Chapter 4

Chemical Energy

Perhaps the most convenient form in which to *store* energy is chemical energy. The foods we eat, combined with the oxygen we breathe, store energy that our bodies extract and convert into mechanical and thermal energy. The fuels that we burn in our automobiles, furnaces, and campfires also store energy in chemical form. Batteries store chemical energy as well, for later retrieval in the form of electrical energy.

A chemical reaction is simply a rearrangement of the way that atoms are connected together to form molecules. A simple example is

$$H_2 + \tfrac{1}{2}O_2 \longrightarrow H_2O, \tag{4.1}$$

the combination of hydrogen gas with oxygen gas to form water. Initially, the hydrogen atoms are connected together in pairs to form H_2 molecules, and the oxygen atoms are similarly connected in pairs. To form a water molecule (H_2O), we need two hydrogen atoms but only one oxygen atom, or equivalently, an entire H_2 molecule but only half an O_2 molecule. (In practice, the other half of the O_2 will usually combine with another H_2, to form a second water molecule.)

Many chemical reactions, especially those that consume oxygen, give off energy. For example, each time reaction 4.1 occurs, a very small amount of energy, roughly half a billionth of a billionth of a joule, is given off. If the reaction occurs by ordinary combustion (burning), then the energy is given off as thermal energy, but sometimes the energy released can take on other forms, such as electrical or even mechanical energy.

Because the amount of energy released by a single molecular reaction is so tiny, we often refer to the energy given off when a large number of identical molecular reactions take place. The standard number used by scientists for this purpose is called **Avogadro's number**,

$$N_A = 602,000,000,000,000,000,000,000 = 6.02 \times 10^{23}. \tag{4.2}$$

This number was chosen for convenience, because it happens to equal the number of hydrogen atoms in one gram of hydrogen. When we have this many of something, we say that we have one **mole** of that substance.

A mole of hydrogen *atoms* has a mass of one gram, so a mole of H_2 *molecules* has a mass of two grams. Hydrogen is the simplest and lightest element. The other two elements that commonly occur in chemical fuels are also among the lightest: carbon, which is 12 times heavier (per atom) than hydrogen, and oxygen, which is 16 times heavier than hydrogen. So a mole of carbon atoms has a mass of 12 grams, while a mole of oxygen atoms would have a mass of 16 grams. (A mole of O_2 *molecules* is twice as heavy, 32 grams.) Although chemical reactions can involve any of the other hundred-plus elements on the periodic table, it turns out that most of the familiar energy-releasing reactions involve only hydrogen, carbon, and oxygen.

Consider again reaction 4.1, which involves only hydrogen and oxygen atoms. When one mole of hydrogen molecules (two grams) combines with half a mole of oxygen molecules (16 grams) to form one mole of water molecules (18 grams), the energy given off turns out to be 242,000 joules, assuming that the water comes out as a gas rather than as a liquid. (If the water comes out as a liquid, then the total energy released is greater: 286,000 joules.) Since there are 500 moles of hydrogen gas in a kilogram, this means that burning a kilogram of hydrogen gas releases 500 times as much energy, or 121 million joules (again, assuming that the water comes out as a gas, as is usually the case in a combustion process).

Exercise 4.1. What is the mass of one mole of carbon dioxide (CO_2) molecules?

Exercise 4.2. What is the mass of one mole of ethyl alcohol (C_2H_5OH) molecules?

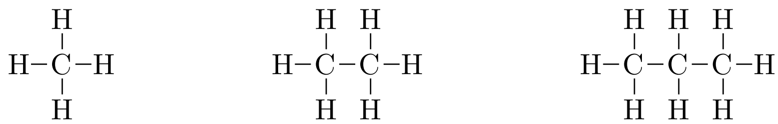
Fossil Fuels

Hydrogen gas is rather uncommon on earth, so it is not widely used as a fuel. Instead, we tend to burn carbon-containing fuels, the remains of organic material, mostly plants, that have been buried, compressed, and chemically altered over hundreds of millions of years. The most common of these "fossil fuels" is **coal**, a complex mineral composed mostly of carbon, with small amounts of hydrogen, oxygen, and other elements mixed in. To a first approximation, we can pretend that coal is pure carbon, in the form of graphite. (The other familiar form of pure carbon is diamond, usually considered much too valuable to burn as a fuel!) The combustion of graphite results in the gas carbon dioxide, according to the reaction equation

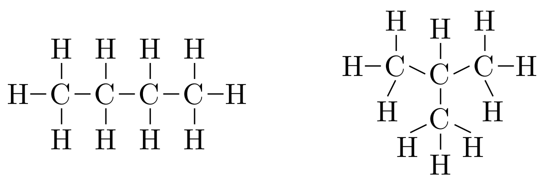
$$C + O_2 \longrightarrow CO_2. \tag{4.3}$$

For each mole of graphite burned, or equivalently, each mole of carbon dioxide produced, this reaction yields 394,000 J of energy. A mole of graphite is 12 grams, so a kilogram of graphite contains $1000/12 = 83$ moles. Burning one kilogram of graphite therefore yields $83 \times 394,000$ J, or 33 MJ. High-grade coal generally yields a little less energy, roughly 29 MJ per kilogram.

The other important fossil fuels are oil (petroleum) and natural gas, both of which are composed mostly of **hydrocarbons**: molecules containing both hydrogen and carbon. The simplest and most common hydrocarbons are the so-called **alkanes**, in which each carbon atom is linked to four other atoms, each of which can be either another carbon or a hydrogen. The simplest alkane is methane, CH_4, the principal component of natural gas. Next-simplest is ethane (C_2H_6), then propane (C_3H_8), then butane (C_4H_{10}), then pentane, hexane, heptane, octane, nonane, decane, and so on, where each molecule in the list is bigger than the previous one by a single carbon atom and two hydrogens. Figure 4.1 shows abstract drawings of the simplest alkane molecules.



Methane Ethane Propane



Butane (two varieties)

Figure 4.1. Structure of the simplest alkane molecules. Note that each carbon atom is linked to four other atoms, while each hydrogen is linked only to a single (carbon) atom. Given these rules, you can draw similar diagrams for larger alkane molecules (pentane, hexane, heptane, octane, and so on). Note that butane has two possible structures. The larger alkanes similarly have multiple structures, varying from linear to multi-branched.

Because hydrocarbon molecules contain both hydrogen and carbon, their combustion in oxygen yields both water (as in reaction 4.1) and carbon dioxide (as in reaction 4.3). You can figure out how much of each simply by requiring that the number of atoms of each type "balance" on both sides of the reaction equation. Consider the combustion of methane, for instance. Since methane contains a single carbon atom, its combustion must produce a single CO_2 molecule. Meanwhile, to account for the four hydrogen atoms in methane, we must end up with two H_2O molecules. Finally, the CO_2 and the two H_2O's contain a total of four oxygen atoms, so the number of O_2 *molecules* that go into the reaction must be 2. The full reaction equation is therefore

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O. \tag{4.4}$$

Determining the correct numerical coefficients in a reaction equation is called "balancing" the equation. Using the technique just described, you should be able to balance the reaction equation for the combustion (in oxygen) of any fuel containing hydrogen, carbon, and/or oxygen. (The end-products will always be water and carbon dioxide.)

The energy given off by the combustion of one mole of methane turns out to be 802 kJ. From this information, you should be able to show that the combustion of one *kilogram* of methane releases 50 MJ. Heavier hydrocarbons generally yield more energy per mole, but approximately the same energy on a per-kilogram basis. Gasoline, a mixture of hexane, heptane, octane, and various other hydrocarbons, yields about 44 MJ per kilogram, or about 124,000 Btu's per gallon.

Exercise 4.3. In analogy with Figure 4.1, draw the structure of pentane, the alkane containing five carbon atoms. Try to draw at least two different varieties, analogous to the two varieties of butane. What is the chemical formula for pentane?

Exercise 4.4. Write down the reaction equation for the combustion of ethane, and balance it with the correct coefficients.

Exercise 4.5. Write down the reaction equation for the combustion of propane, and balance it with the correct coefficients.

Exercise 4.6. The combustion of one mole of ethane releases 1430 kJ of energy. How much energy is released by the combustion of one *kilogram* of ethane? (Hint: First compute the mass in grams of one mole of ethane.)

Exercise 4.7. The combustion of one mole of propane releases 2040 kJ of energy. How much energy is released by the combustion of one *kilogram* of propane? (Hint: First compute the mass in grams of one mole of propane.)

Exercise 4.8. In a previous exercise you estimated the kinetic energy of a typical car traveling at highway speed. How much gasoline would be needed to provide this amount of energy, if we unrealistically assume 100% efficiency in the conversion of chemical to kinetic energy?

Exercise 4.9. In a previous exercise you estimated the gravitational energy gained by a typical car climbing a 12,000-foot mountain pass. How much gasoline would be needed to provide this amount of energy, if we unrealistically assume 100% efficiency in the conversion of chemical to gravitational energy?

Energy from Food

Although fossil hydrocarbons are convenient chemical fuels, our bodies cannot digest them. Instead we obtain our energy from the more complicated molecules found in living plants and animals: sugars, fats, and proteins.

The "simple" sugars, such as glucose, have the chemical formula $C_6H_{12}O_6$. The balanced reaction equation for "combustion" of glucose (whether by literal burning or the more subtle processes of metabolism) is

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O. \tag{4.5}$$

For one mole of glucose, this reaction releases 2.80 MJ of energy (assuming now that the water ends up in liquid form). That translates to 15.6 kJ, or 3.7 kcal, per gram. Dieticians normally just round this number to 4 kcal/gram, the "caloric content" of sugar. Sucrose, or table sugar, is basically just two simple sugar molecules linked together, while starch is a long chain of many simple sugar links. All such molecules are referred to as **carbohydrates**. On a per-gram basis, all carbohydrates yield essentially the same amount of energy as glucose. The cellulose ("fiber") that makes up the cell walls of plants is also made of linked sugar molecules, but the human digestive tract is unable to unlink them, so their caloric value to us is zero. That doesn't prevent us from burning wood for fuel, however, and in this case, the energy released is again a little under 4 kcal per gram (not counting the mass of any moisture in the wood).

Notice that carbohydrates yield considerably less energy than hydrocarbons, on a per-gram basis. This is because they already contain quite a bit of oxygen, so there is less opportunity for them to combine with oxygen to release energy. In a sense, they are already "partially burned."

Fat molecules, on the other hand, are made almost entirely of carbon and hydrogen, with very little oxygen. Their metabolism in our bodies yields approximately 9 kcal/gram, more than twice the energy released from carbohydrates. Our bodies also produce fats as a way of storing energy for later use.

Proteins are very complex molecules containing considerable nitrogen in addition to carbon, hydrogen, and oxygen. They serve a variety of nutritional needs, but our bodies can metabolize them for energy when necessary, extracting approximately 4 kcal/gram, the same as from carbohydrates.

Exercise 4.10. A standard jelly donut (JD) provides 1 MJ of energy, or approximately 250 kcal. What would be its mass in grams, if all this energy comes from carbohydrates? What if all the energy comes from fats?

Exercise 4.11. Many dieticians recommend that you limit your fat intake to no more than 25% of your total calories. If your daily diet amounts to 2000 kcal, of which 25% comes from fats and the rest from carbohydrates and proteins, what is the total mass of the energy-providing components of the foods you eat in one day? If you're hungry and wish to eat more food than this, how can you do so without getting more calories?

Exercise 4.12. Suppose you attempt to lose weight by going on a crash diet (not recommended!), eating no energy-providing foods at all while maintaining your usual level of activity. Suppose further that your body's internal energy needs remain unchanged—that is, your basal metabolic rate isn't affected by the lack of food (probably an unrealistic assumption). Your body will then obtain the energy it needs by consuming stored fat (at least until you run out of fat, when it would start consuming protein, endangering your health). Roughly how much weight will you lose, in pounds, each day?