1.1.4 The ~~Wh~~ Wh Energy Unit

As we have discussed, energy is an accumulation of power over time, thus a joule would be W × s. This concept is equally applicable by changing the units of time. Indeed, it is common in electrical power systems to use Watts × Hours = Wh as energy units. We refer to this unit as "watt hour." Please note that it would be incorrect to say "watts per hour" when referring to the units of energy, since the term *per* implies a rate quantity and what we really mean is "watts multiplied by hours." Since J and Wh are both units for energy, let us see how many J are in 1 Wh. Simply take the number of seconds in one hour $60\frac{\text{s}}{\text{min}} \times 60\frac{\text{min}}{\text{h}} = 3600\frac{\text{s}}{\text{h}}$ and convert $\cancel{3.27\ \text{W} \times \frac{5\ \text{min}}{60\ \text{min}/\text{h}} = 0.272\ \text{Wh}}$. $1\text{Wh} = 1\frac{\text{J}}{\text{s}} \times 3600\text{s} = 3600\text{J}$ Therefore, 1 Wh = 3600 J, a convenient number to remember.

Example 1.5

Calculate work in Example 1.3 giving the value in Wh. Answer: We know that work is 980 J. Converting $W = \frac{980 \quad \text{J}}{3600\ \text{J/Wh}} = 0.272$ Wh. Alternatively, we know the power is 3.27 W and if we provided it for 5 min, $3.27\ \text{W} \times \frac{5\ \text{min}}{60\ \text{min}/\text{h}} = 0.272$ Wh.

For larger amounts, we use kWh, MWh, and GWh. Examples: A 100 W device used for 10 h, would have consumed 100 × 10 Wh = 1 kWh of energy. A 1 MW electric power plant producing power constantly for one year would produce annual energy of E = 1 MW × 365 d × 24 h/d = 1 MW × 8760 h = 8.76 GWh. The number 8760 of hours in a year is a number worth remembering and we will use it frequently.

1.1.5 Electromagnetic Radiation

An *electric field* emanates from electric charges (e.g., an electron) and exerts a force on other charges, whereas a *magnetic field* results from moving electric charges and moments of particles of magnetic materials. The propagation of fluctuating electric and magnetic field at the speed of light carry *radiant energy*, which can be modeled as an *electromagnetic* (*EM*) *wave* or as a flux of particles called *photons*. We will refer to the waves or photons as *EM radiation*, indistinctly from the model employed to analyze it. Atoms emit EM radiation due to the rearrangement of previously excited electrons. Radiant energy carried by the EM radiation can in turn be transferred as some other form of energy upon interaction with other matter. For example, energy from *fusion* (a form of nuclear energy) in the sun is transferred to EM radiation emitted at its surface, which when reaching the Earth's surface can transfer energy to the land, ocean water, and plants. EM radiation from the sun is referred to as *solar radiation* or commonly as *sunlight*.

Using the wave model, we employ the frequency ν of the EM fluctuation (in Hz or s^{-1}), which when multiplied by the wavelength λ (length of the wave, in m, Figure 1.10) equals the speed of wave propagation, which is the speed of light c (3×10^8 m/s). The relationship is

$$c = \nu\lambda \tag{1.10}$$

In other words, the frequency would be the number of cycles that go through a fixed point in space every second; wavelength is the distance between a point of the wave and a similar point in the next cycle, say, between peaks (Figure 1.10).

The energy E of a photon of an EM wave is directly related to the frequency ν of the EM wave by the Planck's constant $h = 6.6 \times 10^{-34}$ Js

$$E = h\nu \tag{1.11}$$

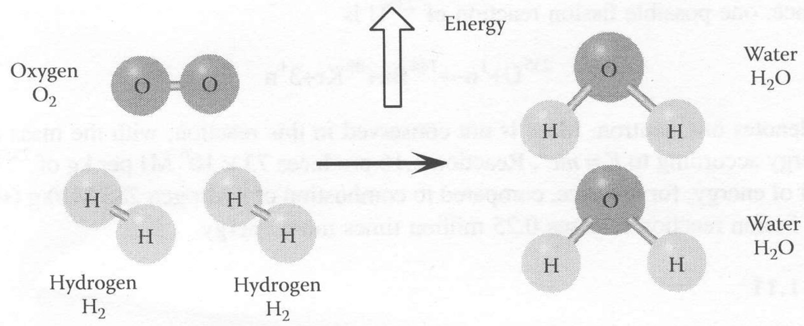

FIGURE 1.12 A simple example of an exothermic reaction. Combustion of hydrogen. Ball-and-stick models are used for illustration only.

kJ/mol to MJ/kg. For instance, one kg of H_2 yields $\frac{285.8\ \text{kJ/mol}}{2\ \cancel{1}\ \text{g/mol}} = \overset{142.9}{\cancel{285.8}}\ \text{MJ/kg}$. When using hydrogen to obtain electricity in a fuel cell, the $H_2(g)$ flow rate in g/s will determine the power output. First convert g/s to mol/s, then calculate power taking into account that only part of the energy released can be converted to work.

Example 1.9

Assume a $H_2(g)$ flow rate of 1 mg/s and that 40% of the energy can be converted to electricity. What is the power output? Answer: The flow rate in mol/s is $\frac{1\ \text{mg/s}}{2\ \cancel{1}\ \text{g/mol}} \simeq \overset{0.5\times10^{-3}}{\cancel{0.001}}$ mol/s. Then, dH/dt = $-\overset{0.5\times10^{-3}}{\cancel{0.001}}$ mol/s × 285.8 kJ/mol = $\overset{-0.1429}{\cancel{-0.2858}}$ kJ/s = $\overset{-142.9}{\cancel{-285.8}}$ W. But the power output is only 40% of this total energy released or $P \simeq \underset{142.9}{\cancel{285.8}} \times 0.4 = \underset{57.16}{\cancel{114.32}}$ W.

1.1.8 Nuclear Energy

Nuclear energy is derived for the binding force that holds the nucleons (subatomic particles) of the atomic nucleus [2,4]. For our purposes, we will only consider protons and neutrons as nucleons. Elements have an *atomic number* Z (number of protons) and mass number or *atomic mass* A (sum of number of protons and neutrons). Each proton or neutron has 1 atomic mass unit (1 amu). One amu is 1.66×10^{-27} kg or 1/12 of the mass of the most common carbon atom, with 6 protons and 6 neutrons. An element's nucleus corresponds to the number of protons Z, for instance, carbon has 6 protons. But many elements have several forms or *isotopes* according to the atomic mass [11]. For instance, ^{12}C has 6 neutrons ($A = 6 + 6 = 12$), but ^{14}C has 8 neutrons ($\underset{A}{\cancel{Z}} = 6 + 8 = 14$).

Example 1.10

Uranium has 92 protons. How many neutrons are in the ^{235}U isotope and in the ^{238}U isotope? Answer: 235 − 92 = 143 neutrons in the ^{235}U isotope and 238 − 92 = 146 neutrons in the ^{238}U isotope.

Binding energy per nucleon is greater for the middle value of atomic mass; it decreases for larger A or smaller $\cancel{Z}$. Binding energy increases when lighter nuclei undergo *fusion* or when heavier nuclei undergo *fission*. Currently, nuclear-fueled power plants are based on fission (Chapter 7). To initiate a fission *chain reaction* using uranium, ^{235}U is bombarded by a neutron producing fission into two lighter and unstable isotopes (which decay later to stable forms emitting β and γ rays), and two or three neutrons that can bombard other nuclei and sustain a chain reaction. Several different fission reactions could occur.

For instance, one possible fission reaction of ^{235}U is

$$^{235}U + {}^{1}n \rightarrow {}^{144}Ba + {}^{89}Kr + 3\,{}^{1}n \tag{1.16}$$

Here ^{1}n denotes one neutron. Mass is not conserved in this reaction, with the mass deficit converted to energy according to $E = mc^2$. Reaction 1.16 produces 73 × 10^6 MJ per kg of ^{235}U. This is a great amount of energy, for instance, compared to combustion of hydrogen ~~285~~ 142.9 MJ/kg (see previous section), the fission reaction releases ~~0.25~~ 0.6 million times more energy.

Example 1.11

Calculate the energy released by reaction 1.16 [2]. Answer: Calculate atomic mass deficit; the left side is 235.04 + 1.00 = 236.04 amu and the right-hand side is 143.92 + 88.91 + 3.02 = 235.85 amu. The mass deficit is 0.19 amu. Now convert to kg $m = 0.19 \times 1.66 \times 10^{-27}$ kg $= 0.315 \times 10^{-27}$ kg and use $E = mc^2 = 0.315 \times 10^{-27}$ kg $\times (3 \times 10^8$ m/s$)^2 = 2.84 \times 10^{-11}$ J per atom of ^{235}U. Multiply by Avogadro's number to get moles $2.84 \times 10^{-11} \times 6.023 \times 10^{23} = 17.10 \times 10^{12}$ J/mole and now use molar mass to $\dfrac{17.10 \times 10^{12}\ \text{J/mol}}{235\ \text{g/mol}} = 73 \times 10^6$ MJ/kg.

1.2 CARBON-BASED POWER SYSTEMS

1.2.1 Energy from Hydrocarbon Combustion

As we will see shortly, the major source of today's generation of electricity is burning *fossil fuels*. These fuels are derived from oil, coal, and gas that formed from algae, trees, and other living organisms in the geologic past (we will study this with detail in Chapter 2). The fossil matter contains energy stored in chemical form that can be converted to heat by combustion. The common type of compound found in fossil fuels are *hydrocarbons*, which are organic compounds formed entirely by *carbon* C and hydrogen H. For example, *methane* CH_4 has one carbon atom and four hydrogen atoms; *ethane* C_2H_6 has two carbon atoms and six hydrogen atoms (Figure 1.13). Note that as you increase the number of C atoms and keep all bonds single and occupied with hydrogen (a *saturated* hydrocarbon), you form chains with the general formula C_nH_{2n+2}, because there are two H atoms per C atom and one at each end of the chain. These are called *alkanes*.

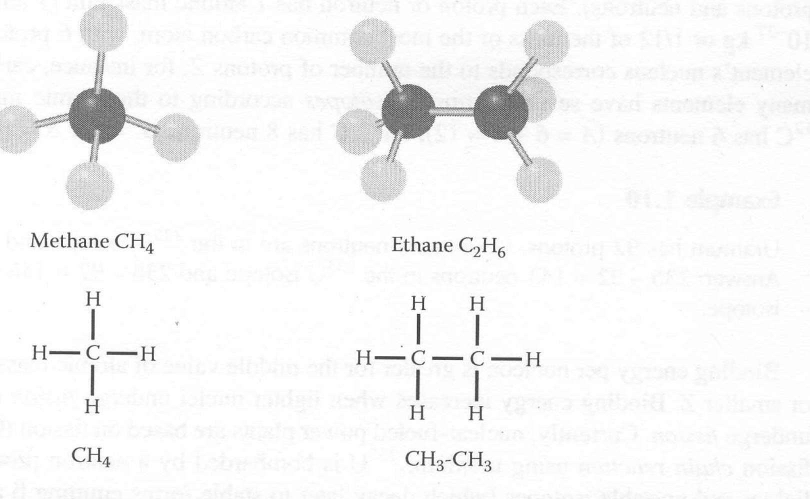

FIGURE 1.13 Examples of simple hydrocarbons: methane and ethane. Ball-and-stick molecular model, name and formula, structure, and condensed structure.

When burned in the presence of sufficient oxygen, the *combustion* products are water vapor H_2O, carbon dioxide CO_2, and energy. As the simplest illustration of hydrocarbon combustion, consider burning methane CH_4 with sufficient oxygen. The reaction is exothermic and given by

$$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$$

It releases $dH = -802.3$ kJ/mol (Figure 1.14). If there is not enough oxygen, the reaction yields carbon monoxide CO rather than CO_2.

Example 1.12

What is the energy produced by burning one gram of methane? What is the mass of CO_2 produced?

Answer: First, convert grams of methane to moles using the molecular formula for methane. Use ≈ 12 g/mol for C and ≈ 1 g/mol for H. Then $\frac{1\cancel{\text{g}}}{(1 \times 12 + 4 \times 1)\text{g/mol}} = \frac{1}{16}\,\cancel{\text{mol.}}$ mol/g Then multiply by $dH = -802.3$ kJ/mol to get $dH = -802.3 \text{ kJ/mol} \times \frac{1}{16}\,\cancel{\text{mol}}$ mol/g $= -50.14$ kJ/g. Now, convert moles of CO_2 to mass, use ≈ 12 g/mol for C and ≈ 16 g/mol for O, so that $\frac{1}{16}\text{mol} \times (1 \times 12 + 2 \times 16)\text{g/mol} = \frac{44}{16}\text{g} = 2.75$ g.

This example illustrates the importance of the relationship between energy released and the mass ratio of CO_2 to fuel combusted. In this simple case, there is a ratio of 2.75:1 of CO_2 to fuel. You can build an intuitive reasoning for this ratio by realizing that in an alkane the C atoms are bonded with H, and when combusted the C atoms are oxidized (bonded with O). Since oxygen weighs more than H (16 to 1), then the mass of CO_2 emitted is larger than the mass of methane combusted. The ratio of CO_2 to fuel allows an estimate of how much CO_2 is emitted per J of energy. Since it takes 1 g of methane to produce 50.14 kJ while releasing 2.75 g of CO_2, we can conclude that there is ~0.055 g CO_2 produced per kJ of energy. Converting to kg and kWh, we obtain ~0.20 kg CO_2 per kWh.

This is an idealized estimate not applicable to real electricity production since we are considering the simplest alkane and ignoring inefficiencies. In general, the combustion of an alkane containing n C atoms is

$$C_nH_{2n+2} + \frac{3n+1}{2}O_2 \rightarrow (n+1)H_2O + nCO_2$$

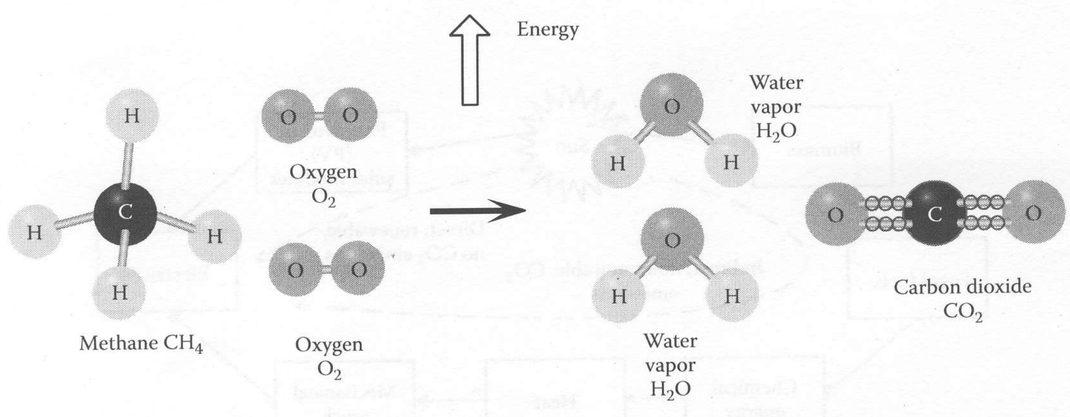

FIGURE 1.14 Simple example of hydrocarbon combustion: methane combustion with sufficient oxygen. Ball-and-stick models are used for illustration only.