SECTION 7: THE CALORIE – THE ALTERNATE ENERGY UNIT

During the early development of the science of thermodynamics in the 18th and 19th centuries, heat and work were believed to be two entirely separate phenomena. It was not until the mid-19th century that their equivalence was widely accepted. As a result, heat was originally measured using a different unit: the calorie. A calorie is defined to be the amount of heat required to raise the temperature of 1 gram of water by 1°C (from 20°C to 21°C, for instance). This definition was refined a bit over the years, as people realized that the specific heat of water depends on the temperature: the calorie was later defined to be the heat required to raise the temperature of 1 gram of water from 17°C to 18°C. This amount of heat is equal to 4.184 joules, so the modern definition of the calorie is simply:

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1 \text{ calorie} = 4.184 \text{ joules}
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For many years, all thermochemical measurements were made and reported in calories. Only in the past 20 years have chemists shifted to joules as the standard unit for energy. Organic chemists have been the slowest to make the change: most organic chemistry textbooks still quote reaction energies and bond energies in kilocalories, for no better reason than “it’s always been done that way….”

Incidentally, the “calories” that you see on food packaging and nutritional guidebooks are actually kilocalories. They represent the potential (chemical) energy of the food. When you consume a candy bar whose label says “100 calories,” you are actually consuming the chemical equivalent of 100,000 calories = 418,400 joules. A typical adult consumes 2000 nutritional “calories,” which is enough energy to lift a 120-pound person over 9 miles in the air (that’s about 50,000 feet above the ground, well into the stratosphere) if it were used entirely to do work.