As indicated in our course introduction, much of Biology 101 emphasizes the study of cells, which structurally and functionally are an aggregate of atoms and molecules (chemicals) working together, and which require the **energy** of these chemicals to stay alive and to function.

Atoms and molecules combine in various ways (to be discussed) to form the structures of the cells and tissues of which living organisms are composed, and provide the energy to sustain these cells and tissues.

In our first unit of lectures we will discuss the basic structure of atoms and molecules to help us understand the biological concepts that will follow in this course.

Atoms

The atom is the fundamental unit of matter. And matter is any substance of the universe (gas, liquid, solid, plasma). More simply, matter is stuff or anything that has mass and occupies space. Living organisms are composed of matter. Matter can be changed from one form to another in a chemical reaction, a process in which different forms of matter combine or break apart.

There are 92 naturally occurring atoms on earth. We have about 108 total different kinds of atoms, because humans have been able to make atoms through nuclear reactions.

Each type of atom is composed of hundreds of smaller, subatomic particles, three of which we shall discuss. The proportion of these subatomic particles in any given atom identifies the kind of atom.

An **element** is a substance composed exclusively of one kind of atom. An element is a pure chemical that cannot be separated into or converted into a simpler substance. An atom is the smallest portion of an element that retains the properties of the element.

Each element (atom) has a name and a one- or two-letter abbreviation. The Periodic Table of Elements shows these, along with other useful information about each element. (See Appendix D of your Biology 101 Handbook.)

Although we have 92 different elements (formed from the 92 different kinds of atoms, there are just a few elements from which we are organized, and even fewer that are abundant in living organisms. We shall get acquainted with some of these.

The Arrangement and Properties of Atoms
Atoms are composed of hundreds of smaller **subatomic particles.** Fortunately, to understand Biology 101, we need only look at three of these basic particles, which are located in two regions of the atom: the **nucleus** and the **surrounding electron orbitals.** The "force" that holds these particles "together" forming the atom is an electrical charge (positive and negative) between subatomic particles.

The nucleus of an atom contains two sub-atomic particles: **protons** and **neutrons**. A third type of particle, **electrons,** are found in orbitals in motion surrounding the nucleus of the atom.

Let's look a little at the structure of the atom.

The Nucleus

- The central portion of the atom
- Contains two types of subatomic particles
- Contains the bulk of the atom's mass, but very little of the space of an atom

Particles in the Nucleus
Proton

- Positively charged particle ($^+$)
- Number of protons in the nucleus for a particular element is always constant (and fixed)
- The number of protons ranges from 1 ---> 108 from which the **atomic number** is assigned.
- An element is defined by its number of protons.
- Each proton has a relative mass of "1" (For biology purposes, this is fine. It's actually 1 dalton, which is a relative measure of the mass of the hydrogen atom).

Neutron

- No electrical charge
- Mass of "1" (just a tiny bit heavier than the proton; not enough for us to be concerned about).
- About the same number of neutrons as protons/atom, although the ratio changes with larger atoms...
- The number of neutrons plus the number of protons gives the **atomic mass** (or atomic weight) of an atom.

Isotopes

- While a typical element has a fixed number of protons and neutrons; many elements have forms with a number of neutrons, which differs from the typical number. These are called **isotopes**

e.g. Carbon (6p, 6n)
Carbon (6p, 7n)
Carbon (6p, 8n)

- Isotopes will have the same chemical properties.
- Since mass varies, some slight physical property differences occur.

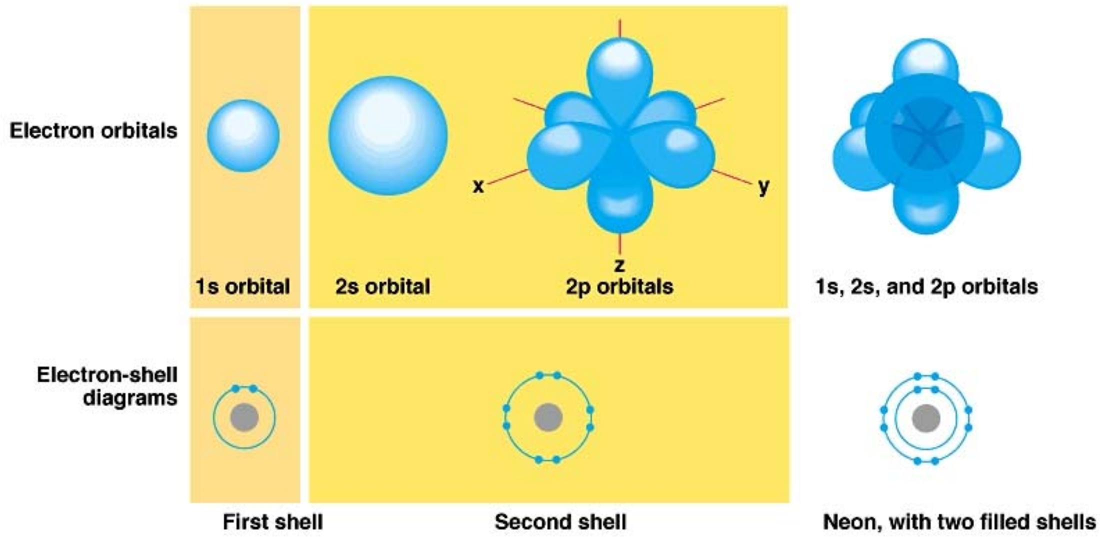
Unstable Isotopes

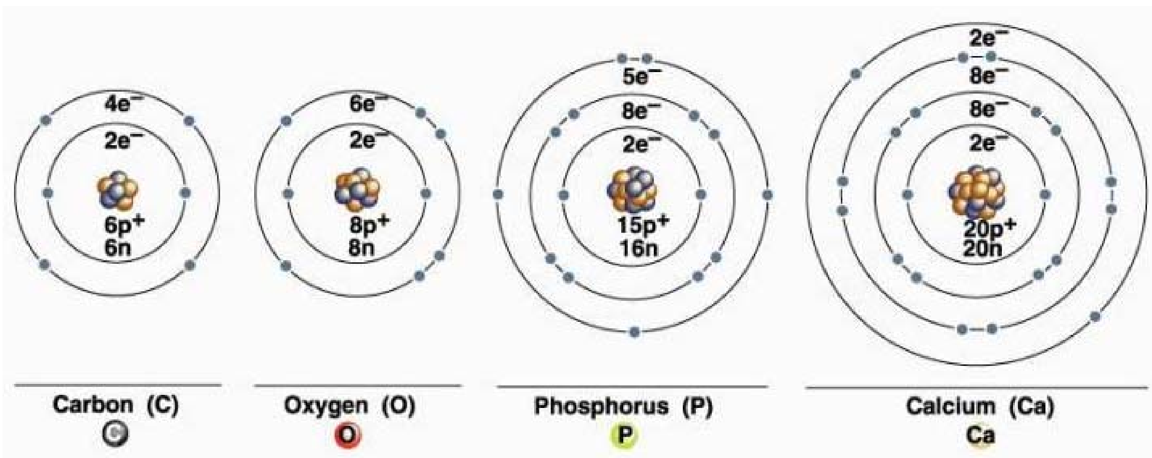
- Some isotopes have unstable nuclei, where one or more of the neutrons slowly decay, releasing energy in the form of beta particles. Eventually these neutrons decay into a proton (and electron, released as the beta particle).
- Such isotopes are called **radioactive isotopes** and the energy released is radiation.
  - The rate of decay varies for different isotopes.
  - Radioactive isotopes are important in biological research.
  - The rate of decay is measured by the half-life, the amount of time needed for 1/2 of the radioactive isotope to decay.

Electron Properties

The chemical properties of an atom are determined by the arrangement of its **electrons**.

- Negatively charged particle ($^-$)
- Very small mass, (1/2000 the mass of a proton) so small that we can think of an electron as having essentially no mass
- The number of electrons in a pure atom is the same as the number of protons
- Electrons are always in motion, found in **orbitals** located at fixed distances outside of the nucleus called **electron shells**, that correspond to different **energy levels**.
- Each electron orbital holds a maximum of 2 electrons. Each energy level or electron shell has a fixed number of orbitals, and each orbital has a precise pattern. For example, the first electron shell, closest to the nucleus, has one orbital, and it has a spherical pattern.
- The **orbital** of an electron is the most probable location where the electrons might be found. (Orbitals are very important in biology.)





As we look at a hypothetical atom, the electron fill pattern would be:

Shell	orbital(s)	# Electrons
1st	$(1)^2$ one	2
2nd	$(2)^2$ four	8
3rd	$(3)^2$ nine	18

Energy and Electrons

An electron in a given orbital has a characteristic energy. Each electron orbital has a different energy, so that electrons in different shells have different amounts of energy. The **energy level** and motion of electrons has much to do with the stability and reactivity of a given atom.

With reference to the energy of electrons, one might think of electron orbitals and shells like stairs. Electrons can be raised to a higher energy level with additional energy (climbing upstairs), and can release energy when (if) they fall to a lower energy level (going down the steps).

Electron Orbitals and the Stability of Atoms

Generally there are two rules for finding electrons in orbitals

1. Each electron orbital holds 2 electrons
2. Electrons occupy orbitals with the lowest energy level possible

Again, there are a variety of possible energy levels (shells) for the electrons of an atom to occur in, and within each energy level, a set number of **electron orbitals**. The energy levels and orbital patterns are specific characteristics of each type of atom, and are best left to chemistry classes for detailed discussion. But we can have a little lesson.

Electrons fill lower energy level shells first, and then progress to higher shells.

Interactions Between Atoms

The nucleus of the atom tends to provide stability, while electron shells permit interactions between atoms, called bonds. Nuclei of atoms are not affected by normal energy sources, whereas electrons are dynamic; bonds form when electrons from one atom are gained, lost or shared with other atoms. Such interactions are called chemical bonds.

When two or more atoms join together in a chemical bond they form a **molecule**. Chemical bonds are interactions that occur between the outermost energy level electrons of different atoms, details to follow. When molecules are formed from different atoms, they form a **compound**.

As mentioned, different electron orbitals have patterns that are unique and identified by shape. Further, an atom is most stable when its electron shells have **pairs of electrons** in each of the orbitals, and when the orbitals of its outermost energy level are filled.

Atoms that naturally have filled outer energy level orbitals are **non-reactive** and are found in nature as pure elements. (The noble gases were named that way because they were always found in the pure element state. It is not coincidence that the noble gases have filled outer energy levels.)

Most of the atoms naturally have numbers of electrons that do **not** result in filled outer energy level orbitals. These atoms are not stable and tend to undergo chemical reactions, or bonding, with other atoms to form **molecules** or **compounds**, which are more stable. Atoms that have similar electron configurations have similar properties and undergo similar chemical bonding. (The periodic table is organized according to these similar properties.)

By the way, there are some nice "rules" which help determine how an atom will bond. One of these is the **octet** rule. When the outer energy level of an atom has a total of 8 electrons it is especially stable, so most bonds take place to obtain eight electrons in the outer energy level.

Chemical Bonding
When an atom lacks a stable number of electrons In its outer energy level, it will share or transfer electrons to or from other atoms in very precise ways to achieve a stable number of electrons (again, generally 8) in its outer shell. This is the subject of **chemical bonding.** Note that a chemical bond is an energy relationship, involving electrons, and the energy that each electron has.

The energy all living organisms need to sustain life comes from making and breaking chemical bonds. It's vital that we understand how this works!

Types of Bonds
There are two types of common or strong bonds that occur between atoms:

Ionic bonds

One or more electrons are transferred from one atom to a second atom.

When an atom gains or loses one or more electrons, it becomes a charged atom, or an **ion**, because its number of electrons relative to the number of protons has changed.

Covalent bonds

Electron(s) of one atom are shared with electron(s) of a second atom.

Atoms that form covalent bonds do not carry a charge.

Let's looks a little at these two types of bonds.

Ionic Bonds

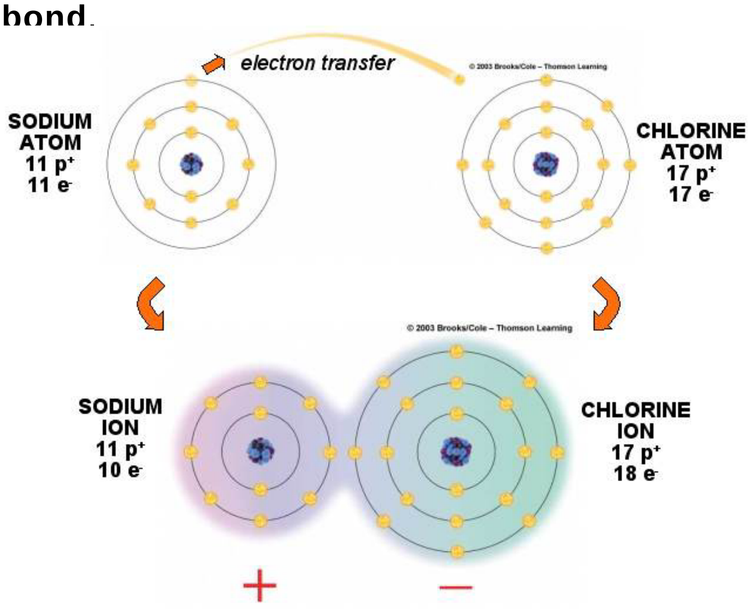
- An atom which has 1 or maybe 2 electrons in its outermost energy level (shell) may donate (give up) these electrons, which results in an outer energy layer with full orbitals.
- This action results in an atom with more protons than electrons, which results in a charged atom, an **ion.**

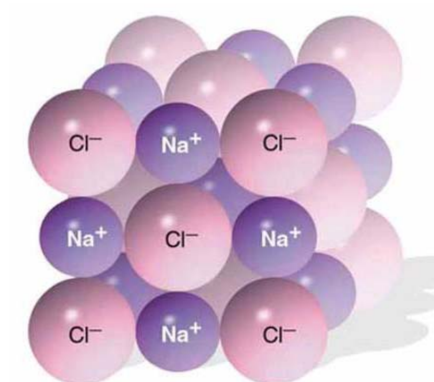
Example: $Na^{11p\ \ 11e}$ ---> $Na^{11p\ \ 10e}$ = $Na^{(+1)}$ ion

- A second atom may have 7 electrons in its outer shell, and may take on a donated electron to complete its orbitals.
- This too, produces a charged atom, or ion, but now one which is negatively charged.

Example: $Cl^{17p\ \ 17e}$ ---> $Cl^{17p\ \ 18e}$ = $Cl^{(-1)}$ ion

- The bond which forms occurs between the **charges** of the respective ions, forming a complex of positive and negative ions. This bond is called the **ionic bond**.





- Ions can be formed from atoms: or from molecules:

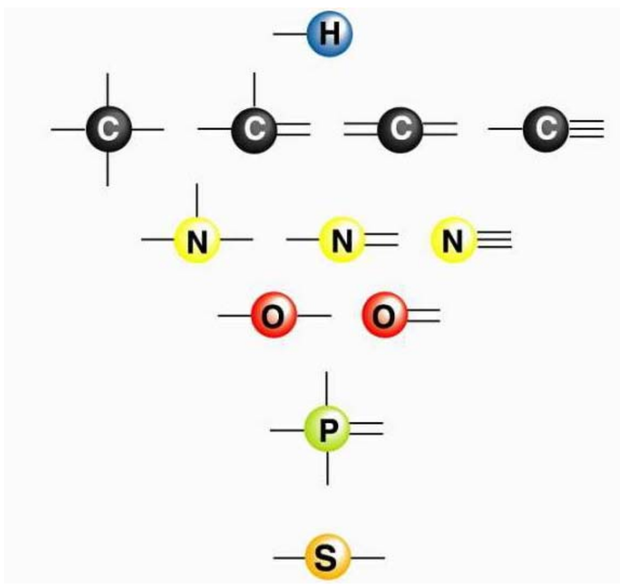
Na^{+}	$CO_3^{=}$
Cl^{-}	NH_4^{+}

Covalent Bonds

The major elements found in living organisms, **carbon, hydrogen, oxygen,** and **nitrogen**, tend to form covalent bonds. Our carbohydrates, proteins, lipids and genetic molecules are formed from these elements, as are the vitamins needed for living organisms. The exception to covalent bonding in the compounds found in living organisms are the minerals, or salts, which form ionic bonds.

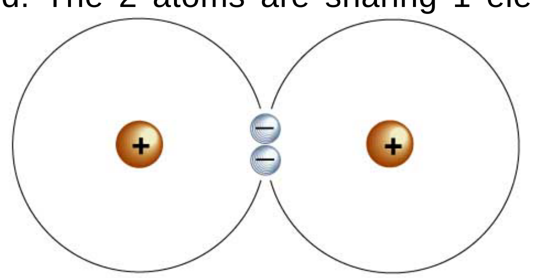
Covalent Bond Characteristics

- A covalent bond is formed when different atoms share one or more of their electrons with each other.
- This sharing causes each atom's orbitals to expand and alter its pattern to include the other atom (more or less). This determines the shape (bonding angles) of the molecule which forms.
- This collectively results in more stable orbitals for both atoms
- Covalent bonds form generally when a transfer of electrons would result in too great of a charge imbalance between the atom's positively charged nucleus and it's negatively charged electron field.
- Depending on the total number of electrons an atom has to share, an atom may covalently bond with one, or with several other atoms.



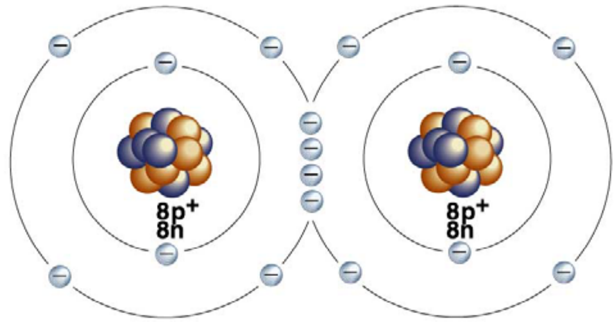
Variations in covalent bonds

Single covalent bond: The 2 atoms are sharing 1 electron pair



Hydrogen (H–H or H_2)

Double covalent bond: The 2 atoms are sharing 2 electron pairs



Oxygen (O = O or O_2)

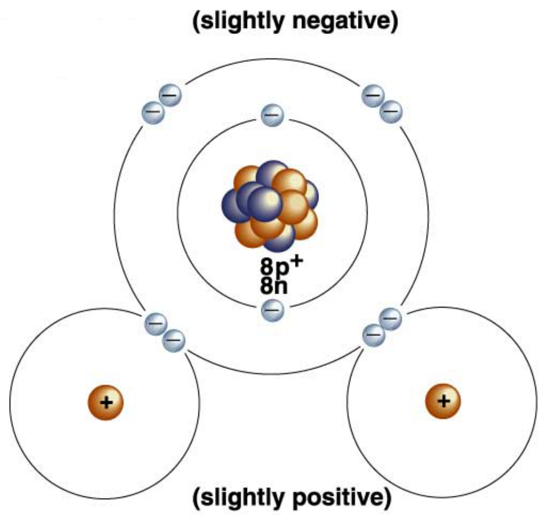
Triple covalent bond: The 2 atoms are sharing 3 electron pairs

Electronegativity and Polar and Non-Polar Bonds
As stated, in a covalent bond, the electrons of the bonding atoms are shared rather than transferred, so that no charged atoms or ions are formed. We also mentioned that atoms tend to form covalent bonds when a transfer of electrons would result in too great of a charge imbalance between the atom's positively charged nucleus and it's negatively charged electron field. All atoms, however, have a property, called **electronegativity**, which is a measure of how strongly the protons of the atom's nucleus attract and hold electrons. The electronegativity of all atoms is not the same.

Atoms that form ionic bonds typically have a reasonably high electronegativity. When covalent bonds are formed between atoms which have similar electronegativity, the electrons of the atoms are equally shared and the bond is said to be nonpolar (with no charge attraction).

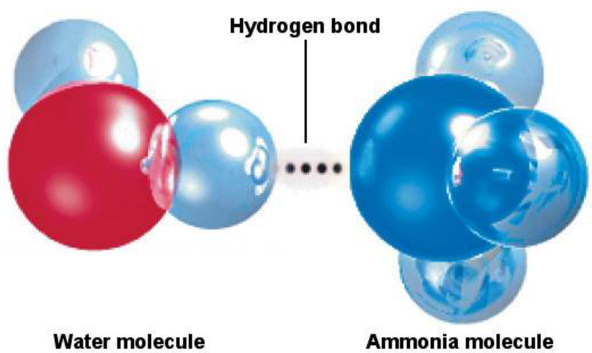
When covalent compounds are formed between atoms which have very different electronegativities, the electrons of the two atoms are not shared equally. The electrons that are spinning around both atoms will be attracted to the atom which has the stronger electronegativity, so that the electrons spend more time closer to the nucleus of that atom than in the field of the atom with a weaker electronegativity.

A compound (or molecule) formed by the unequal sharing will have one end functionally slightly **negative** (the end with the atom having a strong attraction for electrons) and the other end of the molecule slightly **positive,** resulting in a **polar covalent bond** and molecule.



Water (H – O – H or H_2O)

These slightly charged polar covalent molecules are attracted to other polar molecules (+ attracts -) and form very weak polar bonds. These bonds are weak because the charges are weak and because the electrons are always in motion. (Some of the time anyway, the end which is primarily positive will have the electrons spinning there and lose its polarity for that instant.) The bond that forms between adjacent polar covalent molecules is called a **Hydrogen bond**.



Hydrogen Bonding

- Attraction between adjacent polar molecules
- Common with the element hydrogen (which has a very low negativity) and high electronegativity atoms.

  e.g..: H-N

  H-O
- Hydrogen bonding is very important in the structure of:

  Water

  Proteins

  Nucleic acids

Other Weak Bonds
Hydrophilic (polar) and Hydrophobic (non-polar) molecules separate from each other in aqueous solutions because water is a polar molecule, and water molecules are attracted to other polar substances and repelled by non-polar molecules. Likewise, non-polar molecules are attracted to other non-polar substances and repelled by polar molecules. (Oil and water truly do not mix.)

The hydrophobic interactions of non-polar molecules in solution are reinforced by very weak interactions called van der Waals forces. These interactions are the result of electron motions and the brief attraction a polar molecule will have for another atom's electrons when adjacent to it. This attraction causes a brief charge reaction between the two molecules

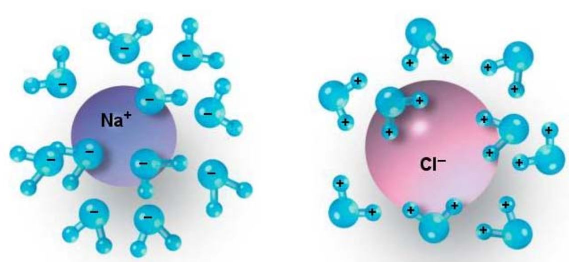
Water and Life on Earth
As we know, life on earth is based on the substance, **water**. Water is the most abundant compound found in living organisms (about 80%).

We have just seen that polar and non-polar molecules act differently in water. We have seen too, that water is one of the polar covalent molecules, and hydrogen bonds form between adjacent water molecules.

Lets look now at some properties of water, especially as these properties relate to water's polar nature and the phenomenon of hydrogen bonding.

Properties of Water
Solvent Properties

- Water is an excellent solvent (a fluid in which something can be dissolved) for many substances because of its polar nature.
- Many covalent molecules dissolve in water because covalent molecules often mix with other covalent substances.
- Polar substances and ions dissolve in water because opposite charges are attracted. Ions are attracted to the appropriate ends of water molecules which keep the ions dispersed in the water - or - dissolved.



- Strictly hydrophobic molecules, including most lipids, do not mix well with water
- Some molecules have both hydrophobic and hydrophilic ends. Such molecules are said to be **amphipathic**. Amphipathic molecules make good emulsifiers because they can attract both hydrophobic substances and hydrophilic substances to them.
- Substances dissolved in a solvent are called **solutes.**

Water in Biochemical Reactions

- The breakdown and assimilation of many molecules of living organisms involves water. Water is needed to breakdown carbohydrates, lipids and proteins during digestion. The formation of large biological molecules from smaller building blocks releases water.

Cohesion property of water

- A substance placed under sufficient tension will eventually rupture or be broken apart. Cohesion is the ability to resist rupture.
- The hydrogen bonds of water provide for good cohesion which results in a high surface tension for water.
- Cohesion allows for organisms to "walk on water surfaces" and allows for water to be drawn up to the tops of plants, even the tallest trees.


Adhesion property of water

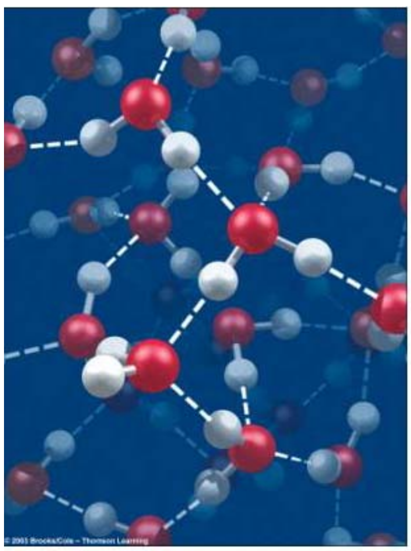
- Water molecules are often attracted to other polar substances, a property known as adhesion. This property provides for capillary action, the manner in which water "creeps" up the surfaces of tiny tubes, or "wicks" up paper and other surfaces.

Temperature moderating property of water

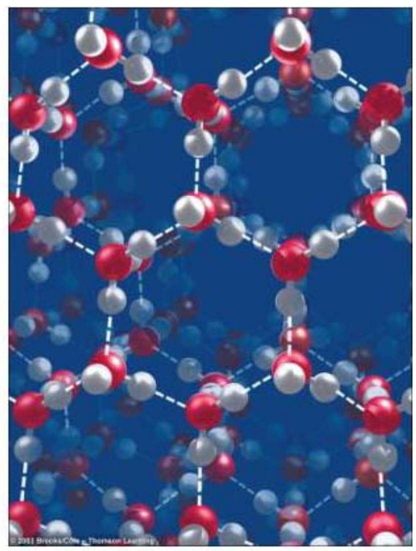
- Temperature is a measure of the rate of molecule movement of some piece of matter
- Heat is a measure of the amount of energy which results from the movement of molecules in some piece of matter.
- Because water has hydrogen bonding it has:
  - A high specific heat, the amount of energy needed to increase the temperature of 1 gram of a substance 1 degree Celsius
  - A high heat of vaporization, the amount of heat energy needed to change a liquid to a gas (or evaporate)
  - A high heat of fusion, the energy needed to convert a substance from a liquid to a solid.

  For example:

- Molecules of liquid water, when heated, can not move faster until the hydrogen bonds are broken; therefore it takes more heat energy to increase the temperature of water.
- Because of hydrogen bonding, it takes more energy to convert water to steam at its boiling point. Since it takes lots of energy to evaporate, and that energy must come from the surrounding liquid water, evaporation is a cooling process.
- The constant forming and breaking of hydrogen bonds in water provides for greater movement of the water molecules at lower temperatures. Water molecules do not form a rigid structure until 0 degrees Celsius is reached.
- Water molecules are densest at 4 degrees Celsius, which means that solid water is less dense than liquid water, or ice floats. (This is also because of hydrogen bonding.)



Liquid water



Ice

What do these temperature properties mean for living organisms?

- Water provides a more stable environment, in terms of temperature, both inside and outside of cells, in a world of fluctuating temperatures.
- Life can occur in lakes in temperate areas, since the frozen ice at the top of the lake in winter serves to insulate the water below. Also, if ice sunk, the organisms might find themselves out of a home, flopping on the surface of a solid block of ice...
- The process of evaporation is used by many organisms to maintain appropriate body temperature. (Sweating, for example)

Some special properties of water: Acids, Bases, Salts and Buffers
Recall that ionic bonds dissociate in solution forming ions, but covalent compounds resist dissociation. This is essentially true. However, some polar covalent compounds, including water, do dissociate, and it is important to discuss the meaning of this as it relates to living organisms, and their chemical environment.

A molecule of water is formed by the bonding of 1 oxygen with 2 hydrogen atoms. Oxygen has a very high electronegativity so that water is very polar, so polar that at any given instant, the attraction of the oxygen atom for electrons will literally draw an electron away from a hydrogen, or in fact, the **water molecule ionizes.**

At any time, a fixed proportion (specifically 10^{-7}) of a volume of water will be:

H^+ (Hydrogen ions)

OH^- (Hydroxide ions)

and the rest of the water will be: H_2O (Water molecules)

This phenomenon of water dissociation has bearing on a whole class of substances which contain hydrogen or hydroxide ions:

Acids and **Bases,** Plus **Salts.**

Acid

A substance which liberates H^+ in solution

Base or alkaline substance

A substance which combines with H^+ in solution

A substance which donates OH^- to solution

Salt

- An ionic compound which can be formed which an acid and base react. (Water is also formed along with the salt.)
- Salts dissociate into ions in solution.
- Many minerals needed for living organisms are salts. Some examples:
  - Ca^{++}
  - Fe^{++} or Fe^{+++}
  - K^+
  - Na^+
  - Cl^-

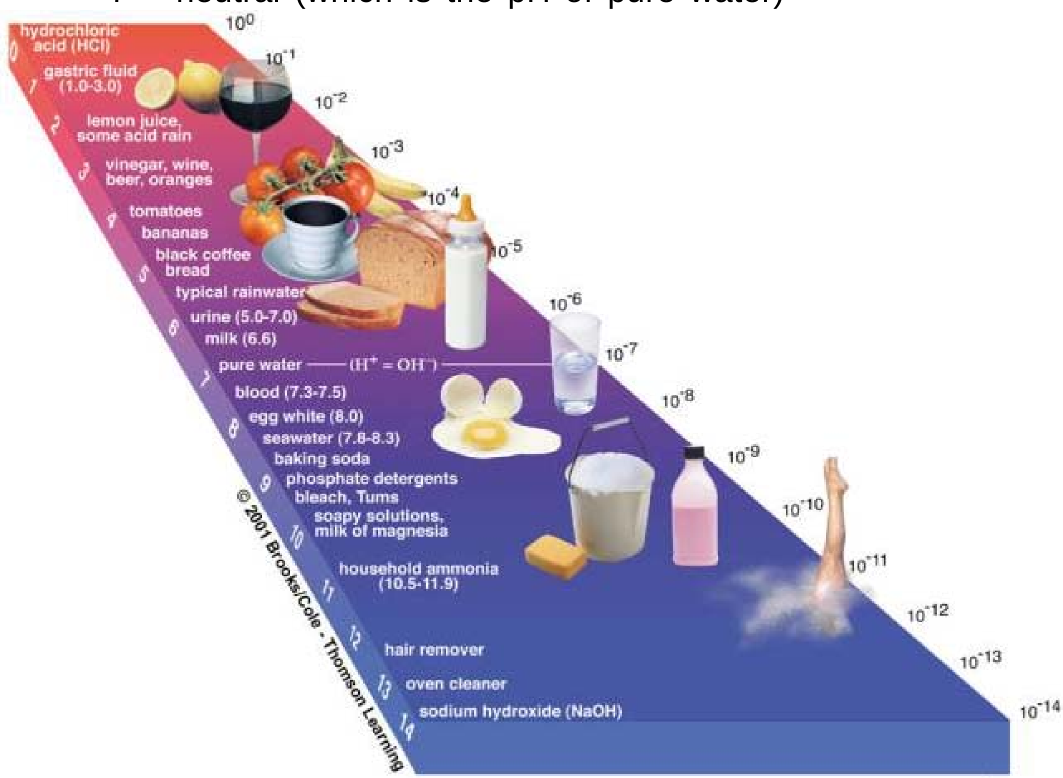
pH, Acids, Bases and Water
How do acids and bases relate to water and its dissociation, and what does this mean to living organisms? We'll answer that in just a minute, after discussing the phenomenon of pH (which is related to H^+ concentration).

Recall that at any given moment, 10^{-7} of any H_2O (Water) will be H^+ and OH^-. This ratio of H^+ to OH^- ions in water is used as a standard to measure the acidity of a substance.

This measurement of H^+ ions has been "translated" to a scale of pH (power of hydrogen).

NOTE: pH = the negative log of H^+ in the substance so that on the pH scale:

1 = maximum H^+, or the most acid.
14 = the most basic
7 = neutral (which is the pH of pure water)

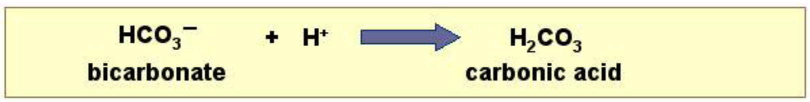


Now, why do we care?

Reactions of living organisms are very sensitive to levels of pH. It is therefore critical to maintain proper pH in an environment where cells and tissues are exposed to much variation in H^+ concentration.

How is this resolved? **Buffers**

- Group of materials (generally salts of weak acids and bases) that absorb or release H^+ depending on the condition so that proper pH is maintained.
- One of the most common buffers is the bicarbonate buffer.
    - If a solution is too acidic, bicarbonate can combine with H^+ to form carbonic acid
      ($HCO_3^- + H^+ \rightarrow H_2CO_3$)
    - If a solution is too basic, carbonic acid can combine with OH- to form bicarbonate + water
      ($H_2CO_3 + OH^- \rightarrow HCO_3^- + H_20$)



- Buffers are critical to the maintenance of life. Buffered systems mean that organisms can maintain a suitable pH environment in their cells and tissues.