SBI4U-C

Lesson 1

Introducing Biochemistry

Introduction

If you think about it, living things are made up of chemicals. The cells, tissues, organs, and organ systems that together make up a functioning organism are all the product of interacting chemicals. In order to understand how the organism functions, you must also recognize how these chemicals function. The study of the chemicals that are involved in living things is known as biochemistry. Many of the chemistry concepts that you have already learned will be applied in your study of biochemistry. So, in this first lesson, you will review some basic chemistry.

Planning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Chemistry Basics	¾
Radioisotopes	¾
Chemical Bonding	1
Types of Biochemical Reactions	1
Key Questions	1

What You Will Learn

After completing this lesson, you will be able to

- use appropriate terminology related to biochemistry
- identify and describe the four main types of biochemical reactions
- evaluate technological applications related to the use of isotopes in biology and medicine

Chemistry Basics

In this lesson, you will begin to explore the structure and function of the chemical components of living organisms. You will start by reviewing concepts of atomic structure, isotopes, and chemical bonding, as well as the main types of chemical reactions that occur. Throughout the lesson you will learn why these concepts are important in understanding biochemistry and living systems.

Matter Is Made Up of Elements

Matter is anything that has mass and takes up space. Matter is made up of tiny particles called atoms. Atoms consist of three subatomic particles. The nucleus of an atom contains the positively charged proton(s), and the neutral neutron(s). Orbiting the nucleus are the negatively charged electron(s). Although electrons have a charge, they have a relatively tiny mass, so the weight of an atom is determined by its protons and neutrons. Atoms are normally electrically neutral, meaning that the number of protons in an atom equals the number of electrons.

The number of protons in the nucleus is called the atomic number. This number determines many of the atom's properties, which you will learn about later on. Many atoms have multiple forms, called isotopes, which have different numbers of neutrons in their nucleus, but the same atomic number.

Atoms are normally electrically neutral. So, for example, oxygen has an atomic number of 8, which means that it has eight protons in its nucleus and therefore normally eight electrons orbiting around its nucleus.

Atoms are also defined by their atomic mass. The atomic mass is the number of protons plus the number of neutrons in the atom. For example, oxygen has a mass of 16. Since you know from the atomic number that oxygen contains eight protons, then it must contain $16 - 8 = 8$ neutrons. The number of protons is usually similar to the number of neutrons, but not always. For example, the element chlorine has an atomic number of 17, indicating that it has 17 protons, but its atomic mass is 39, indicating that it contains $39 - 17 = 22$ neutrons.

The different types of atoms are called elements. Elements are organized in a chart called the periodic table of elements. Each element is represented by its one- or two-letter atomic symbol. The elements are organized into vertical columns called groups, and horizontal rows called periods.

Periodic Table

Atomic Number → 6
Symbol → C
Name → Carbon
Average Atomic Mass → 12.011
Common Oxidation States ← 4 −4 2

Period	Group 1 IA	Group 2 IIA	Group 3 IIIB	Group 4 IVB	Group 5 VB	Group 6 VIB	Group 7 VIIB	Group 8	Group 9 VIIIB	Group 10	Group 11 IB	Group 12 IIB	Group 13 IIIA	Group 14 IVA	Group 15 VA	Group 16 VIA	Group 17 VIIA	Group 18 VIIIA
1	1 **H** Hydrogen 1.00794 1 −1																	2 **He** Helium 4.002602
2	3 **Li** Lithium 6.941 1	4 **Be** Beryllium 9.012182 2											5 **B** Boron 10.811 3	6 **C** Carbon 12.011 4 −4 2	7 **N** Nitrogen 14.00674 3 −3 5	8 **O** Oxygen 15.9994 −2	9 **F** Fluorine 18.9984032 −1	10 **Ne** Neon 20.1797
3	11 **Na** Sodium 22.989768 1	12 **Mg** Magnesium 24.3050 2											13 **Al** Aluminum 26.981539 3	14 **Si** Silicon 28.0855 4	15 **P** Phosphorus 30.973762 5 −3 3	16 **S** Sulfur 32.066 6 −2 2	17 **Cl** Chlorine 35.4527 −1 1 3	18 **Ar** Argon 39.948
4	19 **K** Potassium 39.0983 1	20 **Ca** Calcium 40.078 2	21 **Sc** Scandium 44.955910 3	22 **Ti** Titanium 47.88 4 3	23 **V** Vanadium 50.9415 5 4 3	24 **Cr** Chromium 51.9961 3 2 6	25 **Mn** Manganese 54.93805 2 3 4	26 **Fe** Iron 55.847 2 3	27 **Co** Cobalt 58.93320 2 3	28 **Ni** Nickel 58.6934 2 3	29 **Cu** Copper 63.546 2 1	30 **Zn** Zinc 65.39 2	31 **Ga** Gallium 69.723 3	32 **Ge** Germanium 72.61 4	33 **As** Arsenic 74.92159 3 −3 5	34 **Se** Selenium 78.96 4 −2 6	35 **Br** Bromine 79.904 −1 1 5	36 **Kr** Krypton 83.80
5	37 **Rb** Rubidium 85.4678 1	38 **Sr** Strontium 87.62 2	39 **Y** Yttrium 88.90585 3	40 **Zr** Zirconium 91.224 4	41 **Nb** Niobium 92.90638 5 3	42 **Mo** Molybdenum 95.94 6 5 4	43 **Tc** Technetium (97.90772) 7	44 **Ru** Ruthenium 101.07 3 4 6	45 **Rh** Rhodium 102.90550 3 2 4	46 **Pd** Palladium 106.42 2 4	47 **Ag** Silver 107.8682 1	48 **Cd** Cadmium 112.411 2	49 **In** Indium 114.818 3	50 **Sn** Tin 118.710 4 2	51 **Sb** Antimony 121.757 3 −3 5	52 **Te** Tellurium 127.60 4 −2 6	53 **I** Iodine 126.90447 −1 1 5	54 **Xe** Xenon 131.29
6	55 **Cs** Cesium 132.90543 1	56 **Ba** Barium 137.327 2	57 **La** Lanthanum 138.9055 3	72 **Hf** Hafnium 178.49 4	73 **Ta** Tantalum 180.9479 5	74 **W** Tungsten 183.84 6 5 4	75 **Re** Rhenium 186.207 7 −1 6	76 **Os** Osmium 190.23 4 2 3	77 **Ir** Iridium 192.22 4 2 3	78 **Pt** Platinum 195.08 4 2	79 **Au** Gold 196.96654 3 1	80 **Hg** Mercury 200.59 2 1	81 **Tl** Thallium 204.3833 1 3	82 **Pb** Lead 207.2 2 4	83 **Bi** Bismuth 208.98037 3 5	84 **Po** Polonium (208.9824) 4 2	85 **At** Astatine (209.9871) −1 1 3	86 **Rn** Radon (222.01761)
7	87 **Fr** Francium (223.0197) 1	88 **Ra** Radium (226.0254) 2	89 **Ac** Actinium (227.0278) 3	104 **Rf** Rutherfordium (261.11)	105 **Db** Dubnium (262.114)	106 **Sg** Seaborgium (263.118)	107 **Bh** Bohrium (262.12)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (266)	110 **Ds** Darmstadtium (269)	111 **Rg** Roentgenium (272)	112 **Uub** Ununbium (275)	113 **Uut** Ununtrium (278)	114 **Uuq** Ununquadium (282)	115 **Uup** Ununpentium	116 **Uuh** Ununhexium	117 **Uus** Ununseptium	118 **Uuo** Ununoctium

58 **Ce** Cerium 140.115 3 4	59 **Pr** Praseodymium 140.90765 3 4	60 **Nd** Neodymium 144.24 3	61 **Pm** Promethium (144.9127) 3	62 **Sm** Samarium 150.36 3 2	63 **Eu** Europium 151.965 3 2	64 **Gd** Gadolinium 157.25 3	65 **Tb** Terbium 158.92534 3 4	66 **Dy** Dysprosium 162.50 3	67 **Ho** Holmium 164.93032 3	68 **Er** Erbium 167.26 3	69 **Tm** Thulium 168.93421 3 2	70 **Yb** Ytterbium 173.04 3 2	71 **Lu** Lutetium 174.967 3
90 **Th** Thorium 232.0381 4	91 **Pa** Protactinium 231.03588 5 4	92 **U** Uranium 238.0289 6 5 4	93 **Np** Neptunium (237.0482) 5 4 3	94 **Pu** Plutonium (244.0642) 4 3 5	95 **Am** Americium (243.0614) 3 4 5	96 **Cm** Curium (247.0703) 3	97 **Bk** Berkelium (247.0703) 3 4	98 **Cf** Californium (251.0796) 3	99 **Es** Einsteinium (252.083)	100 **Fm** Fermium (257.0951)	101 **Md** Mendelevium (258.10)	102 **No** Nobelium (259.1009)	103 **Lr** Lawrencium (262.111)

Figure 1.1: The periodic table of elements. The number above each symbol is its atomic number and the number below it is its average atomic mass.

You will notice that many elements have an atomic mass number that is not a whole number. For example, nitrogen (N) has a mass number of 14.00674 and carbon (C) has a mass number of 12.011. This does not mean, however, that the number of protons or neutrons in individual atoms will be in fractions. These atomic mass numbers are the weighted averages of the different types of isotopes. For example, most carbon atoms have six neutrons, while a tiny percentage has seven or eight. That's why their average atomic mass is slightly more than 12. However, no single carbon atom would ever have a mass of 12.011—it would either be 12 (six protons and six neutrons), 13 (six protons and seven neutrons) or 14 (six protons and eight neutrons).

Isotopes are named using a system that includes the element name and its mass. For example, the isotope of carbon that contains eight neutrons is called "carbon-14" to indicate that it has a different nucleus to the normal form of carbon with six neutrons ("carbon-12"). The short form of the name uses the element symbol with the atomic mass given in superscript on the upper left, so that carbon-14 becomes ^{14}C. Generally, you would only add the number to the element name or symbol when you are talking about the less common isotope forms of an element, so carbon-12 is usually just known simply as carbon. Commonly studied isotopes in biological systems are carbon-14 (^{14}C) and oxygen-18 (^{18}O). You will learn more about the importance of isotopes later on in the lesson.

The following example shows how you can use the periodic table to retrieve information about subatomic particles.

Example

Determine the number of protons, neutrons, and electrons found in the element potassium, which is represented by the symbol "K."

Solution

Data from the periodic table for the element potassium are shown below. The atomic number is shown above the atomic symbol and the mass number is shown below it.

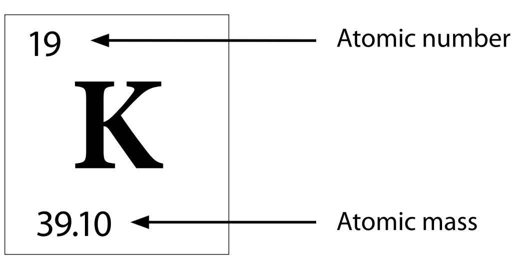


Figure 1.2: The element potassium, with its atomic number and atomic mass

The atomic number of potassium is 19. Since the atomic number is equal to the number of protons, a potassium atom has 19 protons. Since the number of protons equals the number of electrons, the potassium also has 19 electrons.

The atomic mass is 39.10. By rounding to the nearest whole number, you know that the most common isotope of this element has a mass of 39. Therefore, for the most common isotope, the number of neutrons is 39 – 19 = 20 neutrons.

Structure of the Atom

For most applications in biochemistry, a simplified version of the atomic structure is usually sufficient for understanding chemical reactions. Bohr-Rutherford diagrams, or Bohr diagrams as they are commonly called, are simple models that represent the arrangement of the subatomic particles in an atom. Figure 1.3 shows the Bohr diagram for carbon, which is the element upon which all life is built. It shows that carbon has six protons, six neutrons, and six electrons arranged in orbits around the nucleus.

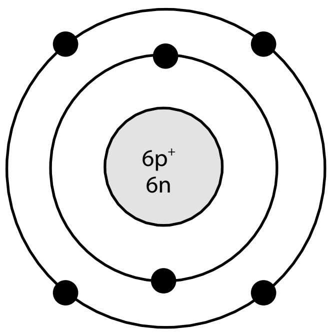


Figure 1.3: Bohr diagram for the element carbon. The electrons are the black circles orbiting the grey nucleus.

You can think of electrons moving in orbits around the nucleus like planets revolving around the sun. However, their orbits are not as simple as those of the planets, so you can't tell exactly where the electrons will be at any given time. In fact, because the electrons seem to be everywhere at once within their orbits, they appear to form a solid shell around the nucleus. That is why the orbits are called electron shells.

Electrons cannot exist between their orbits, but can move up or down from one orbit to another. Each orbit has a maximum number of electrons that it can hold. Electrons are more stable when they are at lower energy, which occurs when they are closer to the nucleus. This means that the electrons are always trying to occupy the lowest orbits they can find, so they fill up the orbits starting from the nucleus and working outwards.

Orbits generally hold a maximum of eight electrons. However, the first shell—the one closest to the nucleus—can only hold a maximum of two. For example, carbon contains six electrons in total, which means that it has two electrons filling its first orbital shell and four electrons occupying its second (and last) orbital shell. Potassium, with 19 electrons, would have two in its first shell, eight in its second, eight in its third, and only one in its last shell. The electron configuration for potassium can be written in a short form of the Bohr diagram as follows:

K)2e–)8e–)8e–)1e–

with the element symbol ("K") written first, followed by its electrons (represented by "e-"), shown according to their abundance in each shell, moving out from the centre. The shells are represented by closed brackets.

Valence Electrons

The number of electrons occupying the outermost (that is, the last) shell will determine how the element reacts with other atoms to form chemical bonds in biological molecules. Because of their importance, the electrons in the outer shell are given a special name: valence electrons. In the examples you have looked at so far, carbon has four valence electrons and potassium has one.

The maximum number of valence electrons in any shell (except the first shell) is eight. When an atom gets eight valence electrons it becomes optimally stable and chemically inert, meaning that it does not want to take part in chemical reactions. All atoms are trying to get the maximum number of valence electrons in their outer shells (usually eight).

You can determine how many valence electrons an element has by seeing which vertical column it occupies in the periodic table. Carbon is in group IV, which means that it has four valence electrons, as do other members of group IV, including silicon (Si) and lead (Pb).

Lewis Diagrams

Another useful way to represent the electron configuration of an element is to draw its Lewis structure or Lewis diagram. Lewis diagrams depict only the valence electrons and their relative position. This makes it easier to see how they might react with other atoms. Figure 1.4 below depicts the Lewis diagram for the element chlorine (Cl). This is an important element found in salt (NaCl), and is an essential molecule for nerve function.

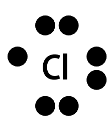


Figure 1.4: Lewis diagram for chlorine

The Lewis diagram shows the element symbol surrounded by its valence electrons, each represented by a dot (or sometimes also by an "X," when several elements are being combined and you want to keep track of where the different electrons go). In this example, the seven valence electrons of chlorine are represented by dots. By convention, a maximum of two electrons are placed on each of the four sides of the symbol. In this case, this helps you to quickly see that chlorine is missing one valence electron, in order to reach its stable configuration of eight valence electrons. You will use Lewis diagrams later on in this lesson, when you review chemical bonding.

Ions

The most chemically stable elements in the periodic table are called the noble gases, which are located in group VIII along the far right-hand column of the periodic table. They include helium (He), neon (Ne), and argon (Ar), among others. All of these elements naturally possess their maximum number of valence electrons. Note that helium only has two electrons in total so it has only two in its outer shell, while all the others have eight in their outer shell. Because their outer shells are full, noble gases are extremely non-reactive and tend not to form compounds.

All elements are most stable when they have the maximum number of valence electrons. In most cases, this means having eight electrons in the outer shell. This is called the octet rule. According to this rule, atoms bond in order to achieve the same electron configuration as a noble gas and become chemically stable. Atoms bond with other atoms in order to either gain or lose electrons in their outer shell so that they will end up with eight electrons in their outer shell.

Once atoms have either gained or lost electrons, they no longer have a neutral charge because the number of electrons no longer equals the number of protons. They are now called ions. An ion is an atom in which the total number of electrons is not equal to the total number of protons. An atom that has gained electrons to become stable is now negatively charged and is called an anion. An atom that has lost electrons becomes positively charged and is called a cation.

For example, consider the formation of the calcium cation. Calcium cations are particularly important for muscle contractions in your body. To find out the number of valence electrons in calcium, you can write out its electron configuration. Start by determining how many electrons calcium has. From the periodic table, you can see that calcium has an atomic number of 20, so it has 20 electrons. Its electron configuration using the short form of the Bohr diagram is:

Ca)2e–)8e–)8e–)2e–

Since calcium has two electrons in its outer shell, you know that it has two valence electrons.

The Lewis diagram for calcium is:

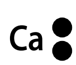


Figure 1.5: Lewis diagram for calcium

To achieve the same electron configuration as a noble gas, this atom could either lose two electrons in its outer shell (and thus lose this shell and so have eight electrons in its new outer shell, one orbit below) or gain six. Atoms tend to do whatever is easiest, so in this case, it is easier to lose two electrons, rather than gain six. The electron configuration for a calcium cation that has lost the two electrons in its outer shell is written as:

[Ca)2e–)8e–)8e–]$^{+2}$

As the calcium cation has lost two electrons, it now has a charge of positive two (+2). Notice how the ion is written with square brackets, and that the overall charge is indicated on the upper right-hand side.

Example

Show the electron configuration of the nitrogen anion.

Solution

First, work out the number of valence electrons in nitrogen. From the periodic table, you can see that the atomic number for nitrogen is 7. The electron configuration is then:

N)2e–)5e–

In this case, nitrogen has five valence electrons, so it needs to gain three to obtain a full octet (eight valence electrons). Thus, nitrogen will form a negatively charged ion or anion. Since the ion gained three electrons, it now has a charge of negative three (–3), so you draw the revised electron configuration as:

[N)2e–)8e–]$^{-3}$

Support Questions

Be sure to try the Support Questions on your own before looking at the suggested answers provided. Click on each "Suggested answer" button to check your work.

1. Describe the structure of an atom and compare this to the structure of an ion.

Suggested answer

2. Print out and complete the following table by filling in the missing information in each column. One piece of data is given for each element, to get you started.

Element name	Boron			
Symbol		O		
Atomic #				11
# Protons			4	
# Neutrons				
Total # electrons				
# Electrons in 1st shell				
# Electrons in 2nd shell				
# Electrons in 3rd shell				
# Valence electrons				

Suggested answer

3. Write the short form of the Bohr diagram for the following atoms:

a) Beryllium **b)** Sulphur

Suggested answer

4. Draw a Lewis diagram to show the arrangement of electrons in the following atoms:

a) Phosphorus **b)** Boron

Suggested answer

Radioisotopes

Isotopes, in particular, are used in a wide variety of important biotechnologies. Recall that isotopes are atoms that contain the same number of protons as the common form of the element, but have a different number of neutrons. A radioisotope is a special type of isotope that has an unstable nucleus. As it breaks down in order to contain fewer neutrons during a process called radioactive decay, it emits radiation. The time it takes for the isotope to complete half of its decay process is called its half-life. Radioisotopes are found naturally in very low abundance, but many can now be made using small nuclear reactors. Ontario has been a world leader in producing and supplying medical radioisotopes for many years.

Many common elements in biology also come in radioisotope forms. For example, as stated earlier, carbon normally contains six neutrons (^{12}C), but one isotope form contains eight neutrons (^{14}C). The ^{14}C isotopes are unstable and will gradually decay, so the extra neutrons eventually will be lost and the atom will become ^{12}C. There are many other examples of radioisotopes used in biology and medicine, including isotopes of oxygen, nitrogen, phosphorus, iodine, barium, and strontium.

The radioactive decay of the isotopes makes them valuable as diagnostic and treatment tools in medicine and other fields, for three main reasons:

- The decay process occurs at a constant rate; it can therefore act as a clock, to help determine the age of materials.
- The radiation emitted can be detected as a form of light, and can therefore serve as a marker for the progress of biochemical reactions.
- The radiation emitted can, in some cases, be strong enough to kill nearby cells, which can be helpful in cancer treatments.

Using Isotopes as Clocks: Ratios Can Tell You about Age

Isotopes get incorporated into living tissues along with the regular forms of the element at ratios that reflect their abundance in the environment. Radioisotopes also decay at a constant rate, with time scales varying from days to hundreds of millions of years, depending on the isotope. This means that the ratio of different isotopes in tissue can potentially provide clues as to when an organism lived. In most cases, the isotope ratios in living and dead tissue will be the same, because the isotope formed millions of years ago and experiences the same rate of radioactive decay in both living and dead tissue. However, carbon is an exception. The main isotope of carbon, carbon-14 (^{14}C), is constantly being created by a bombardment of cosmic rays in the upper atmosphere.

Living plants and animals incorporate this new ^{14}C in their tissues through the nutrients they consume and the air they breathe. However, when they die, they stop incorporating new ^{14}C into their tissues and the existing ^{14}C decays at a constant rate, with a half-life of 5730 years. The decrease in the ratio of ^{14}C to normal carbon (^{12}C) after death acts as a clock to tell us when the organism died. The lower the ratio of ^{14}C in the sample, the longer the tissue has been dead. However, after 40 000 years, less than 1% of the ^{14}C is left in the tissue, so this isotope is not useful for dating very old material.

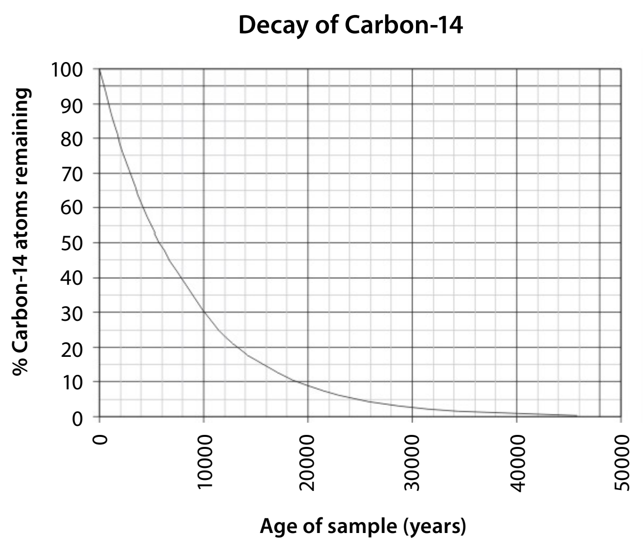


Figure 1.6: Decay of carbon-14

This technique is called radiocarbon dating. It has been widely used, for example, to date the wood in archeological sites, samples of clothing and food found in Egyptian tombs, and objects made from animal tissues like leather and bone. It has even been used to detect art forgeries.

Using Radioactive Tracers

Radioisotopes give off radiation as they decay. This radiation can be detected as a form of X-ray light. This means that molecules containing radioisotopes literally can be watched as they move through a person's body, as they look almost as if they are carrying little flashlights.

Radioisotopes used to follow chemicals through many chemical reactions are called radioactive tracers. They have been used to help us understand the steps in biochemical reactions, such as those that occur in cellular respiration or photosynthesis. They have also been used to "tag" protein and DNA molecules to help us understand the role of DNA in heredity.

Nuclear Medicine

Radioisotopes are also used in medicine for diagnoses and treatments. This field of medicine is called nuclear medicine.

Radioisotopes are often used to diagnose medical conditions. For example, a radioisotope of iodine is used to detect abnormalities in the thyroid gland. The thyroid produces hormones for growth and metabolism. Thyroid disorders affect about one in three Canadians. To determine thyroid function, doctors can inject radioactive iodine (^{131}I) into the gland. Since the thyroid readily absorbs iodine, an X-ray image can then be taken to detect the size of the thyroid based on the amount of (^{131}I) it has absorbed. An enlarged thyroid can be a sign of disease or cancer.

Since radioisotopes give off small doses of radiation, they can also be used to treat some cancers and other medical conditions that require the destruction of harmful cells. Radioisotopes with very short half-lives (of hours to days) can be inserted into molecules targeted for absorption by specific cells or tissues. The low level of radiation they give off locally is intense enough to kill the nearby cancerous cells. For example, strontium-89 is widely used to destroy cancer cells lodged in bone.

Support Questions

5. How can radioisotopes be used to determine the age of archeological material?

Suggested answer

6. Describe one example where radioisotopes can be used as a diagnostic tool in medicine.

Suggested answer

Chemical Bonding

Life is possible because atoms join up with other atoms in a process called chemical bonding. You will review two broad categories of chemical bonding in this section: bonds within molecules (*intra*molecular) and bonds between molecules (*inter*molecular).

Bonds within Molecules: Intramolecular Bonds

The bonding between atoms results in a compound called a molecule. Because these bonds occur within the molecule, they are called intramolecular bonds ("intra" means "within").

There are three main types of intramolecular bonds, which you will briefly review in this lesson:

1. Ionic bonds
2. Covalent bonds
3. Polar covalent bonds

Ionic Bonds

An ionic bond forms when one atom transfers electrons to another atom. As a result, the atoms end up with a stable electron arrangement in their outer orbit, similar to that of a noble gas.

Ionic bonds usually form between metals and non-metals. Opposite charges attract and the charge attraction is what forms the ionic bond.

The elements potassium and fluorine, for example, will form an ionic bond between them. Lewis diagrams can be useful in illustrating how the ionic bond forms. Figure 1.7 below shows how the electrons arrange themselves to form the ionic bond.

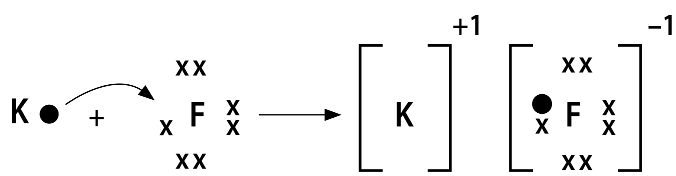


Figure 1.7: Formation of an ionic bond between potassium (K) and fluorine (F)

Potassium (K) loses its one valence electron, becoming a cation with a charge of +1. Fluorine (F) gains one valence electron, becoming an anion with a charge of –1. The positive potassium cation (+) is attracted to the negative fluorine anion (–). This is what forms the ionic bond. When using Lewis diagrams, notice that the ions are written with square brackets and the overall charge is indicated on the upper right-hand side.

Since the valence numbers summarize the dot patterns, it is often easier just to write the Lewis diagram without adding the dots representing the electrons. The short form of the Lewis diagram for this reaction is then:

$[K]^{+} [F]^{-}$

If the valences between the atoms are not of equal size, then the stable molecule that forms from the ionic bonds will have to involve more than one atom of each element, in order to balance out. For example, the ionic bond between magnesium and phosphorus is:

$[Mg]^{+2} [P]^{-3}$,

which means that three atoms of magnesium have to bond with two molecules of phosphorus, in order for each to form stable ionic bonds within a molecule. The resulting molecule is Mg_2P_3.

Covalent Bonds

A covalent bond forms when two or more non-metals share one or more pairs of electrons. As a result, the atoms end up with a stable electron arrangement in their outer orbit, similar to that of a noble gas. The simplest example of a covalent bond is between two hydrogen atoms, as shown below.

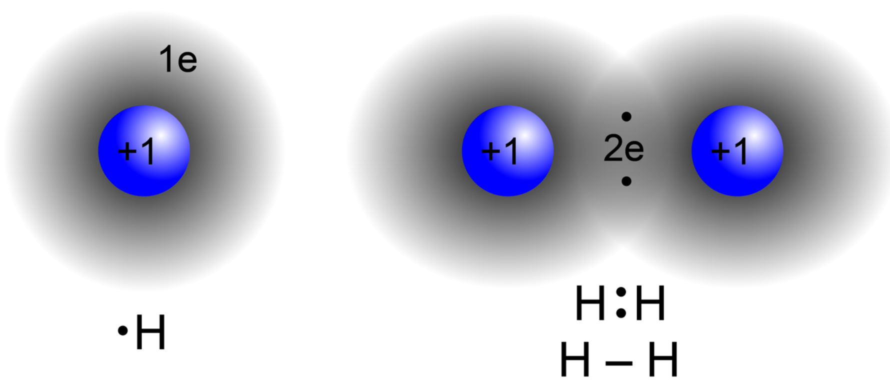
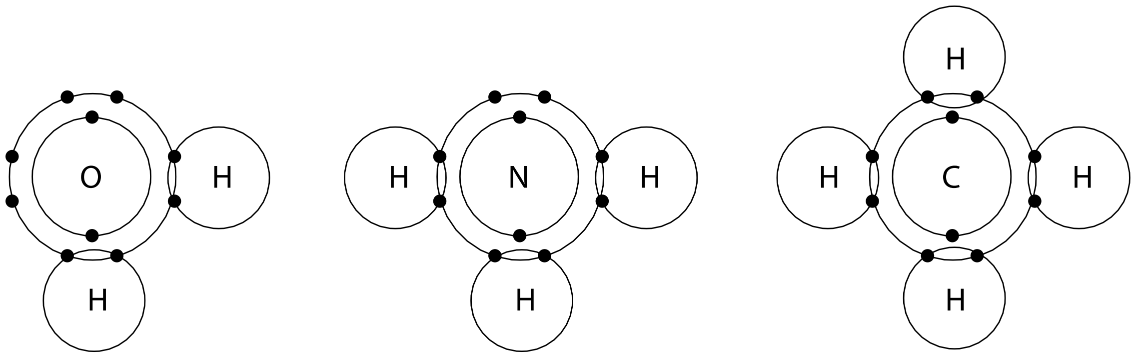


Figure 1.8: Formation of a covalent bond between two hydrogen (H) atoms

Each hydrogen atom only has one electron in its outer shell, but is most stable when this shell is full, with two electrons in it. When two hydrogen atoms meet, they can both share a pair of electrons, so both become stable.

Chlorine gas (Cl_2) is another example of a molecule that is formed using covalent bonds. A chlorine atom has seven electrons in its outer orbit, so it needs to gain one electron in order to become a stable octet. When two chlorine atoms meet, they can share a pair of electrons to form a covalent bond. The result is that each chlorine atom now has eight electrons in its outer orbit. This forms a stable octet.

Examples of covalently bonded molecules that are important for biology include water, ammonia, and methane, as shown in Figure 1.9, which follows. Notice how all of the hydrogen atoms in the covalent bond now have two electrons filling their outer shells, while the oxygen, nitrogen, and carbon atoms all have eight electrons in their outer shells.



H_2O NH_3 CH_4

Figure 1.9: Examples of covalent bonds in important biological molecules are water (H_2O), ammonia (NH_3), and methane (CH_4).

Polar Covalent Bonds

A polar covalent bond is a special type of covalent bond. It is formed when there is an unequal sharing of valence electrons between atoms in the pair, with one atom pulling harder on the electrons than the other. This results in different ends of the molecule having different electric charges. Because this results in two "poles," the molecule is said to be a dipole, with one end being slightly positive and the other, slightly negative. Whichever atom pulls the hardest on the electrons will become the slightly more negative pole, while the other one will become more positive. This happens because atoms differ in their ability to attract electrons. Electronegativity is a measure of an atom's ability to attract electrons in a chemical bond. This is a fundamental property of the atom. Table 1.1 on the next page summarizes the electronegativity values for several important biological elements.

Table 1.1: Electronegativity of some important biological elements

	Element	Electronegativity
	H	2.1
Metals	Na	0.9
	Mg	1.2
	K	0.8
	Ca	1.0
Non-metals	C	2.5
	N	3.0
	O	3.5
	P	2.1
	S	2.5
	Cl	3.0

The more electronegative atom in the pair will pull on the electrons with greater force, and so will form a negative region in the molecule.

Water (H_2O) is one of the most important polar covalent molecules in biological systems. The electronegativity value for oxygen is 3.5, while for hydrogen it is 2.1. So, in a water molecule, the oxygen is pulling harder on the electrons than each hydrogen atom. This brings the electrons closer to the oxygen nucleus, thus making the oxygen end slightly negative (δ–) and the hydrogen regions slightly positive (δ+), as depicted in the ball-and-stick model for water shown in Figure 1.10 below.

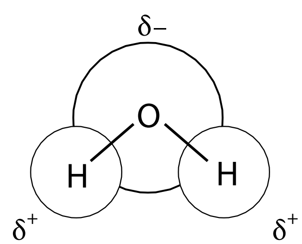


Figure 1.10: Water is a dipole molecule with its oxygen side more negative than its hydrogen side.

The polarity of water has important implications for all sorts of biochemical processes, as you will see throughout this course.

Bonds between Molecules: Intermolecular Bonds

Intermolecular bonds are the chemical bonds between molecules ("inter" means "between"). Bonds between molecules are much weaker than the bonds within molecules. Intermolecular bonds determine the physical state of molecular substances. These bonds are broken as a substance undergoes a change of state, such as from a solid to a liquid or from a liquid to a gas.

There are three types of intermolecular bonds:

1. London forces
2. Dipole-dipole forces
3. Hydrogen bonds

These intermolecular forces are collectively called van der Waals forces. Table 1.2 summarizes the main features of each of these types of bond.

Table 1.2: Intermolecular bonds listed in order of relative strength

Force	Description	Relative strength	Example
London forces	• Hold non-polar molecules together • Very weak forces of attraction • Momentary dipoles are created by the electrons contained within the compound, which are constantly in motion.	Weakest	CH_4
Dipole-dipole forces	• Hold polar molecules together • These forces are stronger than London forces.	Medium	HCl
Hydrogen bonding	Is formed between the electropositive hydrogen dipole and an electronegative dipole of oxygen, chlorine, or fluorine	Strongest	Pure distilled water

The different types of intermolecular bonds give rise to a number of important structural and functional properties of living organisms. For example, water is a polar molecule that has many unique properties. It is often called the universal solvent because of its ability to dissolve many ionic and polar compounds. Life as we know it could not exist without liquid water, since nearly all biochemical reactions take place within a water medium.

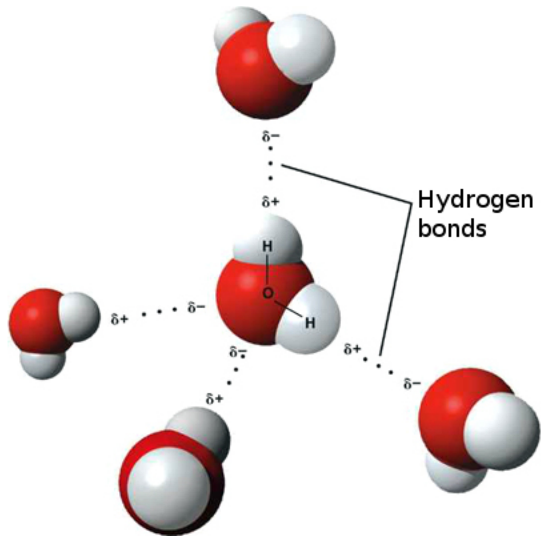


Figure 1.11: Hydrogen bonding in water

Hydrogen bonding (the strongest of the intermolecular forces) gives water its physical properties and also helps to support life on earth. For example, water has a relatively high boiling point (100°C) and a low freezing point (0°C). This allows water to be an excellent transport medium in our bodies, since it will not boil or freeze at temperatures optimal for biochemical reactions. In the external environment, the physical properties of water provide habitat and protection for many life forms. For example, because frozen water is less dense than liquid water, ice floats and can act like an insulating blanket to protect the life below it in the polar oceans. The surface tension of liquid water allows several insects and other life forms, such as aquatic plants, to live on the water's surface. More properties of water will be discussed in the next section.

Support Questions

7. Print out and complete the following table. Use the short form of Lewis diagrams to depict the bond formation. The first one has been done for you.

	Bond formation	Name of compound	Chemical formula	Anion	Cation
a) Sodium and chlorine	$[Na]^+ [Cl]^-$	Sodium chloride	NaCl	Cl^-	Na^+
b) Lithium and fluorine					
c) Calcium and phosphorus					

Suggested answer

8. Explain the polar covalent nature of the water molecule.

Suggested answer

Types of Biochemical Reactions

There are four main reaction types that are important in biochemistry:

1. Hydrolysis
2. Condensation
3. Oxidation-reduction (redox)
4. Neutralization reactions

You will review each of these in more detail in the sections that follow.

Hydrolysis

In the previous section, you learned that water is a biologically important molecule. You learned that water was polar covalent because it had "poles" or regions that had opposite charges. This polar nature of water allows it to stick together, as well as stick to other structures in your body. Water is also important for biochemical reactions. Water is often required to break down and join various molecules together. For example, when you consume food, you require water and enzymes to break apart the food you eat into smaller, simpler units. This reaction is called hydrolysis because it uses water to help break down molecules ("hydro" means "water" and "lysis" means "to break down").

Here's a simple hydrolysis reaction that occurs every second in your body, where water is used to break down a protein into its component amino acids:

Protein + water → amino acid + amino acid

Recall that in a chemical reaction, the *reactants* are written on the left (in this example, the protein and water), while the *products* (amino acids) are written on the right.

Notice how the atoms of the water molecule are rearranged in a hydrolysis reaction, as shown in this simple reaction of:

Dimethyl ether + water → methanol + methanol

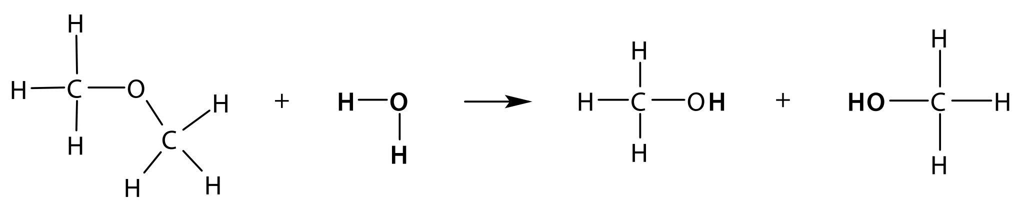


Figure 1.12: Hydrolysis reaction of dimethyl ether. The atoms in the water molecule are shown in bold so that you can follow where they appear within the products.

The water molecule is split into H and OH, and each part is added to the product molecules.

Condensation

Condensation reactions occur when two molecules combine to form one molecule. They are the opposite of hydrolysis reactions. In most biochemical condensation reactions, water is produced when two smaller molecules join to produce a larger molecule. This is also known as dehydration synthesis because water is removed ("dehydrated") in the synthesis of the new molecule.

Here's a simple condensation reaction that occurs all the time in your body:

Amino acid + amino acid → protein + water

In this case, water is a product, rather than a reactant. This condensation reaction occurs in your body when you are building muscle tissue.

Notice how the atoms of the water molecule are rearranged in a condensation reaction, as shown in this simple reaction of:

Methanol + methanol → dimethyl ether + water

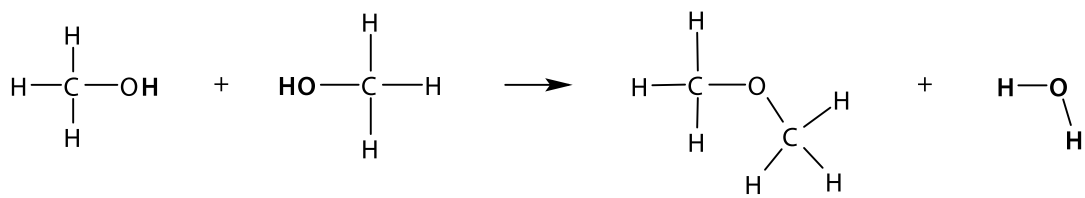


Figure 1.13: Condensation reaction of two methanol molecules to produce dimethyl ether and water. The atoms used to form the water molecule are shown in bold, so that you can follow where they appear within the reactants and products.

Oxidation and Reduction (Redox)

Many chemical reactions involve only the transfer of one or more electrons from one reactant to another, rather than the transfer of atoms or molecules. The process of losing electrons is called oxidation and the process of gaining electrons is called reduction. An electron transfer between two substances always involves a reduction of one and an oxidation of the other, and so is commonly called a redox reaction.

Consider the following redox reaction between zinc and copper:

$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$

Notice what happens to the reactants in this equation. The zinc atoms in solid form *lose* two electrons to form zinc ions in aqueous form ($Zn_{(s)} \rightarrow Zn^{+2}_{(aq)}$). Meanwhile, the copper ions in aqueous form *gain* two electrons to form copper atoms in solid form ($Cu^{2+}_{(aq)} \rightarrow Cu_{(s)}$).

One easy way to remember what happens in a redox reaction is to learn the expression:

"LEO the lion says, 'GER',"

which stands for:

"**L**ose **E**lectrons **O**xidation" and "**G**ain **E**lectrons **R**eduction."

Cellular respiration is an important example of a redox reaction in biology, as you will learn in a later lesson. The overall reaction combines glucose and oxygen to produce carbon dioxide, water, and energy:

$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O + \text{energy}$

The transfer of electrons from glucose helps the cell to make high-energy molecules (ATP), which it uses to do work.

Neutralization

Neutralization reactions involve the reaction of an acid and a base to produce water and a salt. These reactions occur continually within your cells, as well as in your stomach and intestines, for example. Before examining the details of neutralization reactions, it is helpful to review the basic chemistry of acids and bases.

Properties of Acids and Bases

Pure water contains H_2O molecules, as well as a tiny number of H^+ and OH^- ions, through a natural process of ionization.

$H_2O \rightarrow OH^- + H^+$

The ions in the water create the properties of an acid and a base.

The hydrogen ion (H^+) gives rise to the following characteristics of an acidic solution:

- It has a sour taste.
- It conducts electricity.
- It turns litmus paper red.
- It has a pH of below 7.

The hydroxide ion (OH^-) gives rise to the following characteristics of a base:

- It has a bitter taste.
- It has a slippery feel.
- It turns litmus paper blue.
- It has a pH of above 7.

Acids are substances that, when dissolved in water, increase the concentration of the hydrogen ions. Bases are substances that, when dissolved in water, increase the concentration of hydroxide ions. Pure water is considered neutral because it contains an equal number of hydroxide and hydrogen ions.

pH Scale

The acidity of an aqueous solution may be expressed in terms of hydrogen ion concentration [H^+] in mol/L. The pH scale is a convenient method of expressing hydrogen ion concentration. The pH is measured on a logarithmic scale so that each increment in pH unit represents a tenfold difference:

$pH = -\log_{10}[H^+]$

The pH is measured on a scale that ranges from a low of 0 to a high of 14. Pure water is defined as having a pH of 7—exactly in the middle of the scale—because it is neither acidic nor basic; it is neutral. Acids have pH values of below 7, while bases have pH values of above 7. The stronger the acid, the lower its pH reading, and the stronger the base, the higher its pH reading. The scale shown in Figure 1.14 depicts the pH values of some common substances.

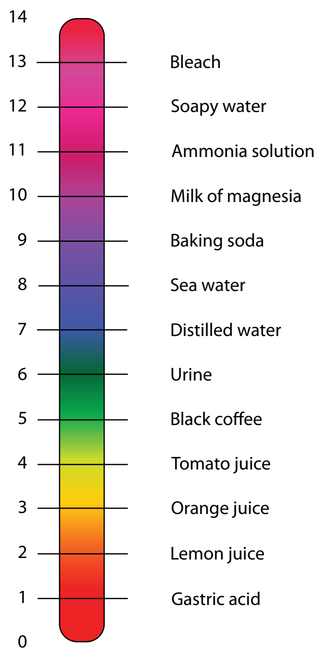


Figure 1.14: pH scale showing values for common household chemicals

Strong and Weak Acids and Bases

Acids and bases may be classified as strong or weak, depending on the degree to which they ionize when dissolved in water. Strong acids (for example, hydrochloric acid, HCl) and strong bases (for example, sodium hydroxide, NaOH) ionize completely when dissolved in water:

$$HCl_{(aq)} + H_2O \rightarrow H_3O^+ + Cl^-$$

$$NaOH_{(aq)} \rightarrow Na^+ + OH^-$$

Weak acids (for example, acetic acid, such as vinegar, CH_3COOH) and weak bases (for example, ammonia, NH_3) ionize only partially in water:

$$CH_3COOH_{(aq)} + H_2O \leftrightarrow CH_3COO^- + H_3O^+$$

$$NH_{3(aq)} + H_2O \leftrightarrow NH_4 + OH^-$$

Notice that double arrows are used to represent the chemical equations of weak acids and bases because these reactions are reversible, and so may proceed in both directions. This results in an equilibrium concentration of hydrogen or hydroxide ions. The pH of the solution they are in will determine how much the weak acids or bases will ionize, and so determine the equilibrium concentration. Most organic acids and bases are weak.

Neutralization Reaction

When an acid reacts with a base, a neutralization reaction occurs because water is produced from the H^+ and OH^- ions in solution. For example, sodium hydroxide (NaOH) and hydrochloric acid (HCl) react together to form the salt, sodium chloride (NaCl), and water:

$NaOH + HCl \rightarrow NaCl + H_2O$

The neutralization reaction is very important in biochemistry because it is used to regulate the pH of the internal environment, which determines the speed and direction of many critical biological reactions such as DNA synthesis. One way in which organisms use the neutralization reaction to regulate internal pH using a system of acid-base buffers, is described in the next section.

Acid-Base Buffers

Buffers are solutions that are able to maintain nearly constant pH levels, despite fluctuating environmental conditions. They do this by taking up excess hydrogen or hydroxide ions, thus neutralizing excess acid or base. There are many chemical combinations that act as buffers in the body, including proteins, amino acids, and carbonic acid. The example of the carbonic acid buffer that follows will be used to illustrate how buffers work.

Example

An important buffer in the human body is the carbonic acid-bicarbonate buffer. This buffer operates both in the blood and the extracellular fluid, to maintain an optimal pH of around 7.4. If the blood pH changes by more than about 0.2 to 0.4 pH units, it can be fatal, so this buffer is extremely important. This buffer contains both an acid and a base, so it can neutralize conditions that are either too acidic or too basic.

The buffer is formed by the reaction of carbon dioxide with water to form carbonic acid, which then ionizes to form bicarbonate and a hydrogen ion:

$H_2O + CO_2 \leftrightarrow H_2CO_3 \leftrightarrow HCO^-_3 + H^+$

Notice that all of the reactions are reversible, since they are represented by double arrows. All of these molecules are present in the blood in equilibrium concentrations that are determined by the pH.

If your blood pH were to drop a bit because you ingested some acid in the form of juice or vinegar, the buffer would react to remove the extra H^+ ions that are in your bloodstream by bonding the bicarbonate ions (HCO^-_3) to the H^+ ions to form more carbonic acid (H_2CO_3). Because carbonic acid is a weak acid, it tends to hold on to its H+ ions, thus lowering the concentration of H+ ions in the blood and raising its pH. If the blood pH were to increase because OH^- ions were added, then the H^+ ions would bond with the OH^- ions to form water, and the reduction in H^+ ions would be counterbalanced by a greater ionization of H_2CO_3.

Buffers are very effective at regulating pH levels, but they can become overwhelmed if too much acid or base were to enter the cell or bloodstream.

Support Questions

9. For each of the following, state the type of biochemical reaction (redox, condensation, hydrolysis, or neutralization) involved.

a) Glucose + galactose → lactose + water

b) 2 H^+ + 2 electrons → H_2

c) Pyruvate + $2H^+$ + 2 electrons → lactate

d) $HCl + NaOH \rightarrow NaCl + H_2O$

e) $Cu \rightarrow Cu^{2+}$ + 2 electrons

f) $KOH + HF \rightarrow H_2O + KF$

g) $CH_3COOCH_3 + H_2O \rightarrow CH_3COOH + CH_3OH$

h) Glycogen + water → many glucoses

i) Many amino acids → polypeptide + water

Suggested answer

10. For the following chemical equation:

$Fe + Cl_2 \rightarrow Fe^{2+} + 2\,Cl^-$

identify

a) what is being oxidized.

b) what is being reduced.

Suggested answer

Key Questions

Now work on your Key Questions in the online submission tool. You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

Total: 16 marks

1. How many electrons does Mg^{2+} have? **(1 mark)**

2. Given that the half-life of the radioisotope carbon-14 is 5730 years, how useful do you think this isotope would be for dating bones that are over a million years old? **(2 marks)**

3. Summarize the four major types of biochemical reactions studied in this lesson. For each type give its name, a word summary of what happens during the reaction, and an example of where the reaction might be biologically important. **(12 marks)**

4. For the following chemical reaction:

$$C_6H_{12}O_6 + 6\,O_2 \rightarrow 6\,CO_2 + 6\,H_2O + \text{energy}$$

identify

a) the substance(s) being oxidized.

b) the substance that is reduced.

(1 mark: ½ mark for each)

Now go on to Lesson 2. Send your answers to the Key Questions to the ILC when you have completed Unit 1 (Lessons 1 to 4).