Chapter 6 - Thermochemistry

Section 6.1 - The Nature of Energy

1. **Energy** - the capacity to do work or to produce heat.
2. **Law of conservation of energy** - energy can be converted form one form to another but can be neither created nor destroyed.
3. **Potential Energy** - energy due to position or composition.
4. **Kinetic Energy** - energy due to the motion of the object and depends on the mass of the object m and its velocity v :

$$KE = \frac{1}{2}mv^2$$

5. **Heat (q)** - involves the transfer of energy between two objects due to a temperature difference.
6. **Work (w)** - a force acting over a distance.
7. **Pathway** - the way that energy is transferred is divided between work and heat depends on specific conditions.
8. **State function (property)** - a property of the system that depends only on its present state. A change in this function in going from one state to another state is independent of the particular pathway taken between the two states.
9. **Chemical Energy**

   A. **System** - the part of the universe on which we with to focus attention.
   B. **Surroundings** - includes everything else in the universe.
   C. **Exothermic** - energy flows out of the system.
   D. **Endothermic** - energy flows into the system.

10. **Thermodynamics** - study of energy and its interconversions.
11. **First Law of Thermodynamics** - (law of conservation of energy) the energy of the universe is constant.

Internal energy - (**E** of a system) sum of the kinetic and potential energies of all the "particles" in the system.

$$\Delta E = q + w$$

where ΔE represents the change in the system's internal energy, q represents heat and w represents work.

- if q is equal to +x, the system's energy is increasing.
- if q is equal to -x, the system's energy is decreasing.

Note that pressure is a force divided by an area and that the change in volume is a distance. These can be used to define work (w) as $w = -P\,\Delta V$

For an expanding gas, ΔV is a positive quantity because the volume is increasing. Thus ΔV and w must have opposite signs, so we have a negative sign in front of our equation, because it is doing work on our surroundings.

Section 6.2 - Enthalpy and Calorimetry

1. **Enthalpy** - at constant pressure the change in enthalpy equals the energy flow as heat (q).

$$\Delta H = \Delta E + P\,\Delta V = q_P$$

Enthalpy is a state function.
Since enthalpy at constant pressure is the measurement of heat, we can use the term heat of reaction to also define enthalpy.

$$\Delta H = H_{products} - H_{reactants}$$

2. **Calorimetry** - is the science of measuring heat, and is based on the temperature change when a body absorbs or discharges energy as heat.

3. **Heat capacity (C) -** is a measurement of how a substance responds to being heated.

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

4. **Specific heat capacity (s) -** the energy required to raise the temperature of one gram of a substance by one degree Celsius
5. **Molar heat capacity -** the energy required to raise the temperature of one mole of a substance by one degree Celsius.
6. **Energy released** = specific heat capacity × mass ×change in temperature

$$\mathbf{q = q_P = \Delta H = s \times m \times \Delta T}$$

heat of reaction - is an **extensive property**; it depends directly on the amount of the substance.
intensive property - is not related to the amount of the substance.

Calorimetry can also be done for constant volume. In this case our definition of **q** changes:

$$\mathbf{q = q_V = \Delta E, \text{ work (w) will be zero}}$$

Section 6.3 - Hess's Law

Hess's Law - 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.

$$\Delta H_1 = \Delta H_2 + \Delta H_3$$

Characteristics of Enthalpy Changes:

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.

Section 6.4 - Standard Enthalpies of Formation

Standard enthalpy of formation ($\Delta H_f°$) - the change in enthalpy that accompanies the formation of one mole of acompound from its elements with all substances in their standard states. **(see purple box on p. 258)**

The **enthalpy change** for a given reaction can be calculated by subtracting the enthalpies of formation of the reactants from the enthalpies of formation of the products.

$$\Delta H^\circ_{\text{reaction}} = \sum n_p \Delta H^\circ_f (\text{products}) - \sum n_r \Delta H^\circ_f (\text{reactants})$$

Elements are not included in the calculation because elements require no change in form.

Read Sections 6.5 and 6.6

****Notes have been derived from Zumdahl 4th ed. - All page and table references are made to this edition.**