongly acidic solutions like those with pH 1 have mostly water molecules and not ions, $10^{-1}/55.56$ ions of H^+ , or just one H^+ for every 556 water molecules!
 - However, this concentration of H^+ can have a powerfully corrosive effect on a variety of materials including metals, polysaccharides, proteins, nucleic acids, bone and others.
- It is the $[H^+]$ that determines how acidic a solution is.
 - The pH 7 is neutral because $[OH^-] = [H^+]$
 - The pH 6 is acid because $[OH^-] < [H^+]$
 - The pH 5 is even more acidic.
 - The pH 8 is basic because $[OH^-] > [H^+]$.

D. Buffers Minimize pH Change

- Buffers are acids or bases that exist in a solution in uncharged molecular form or as ions, depending on the pH.
- Because these compounds change forms as pH changes they make solution more resistant to pH change. (See Figure 2.19)

- A buffer might be mostly uncharged at pH 2, half in the uncharged and half in the ionic form at pH 4, and mostly in the ionic form at pH 6. As the pH moves through the range of pH 2 through 6, H^+ are released from the uncharged molecules, as these molecules are converted to the charged or ionic form.
- The contribution of the H^+ from these molecules tends to make the solution resist the movement into higher pH ranges.

- Different buffers transition to and from ionic form at their particular characteristic pH ranges.
- Buffers are extremely important and common in biology.
- The many important biological buffers transition around pH 7, helping to keep the pH near neutral.
- Law of Mass Action: Addition of components to one side of a reaction drives the reaction in the direction that uses that component. As acid or base is added to a solution the buffer will change form, transitioning between ionic and non-ionic forms.

VI. The Properties of Molecules

- Molecules range in size and molecular weight from the very smallest, H_2 , to massive, like the DNA molecule that makes up the length of a chromosome, and contains millions of atoms.
- Carbon containing molecules are common in life forms.

- Organic molecules *are* carbon-containing molecules.
- Most organic molecules also contain hydrogen and oxygen. Many contain nitrogen and phosphorus.
- Nitrogen, oxygen and carbon are also all found in air.
- Hydrogen and oxygen are found in water.
- Molecules have three-dimensional shapes.
 - CH₄ is tetrahedral
 - Large molecules have complex shapes that are the result of the atoms and the location of atoms found within them.
 - Some molecules are long and ropelike; some are ball-shaped.
 - Shape influences the behavior or function of molecules.
- Chemists use the characteristics of composition, structure, reactivity and solubility to help classify them.

A. Function Groups of Organic Molecules

- Some biologically important functional groups are hydroxyl, carbonyl, carboxyl, amino, phosphate and sulfhydryl. (*See Figure 2.20 (the one on page 31)*) <<You currently have two figures with the number 2.20, one on 31 and the other one on 32.>>
- Functional groups are covalently bonded to organic molecules.
 - Carbonyl groups, which are double-bonded oxygen to a carbon, found at the end of organic molecules, are aldehydes.
 - A carbonyl in the middle of a chain of carbons is called a ketone.

B. Isomers Have Different Arrangements of the Same Atoms.

- Iso means same.
- Structural isomers differ in how atoms are joined together.
- Optical isomers can occur whenever a carbon has four different atoms or groups attached. (*See Figure 2.20 (the one on page 32)*)
 - This carbon is called asymmetric.
 - Optical isomers are mirror images of each other.
 - An example of this type of symmetry would be your left and right hand.
 - Amino acids (except glycine) exist in two optical isomeric forms called D- and L- amino acid. L-amino acids are the ones commonly found in most organisms.