Notes—Review of the First Half of Ch. 6: Enthalpy and Calorimetry

(You already have notes over this first part of the chapter. I will give you an overview of the information. The first "Lecture video" you will watch covers Enthalpy and Calorimetry and is a review.)

$\Delta E = q + w$ q is heat but is also equal to ΔH under constant pressure (q_p) when no work is done.
w is work and is limited to PΔV work (Pressure in atm, Volume in L)
(Remember to multiply PΔV by 101.3 J/(atm) (L). 1 (L)(atm) = 101.3 J)
"w" will then be in Joules.

$q_{reaction} = -q_{solution}$ Heat given off by the reaction (-) is absorbed by the solution (+) or visa/versa. When calculating the heat either absorbed or released by a solution, you need to change the sign to get the sign of the reaction.

ΔH is Enthalpy change.
It is equal to q_p under constant pressure. However, the value of q must be changed to kJ and divided by the number of moles uses in the reaction.
In a reaction, $\Delta H_{rxn} = H_{products} - H_{reactants}$. (We will discuss this in depth later)

Calorimetry

$q = s\ m\ \Delta T$ "q" is heat, "s" is specific heat capacity, "m" is mass in grams, ΔT is $(T_2 - T_1)$

(another form of this equation is **$q = c\ n\ \Delta T$**) This equation uses moles (n) instead of mass (m).
"c" is the molar heat capacity of the substance.
If you know "s", you can get "c" by multiplying the value of "s" by the molar mass of the substance. For example, "s" of water is 4.184 J/g^{0}C. "c" would be (4.184 J/g^{0}C) (18.0 g/mol$_{water}$) = 75.3 J/mol^{0}C.

Specific heat capacity (s): The energy required to raise the temp of 1 gram of a substance, 1^{0}C.
Molar heat capacity (c): The energy required to raise the temp of 1 mole of a substance, 1^{0}C.

Coffee-cup Calorimeter: (Most of the time with a coffee-cup calorimeter, we are finding the heat of the solution so the reaction will be the opposite sign.)

Remember $q_p = \Delta H$ (To get ΔH from q_p, convert the Joules of q to kJ then divide by the number of moles of the substance in the reaction.)

Example: If $q = +2.6 \times 10^4$ J then q = 26 kJ. The reaction used 105 g
NH_4NO_3. 105g/88 g/mole for the ammonium nitrate will
Equal 1.2 moles.
$\Delta H = 26$ kJ/1.2 moles = 22 kJ/mol.

Properties of Enthalpy

If a reaction has a ΔH (+), then the reaction will go forward better under high heat conditions since it absorbs heat as it reacts.
If a reaction has a ΔH (-), then the reaction will go forward better under low heat (cooling) conditions since it releases heat.

Breaking bonds requires energy (heat)—endothermic.
Change of state that goes s→ l → g requires energy (heat)—endothermic.

Making bonds releases energy (heat)—exothermic.
Change of state that goes g → l → s releases energy (heat)—exothermic.

In a reaction such as:

$$\mathbf{N_2\,(g) + O_2\,(g) \rightarrow 2\,NO_2\,(g) \qquad \Delta H = +180\ kJ}$$

The coefficients are the number of moles.
1 mol N_2, 1 mol O_2, and 2 mol NO_2 for this reaction.

If I asked how much energy is absorbed for **1 mol of NO_2**, then I would set up a ratio of mol/kJ:

$$\frac{2\text{ mol }NO_2}{180\text{ kJ}} = \frac{1\text{ mol }NO_2}{X}$$

Cross multiply and solve for X.
X = 90 kJ

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