CHAPTER 2
THE CHEMICAL CONTEXT OF LIFE

OUTLINE

I. Chemical Elements and Compounds
   A. Matter consists of chemical elements in pure form and in combinations called compounds
   B. Life requires about 25 chemical elements

II. Atoms and Molecules
   A. Atomic structure determines the behavior of an element
   B. Atoms combine by chemical bonding to form molecules
   C. Weak chemical bonds play important roles in the chemistry of life
   D. A molecule's biological function is related to its shape
   E. Chemical reactions make and break chemical bonds

OBJECTIVES

After reading this chapter and attending lecture, the student should be able to:

1. Define element and compound.
2. State four elements essential to life that make up 96% of living matter.
3. Describe the structure of an atom.
4. Define and distinguish among atomic number, mass number, atomic weight, and valence.
5. Given the atomic number and mass number of an atom, determine the number of neutrons.
6. Explain why radioisotopes are important to biologists.
7. Explain how electron configuration influences the chemical behavior of an atom.
8. Explain the octet rule and predict how many bonds an atom might form.
9. Explain why the noble gases are so unreactive.
10. Define electronegativity and explain how it influences the formation of chemical bonds.
11. Distinguish among nonpolar covalent, polar covalent and ionic bonds.
12. Describe the formation of a hydrogen bond and explain how it differs from a covalent or ionic bond.
13. Explain why weak bonds are important to living organisms.
14. Describe how the relative concentrations of reactants and products affect a chemical reaction.

KEY TERMS

matter	atomic weight	valence electron	polar covalent bond
element	isotope	valence shell	ion
trace element	radioactive isotope	chemical bond	cation
atom	energy	covalent bond	anion
neutron	potential energy	molecule	ionic bond
proton	energy level	structural formula	hydrogen bond
electron	energy	molecular formula	chemical reactions
atomic nucleus	potential energy	double covalent bond	reactants
dalton	energy level	valence	products
atomic number	electron shell	electronegativity	chemical equilibrium
mass number	orbital	nonpolar covalent bond	

LECTURE NOTES

I. Chemical Elements and Compounds

A. Matter consists of chemical elements in pure form and in combinations called compounds

Chemistry is fundamental to an understanding of life, because living organisms are made of matter.

Matter = Anything that takes up space and has mass.

Mass = A measure of the amount of matter an object contains.

> You might want to distinguish between mass and weight for your students. *Mass* is the measure of the amount of matter an object contains, and it stays the same regardless of changes in the object's position. *Weight* is the measure of how strongly an object is pulled by earth's gravity, and it varies with distance from the earth's center. The key point is that the mass of a body does not vary with its position, whereas weight does. So, for all practical purposes—as long as we are earthbound—weight can be used as a measure of mass.

B. Life requires about 25 chemical elements

Element = A substance that cannot be broken down into other substances by chemical reactions.

- All matter is made of elements.
- There are 92 naturally occurring elements.
- They are designated by a symbol of one or two letters.

About 25 of the 92 naturally occurring elements are essential to life. Biologically important elements include:

$$\left.\begin{array}{lll} C & = & \text{carbon} \\ O & = & \text{oxygen} \\ H & = & \text{hydrogen} \\ N & = & \text{nitrogen} \end{array}\right] \text{make up 96\% of living matter}$$

Ca = calcium		
P = phosphorus		
K = potassium		
S = sulfur		make up remaining 4% of an organism's weight
Na = sodium		
Cl = chlorine		
Mg = magnesium		
Trace elements		

Trace element = Element required by an organism in extremely minute quantities.

- Though required by organisms in small quantity, they are indispensable for life
- Examples: B, Cr, Co, Cu, F, I, Fe, Mn, Mo, Se, Si, Sn, V and Zn

Elements can exist in combinations called compounds.

- *Compound* = A pure substance composed of two or more elements combined in a fixed ratio.
- Example: NaCl (sodium chloride)
- Has unique emergent properties beyond those of its combined elements (Na and Cl have very different properties from NaCl). See Campbell, Figure 2.2.

> Since a compound is the next structural level above element or atom, this is an excellent place to emphasize the concept of emergent properties, an integral theme found throughout the text and course.

II. Atoms and Molecules

A. Atomic structure determines the behavior of an element

Atom = Smallest possible unit of matter that retains the physical and chemical properties of its element.

- Atoms of the same element share similar chemical properties.
- Atoms are made up of *subatomic particles.*

1. Subatomic particles

The three *most stable* subatomic particles are:

1. *Neutrons* [no charge (neutral)].
2. *Protons* [+1 electrostatic charge].
3. *Electrons* [-1 electrostatic charge].

NEUTRON	PROTON	ELECTRON
No charge	+1 charge	-1 charge
Found together in a dense core called the *nucleus* (positively charged because of protons)		Orbits around nucleus (held by electrostatic attraction to positively charged nucleus)
1.009 dalton	1.007 dalton	1/2000 dalton
Masses of both are about the same (about 1 dalton)		Mass is so small, usually not used to calculate atomic mass

NOTE: The *dalton* is a unit used to express mass at the atomic level. One dalton (d) is equal to 1.67×10^{-24} g.

If an atom is electrically neutral, the number of protons equals the number of electrons, which yields an electrostatically balanced charge.

2. Atomic number and atomic wight

Atomic number = Number of protons in an atom of a particular element.

- All atoms of an element have the same atomic number.
- Written as a subscript to the left of the element's symbol (e.g., $_{11}Na$)
- In a neutral atom, # protons = # electrons.

Mass number = Number of protons and neutrons in an atom.

- Written as a superscript to left of an element's symbol (e.g., ^{23}Na)
- Is approximate mass of the whole atom, since the mass of a proton and the mass of a neutron are both about 1 dalton
- Can deduce the number of neutrons by subtracting *atomic number* from *mass number*
- Number of neutrons can vary in an element, but number of protons is constant
- Is not the same as an element's *atomic weight*, which is the weighted mean of the masses of an element's constituent isotopes

In a large classroom with up to 300 students, it can be difficult to interact. Try putting examples on an overhead transparency and soliciting student input to complete the information. It is a quick way to check for understanding and to actively involve students.

Examples:

(Mass #) 23
(Atomic #) 11 Na

# of electrons ______

# of protons ______

# of neutrons ______

$^{12}_{6}C$

# of electrons ______

# of protons ______

# of neutrons ______

3. Isotopes

Isotopes = Atoms of an element that have the same atomic number but different mass number.

- They have the same number of protons, but a different number of neutrons.
- Under natural conditions, elements occur as mixtures of isotopes.
- Different isotopes of the same element react chemically in same way.
- Some isotopes are radioactive.

Radioactive isotope = Unstable isotope in which the nucleus spontaneously decays, emitting subatomic particles and/or energy as radioactivity.

- Loss of nuclear particles may transform one element to another (e.g., $^{14}_{6}C \rightarrow ^{14}_{7}N$).
- Has a fixed half life.
    - *Half life* = Time for 50% of radioactive atoms in a sample to decay.

Biological applications of radioactive isotopes include:

a. Dating geological strata and fossils

- Has a fixed half life.
  - *Half life* = Time for 50% of radioactive atoms in a sample to decay.

Biological applications of radioactive isotopes include:

a. Dating geological strata and fossils
   - Radioactive decay is at a fixed rate.
   - By comparing the ratio of radioactive and stable isotopes in a fossil with the ratio of isotopes in living organisms, one can estimate the age of a fossil.
   - The ratio of ^{14}C to ^{12}C is frequently used to date fossils less than 50,000 years old.

b. Radioactive tracers
   - Chemicals labelled with radioactive isotopes are used to trace the steps of a biochemical reaction or to determine the location of a particular substance within an organism (see Campbell, p. XX, Methods: The Use of Radioactive Tracers in Biology).
   - Radioactive isotopes are useful as biochemical tracers because they chemically react like the stable isotopes and are easily detected at low concentrations.
   - Isotopes of P, N, and H were used to determine DNA structure.
   - Used to diagnose disease (e.g., PET scanner)
   - Because radioactivity can damage cell molecules, radioactive isotopes can also be hazardous

c. Treatment of cancer
   - e.g., radioactive cobalt

4. The energy levels of electrons

Electrons = Light negatively charged particles that orbit around nucleus.

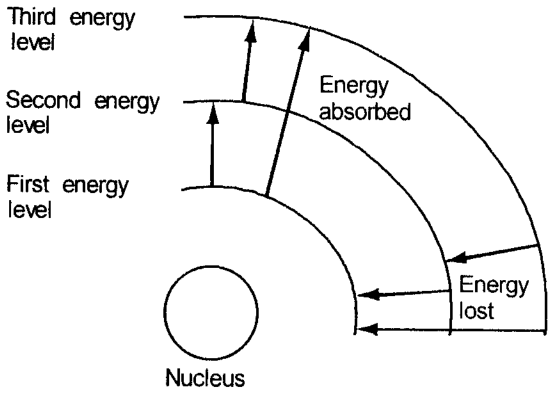
- Equal in mass and charge
- Are the only stable subatomic particles directly involved in chemical reactions
- Have *potential energy* because of their position relative to the positively charged nucleus

Energy = Ability to do work

Potential energy = Energy that matter stores because of its position or location.

- There is a natural tendency for matter to move to the lowest state of potential energy.
- Potential energy of electrons is not infinitely divisible, but exists only in discrete amounts called *quanta.*
- Different fixed potential energy states for electrons are called *energy levels* or *electron shells* (see Campbell, Figure 2.7).
- Electrons with lowest potential energy are in energy levels closest to the nucleus.
- Electrons with greater energy are in energy levels further from nucleus.

Electrons may move from one energy level to another. In the process, they gain or lose energy equal to the difference in potential energy between the old and new energy level.



5. **Electron orbitals**

Orbital = Three-dimensional space where an electron will most likely be found 90% of the time (see Campbell, Figure 2.8).

- Viewed as a three-dimensional probability cloud (a statistical concept)
- No more than two electrons can occupy same orbital.

First energy level:

- Has one spherical *s* orbital (1*s* orbital)
- Holds a maximum of two electrons

Second energy level

- Holds a maximum of eight electrons
- One spherical *s* orbital (2*s* orbital)
- Three dumbbell-shaped *p* orbitals each oriented at right angles to the other two ($2p_x$, $2p_y$, $2p_z$ orbitals)

Higher energy levels:

- Contain *s* and *p* orbitals
- Contain additional orbitals with more complex shapes

6. **Electron configuration and chemical properties**

An atom's electron configuration determines its chemical behavior.

- *Electron configuration* = Distribution of electrons in an atom's electron shells

The first 18 elements of a periodic chart are arranged sequentially by atomic number into three rows (periods). In reference to these *representative* elements, note the following:

- Outermost shell of these atoms never have more than four orbitals (one *s* and three *p*) or eight electrons.
- Electrons must first occupy lower electron shells before the higher shells can be occupied. (This is a reflection of the natural tendency for matter to move to the lowest possible state of potential energy—the most stable state.)
- Electrons are added to each of the *p* orbitals singly, before they can be paired.
- If an atom does not have enough electrons to fill all shells, the outer shell will be the only one partially filled. Example: O_2 with a total of eight electrons:

Two electrons have the 1*s* orbital of the first electron shell.

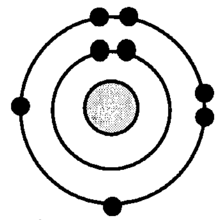
First two electrons in the second shell are both in the 2*s* orbital.

Next three electrons each have a *p* orbital ($2p_x$, $2p_y$, $2p_z$).

Eighth electron is paired in the $2p_x$ orbital.

OXYGEN

$_8O$



1*s*
2
2*s* $2p_x$ $2p_y$ $2p_z$
2 2 1 1

Chemical properties of an atom depend upon the number of valence electrons.

- *Valence electrons* = Electrons in the outermost energy shell (valence shell).

Octet rule = Rule that a valence shell is complete when it contains eight electrons (except H and He).

- An atom with a complete valence shell is unreactive or *inert.*
- Noble elements (e.g., helium, argon, and neon) have filled outer shells in their elemental state and are thus inert.
- An atom with an incomplete valence shell is chemically reactive (tends to form chemical bonds until it has eight electrons to fill the valence shell).
- Atoms with the same number of valence electrons show similar chemical behavior.

NOTE: The consequence of this unifying chemical principle is that the valence electrons are responsible for the atom's bonding capacity. This rule applies to most of the representative elements, but *not all.*

B. Atoms combine by chemical bonding to form molecules

Atoms with incomplete valence shells tend to fill those shells by interacting with other atoms. These interactions of electrons among atoms may allow atoms to form chemical bonds.

- *Chemical bonds* = Attractions that hold molecules together

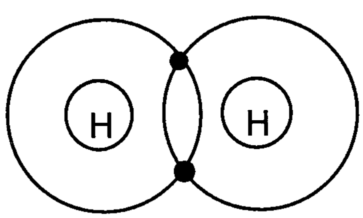
Molecules = Two or more atoms held together by chemical bonds.

1. Covalent bonds

Covalent bond = Chemical bond between atoms formed by *sharing* a pair of valence electrons.

- Strong chemical bond
- Example: molecular hydrogen (H_2); when two hydrogen atoms come close

H_2



H-H

Structural formula = Formula which represents the atoms and bonding within a molecule (e.g., H-H). The line represents a shared pair of electrons.

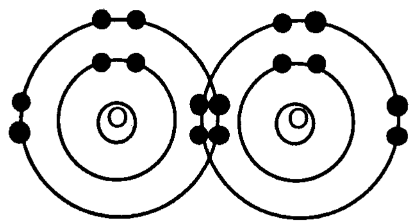
Molecular formula = Formula which indicates the number and type of atoms (e.g., H_2).

Single covalent bond = Bond between atoms formed by sharing a single pair of valence electrons.

- Atoms may freely rotate around the axis of the bond.

Double covalent bond = Bond formed when atoms share *two* pairs of valence electrons (e.g., O_2).

O_2



O=O

Molecules = Two or more atoms held together by chemical bonds.

Triple covalent bond = Bond formed when atoms share three pairs of valence electrons (e.g., N_2 or N°N).

NOTE: Double and triple covalent bonds are rigid and do not allow rotation.

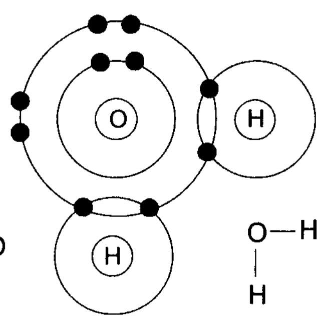
Valence = Bonding capacity of an atom which is the number of covalent bonds that must be formed to complete the outer electron shell.

- Valences of some common elements: hydrogen = 1, oxygen = 2, nitrogen = 3, carbon = 4, phosphorus = 3 (sometimes 5 as in biologically important compounds, e.g., ATP), sulfur = 2.

Compound = A pure substance composed of two or more elements combined in a fixed ratio.

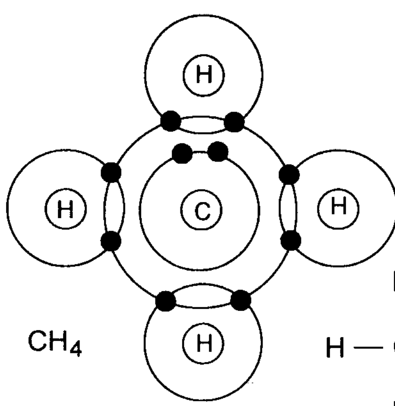
- Example: water (H_2O), methane (CH_4)
- Note that two hydrogens are necessary to complete the valence shell of oxygen in water, and four hydrogens are necessary for carbon to complete the valence shell in methane.

H_2O



O—H
|
H

CH_4



H
|
H—C—H
|
H

2. Nonpolar and polar covalent bonds

Electronegativity = Atom's ability to attract and hold electrons.

- The more electronegative an atom, the more strongly it attracts shared electrons.
- Scale determined by Linus Pauling:

  O = 3.5

  N = 3.0

  S and C = 2.5

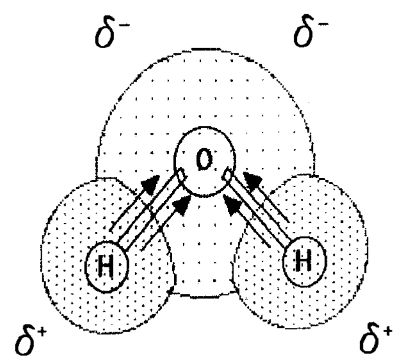
  P and H = 2.1

Nonpolar covalent bond = Covalent bond formed by an equal sharing of electrons between atoms.

- Occurs when electronegativity of both atoms is about the same (e.g., CH_4)
- Molecules made of one element usually have nonpolar covalent bonds (e.g., H_2, O_2, Cl_2, N_2).

Polar covalent bond = Covalent bond formed by an unequal sharing of electrons between atoms.

- Occurs when the atoms involved have different electronegativities.
- Shared electrons spend more time around the more electronegative atom.
- In H_2O, for example, the oxygen is strongly electronegative, so negatively charged electrons spend more time around the oxygen than the hydrogens. This causes the oxygen atom to have a slight negative charge and the hydrogens to have a slight positive charge (see also Campbell, Figure 2.11).



3. Ionic bonds

Ion = Charged atom or molecule.

Anion = An atom that has gained one or more electrons from another atom and has become negatively charged; a negatively charged ion.

Cation = An atom that has lost one or more electrons and has become positively charged; a positively charged ion.

Ionic bond = Bond formed by the electrostatic attraction after the complete transfer of an electron from a donor atom to an acceptor.

- The acceptor atom attracts the electrons because it is much more electronegative than the donor atom.
- Are strong bonds in crystals, but are fragile bonds in water; salt crystals will readily dissolve in water and dissociate into ions.
- Ionic compounds are called salts (e.g., NaCl or table salt) (see Campbell, Figure 2.13).

NOTE: The *difference* in electronegativity between interacting atoms determines if electrons are shared equally (nonpolar covalent), shared unequally (polar covalent), gained or lost (ionic bond). Nonpolar covalent bonds and ionic bonds are two extremes of a continuum from interacting atoms with similar electronegativities to interacting atoms with very different electronegativities.

C. Weak chemical bonds play important roles in the chemistry of life

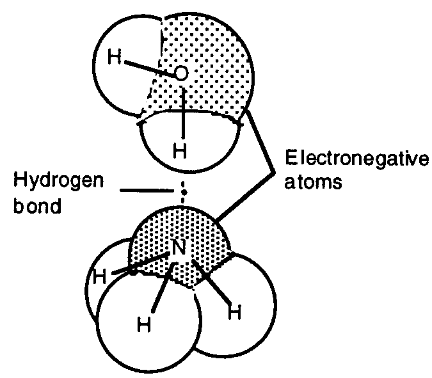
Biologically important weak bonds include the following:

- Hydrogen bonds, ionic bonds in aqueous solutions, and other weak forces such as Van der Waals and hydrophobic interactions
- Make chemical signaling possible in living organisms because they are only temporary associations. Signal molecules can briefly and reversibly bind to receptor molecules on a cell, causing a short-lived response.
- Can form between molecules or between different parts of a single large molecule.
- Help stabilize the three-dimensional shape of large molecules (e.g., DNA and proteins).

1. Hydrogen bonds

Hydrogen bond = Bond formed by the charge attraction when a hydrogen atom covalently bonded to one electronegative atom is attracted to another electronegative atom.

- Weak attractive force that is about 20 times easier to break than a covalent bond
- Is a charge attraction between oppositely charged portions of polar molecules
- Can occur between a hydrogen that has a slight positive charge when covalently bonded to an atom with high electronegativity (usually O and N)
- Example: NH_3 in H_2O (see Campbell, Figure 2.14)



2. Van der Waals interactions

Weak interactions that occur between atoms and molecules that are very close together and result from charge asymetry in electron clouds.

D. A molecule's biological function is related to its shape

A molecule has a charasteric size and shape.

The function of many molecules depends upon their shape

Insulin causes glucose uptake into liver and muscle cells of veterbrates because the shape of the insulin molecule is recognized by specific receptors on the target cell.

- Molecules with only two atoms are linear.
- Molecules with more than two atoms have more complex shapes.

When an atom forms covalent bonds, orbitals in the valence shell rearrange into the most stable configuration. To illustrate, consider atoms with valence electrons in the *s* and three *p* orbitals:

- The *s* and three *p* orbitals *hybridize* into four new orbitals.
- The new orbitals are teardrop shaped, extend from the nucleus and spread out as far apart as possible.
- Example: If outer tips of orbitals in methane (CH_4) are connected by imaginary lines, the new molecule has a tetrahedral shape with C at the center (see Campbell, Figure 2.15).

E. Chemical reactions make and break chemical bonds

Chemical reactions = process of making and breaking chemical bonds leading to changes in the composition of matter.

- Process where *reactants* undergo changes into *products*.
- Matter is conserved, so all reactant atoms are only rearranged to form products.
- Some reactions go to completion (all reactants converted to products), but most reactions are *reversible*. For example:

$$3H_2 + N_2 \rightleftharpoons 2NH_3$$

- The relative concentration of reactants and products affects reaction rate (the higher the concentration, the greater probability of reaction).

Chemical equilibrium = Equilibrium established when the rate of forward reaction equals the rate of the reverse reaction.

- Is a *dynamic* equilibrium with reactions continuing in both directions
- Relative concentrations of reactants and products stop changing.

Point out to students that chemical equilibrium does NOT mean that the concentrations of reactants and products are equal.

REFERENCES

Atkins, P.W. *Atoms, Electrons and Change*. W.H. Freeman and Company, 1991.

Campbell, N., et al. *Biology*. 5th ed. Menlo Park, California: Benjamin/Cummings, 1998.

Weinberg, S. *The Discovery of Subatomic Particles*. New York, San Francisco: W.H. Freeman and Company, 1983.

Brown, T.L., H. E. Le May, Jr., and B. Bursten. *Chemistry: The Central Science*. 7th ed. Upper Saddle River, New Jersey: Prentice Hall, 1997.