Thermochemistry Study Notes (chapter 5)

You should be familiar with the sample calculations in chapter 5 (especially examples 5.4, 5.7, and 5.8).

5.1-5.2

How many Joules are in one calorie?	1 cal = 4.2 J
Give water's specific heat (with units)	4.2 J/(g °C)
Define "thermal property". Give three examples.	The ability of a substance to absorb heat without changing chemically. E.g. heat capacity, molar heat capacity, specific heat (capacity).
Distinguish between heat capacity, molar heat capacity, and specific heat capacity.	Heat capacity is the amount of heat that must move into or out of an object to change its temperature by 1°C. It has units of J/°C. Specific heat is similar, except it is expressed on a per gram basis (J/(g·°C)). Molar heat capacity has units of J/(mol·°C).
How can one calculate the energy that is absorbed or released in a reaction?	energy change = cmΔT or q = cmΔT, where c is specific heat, m is mass, and ΔT is change in temperature.

5.3

What does enthalpy mean? What is its symbol?	Enthalpy (symbolized H) is the total energy of a system including its total kinetic and potential energy.
What does ΔH mean? By definition, what is the equation for ΔH?	ΔH is the change in enthalpy when reactants go to products: $\Delta H = H_{products} - H_{reactants}$ (same as: $\Delta H = H_{final} - H_{initial}$)
What is the limitation of this equation?	We cannot calculate ΔH from this equation because it is impossible to know the initial enthalpy ($H_{initial}$ or $H_{reactants}$) or the final enthalpy (H_{final} or $H_{products}$)
What is the benefit of this equation?	It helps us define the meaning of negative and positive values for ΔH.
What sign can ΔH have? Explain.	ΔH can be either positive or negative. +ΔH indicates an endothermic reaction (i.e. $H_{products} > H_{reactants}$). –Δ$H$ indicates an exothermic reaction (i.e. $H_{products} < H_{reactants}$).
In summary, what happens to the enthalpy of a system in endothermic vs. exothermic reactions?	Enthalpy increases in endothermic reactions. Enthalpy decreases in exothermic reactions.
If ΔH cannot be calculated from the above equation, how is it measured?	ΔH = q at constant pressure. Thus, if we determine q (remember q=cmΔt) at constant pressure then we will have the value of ΔH.
What two conditions must be true for ΔH to equal q?	As mentioned the pressure must be constant. Also, all of the enthalpy change must be in the form of heat energy.
What law allows us to use the equation ΔH = q provided all of the change in energy occurs as heat?	The law of conservation of energy. This implies, for example, that the amount of energy lost from reacting chemicals is exactly the same as that gained by the surroundings. Thus, we are justified in using ΔH = q.

5.4

How is q measured experimentally?	With a calorimeter.
What are two common types of calorimeters? What are the advantages of each?	A bomb calorimeter is a very accurate way of measuring q (and therefore ΔH). A "coffee cup" calorimeter is a low tech and cheap method of measuring q.
What are the components of a bomb calorimeter? What are their functions?	The "bomb" is a chamber where a substance is burned. The valve and tube to the bomb provide a gas (usually O_2) under pressure. The wires to the bomb provide a spark to initiate burning. The water (surrounded by a heavily insulated vat) will gain all of the heat energy lost by the burning substance. The stirrer will ensure that the water is uniform in temperature. The thermometer measures the temperature of the water (fig. 5.5, pg. 161).
How is q calculated from a calorimeter?	q=cmΔT. c refers to the water being heated; c = 4.2 J/g°C. m is the mass of water (which will vary). ΔT is calculated from the initial and final temperature readings from the thermometer. Note: for a bomb calorimeter c and m may be combined; for example instead of 4.2 J/g°C and 100 g, you might be given 420 J/°C (i.e. you could be given heat capacity instead of specific heat capacity).