

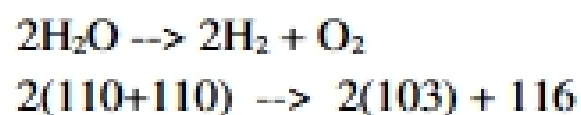
We have discussed several forms of energy - kinetic, potential, mechanical, thermal. Now we will add another to our list: chemical energy. You utilize chemical energy whenever you use a battery-operated device, burn fuel, eat food, or breathe oxygen.

The amount of energy stored in a molecule depends on the kinds of atoms in that molecule as well as their arrangement with respect to each other. For example, we can pass an electric current through water and break the water molecules into hydrogen and oxygen gas.



We can experimentally determine that this process requires a net input of energy equal to 118 kcal. This is because a system containing two molecules of hydrogen gas and one molecule of oxygen is more energetic (or less stable) than a system containing just two molecules of water. Not only does the system with hydrogen and oxygen gas contain three molecules compared to only two water molecules, a system containing molecules in the gas phase requires more energy than a system containing molecules in a "cooler" liquid phase. The reverse reaction - the formation of water from hydrogen and oxygen gas - will release 118 kcal of energy because that reaction proceeds from a more energetic to a less energetic state.

Why is 118 kcal the magic number for water? What is the number for other substances? By calculating exactly how much energy is required to decompose many different substances (as well as other experiments), researchers began to build a model of how molecules are able to store this energy. The result was the chemical bond model which states that the specific interactions and associations among atoms within a molecule have specific bond energies associated with them. In the case of water, the energy of each H-O bond is 110 kcal, each H-H bond is 103 kcal, and each O=O bond is 116 kcal. These numbers refer to the amount of energy required to break the bond and also the amount of energy released when the bond is formed. Two hydrogen atoms are much more energetic and less stable than one hydrogen gas molecule, so 103 kcal of energy is released when the two hydrogen atoms combine to form one molecule. Thus the balance sheet for the decomposition of water is

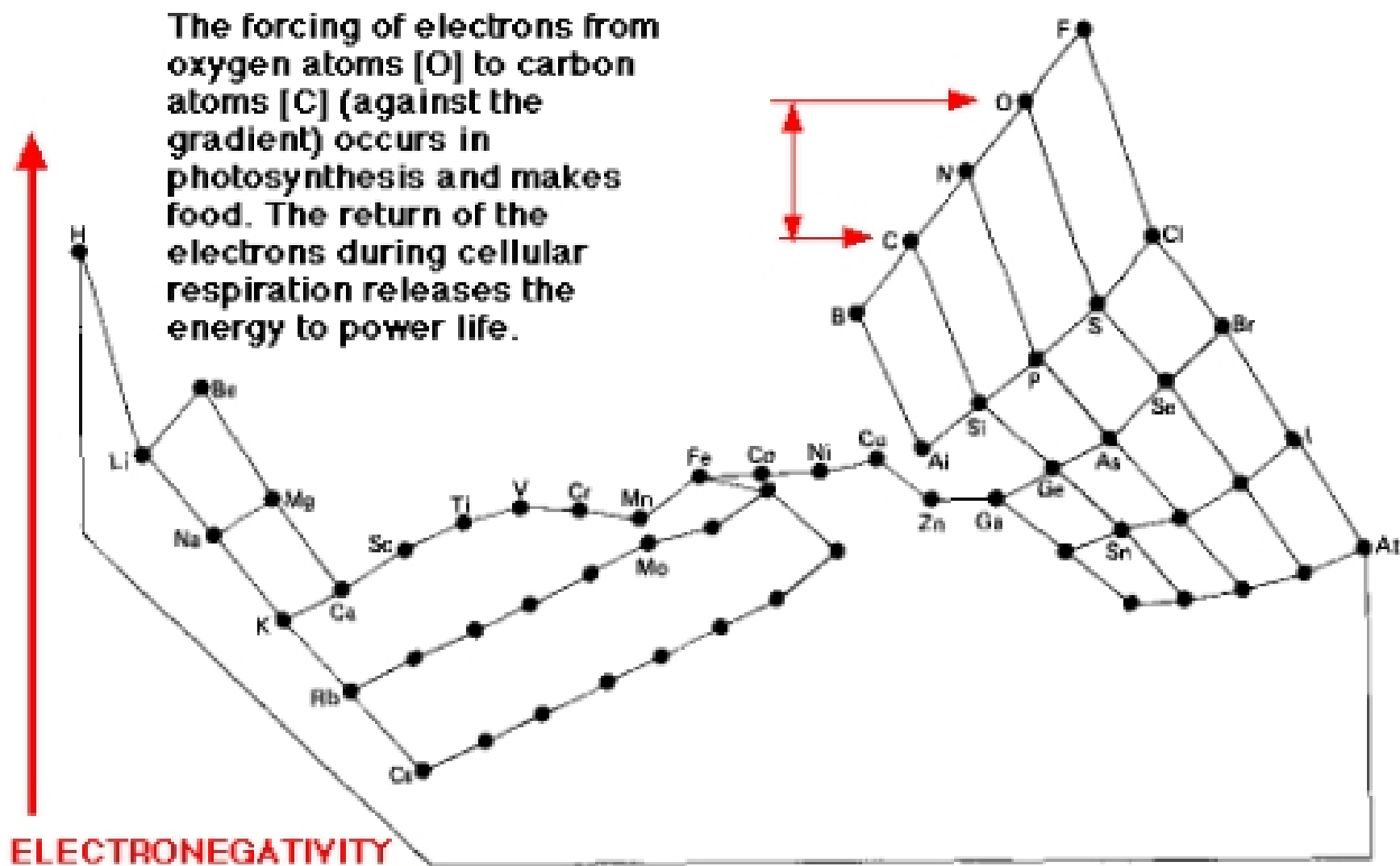


The total on the left side (the total input of energy required) is 440 kcal; the total on the right side (the total release of energy) is 322 kcal. The difference of these two numbers is 118 kcal, the net input of energy required for the reaction. If there is a net release of energy, such as in the reverse reaction, we call this free energy because it is available to do work. If you combined hydrogen gas with oxygen gas, the result will be an explosion releasing 118 kcal of

energy, which, theoretically, one could harness to do work. In real life, however, the conversion of chemical energy to work is very inefficient, as we will discuss later.

What exactly is a chemical bond? Nobody knows the answer to this question. For now, the concept of chemical bonds is a model to help us quantitatively explain the energy difference between a molecule and a set of free-floating atoms. We can develop the model in much more detail, however.

First of all, it is important to know that there are two main kinds of chemical bonds: ionic and covalent. Ionic bonds result from two atoms having very different electronegativities. You can think of electronegativity as the ability of an atom in a molecule to attract electrons to itself. See the table below, which is a modified version of the periodic table of elements:



<http://users.rcn.com/jkimball.ma.ultranet/BiologyPages/E/Electronegativity.html>

The most electronegative element is fluorine (F) followed by oxygen (O). This means that fluorine and oxygen atoms will attract electrons to themselves and away from other atoms in the molecule. If two atoms having very different electronegativities, for example fluorine and sodium (Na), interact, they will form an ionic bond. If the electronegativities are similar, for example hydrogen (H) and carbon (C), they will form a covalent bond, meaning that the electrons are shared equally between the two atoms.

If there is only a moderate difference in electronegativity, as occurs between hydrogen and oxygen, a polar covalent bond may form. This is well illustrated in the water molecule, where there are two H-O covalent bonds, but the electrons spend more time around the oxygen atom than the hydrogen atoms, resulting in a polar molecule having one positive end (or pole) and one negative end. If you have many water molecules, then the positive pole of one water molecule will interact with the negative pole of another molecule - this interaction is what is known as a hydrogen bond. I will draw this on the board. The hydrogen bonding of water molecules is responsible for many of water's unusual behaviors that make it absolutely essential to life on Earth.

One more concept you should be aware of is that of double or triple bonds. A single covalent bond involves the sharing of one pair of electrons. In the methane (natural gas) molecule, for example, 4 hydrogen atoms share one pair of electrons each with one carbon atom. This is the most stable arrangement of atoms because the carbon atom by itself has only 4 electrons and needs 4 more electrons from other atoms in order to get to its desired stable state of 8 electrons. Most of the atoms that we will be dealing with like to have 8 total electrons (except hydrogen which only needs two electrons). In the carbon dioxide molecule, carbon needs to get four extra electrons from just two oxygen atoms. The only way this can occur is if each oxygen atom shares two pairs of electrons with carbon so that each atom in the molecule ends up with 8 total electrons. Since two pairs of electrons are being shared, each carbon-oxygen bond in carbon dioxide is called a double bond and denoted with a double bar: C=O.

Below is a table of bond energies for common covalent bonds.

C-H	98 kcal/mole
O-H	110 kcal/mole
C-C	80 kcal/mole
C-O	78 kcal/mole
H-H	103 kcal/mole
C-N	65 kcal/mole
O=O	116 kcal/mole
C=O	187 kcal/mole