10.2.1   ENERGY

10.2.1.1   The Nature of Energy

Energy is the ability to do word or produce heat.

Energy can be either potential or kinetic energy. **Potential energy** is due to position or composition. The **kinetic energy** of an object is the energy due to the motion of the object and depends on the mass of the object.

The **Law of Conservation of Energy** states that *energy can be converted from one form to another but can be neither created nor destroyed*. In other words, the energy of the universe is constant.

10.2.1.2   Temperature and Heat

Temperature is a measure of the random motions of the components of a substance.

Heat can be defined as the flow of energy due to a temperature difference.

10.2.1.3   Exothermic and Endothermic Processes

When a process results in the evolution of heat, it is said to be **exothermic**—energy flows out of the system. A common example of an exothermic reaction is the combustion that occurs when a match is struck.
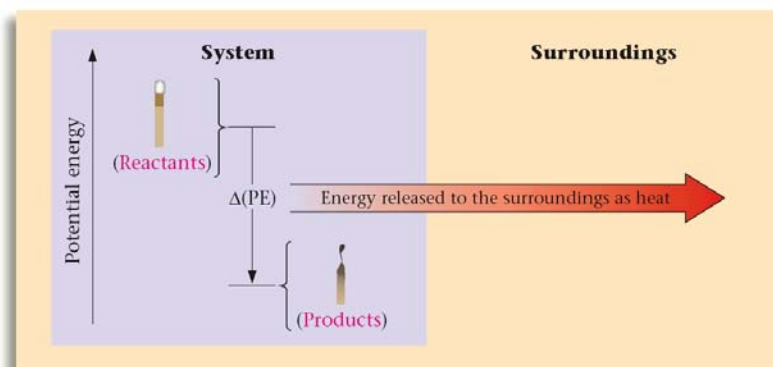
Processes that absorb energy from their surroundings are said to be **endothermic**—energy flows into the system. A common example of an endothermic process is the boiling of water to form steam.

10.2.1.4   Thermodynamics

The study of energy is called **thermodynamics**, and the Law of Conservation of Energy is often referred to as the **First Law of Thermodynamics**, and is stated as follows:

> *The energy of the universe is constant.*

The **internal energy**, E, of a system can be defined as the sum of the kinetic and potential energies of all 'particles' in the system. The internal energy of a system can be changed by a flow of work, heat, or both. That is:

$$\Delta E = q + w$$

where:

ΔE is the change in energy
q is the input heat, and
w is the work done on the system.

By convention, the sign of q is positive when heat energy flows into the system (*i.e.* for an endothermic process), and negative when heat energy flows out of the system (*i.e.* for an exothermic process). In the present context, the same convention applies to the flow of work, so that ΔE will be negative if the energy of the system is decreasing, and positive if the energy of the system is increasing.
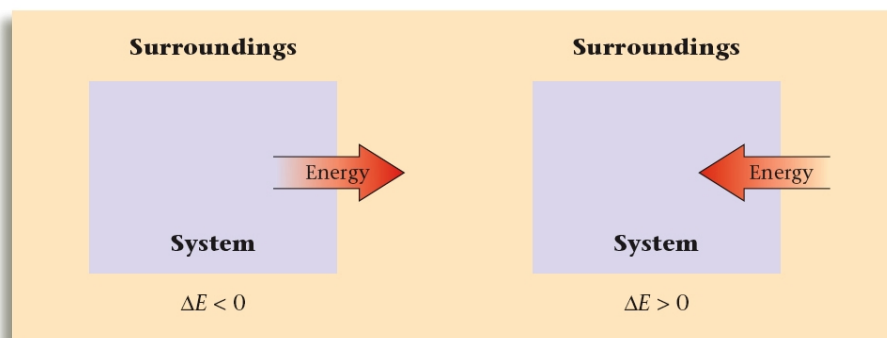
10.2.1.5 Measuring Energy changes

In the metric system, the **calorie** is defined as the amount of energy (heat) required to raise the temperature of one gram of water by one degree Celsius. The calorie that is used to measure the energy content of food is actually a kilocalorie, written with a capital "C" (Calorie) to distinguish it from the calorie used in chemistry.

The **joule** is the SI unit of energy, and is defined in physical terms as energy exerted by a force of one newton acting to move a one kilogram object through a distance of one metre.

The relationship between calories and joules is given in the following equation:

1 calorie = 4.184 joules

Different substances respond differently to being heated. We have noted that 4.184 J (or 1 calorie) of energy raises the temperature of 1 g of water by 1 °C. In contrast, the same amount of energy applied to 1 g of gold would raise its temperature by approximately 32°C!

We describe this difference by saying that different substances have different heat capacities. In general, the heat capacity (c) of a substance is given by:

$$c = \frac{\text{heat absorbed}}{\text{increase in temperature}}$$

The (specific) amount of energy required to change the temperature of one gram of a substance by one degree Celsius, under constant pressure, is called its *specific heat capacity*, or simply its **specific heat** (c_p).

Specific Heat Capacities of Some Common Substances

Substance	Specific Heat (J/g °C)
Water (l)	4.184
Water (s) (ice)	2.03
Water (g) (steam)	2.0
Aluminium (s)	0.89
Iron (s)	0.45
Mercury (l)	0.14
Carbon (s)	0.71
Silver (s)	0.24
Gold (s)	0.13

10.2.1.6   Thermochemistry (Enthalpy)

In order to measure exactly how much energy is produced or absorbed in a particular reaction, we use specific quantity known as the **enthalpy**, which is designated by the symbol H. For a reaction occurring under conditions of constant pressure, the change in enthalpy (ΔH) is equal to the energy that flows as heat. *i.e.*

$$\Delta H_p = \text{heat}$$

where the subscript "p" indicates that the process has occurred under conditions of constant pressure.

At constant pressure, an exothermic process has a negative ΔH, while an endothermic process has a positive ΔH.

10.2.1.6.1   Calorimetry

Calorimetry is the science of measuring heat, and the device used to determine experimentally the heat associated with a chemical reaction is called a **calorimeter** (Figure 10.2.1). The reaction is run in the calorimeter, and the temperature change of the calorimeter is recorded. Knowing this temperature change, and the heat capacity of the calorimeter enables the calculation of the heat released or absorbed by, and thus ΔH for, the reaction.
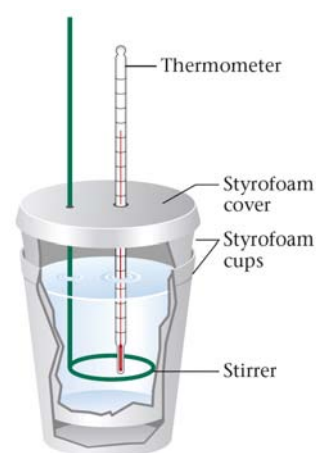

Figure 10.2.1
'Styrofoam cup'
calorimeter

10.2.1.7   Hess's Law

One of the most important characteristics of enthalpy is that it is a **state function**. That is, the change in enthalpy for a given process is independent of the pathway for the process. Thus, Hess's Law states that, *in going from a particular set of reactants to a particular set of products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.*

10.2.1.7.1   Characteristics of Enthalpy Changes

To use Hess's Law to compute enthalpy changes for reactions, it is important to understand two characteristics of ΔH for a reaction:

1. If a reaction is reversed, the sign of ΔH is also reversed;

2. The magnitude of ΔH is directly proportional to the quantities of reactants and products in a reaction. If the coefficients in a balanced reaction are multiplied by an integer, the value of ΔH is multiplied by the same integer.

10.2.1.7.2   Applying Hess's Law

Calculations involving Hess's Law typically require that several reactions be manipulated and combined to finally give the reaction of interest. In executing this procedure, you should:

*   Work backward from the required reaction, using the reactants and products to decide how to manipulate the other given reactions at your disposal;

*   Reverse any reactions as needed to give the required reactants and products;

*   Multiply reactions top give the correct numbers of reactants and products.

This process often involves some trial and error, but it can be very systematic if you allow the final reaction to guide you.

10.2.1.8   Entropy

Entropy (S) is a measure of disorder in a system. As disorder increases, so does entropy. The molecules in ice, for example, are more ordered than those in water or steam, and so ice has a lower value of S (*i.e.* lower entropy).

The **Second Law of Thermodynamics** states that:

The entropy of the universe is always increasing.

A spontaneous process is one that occurs in nature without outside intervention—it happens "on its own", or spontaneously. A process is spontaneous only if the entropy of the universe increases as a result of the process.

References

Introductory Chemistry—A Foundation (6th Ed), Zumdahl, S.S. and DeCoste, D.J. (Houghton Mifflin, 2009) [ISBN 13: 978-0-618-80327-9]

Work directly from text, with exercises:

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