

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 $P\Delta V$ work (Pressure in atm, Volume in L)
(Remember to multiply $P\Delta V$ by 101.3 J/(atm) (L). 1 (L)(atm) = 101.3 J)
“ w ” will then be in Joules.

$q_{\text{reaction}} = -q_{\text{solution}}$ 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, $\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$. (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/g⁰C. “ c ” would be (4.184 J/g⁰C) (18.0 g/mol_{water}) = 75.3 J/mol⁰C.

Specific heat capacity (s): The energy required to raise the temp of 1 gram of a substance, 1⁰C.

Molar heat capacity (c): The energy required to raise the temp of 1 mole of a substance, 1⁰C.

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

Properties of Enthalpy

If a reaction has a $\Delta H (+)$, then the reaction will go forward better under high heat conditions since it absorbs heat as it reacts.

If a reaction has a $\Delta H (-)$, then the reaction will go forward better under low heat (cooling) conditions since it releases heat.

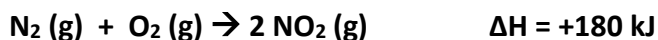
Breaking bonds requires energy (heat)—endothermic.

Change of state that goes $s \rightarrow l \rightarrow g$ requires energy (heat)—endothermic.

Making bonds releases energy (heat)—exothermic.

Change of state that goes $g \rightarrow l \rightarrow s$ releases energy (heat)—exothermic.

In a reaction such as:



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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