Appendix C: Basic Chemistry

Elements and Molecular Bonds

Elements in the body

The major elements in the body are summarized in Table C.1. Just six elements (hydrogen, oxygen, carbon, nitrogen, calcium, and phosphorus) make up about 99.5% of the elements in the body, and 98.5% of the weight of the body. A large percentage of the hydrogen and oxygen atoms combine to form water. Hydrogen and oxygen, together with carbon, and nitrogen combine to form most of the organic molecules (AKA biomolecules) of the body. Other elements, such as calcium and phosphorus combine to form bone. These and other elements, such as sodium, potassium, and chlorine, function in critical cellular processes, including communication and metabolism.

Table C.1. Major elements in the body (% of elements; % of mass)

Hydrogen (H)	(62%; 10%)	Calcium (Ca)	(0.22%; 1.4%)	Sulfur (S)	(0.038%)
Oxygen (O)	(24%; 65%)	Phosphorus (P)	(0.22%; 1.1%)	Magnesium (Mg)	(0.007%)
Carbon (C)	(12%; 18%)	Sodium (Na)	(0.037%)	Iron (Fe)	(0.0007%)
Nitrogen (N)	(1.1%; 3%)	Potassium (K)	(0.033%)	Iodine (I)	(0.0000007%)
		Chlorine (Cl)	(0.024%)		

As shown in Figure C.1, hydrogen (H), carbon (C), nitrogen (N), and oxygen (O) are found in the first two periods of the periodic table. Most of the other critical biological elements are found in the third and fourth period.

The property of an element depends on the number of protons and neutrons, and on the number and distribution of electrons in the particular orbitals of that element. The orbitals play a major role in determining the shape and characteristics of an element and of molecules.

The molecular weights (molar masses) for the elements in the first through fourth periods (rows) are also shown in Figure C.1

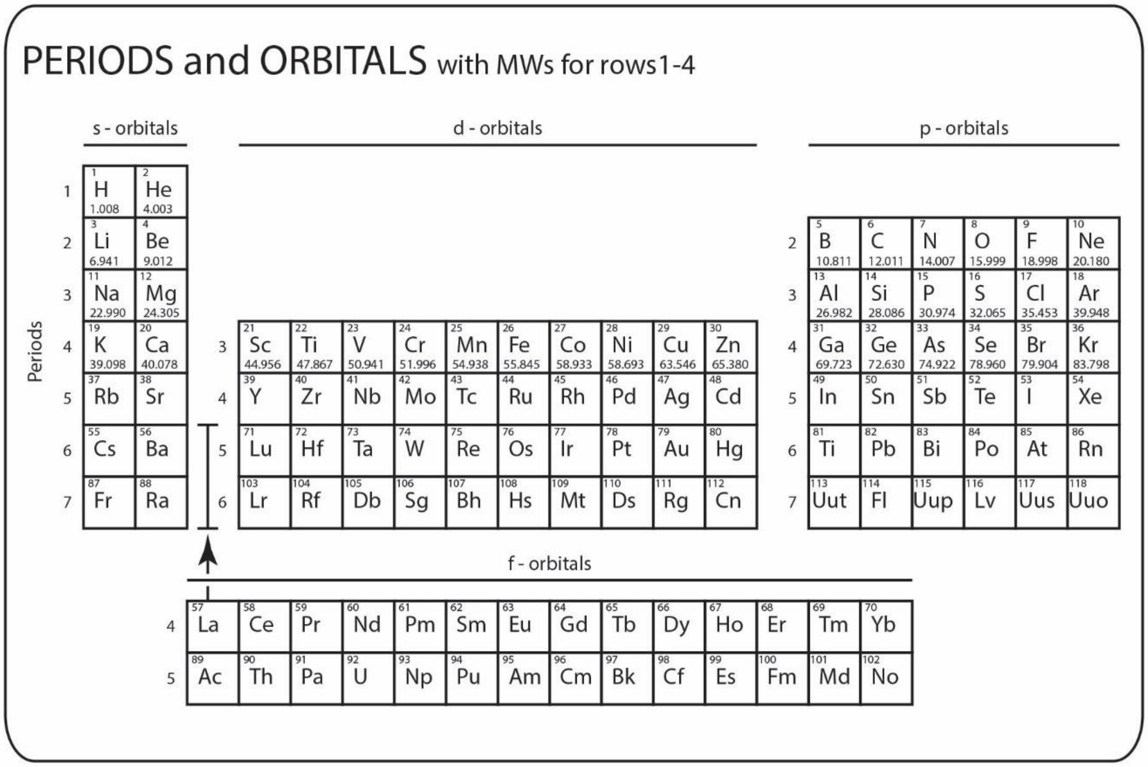


Figure C.1 © 2016 David G. Ward, Ph.D.

A brief explanation of orbitals is appropriate at this time. Orbitals are identified by their energy level (quantum number 1-7) and by the distribution of electrons (s, p, d, f). As shown in Figure C.2:

- The s-orbitals exist in <u>one</u> orientation and exhibit a spherical distribution of electrons. Each s-orbital holds 0, 1, or 2 electrons.
- The p-orbitals exist in <u>three</u> orientations and exhibit "dumbbell" distributions of electrons. Each orientation of a p-orbital holds 0, 1, or 2 electrons, for a total of up to 6 electrons (3x2).
- The d-orbitals exist in <u>five</u> orientations and exhibit "double dumbbell" distributions of electrons. Each orientation of a d-orbital holds 0, 1, or 2 electrons,, for a total of up to 10 electrons (5x2)

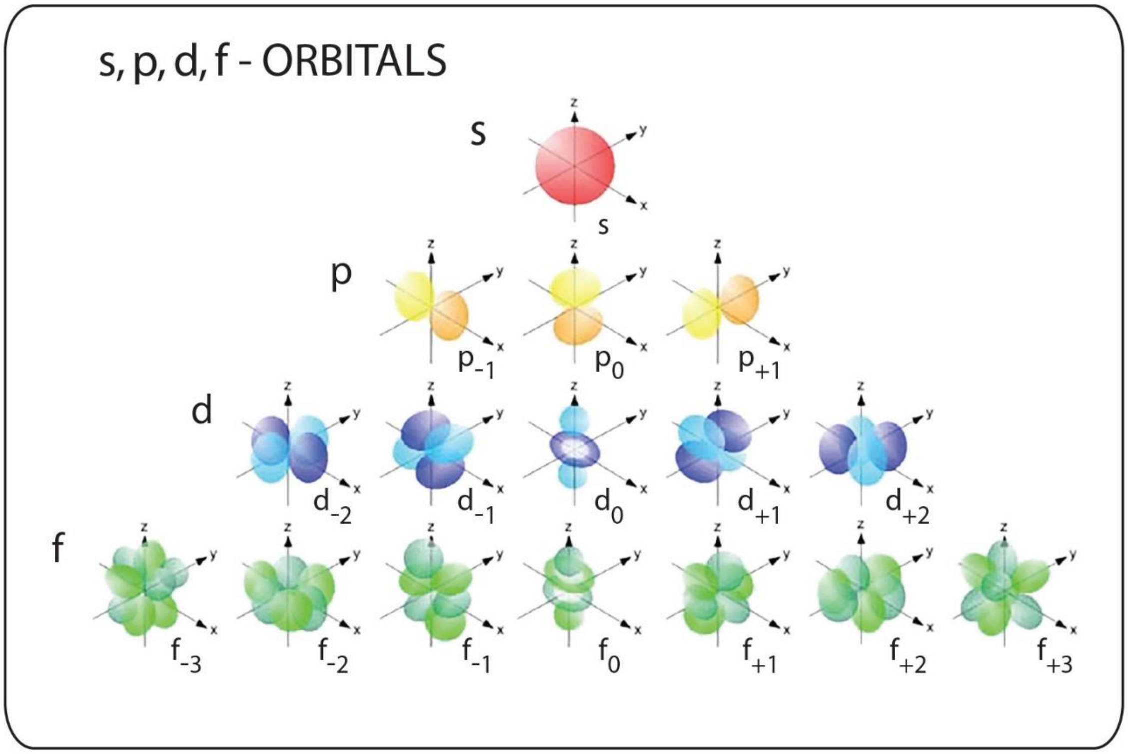


Figure C.2 © 2014 David G. Ward, Ph.D.

Going back to Figure C.1, we can see that:

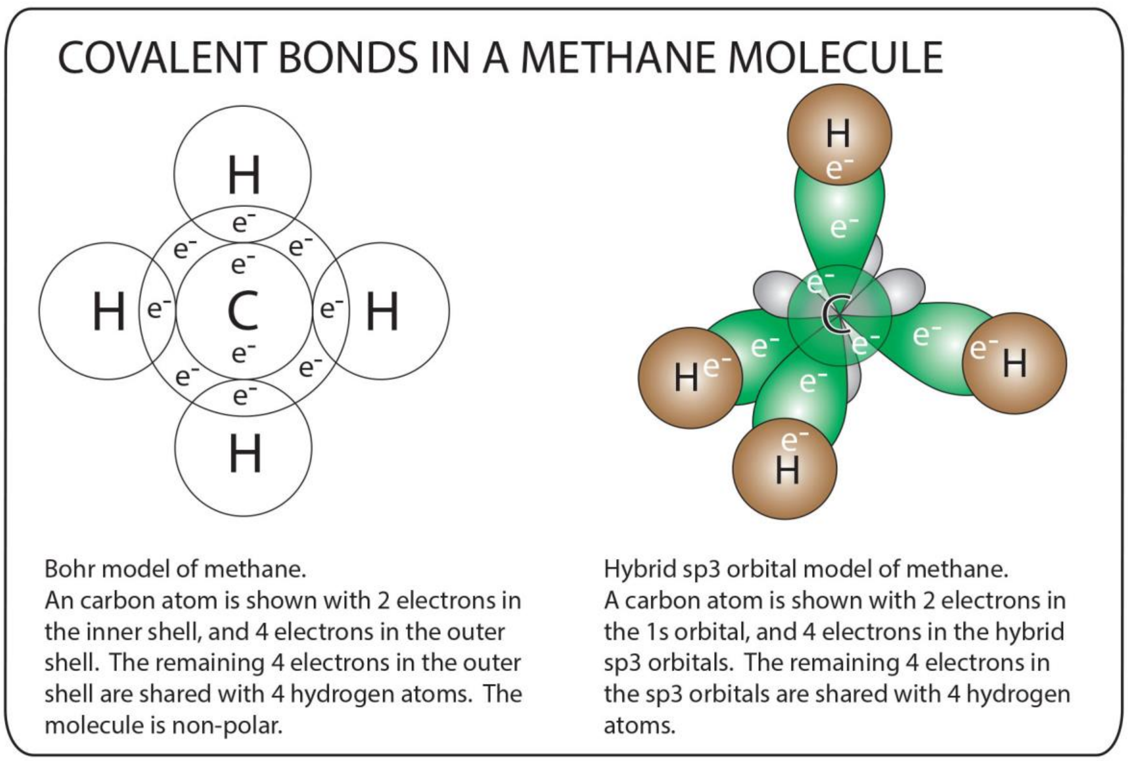
- Period 1 has just a 1s-orbital.
- Period 2 adds a 2s-orbital and three 2p-orbitals.
- Period 3 adds a 3s-orbital and three 3p-orbitals.
- Period 4 adds a 4s-orbital, five 3d-orbitals, and three 4p-orbitals

Molecular bonds

Two or more atoms bond together to form molecules. The bonding is determined generally by the number of electrons in the outer orbitals. Furthermore, the shapes of molecules is determined by the distribution of electrons in the orbitals. Finally, orbitals often hybridize in the process of bonding.

Covalent bonds form when electrons in the outer orbitals are shared between atoms. Sharing of electrons generally occurs between atoms with the same or similar electronegativity. Each participating atom gains filled orbitals.

When the electrons are shared equally, the atoms have no charge (non-polarized bonds). Molecules without polarized bonds are non-polar and **hydrophobic** (water insoluble, or only very sparingly soluble). Figure C.3 shows an example of non-polar covalent bonds within a methane molecule.



Bohr model of methane.
An carbon atom is shown with 2 electrons in the inner shell, and 4 electrons in the outer shell. The remaining 4 electrons in the outer shell are shared with 4 hydrogen atoms. The molecule is non-polar.

Hybrid sp3 orbital model of methane.
A carbon atom is shown with 2 electrons in the 1s orbital, and 4 electrons in the hybrid sp3 orbitals. The remaining 4 electrons in the sp3 orbitals are shared with 4 hydrogen atoms.

Figure C.3 © 2014 David G. Ward, Ph.D.

When electrons in a bond are shared unequally, the atom(s) with the greater attraction for electrons bears a partial negative charge, whereas the other atom(s) bears a partial positive charge (polarized bonds). Molecules with polarized bonds are polar and **hydrophilic** (water soluble). Polar molecules of biological importance contain one or more electronegative atoms, usually oxygen (O), nitrogen (N), sulfur (S), or phosphorus (P). Figure C.4 shows an example of polar covalent bonds within a water molecule.

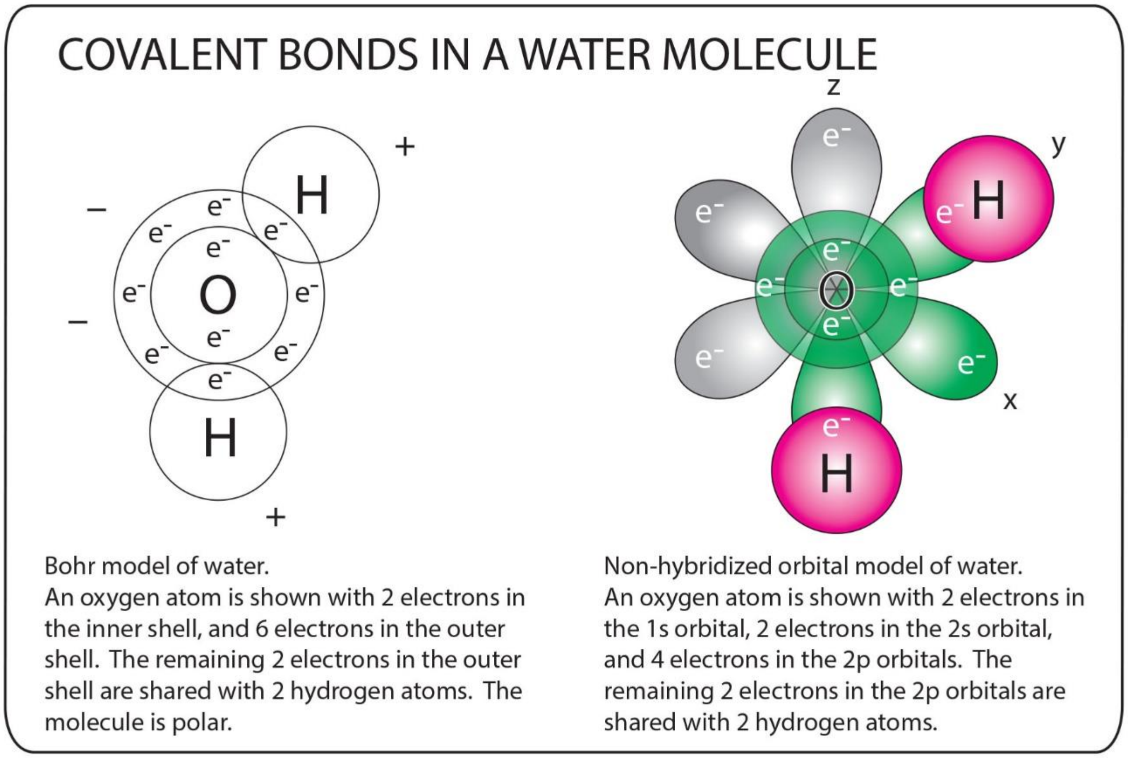


Figure C.4 © 2014 David G. Ward, Ph.D.

- **Covalent bonds** connect atoms through sharing of electrons between atoms.
  - Non-polar bonds – electrons are shared equally, leading to molecules that are hydrophobic (not water soluble, or only sparingly water soluble)
  - Polar bonds – electrons are not shared equally, leading to molecules that are hydrophilic (water soluble)

Ionic bonds form between atoms when electrons in the outer orbitals are completely transferred to another atom. Ionic bonding is an electrical attraction between molecules Transfer of electrons generally occurs when the electronegativity of the atoms is much different. Ionic bonds require at least one electron donor and one electron acceptor. Molecules with ionic bonds are polar and **hydrophilic** (water soluble). Figure C.5 shows an example of ionic bonds in sodium chloride.

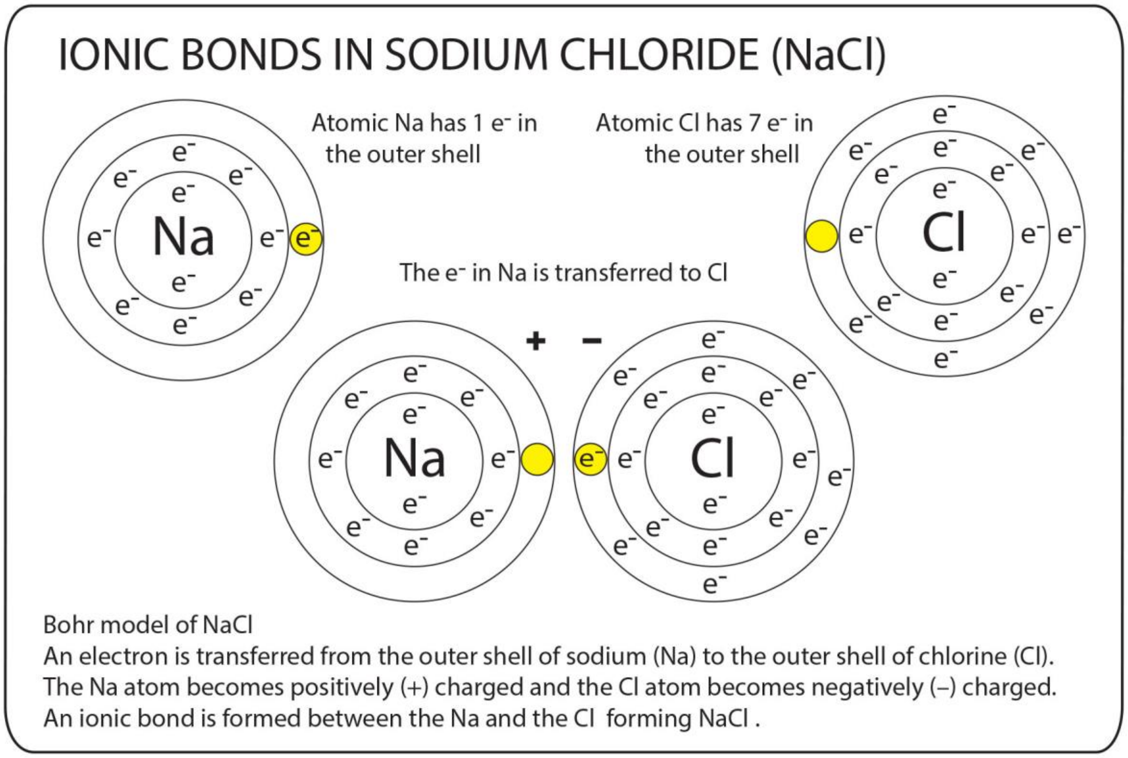


Figure C.5 © 2014 David G. Ward, Ph.D.

- **Ionic bonds** connect atoms through electrical attraction after one atom transfers one or more electrons to another atom, leading to molecules that are hydrophilic (water soluble).

Hydrogen bonds form between molecules that are covalently bonded together. Hydrogen bonding is an attraction that occurs between a hydrogen atom of one molecule (which bears a partial positive charge) and a nitrogen, oxygen or fluorine atom of another molecule (which bears a partial negative charge). Hydrogen and ionic bonds play a critical role in maintaining the structure of biological molecules and for allowing changes in the shape of the molecules. Figure C.6 shows an example of hydrogen bonds between water molecules.

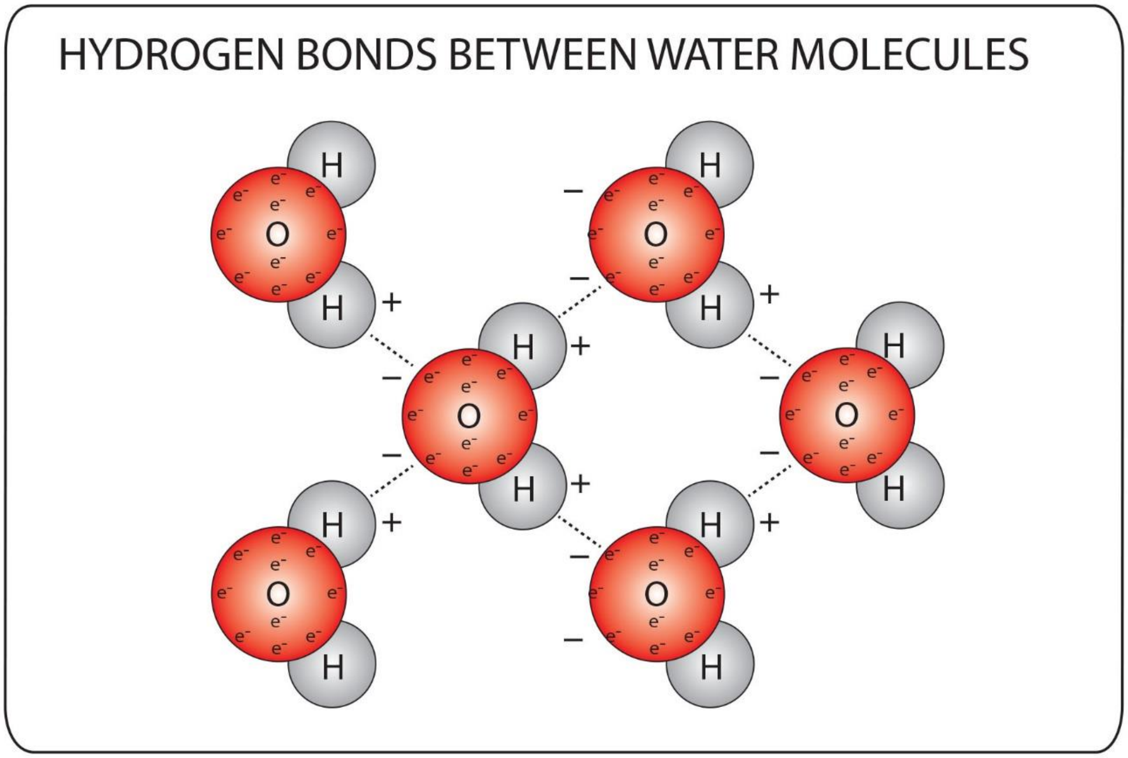


Figure C.6 © 2014 David G. Ward, Ph.D.

- **Hydrogen bonds** connect atoms through attraction between a hydrogen atom of one molecule and a negatively charged site of another molecule, such as with an oxygen, nitrogen, or fluorine atom.

Water and Electrolytes

Water

The covalent bonds of water molecules are highly polarized (refer to Figure C.6). As a result water is an excellent solvent capable of forming hydrogen bonds with virtually all polar molecules. Water is also a major determinant of the structure of biological molecules.

- The solvent property of water is due to the presence of polar molecules that lead to hydrogen bonds and that disrupt ionic bonds of solutes.
  - Substances that dissolve in water are **hydrophilic**.
  - Substances that do not dissolve in water are **hydrophobic**.
- The fluidity of water is due to the hydrogen bonds between water molecules that hold water together.

Electrolytes

Electrolytes are molecules that are held together by ionic bonds and dissociate in water. They are called electrolytes because they lead to electric charges in water. Table salt (sodium chloride, NaCl) is an example of an electrolyte. When electrolytes dissociate in water they give rise to charged ions. For example, sodium chloride dissociates into positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-). Positively charged ions are called **cations**, and negatively charged ions are called **anions.** Some major electrolytes and ions in the body are summarized in Table 1.3.

Table 1.3. Major electrolytes and ions in the body

Electrolytes	Corresponding Ion	Corresponding Ion
NaCl (sodium chloride)	Na^+	Cl^-
KCl (potassium chloride)	K^+	Cl^-
$CaCl_2$ (calcium chloride)	Ca^{2+}	$2Cl^-$
$NaHCO_3$ (sodium bicarbonate)	Na^+	HCO_3^-
$MgCl_2$ (magnesium chloride)	Mg^{2+}	$2Cl^-$
Na_2HPO_4 (disodium phosphate)	$2Na^+$	HPO_4^{2-}
Na_2SO_4 (sodium sulfate)	$2Na^+$	SO_4^{2-}

Molarity and Osmolarity

Molarity

Molarity (M) is defined as moles of solute ***molecules*** per liter of solution.

$$M = \text{molarity} = \frac{\text{moles of solute molecules}}{\text{liter of solution}}$$

A solution that is 1.0 molar (1.0 M) contains 1.0 mole of solute molecules per liter of solution.

The concentration of a solution is always given in terms of the form of the solute *before* it dissolves. For example, when a solution is described as 1.0 M NaCl, this means that the solution was prepared by dissolving 1.0 mole of NaCl molecules in enough water to make 1.0 liter of solution.

Osmolarity

Osmolarity (Osm) is defined as moles of *dissolved solute particles* per liter of solution.

$$\text{Osm} = \text{osmolarity} = \frac{\text{moles of dissolved solute particles (ions)}}{\text{liter of solution}}$$

When a solution is described as having a 2.0 molar osmolarity (2.0 Osm), this means the solution contains 2.0 moles of dissolved solute particles per liter of solution.

Many molecules will dissolve in water. However, ionic molecules, for example sodium chloride, *dissociate in water.* Therefore, one mole of NaCl molecules dissolved in water will produce a solution with 1.0 mole of Na^+ ions (particles) and 1.0 mole of Cl^- ions (particles).

$$1.0 \text{ M } Na^+ + 1.0 \text{ M } Cl^- = 2.0 \text{ M ions (particles)} = \text{osmolarity}$$

In contrast, covalent molecules, for example glucose, *do not dissociate in water*. Therefore, one mole of glucose molecules dissolved in water will produce a solution with 1.0 mole of glucose molecules (particles).

$$1.0 \text{ M glucose}^- = 1.0 \text{ M glucose (particles)} = \text{osmolarity}$$

The concentration of dissolved particles is often referred to as ***osmolarity.*** In human plasma the concentration of dissolved particles is about 0.290 M osmolarity (0.290 Osm). O*smolarity* of body fluids is often expressed in millimoles (mM) rather than in moles (M). Thus the osmolarity of blood plasma is expressed as 290 mM (290 mOsm).

pH and Buffers

pH

The **acidity** of a solution is determined by the hydrogen ion concentration ($[H^+]$). As the $[H^+]$ increases the solution becomes more acidic. The relationships among $[H^+]$ in moles, pH, and acidity are shown in Table 1.4. Using the molar $[H^+]$ is awkward. By convention, pH is expressed as the negative logarithm of the H^+ ion concentration in moles. As the $[H^+]$ increases, the solution becomes more acidic, and the pH number decreases.

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

Table 1.4. Relationships among H^+ ion concentration, pH and acidity

H^+ concentration ($[H^+]$)	pH	acid / base
0.1 M	1	strongly acidic
0.01 M	2	
0.001 M	3	
0.0001 M	4	
0.00001 M	5	
0.000001 M	6	weakly acidic
0.0000001 M	7	neutral
0.0000000398 M	7.4	blood plasma
0.00000001 M	8	weakly basic

- The pH of the blood must be maintained within a very tight range, between pH 7.35 and pH 7.45.

Buffers

Most biological processes are acutely sensitive to pH because changes in hydrogen ion concentration alter the ionic state of biological molecules. Therefore, buffers are extremely important in biological systems.

- An acid may be defined as a solute that dissociates and releases H^+ ions; as a proton donor; or as an electron pair acceptor.

  - Weak acids partially dissociate in water
  - Strong acids completely dissociate in water

- A base may be defined as a solute that dissociates and releases OH^- ions; as a proton acceptor; or as an electron pair donor.
- A conjugate base is formed by the removal of a H^+ ion (proton) from an acid
- A conjugate acid is formed by the addition of a H^+ ion (proton) to a base
- As a generalization, the stronger the acid or base, the weaker the conjugate. The weaker the acid or base, the stronger the conjugate.

Buffers reduce the magnitude of changes in pH by removing or replacing H^+ ions. Most buffers are composed of weak acids and/or their conjugate bases. Two of the most important biological buffers are the bicarbonate buffers and the phosphate buffers.

- bicarbonate buffers:

$$NaOH \text{ \{strong base\}} + H_2CO_3 \text{ \{weak acid\}} \rightarrow NaHCO_3 \text{ \{conjugate base\}} + H_2O$$

$$NaHCO_3 \text{ \{conjugate base\}} + HCl \text{ \{strong acid\}} \rightarrow NaCl + H_2CO_3 \text{ \{weak acid\}}$$

$$OH^- + H_2CO_3 \text{ \{weak acid\}} \rightarrow HCO_3^- \text{ \{conjugate base\}} + H_2O$$

$$H^+ + HCO_3^- \text{ \{conjugate base\}} \rightarrow H_2CO_3 \text{ \{weak acid\}}$$

- phosphate buffers:

$$OH^- + H_2PO_4^- \text{ \{weak acid\}} \rightarrow H_20 + HPO_4^{2-} \text{ \{conjugate base\}}$$

$$H^+ + HPO_4^{2-} \text{ \{conjugate base\}} \rightarrow H_2PO_4^- \text{ \{weak acid\}}$$

In the first example of a bicarbonate buffer, a sodium ion ($Na^{+)}$ from sodium hydroxide (NaOH, a strong base) is exchanged with a hydrogen ion (H^+) of carbonic acid (H_2CO_3, a weak acid) producing sodium bicarbonate ($NaHCO_3$) and water (H_2O). In the second example of a bicarbonate buffer, a sodium ion ($Na^{+)}$ from sodium bicarbonate ($NaHCO_3$, a weak base) is exchanged with a hydrogen ion (H^+) of hydrochloric acid (HCl, a strong acid) producing sodium chloride (NaCl) and carbonic acid (H_2CO_3), a weak acid). In the third example, a hydroxyl ion (OH^-) combines with H^+ from carbonic acid (H_2CO_3, a weak acid) to form a bicarbonate ion (HCO_3^-) and water (H_2O). In the fourth example, a hydrogen ion (H^+) combines with a bicarbonate ion (HCO_3^-) to form carbonic acid (H_2CO_3, a weak acid).

In the first example of a phosphate buffer, a hydroxyl ion (OH^-) combines with a hydrogen ion (H^+) from dihydrogen phosphate ($H_2PO_4^-$, a weak acid) producing water (H_2O) and hydrogen phosphate (HPO_4^{2-}, a conjugate base). In the second example of a phosphate buffer, a hydrogen ion (H^+) combines with hydrogen phosphate (HPO_4^{2-}, a conjugate base) to produce dihydrogen phosphate ($H_2PO_4^-$, a weak acid).