

Introduction to Chemistry for Exercise Physiology Students*

I. Matter

A. All matter we will consider is made of atoms. Atoms of course have two regions:

1. the **nucleus** which is composed of massive positively charged particles (**protons**) and also (except in the case of hydrogen) even slightly more massive neutrally charged (no charge) particles called neutrons. Nearly all of the mass of an atom is in the nucleus. The nucleus also determines (indirectly) the other characteristics of the atom since the more positive charges it has, the more electrons (see below) that the atom will normally possess.

2. a **cloud of electrons**. These nearly mass-less particles (compared to the nucleus) possess exactly one full negative charge each. Thus, one electron balances the charge on one proton. Atoms are usually neutrally charged overall so the number of electrons normally exactly equals the number of protons. This has several consequences.

a. The more protons, the more electrons and the larger the atom. Thus heavier atoms are larger, primarily because they have more electrons -- the nucleus takes up very little space compared to the electrons.

b. Since charges are usually balanced, atoms do not normally attract each other. From a distance, they have neither positive nor negative charge.

c. However, when we get really close to an atom, the negative charges of the electrons predominate. Since all atoms are surrounded by these electrons, when atoms approach each other they increasingly "see" a negative charge. The atoms increasingly repel each other at close distances. So, unless they approach each other really rapidly, they bounce away from each other without really touching. The atomic world is quite different from what we experience (we think). In fact, when two hockey players crash into each other and to the boards, however, they never really touch atoms!

3. The electron cloud determines the chemical properties of the atom. Chemical properties have to do with forming chemical combinations. Chemical combinations always involve the rearrangement of electrons between two different atoms. The way these re-arrangements can occur have to do with the numbers and arrangements of electrons around an atom.

4. Electrons can store energy. The energy that is stored or released in chemical reactions is carried by electrons. Although this analogy is not really accurate, nevertheless you may find it useful to think of electrons as objects in orbit around the nucleus. To move them further away from the nucleus

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takes energy just as moving a satellite further from the earth takes energy (we would use additional rocket power). On the other hand, moving a particle closer to the earth involves release of energy (the object is essentially falling and picking up speed -- one way to see the release of energy is the heat generated in slowing it down when it enters the atmosphere). Likewise, rearranging an electron's position to a lower orbit also releases energy. So, in chemical reactions, energy is stored or released by changing the orbits of electrons.

4. **elements**: these are atoms with different numbers of protons and therefore different masses and most importantly for us, different sizes and different numbers of electrons (and therefore different chemical properties). The elements we will commonly be concerned with in exphys are:

H -- hydrogen
C -- carbon,
N -- nitrogen
O -- oxygen
Na - sodium
K -- potassium
Ca -- calcium
Cl -- chlorine
Fe --iron
P --phosphorous

Know the symbols and names of each of these

6. **ions**: Under a variety of conditions, it is possible for an atom or molecule (see below) to lose or gain electrons. When this happens we have a **particle with one or more full units of charge**, this particle is called an **ion**. Ions of one type will strongly attract ions of the opposite charge. As always with charge the most stable situation is when all charges of one type (e.g., +) are completely balanced with charges of the opposite type -- so positive ions will tend to attract negative ions to their vicinity. Ions are especially important in biological systems since we will see that they have strong effects on the shape and function of the body's proteins.

a. Perhaps the most potent ion in this regard is the **hydrogen ion, H^+** . This ion is very, very small (essentially a proton). In this small size is one full positive charge. We will see that the small size and great "density" of charge allows it to have profound effects on proteins

b. Other ions we will commonly deal with are:
 Ca^{++} (or Ca^{2+})
 Na^+
 Cl^-

And a number of ionic molecules -- for example, HCO_3^- (bicarbonate), $-COO^-$ (carboxylic acid), $-NH_3^+$ (amine) etc. We'll learn these as needed -- but notice

that all of these are very large compared with hydrogen and have lower "densities" of charge.

B. Molecules and compounds

1. definitions:

a. a molecule is two or more atoms that are bonded together chemically. More about what a chemical bond is below. The formula for any molecule gives the elements present and their proportions. For example, the oxygen we breathe is actually a molecule of two atoms of oxygen and therefore has the formula O_2 .

b. a compound is a molecule that contains at least two different kinds of elements. Thus, oxygen gas, O_2 is not a compound. However, carbon dioxide (CO_2) is a compound.

c. Characteristics. Molecules and compounds all have different:

1. **Sizes** -- determined by molecular weight and numbers of atoms in the molecule.

2. **Shapes** -- determined by number and angles of chemical bonds.

3. **Charge Arrangements**-- Within compounds, it is common for electrons to be shared unevenly. The result is that different areas of the molecule may have positive or negative charges or partial charges. These charged areas have great effects on the molecule bearing them and on other molecules nearby. More about this below.

II. Chemical Bonds

A. The bonds that hold molecules together (chemical bonds) include two distinct types: ionic and covalent. In this section, we'll see what each type is and then we will see that a great many of the chemical bonds found in our bodies share some properties with each type of bond.

B. **Covalent bonds:** in a "pure" covalent bond, the two atoms in the bond share electrons equally. One way to think about this is to imagine the electrons orbiting around the two atoms and spending the exactly same amount of time around each atom. We say these bonds are **non-polar** since they show no hint of electrical charge. We'll come back to this in a moment.

C. **Ionic Bonds:** in contrast to a covalent bond, in an ionic bond one atom has taken one or more electrons from the other atom in the bond. The result is that the atom which lost electrons has one or more positive charges (it is a positive ion) while the one that gained the electrons has picked up one or more negative charges. The atoms are held together not by sharing but instead by the attraction of opposite charges. A good example of a substance made of ionic bonds is salt -- in a crystal of salt, sodium loses an electron to chloride. Thus, the sodium is a positively charged ion (Na^+) and the chlorine a negative ion (Cl^- -- called chloride). The electron taken from sodium spends all its time with the

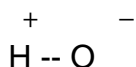
chlorine. Thus, a grain of salt is composed of positive and negative ions and is held together by electrical attraction. Ionic bonds are highly **polar** -- they are due entirely to the attraction of opposite electrical charges.

D. Bonds with covalent and ionic characteristics. Many of the most important bonds are neither truly ionic nor covalent. Instead they show some characteristics of each. In these bonds, electrons are shared but unevenly. Thus, one atom in the bond has the electron more of the time than the other (unlike a true covalent bond). On the other hand, neither atom totally monopolizes the electron (unlike in ionic bonds). Bonds like this are therefore covalent but they atoms in the bond also have **partial charge. These bonds are therefore polar, but not as strongly so as with an ionic bond.** How does this work?

1. The atom which tends to attract the electron most of the time -- called the **electronegative** atom, becomes slightly negatively charged. This is simply a consequence of the fact that an extra negative charge is spending a lot of time in its vicinity.

2. By contrast, the atom where the electron is spending less than half the time (the **electropositive** atom) will take on a partial positive charge. This is because one of its electrons is spending less time around it and more around the other (electronegative) atom.

3. The result is that the bond has a positive and a negative end. But the positive and negative are not a full charge -- this bond is not ionic because some sharing is still going on. Here is an example. When hydrogen bonds to oxygen to form water, oxygen tends to take hold of hydrogen's electrons. They spend more time (but not all time) around oxygen than hydrogen. The result is that the bond has a partial positive charge on the hydrogen end and a partial negative charge on the oxygen end:

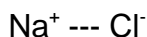


Where the δ^+ means a partial charge. Notice that the bond is **polar** since there is some difference of charge. Again, by contrast, in a covalent bond such as:



The electron is shared exactly equally and so neither atom of carbon is more or less positive or negative than the other. This bond is **not polar**.

And making the contrast with ionic bonds:



Where here there is a full charge difference between the two atoms. It is polar -- far more so than the first example.

E. For our purposes, you should remember that:

1. When atoms **bind with other atoms of their same element (example oxygen with oxygen)** the bond is always truly covalent and therefore **non-polar**.

2. You should also know that when **carbon binds to hydrogen, the bond is non-polar**.

3. Finally, all other bonds will be polar to some degree. Certain elements will tend to be negative in these bonds (since they tend to grab electrons -- they are electronegative) while others will tend to be positive (they tend to lose electrons and are therefore electropositive). Here are examples of each"

nitrogen, oxygen, chlorine are electronegative
hydrogen, carbon and sodium are electropositive

Thus:

C-C is non-polar

C-H is non-polar

O-H is somewhat polar with the oxygen being somewhat negative

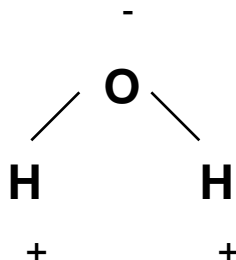
C-O is somewhat polar with the oxygen being somewhat negative

N-H is somewhat polar with the nitrogen being somewhat negative

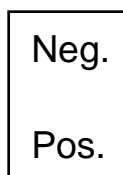
III. About Water and Solutions

A. You are about 75% water -- so its properties are very important in deciding how your body works.

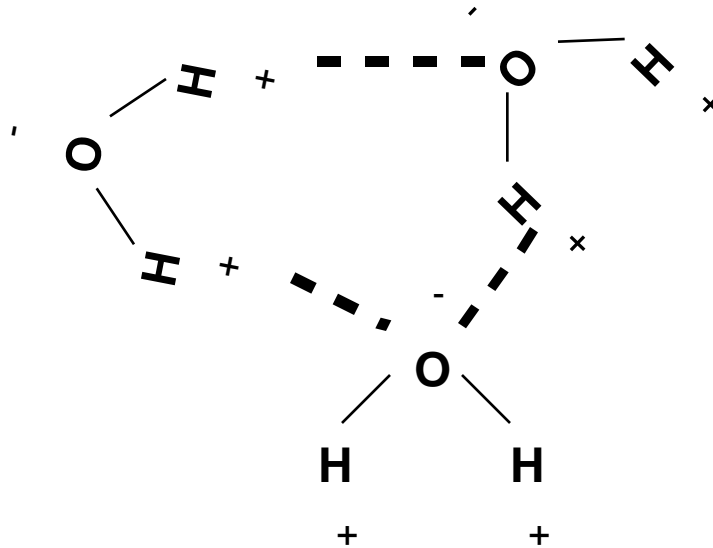
B. The water molecule. As seen previously, water is a polar molecule. It looks like this:



Since the oxygen is more electronegative than the hydrogen, the electrons are not shared equally and the result is a molecule that has a positive and a negative end:



Since each water molecule has a partial charge, it will attract others. They will tend to orient, negative end of one molecule towards the positive end of another:



where the dotted lines represent attractions between the molecules. These are actually weak bonds and are called **hydrogen bonds** or sometimes **weak polar interactions**. Notice that they give the water some structure -- water molecules are not absolutely free to slide past each other. Instead as they move around, they tend to stick to each other for a moment.

C. Interactions between water and non-polar substances. The **hydrophobic interaction**: Contrast the interactions between water molecules and what happens with a non-polar substance such as an oil. In an oil, the molecules do not interact except to bounce off each other. There is no structure.

1. What happens if we add a non-polar substance to water? Well we all know the results of mixing oil and water and so now is good time to explain the separation in terms of polar interactions.

2. When polar and non-polar substances are mixed -- for example, when water and oil are shaken together, the whenever the polar molecules in the mixture (water in this case) come close to each other, they attract each other. As mentioned above, they tend to remain relatively close to each other. On the other hand, there is no such interaction between the non-polar substance (oil in this example) and other non-polar molecules or with the water. These molecules might come close to each other and then "bounce off" (actually recall that they come close enough that their electron clouds repel each other).

3. So, the while the polar molecules tend to linger around each other and briefly form hydrogen bonds (see above), there is no such interaction with the non-polar molecules. The result is that the non-polar oil molecules are

slowly squeezed out the water by the interactions between the water molecules. The oil and water separates and the oil goes to the top because it is less dense than the water and floats on it.

4. This is the same reason that when you shake air into water the bubbles tend to get larger (air is generally made of N_2 and O_2 , both non-polar) and move to the top (air is less dense).

5. We give the lack of interaction between non-polar substances the rather unfortunate name of the **hydrophobic interaction** (why is this a bad name? -- first because it means "water hating" or water fearing and I doubt the non-polar molecule has any hate or fear and also because it is called an interaction when in fact the separation is due to the lack of any interaction).

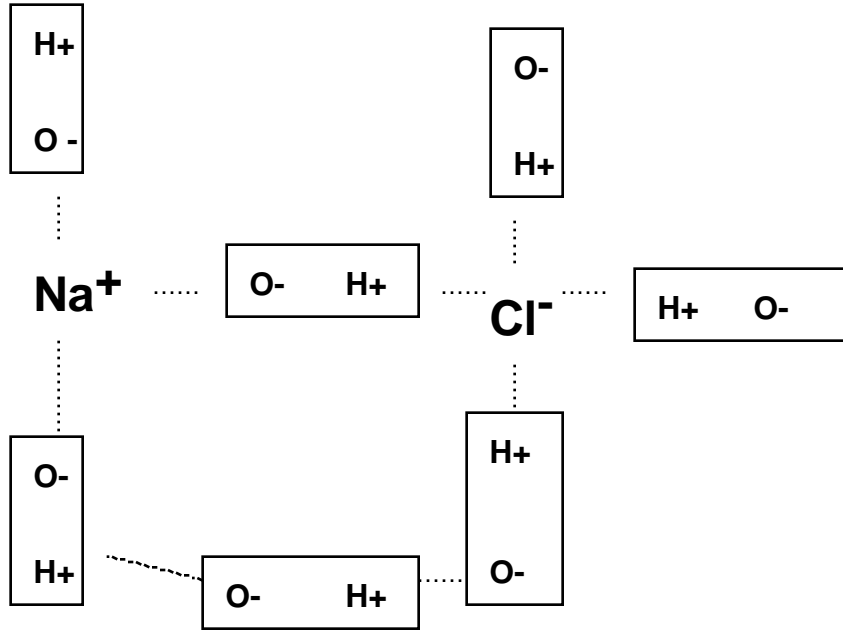
D. Interactions between polar substances and water: the **hydrophilic interaction**.

1. As should be obvious from the earlier discussions, many substances are polar and some are ionic. The way they interact with water is quite different

2. Let's consider an ionic substance such as salt, NaCl. In human physiology such substances are often called **electrolytes**. Recall that a crystal of salt is really a bunch of ions of Na^+ and Cl^- . They are held together by attraction of opposite charge (we called this an ionic bond).

3. What happens when the salt is placed in water. The water molecules are polar with + and - ends. These charged ends will tend to interact with the charges on the Na^+ and Cl^- ions.

4. As the sodium and chloride ions start to interact increasingly with the water molecules, the attraction between them is lessened. A unit of charge is a unit of charge. If some of the unit of positive charge on a Na^+ ion is now interacting with water instead of Cl^- , the attraction between the two is necessarily weakened. As more and more waters interact with the sodium and chloride ions, the interactions between these two ions are destroyed and they are able to float away from the crystal of salt into the water. Each will now be surrounded by water molecules with the appropriate ends facing the ion. The diagram below shows this, please note that the water molecules are stylized into positive and negative ends and do not really look like water molecules (see earlier in these notes).



We call mixtures like the one shown above **solutions**.

Solutions consist of two parts, the **solvent** (water in the body) and the **solute** (whatever is dissolved). Notice that the solute is held in solution by the hydrogen bonds (polar interactions) between it and the water (solvent) molecules.

a. Note: there are other forces we do not need to worry about that are much weaker that also matter in solutions. They are what allow non-polar substances such as O₂ to dissolve (poorly) in water).

b. Polar interactions are extremely important in the body.

They not only determine what is easily carried by water (for instance in blood) but they also have a lot to do with determining the shapes and functions of proteins.

5. What about other substances?

a. any substance with polar groups on it will interact with water. It will tend to dissolve. In fact, there are two simple rules for determining how well something will dissolve in water:

1. What percentage of molecule has polar bonds? -- the greater the percentage, the more soluble in water the molecule should be for the reasons mentioned above.

2. How large is the molecule? -- all else being equal, large molecules do not dissolve as well as small ones.

You should be able to look at the structure of molecules like these and know where the polar and non-polar parts are be able to guess as to whether or not they will dissolve easily.

IV. Counting Atoms and Molecules:

A. When we write the formula for a molecule, it is given as proportions of different types of atoms in the molecule -- H₂O means two atoms of hydrogen and one of oxygen.

B. We will have many occasions to write reactions. In these cases there will often be a number in front of each molecule. For example:



The numbers can be read as proportions (1 glucose molecule plus 6 oxygen molecules reacts to yield 6 each of water and carbon dioxide).

However, it can also be read to read **molar proportions**. A mol is a very large fixed number of particles and is the common "counting unit" in chemistry (equal to Avogadro's number – 6.02 *10²³ particles -- you don't need to know this number for this course). Just realize that generally when we talk about reactions, we will not be concerned with single molecules reacting but instead with large populations of molecules -- often mol amounts.

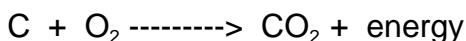
IV Chemical Reactions:

A. Anytime a new chemical bond (such as a covalent or ionic bond) is formed or broken, a chemical reaction has occurred.

B. We write chemical reactions by showing what we start with on one side with an arrow pointing to what we finish with. The starting materials are called **reactants** and the finishing materials are **products**. Thus:

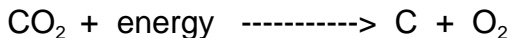
Reactants -----> Products

And for a specific case -- the formation of carbon dioxide by burning carbon (for example, coal) with oxygen:

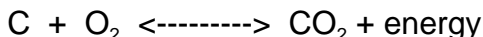


Where C (carbon) and O₂ (oxygen) are the reactants and carbon dioxide (CO₂) and heat are the products.

C. All reactions are in principal reversible. Under the correct conditions we can take the products of a reaction and go back to the original "reactants". In this case what were the products are now the actual reactants and what were the reactants are now the products. Thus, the reverse reaction of the one shown above is:



D. Since all reactions are, in principal, reversible we will usually write them using a double arrow. We generally call the reaction, as read from left to right the **forward reaction** while if we read the reaction in reverse, we have (you guessed it) the **reverse reaction**:



E. You might wonder how we decide which version of the reaction to make the forward reaction and which to make the reverse. The answer is that it is not arbitrary. It is convention to write the spontaneous reaction as the forward reaction. What is the spontaneous reaction? It is the one that will happen without requiring large inputs of energy or other exceptional conditions -- this will be explained in more detail below. Now you might object that burning coal or charcoal does require input of energy -- you need to get it lighted. This is true -- coal or charcoal does not react at an appreciable rate at room temperature. You must supply some energy to get the reaction started. As we will see below, this energy is called **activation energy**. However, what you are actually doing is simply providing the conditions needed for the reaction -- once these conditions are provided (high temperature) the reaction will proceed with no further intervention and will continue to create the conditions necessary for the reaction to continue. The heat released by the reaction keeps it going. On the other hand, the reverse reaction -- the breakdown of CO₂ to elemental carbon and oxygen gas requires a continuous large input of energy to keep it going. This reaction does not release energy to allow it to be self sustaining -- instead it requires a continuous external input of energy. More about this below.

Let's look at the rules for deciding when a reaction will be spontaneous. Let's also learn about what happens to matter and energy in a chemical reaction.

IV. Laws of Thermodynamics

A. We talked earlier about theories. There is something even stronger than a theory – a law. These are so well established that scientists believe that no observations will over turn them. There are two laws we need in our studies -- two of the Laws of Thermodynamics

B. Two of the three Laws of Thermodynamics:

1. **First Law: Conservation of matter and energy: the amount of matter and energy is constant.** In any process, the matter and energy may change -- for instance matter may be rearranged and energy may be stored or released, but no new matter or energy will be made nor will any be lost. This law is important because it allows us to "keep books" on all physical processes and changes.

2. Second Law: In any spontaneous process the overall entropy (disorder of matter/energy) of the universe increases. This is the harder of the two laws to understand, so let's go into a bit more detail:

a. We know that some processes can happen on their own -- they are spontaneous. What this usually means is that they do not require an input of energy to make them happen. Examples of spontaneous processes include objects falling in response to gravity and many chemical processes -- for example, salt dissolving in water or iron (Fe) exposed to oxygen (O₂) turning to rust (FeO₂).

b. On the other hand, many processes are decidedly non-spontaneous. They do not happen without the input of energy. Books laying on the floor do not spontaneously lift themselves up against gravity and take a place on a bookshelf. Nor does rust spontaneously break down into oxygen and iron. What is different about these two types of process?

The next section (section "c") is largely for explanation -- it is more detail than we covered in class.

c. In the case of the spontaneous processes, disorder has increased. Scientists think of disorder of matter and energy in terms of:

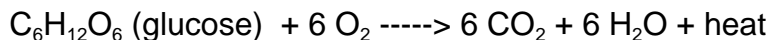
1. **Is energy kept in one place (stored)?** or is it allowed to spread all over the place? The most entropic version of energy is heat which is nothing more than random motion of matter -- the more heat, the greater the motion. Moreover, heat tends to spread out rapidly from places where it concentrated to other areas. Notice that this spreading is a decrease in disorder - - as the heat spreads, more and more matter is in motion and the energy causing this motion is no longer concentrated in one place. Thus, the spread of heat is a spontaneous process. Here are a couple of other examples of increases in entropy for energy:

a. Light is a relatively concentrated form of energy -- it is transmitted as particles called photons. However, when photons hit things and are absorbed (for example sunlight hitting a black object) much of the energy is converted to heat which quickly begins to spread out. Notice that the energy is no longer concentrated in a tiny photon -- it is now spreading over an increasingly large volume of matter. This degradation is spontaneous. On the other hand -- the reverse process is not.

b. The production of rust from pure iron and oxygen. In the process, heat is released. When rust is formed, the iron is essentially being burned by oxygen at relatively low temperatures. Energy was stored in the iron atoms (in their electrons, we will see) and is released as heat in the reaction. Thus, the energy was all stored in one place and after the reaction it is spreading out.

2. **Does the number of particles increase?** Does the process result in more particles than we started out with? If the number of

particles increases, the particle entropy increases. Here is an example. If a single molecule of glucose is burned with 6 molecules of oxygen the result is 6 molecules of carbon dioxide and 6 molecules of water. Also a lot of heat is released. We would write the process like this:



Notice that in this case the number of particles goes from 7 to 12. Thus, the particle entropy has increased.

d. Putting energy and particle entropy together. Some processes only involve changes in energy (for example the book falling). These are easy to understand. However, in many cases changes in both particle and energy entropy occur. When they are both in the same direction (for example the last example where both energy and particle entropy increased) it is easy to see that overall disorder increases and that the process must be spontaneous (according to the second law). However, it is common for processes to involve changes in particle and energy entropy that occur in opposite directions. Let's consider the formation of water:



(The forward reaction is largely what drives the space shuttle into orbit -- lots of energy is released by it).

(1) In this case the number of particles (particle entropy) has decreased while the energy entropy as increased (since stored energy in the hydrogen was released as heat). So how does one know if the reaction is spontaneous?

(2) Physical chemists have equations that take into account both particle and energy entropy (called free energy equations), but we need not consider them.

(3) We'll just make the following rule that will work in nearly every case.

(a) Generally, changes in energy entropy trump particle energy -- so if the reaction releases energy, we will assume that it is spontaneous. Thus, since the carbon to CO_2 reaction yields heat, it is spontaneous (which it is!) even though the particle entropy decreases.

(b) **So although learning about particle entropy was important (mainly because it makes it easy to grasp the relationship between particle number and disorder), we will tend to look at energy changes to see if a process is spontaneous.**

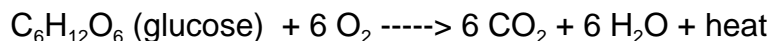
V. Reaction Kinetics

A. The second law of thermodynamics will predict whether or not a reaction is spontaneous (forward). However, it does not predict:

1. Whether or not the reaction will proceed (occur at an observable rate) under a given set of conditions (for example -- we have seen that coal does not normally convert to CO₂ at any appreciable rate at room temperature). Another way of saying this is that it does not predict the rate of reaction -- it just says whether or not the reaction should be able to occur.

2. Nor does it predict the exact path the reaction will take. What does this mean?

(a) It is perhaps easiest to understand if we imagine a reaction that breaks down some large molecule. For example, let's consider the oxidation of glucose:



(b) there are literally hundreds of different sequences in which oxygen could react with glucose to give the products listed. For instance, the first oxygen molecule might interact with one of the end carbons. The next oxygen might react with the new end carbon (after the first one had been removed). Or it could react with one in the middle and split the molecule into two pieces.

(c) Each of these sequences is a **reaction path**. Thermodynamics doesn't specify the particular path taken by a reaction -- it just lists the starting points (C₆H₁₂O₆ (glucose) + 6 O₂ for the forward reaction) and end point (6 CO₂ + 6 H₂O + heat). It does not say how you get there.

(d) When you burn glucose by igniting it with a flame, individual molecules will be destroyed by different sequences of reactions (different paths).

(e) On the other hand, when our bodies oxidize glucose to CO₂ and O₂ we always follow exactly one pathway. More about this later.

3. The study of how chemicals react -- the path the reaction takes and the conditions required to make the reaction occur is called **kinetics** (from the idea of motion or change).

B. Activation Energy:

1. All of the feasible reactions shown above (burning of glucose, burning of coal, burning of hydrogen) do not occur at an appreciable rate at room temperature. Yet they are spontaneous -- they will release tremendous amounts of energy (a characteristic amount per mol or kg burned). However, to get each of them started, we must add some amount of energy. The amount varies in each case. Burning a large piece of coal requires the most, burning hydrogen requires very little (this is partially what makes it an explosive). The amount of energy that must be added to start the reaction is called **activation energy** or "**energy of activation**". We abbreviate it as E_A .

2. What does the activation energy do? In order for molecules to react they must:

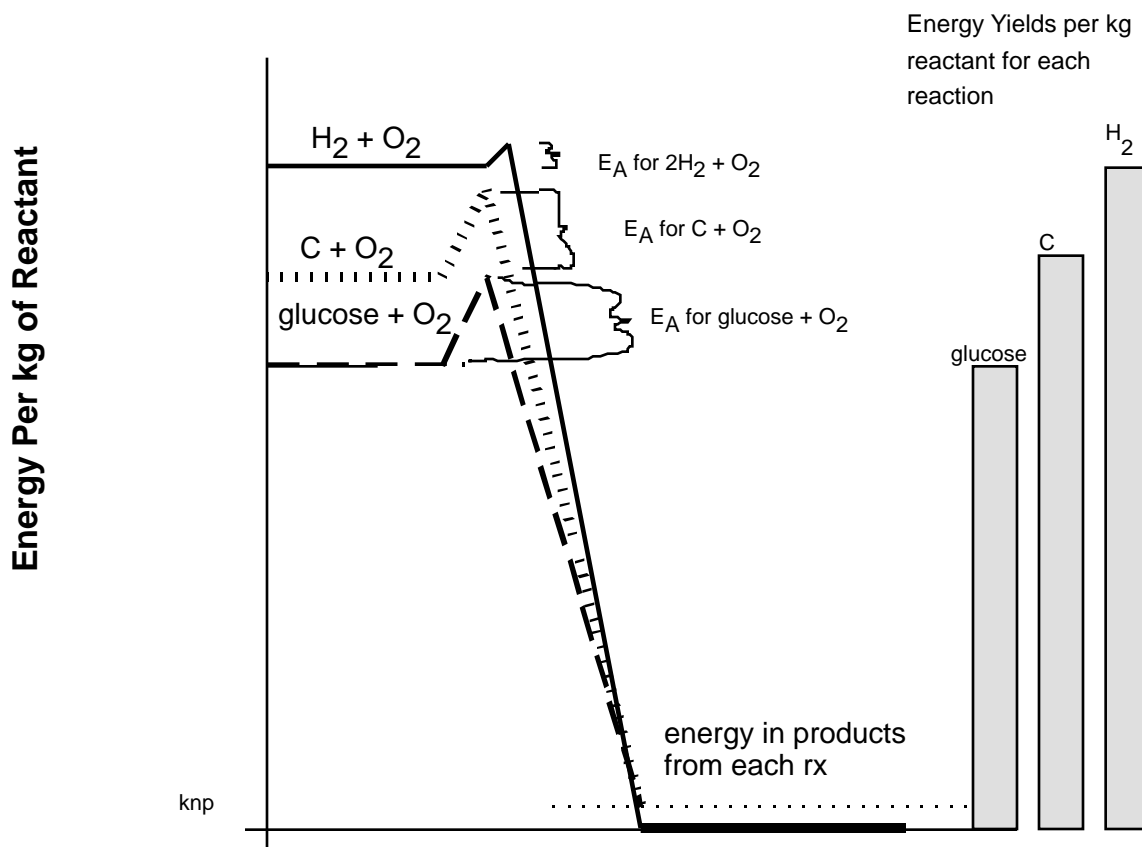
(a) come together with sufficient force to get them close enough to react

(b) they must be oriented correctly -- not all orientations will allow a reaction.

When one provides energy, the particles (reactants) are accelerated. This increases the number and force of collisions between the reactants. Thus, the greater force makes it more likely they get close enough to react and the greater number of collisions per second increases the number of collisions per second that are in the correct orientation.

3. Once the reaction gets started, notice that the energy yielded by the reaction provides E_A to keep the reaction going. You have what is a chemical chain reaction -- some the energy released by one set of reactants serves to provide E_A to other molecules of reactants. They react, releasing more energy some of which provides E_A for other reactants.

4. Let's see a graph that depicts the energy stored in the reactants, the energy of the products, the difference (the energy released) and the activation energy:



The plateaus on the left of the graph show the energy contained in the reactants -- the **potential energy of the reactants**. The single plateau on the right gives the energy contained in the products of the reactions. Here we assume they are all the same and that they equal zero potential energy. The **net energy yield** of the reaction is the difference between the two plateaus for each reaction. The yields are also shown as the bar graphs on the right. Notice that not all substances yield the same amount of energy per kg of reactant. In this case, hydrogen yields the most energy and glucose the least.

The **activation energy** is the amount of energy that is supplied to get the reaction started. Notice that:

- activation energy is equal to the upward tick on the graph from the "reactant plateau",
- it is not the same in each case -- very little is needed in the case of hydrogen and considerably more is needed for glucose
- the activation energy is gotten back -- the net release of energy is the same regardless how much activation energy is added since the net release is simply the difference between the energy of the products and reactants.

Note a couple of things about this graph:

(a) First, not all reactions yield the same amount of energy per amount of reactant. Some compounds contain far more potential energy than others. In the graph above, H_2 contains the most PE.

(b) Second, be aware that the products will not necessarily have zero potential energy. The yield of the reaction, however, will always be the difference between the energy of the reactants and the products.

(c) Third, notice that the activation energy is depicted as an "**energy hill**" that must be overcome to get the reaction going. The height of the energy hill is the amount of activation energy needed. Notice **that in different reactions (and under different conditions) more or less activation energy will be needed.**

(d) Finally note that the activation energy is not part of the energy yield of the reaction -- whatever is put in is also gotten out.

5. Activation energy and our bodies:

(a) The most common way to supply activation energy in the body is increase body temperature. This is essentially what mammals and birds do all the time -- they utilize very large amounts of food, in large part to maintain a higher body temperature. This ranges between about 35 and 41° C (95 and 105°F) for various types of mammals and birds. This gives them a big advantage over other "cold blooded" animal (but it also requires them to eat far more food -- it is easy for a mammal or bird to starve to death, it is much harder for a fish, snake or insect).

(b) in our bodies, we supply additional activation energy in exercise when we heat up. Typically if our body temperature increases from 37 to around 39° C, this will increase reaction rates (and therefore all processes associated with exercise) by about 20% -- a very large increase for such a small change in temperature!

(c) However, the main way that we increase or change rates of reaction in our bodies is not by temperature. Instead, we use catalysts.

C. Catalysts

1. A catalyst has the following properties:

(a) It **accelerates the rate of a specific reaction under a given set of conditions.**

(b) the **acceleration is accomplished by lowering the activation energy** required for the reaction. Thus, a catalyst can be seen as lowering the height of the energy hill in the graph seen previously.

(c) the **amount of acceleration of the reaction is directly proportional to the amount of catalyst that is used.** The more catalyst, the faster the reaction. This is because the catalyst must physically combine with the reactants to have an effect. If there are more catalysts, then more can combine with the available reactant and the overall reaction will be faster.

(d) **Catalysts are not consumed by the reaction.** Once they have combined with the reactants and helped the reaction along, they release the products and are then ready to help the next reaction happen. So, individual catalyst molecules are used over and over.

(e) **Catalysts are highly specific.** They will only catalyze one reaction (or at most a limited number of closely related reactions). This is

because they depend on the fact that they have a specific shape where only the reactants can attach.

(f) Catalysts lower activation energy by bringing the reactants closely together in the correct relative position. Instead of having to rely on chance favorable collisions, catalysts bring the molecules together correctly. Thus, less activation energy is needed for a given rate of reaction.

(g) Since catalysts are so specific about how they bring reactants together, they determine the exact path of the reaction. So catalyzed reactions are very orderly.

2. **In our bodies, ALL REACTIONS ARE CATALYZED** by a class of **proteins called ENZYMES** which are the subject of the next class. A figure worth remembering is that most of the tens of thousands of different types of chemical reactions that occur in our bodies are **accelerated by between 10,000,000 and 100,000,000 times over the rate they would occur at 37° C** (body temperature) **if there were no catalysts involved.**

Review: In the last class we considered that reactions that normally occur only at a very low rate at room conditions can be speeded up by the addition of something known as a catalyst. We finished with the observation that all reactions in our body are catalyzed and that most are accelerated by a factor of ten million to one hundred million times over the rate they would have without the catalyst. Thus, it should be obvious that without these catalysts, life would not be possible for they allow reactions to proceed at a high rate. We will also see that catalysts in the body provide a large measure of control and order to reactions; this sort of control is not seen in non-biological systems.