

AP Chemistry

Chapter 5 Outline

Your emphasis should be on sections 5.3-5.7

- a) The Nature of Energy
 - i) Energy = the capacity to do work or to transfer heat
 - ii) Work = energy used to cause an object with mass to move
 - (1) Heat = the energy used to cause the temperature of an object to increase
 - (a) Heat flows from hotter objects to colder objects, or energy is transferred between system & surroundings as a result in their difference in temperature
 - iii) Kinetic energy = energy of motion
 - iv) Potential energy = energy of position, when a force is operating on an object
 - (a) Force = any kind of push or pull exerted on an object, e.g. GRAVITY
 - v) Forms of energy are interconvertible
 - (1) Electrostatic potential energy (due to separation of charges)
 - (2) Chemical energy (potential energy stored in the arrangement of atoms of the substance)
 - (3) Thermal energy
 - vi) 2 common units of energy
 - (1) Joule, kJ
 - (2) calorie 1 calorie = 4.184 J also kilocalories (kcal)
 - vii) System = the portion of the universe singled out for study
 - (1) Surroundings = everything else
 - (2) Closed system = can exchange ENERGY but not matter with its surroundings

- b) The First Law of Thermodynamics
 - i) Energy is conserved. It cannot be created or destroyed.
 - ii) Internal energy, E = sum of ALL kinetic and potential energies of system components
 - iii) $\Delta E = E_{\text{final}} - E_{\text{initial}}$ This simple relationship is really useful for different state functions!
 - iv) $+\Delta E$ means system has absorbed energy from surroundings (endothermic or endergonic)
 - v) $-\Delta E$ means system has transferred energy to surroundings (exothermic or exergonic)
 - vi) ΔE can be related to work and heat, i.e. $\Delta E = q + w$
 - vii) There is no easy way to know the precise value of E . E is influenced by temperature and pressure, as well as the total amount of matter in the system.
 - (1) E is a **STATE FUNCTION**. This means that this property can be determined by specifying the system's condition.
 - (2) Because E is a state function, ΔE depends only on the initial and final states of the system, **not the pathway** by which the change occurred.

c) Enthalpy, H

- i) Enthalpy is a state function.
- ii) $H = E + PV$, where PV deals with pressure and volume work for the expansion and compression of gases. You don't need to know this in any particular detail.
- iii) For most processes studied, it turns out that $\Delta H =$ heat gained or lost at constant pressure. (Most chemists work at constant pressure, with only P-V work.) This is really useful, because we can measure q changes (remember $Q = mC_p\Delta T$)

d) Enthalpies of Reaction

- i) $\Delta H = H_{\text{products}} - H_{\text{reactants}}$
- ii) ΔH is usually reported at the end of thermochemical equation; + sign for endothermic, -sign for exothermic
- iii) Enthalpy diagrams
 - (1) Endothermic reactions: products have higher enthalpy than reactants
 - (2) Exothermic reactions: reactants have higher enthalpy than products (far more common!)
 - (3) Enthalpy is an extensive property.
 - (a) More reactants used \rightarrow more heat absorbed/lost! In other words, you can treat these problems as a stoichiometry problem, using factor label
 - (b) If you reverse a reaction, the sign of ΔH changes.
 - (c) The enthalpy change for a reaction depends on the state of the reactants and products. They must be specified! For example, it takes more energy to vaporize liquid water, so water vapor has a higher enthalpy than liquid water.

e) Calorimetry

- i) Calorimeter = device to measure heat flow (Styrofoam cups work well to make constant pressure calorimeters!)
- ii) Specific Heat = $C_p = s$ = the amount of energy required to raise the temperature of 1 g of a substance by 1 K (remember that ΔT in K = ΔT in $^{\circ}\text{C}$)
 - (1) $Q = m C_p\Delta T = M s \Delta T$ KEY EQUATION TO MEMORIZE!
 - (2) Simple calorimetry principle: heat lost = heat gained $q_{\text{soln}} = -q_{\text{rxn}}$
 - (3) Specific heat of water = 4.184 J/g K
 - (4) Bomb calorimetry these are actually constant volume conditions!
 - (a) $q_{\text{rxn}} = -C_{\text{cal}} \times \Delta T$ where C_{cal} = heat capacity of calorimeter

f) Hess' Law

- i) Many enthalpies of reaction have been experimentally measured. For some reactions, however, it is impractical or impossible to directly measure ΔH .
- ii) Hess' Law states: If a reaction is carried out in a series of steps, ΔH for the overall reaction will equal the sum of the enthalpy changes for the individual steps.
 - (1) If a reaction needs to be multiplied by a coefficient, ΔH is multiplied by that coefficient.
 - (2) If a reaction is reversed, the sign of ΔH changes.
 - (3) Hess' Law can be represented using enthalpy diagrams.
 - (4) Look at the practice problems! These are critical problem types!

a) Enthalpies of Formation

- i) ΔH_f = the energy change for the formation of a compound from its constituent elements
- (1) depends on conditions of temperature, pressure, and state of reactants and products, so we use enthalpies at “standard state:” (pure form at atmospheric pressure, and 298 K), reported in kJ/mol of compound
 - (2) By definition, the standard enthalpy of formation of the most stable form of any element is zero (Remember this!)
 - (3) We can use these as another way to carry out Hess’ Law
 - (a) $\Delta H_{rxn}^{\circ} = \sum \Delta H_f^{\circ}(\text{products}) - \sum \Delta H_f^{\circ}(\text{reactants})$ See [sample problems](#)
 - (b) In order to do these problems, you need a [table of standard enthalpies of formation](#).

Skip Section 5.8