

Chapter 6 Thermochemistry -homework sets form pages 254 – 263 (Chang)- **Set 1:** 16, 20, 26, 28, 34; **Set 2:** 38, 52, 54, 62, 74; **Set 3:** 80, 88, 98, 100, 108; **Set 4:** 113, 118, 122.

- I. Energy
 - a. The capacity to do work. Work – directed energy change resulting from a process.
 - b. Kinetic – energy of motion $E_k = \frac{1}{2} mv^2$ (m=mass,kg; v=velocity, meters/sec)
 - c. Potential – energy of position
 - d. Radiant – solar ; energy from the sun
 - e. Thermal – energy from the random motion of atoms or molecules
 - f. Chemical - energy stored in chemical structures, from reactions
 - g. Nuclear – energy from the atom
 - h. Law of Conservation of Energy – Total amount of energy in the universe is constant.
- II. Energy in Chemical reactions
 - a. Heat – transfer of thermal energy from high heat to low heat
 - b. Thermochemistry – study of heat change in chemical reactions
 - c. Systems – the part of the universe we are concerned about
 - d. Surroundings - the part of the universe outside of the system
 - i. Open – can exchange mass and energy between system and surroundings
 - ii. Closed – can only exchange energy not mass between system and surroundings
 - iii. Isolated – no exchange of energy or mass between system and surroundings
 - e. Exothermic – energy given off; energy given to surroundings by the system; products of reaction have less energy than reactants
 - f. Endothermic – energy taken in; energy given to the system by surroundings; products of reaction have more energy than reactants.
- III. Thermodynamics
 - a. The study of the interconversions of heat and other kinds of energy
 - b. State functions – Energy, pressure, volume and temperature; properties determined by the state of the system not how the state was achieved.
Example – skiing down a mountain
 - c. First Law – energy can be converted from one type to another but cannot be created or destroyed.
 - i. The energy change of system + the energy change of surroundings = zero.
 - ii. $\Delta E = q + w$; the internal energy (ΔE) of a system is the sum of the heat exchange (q) between the system and the surroundings and the work done (w) on (or by) the system.
 1. $q = +$ when endothermic – heat absorbed by system from surrounding
 2. $q = -$ when exothermic – heat given to surroundings by the system

3. $w = +$ when work is done on the system by surroundings
 4. $w = -$ when work is done on the surroundings by the system
- d. Work is defined as force times distance; $w = f \times d$ which can be changed into pressure and volume. Pressure = force / area (d^2) and volume = (d^3) so if we multiply pressure \times volume = force / $d^2 \times d^3 = f \times d$ so work is now $w = -P\Delta V$ which has units of $L \cdot atm$ and = 101.3 J. Work is NOT a state function.
- e. Heat is NOT a state function.
- IV. Enthalpy of Chemical Reactions
- a. Enthalpy (H) is basically the energy of a system plus the pressure times the volume of that system. $H = E + PV$ and since all are state functions then we can find changes $\Delta H = \Delta E + P\Delta V$, no change in pressure is common.
 - b. Enthalpy of a reaction, $\Delta H = \text{Enthalpy of products (H}_{\text{products}}) - \text{Enthalpy of the reactants (H}_{\text{reactants}})$. If ΔH is positive then endothermic since the products have more enthalpy than reactants which means the reaction absorbed energy. If ΔH is negative then exothermic since the products have less enthalpy than reactants which means the reaction gave off energy.
 - c. For our reality Enthalpy of reactions = Energy of reactions, $\Delta H = \Delta E$.
- V. Calorimetry
- a. Heat Capacity – energy needed to raise the temperature of a substance one Celsius degree. (J / C^0)
 - b. Specific Heat Capacity (C_p) Chang –(s) – energy needed to raise the temperature of one gram of a substance one Celsius degree. (J / gC^0)
 - c. $q = ms\Delta T$ Energy = mass \times Specific Heat Capacity \times change in temperature
 - d. Calorimeter – a special insulated container that does not allow for heat exchange from system to surroundings.
 - e. PRACTICE PROBLEMS!!!!!!
- VI. Enthalpy of Formation (ΔH^0_f)
- a. This is the heat change that results when 1 mole of a substance is formed from its elements at 1 atm.
 - b. ΔH^0_f of elements, in most stable form, is equal to zero
 - c. These can be used to determine the Enthalpy of reaction ΔH^0_{rxn}
 - i. $\Delta H^0_{\text{rxn}} = \Sigma \Delta H^0_f(\text{products}) - \Sigma \Delta H^0_f(\text{reactants})$
 1. use this for direct method
 - d. Can use Hess's Law to determine ΔH^0_f
 - i. The overall Enthalpy of reaction is equal to the sum of its parts the parts being other reaction pathways to get to the final destination. This is possible since ΔH^0_f is a state function.
 - ii. Using a series of reactions, reversing them as needed, reactions are added, reactants and products are added and subtracted as needed. When reactions are reversed the ΔH^0_{rxn} signs are also reversed and if they are multiplied then ΔH^0_{rxn} are also multiplied.
 - iii. PRACTICE PROBLEMS!!!!!!!!!!!!

e. The more negative the ΔH^0_f the more stable the compound

VII. Heat of solution and dilution

a. $\Delta H_{\text{soln}} = U + \Delta H_{\text{hydration}}$

b. U = lattice energy "Break It"

c. $\Delta H_{\text{hydration}}$ = heat of hydration "Surround It"

d. Heat of Dilution is energy associated when more solvent is added to a solution.