

CHAPTER 8

THERMOCHEMISTRY: CHEMICAL ENERGY

We will spend three lecture days on this chapter and will only do selected parts of it. Note that we will breeze through Sections 1 – 6 and then focus on Sections 7 - 12. Here is the plan:

- ❖ Day 1. Briefly cover Section 1 – 5 and then 6 and 9 (units on thermal energy, quantitative energy calculations).
- ❖ Day 2 we do calculations involving calorimetry and heats of formation (Sections 7 and 9).
- ❖ Day 3 we will cover Hess's law and bond energies (Sections 8 and 10).

□ **Read the introductory paragraph to Chapter 8 and Sections 8.1 – 8.5.** We will use the summary for these sections (other side)

- Memorize the definition of the standard state (25 °C and 1 atm). If solutions are involved, standard state is 1 molar.

□ Do

Section 8.6 □ Given a balanced chemical equation and enthalpy change for a chemical reaction, calculate the enthalpy change per mole or per gram of each reactant and product.

- Do Problems: 7, 8, 56, 58, 60,

Section 8.7 □ Perform calculations involving specific heat (or molar heat capacity), heat flow, and temperature change.

Calculate ΔH or ΔE in a calorimetry experiment.

- Do 9 – 11, 62, 64, 66, 68 (Hint: limiting reagent problem: Go Moles!)

Section 8.8 □ Use Hess's law to calculate a standard heat of reaction.

- Do 12 – 14

Section 8.9 □ Use standard heats of formation to calculate a standard heat of reaction.

- Do 15, 16, 70, 74, 76, 78, 80, 82, 84, 86

Section 8.10 □ Use bond dissociation energies to approximate a standard heat of reaction.

- Do 17, 18, 88, 90,

Section 8.11 □ Write balanced equations for combustion reactions.

- Use Hess's Law to calculate enthalpies of combustion.



“Hey! Look what Zog do!”

□ Do 19

Section 8.12 Predict whether entropy increases or decreases for a chemical reaction or physical change.

□ Do 20, 21, 96, 100

Skip Section 8.13 We will do this next semester in huge detail!

Quick & Easy Thermodynamics

1. Internal Energy = Kinetic Energy + Potential Energy: $E = KE + PE$

2. Change in energy:

In general: $\Delta E = E_{\text{final}} - E_{\text{initial}}$

For reactions: $\Delta E = E_{\text{products}} - E_{\text{reactants}}$

3. First Law of Thermodynamics: (Law of Conservation of Energy)

$\Delta E = \text{heat} + \text{work}$ (q = heat and w = work)

4. Work is done when there is a change in volume:

$$w = -P\Delta V$$

$$\Delta E = q - P\Delta V$$

5. Chemists don't want to worry about volume changes (we use open beakers - and don't want to measure the gas volumes produced, etc.) so we invented enthalpy, H:

$$H = E + PV$$

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

$$\Delta H = (q - P\Delta V) + P\Delta V = q$$

6. ΔH and ΔE are very similar in value under normal conditions - and are identical when there is no change in volume.

7. The standard state. When values are reported at 298 K and 1 atm and 1 M, a little "o" is added to ΔH to indicate the values are at standard state. Thus, ΔH° indicates the value is at standard state and ΔH indicates that the conditions were not standard conditions.