AP CHEMISTRY

TOPIC 9: THERMODYNAMICS, PART B, EXAMPLES, PART II Day 109:

- Hess's law
- Enthalpy of Formation
- Spontaneity
- Entropy
- Heat Capacity

HEAT CAPACITY:

Heat capacity, C_p , is a measure of how much the temperature of an object is raised when it absorbs heat. An object with a large heat capacity can absorb a lot of heat without undergoing much of a change in temperature, whereas an object with a small heat capacity shows a large increases in temperature even if a small amount of heat is absorbed.

The equation that appears on your AP Chem equation sheet:

$$C_p = \frac{\Delta H}{\Delta T} \quad \text{units:} \quad \frac{J}{{}^0C}$$

Example #1: Calculate the heat capacity of an unknown substance, with an initial temperature of 23.3 ^{0}C, that was able to absorb 8735 joules of energy and its final temperature was 25.7 ^{0}C.

$$C_p = \frac{\Delta H}{\Delta T} = \frac{8735\ J}{(\ 25.7^0 C - 23.3^0 C\ )} = 3640\ \text{J}\ {}^0\text{C}^{-1}$$

Example #2: . Determine the standard enthalpy of reaction for the combustion of hydrogen sulfide gas, which proceeds according to the reaction shown below:

$$2\ H_2S_{(g)} + 3\ O_{2\,(g)} \rightarrow 2\ H_2O_{(l)} + 2\ SO_{2\,(g)}$$

The standard enthalpies for the constituents are as follows:

Formula	ΔH^0_f (kJ mol^{-1})
$H_2S_{(g)}$	- 20
$H_2O_{(l)}$	- 285.8
$SO_{2\,(g)}$	- 296.8

Answer:

$$\Delta H^0_{rxn} = n\,\Delta H^0_f \text{ (products)} - m\,\Delta H^0_f \text{ (reactants)}$$

$$\Delta H^0_{rxn} = [\ (2)(-285.8\text{ kJ}) + (2)(-2896.8\text{ kJ})\ ] - [\ (2)(20\text{ kJ}) + 0\ ] = -1125.2\text{ kJ}$$

Example #3: Use the information shown above to determine the standard enthalpy change for the formation of nitrogen dioxide shown in the reaction below:

$$2\ NO_{(g)} + O_{2\,(g)} \rightarrow 2\ NO_{2\,(g)}$$

Formula	ΔH^0_f (kJ mol^{-1})
$NO_{(g)}$	90.37
$O_{2\,(g)}$	0
$NO_{2\,(g)}$	33.84

Answer:

$$\Delta H^0_{rxn} = n\,\Delta H^0_f \text{ (products)} - m\,\Delta H^0_f \text{ (reactants)}$$

$$\Delta H^0_{rxn} = [\ (2)(33.84\text{ kJ})\ ] - [\ (2)(90.37\text{ kJ}) + 0\ ] = -113.06\text{ kJ}$$

Example #4: Calculate the enthalpy for the following reaction:

$$N_{2\,(g)} + 2\,O_{2\,(g)} \rightarrow 2\ NO_{2\,(g)} \quad \Delta H^0 = ???\ kJ$$

Using the following two equations:

$$N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,NO_{(g)} \qquad \Delta H^0 = +180\ kJ$$

$$2\ NO_{2\,(g)} \rightarrow 2\,NO_{(g)} + O_{2\,(g)} \qquad \Delta H^0 = +112\ kJ$$

Answer

$$N_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,NO_{(g)} \qquad \Delta H^0 = +180\ kJ$$

$$2\,NO_{(g)} + O_{2\,(g)} \rightarrow 2\ NO_{2\,(g)} \qquad \Delta H^0 = -112\ kJ$$

$$N_{2\,(g)} + 2\,O_{2\,(g)} \rightarrow 2\ NO_{2\,(g)} \quad \Delta H^0 = +68\ kJ$$

Example #5: Calculate the enthalpy for the following reaction:

$$2\,N_{2\,(g)} + 5\,O_{2\,(g)} \rightarrow 2\,N_2O_{5\,(g)} \quad \Delta H^0 = ???\ kJ$$

Using the following two equations:

$$H_{2\,(g)} + \frac{1}{2}O_{2\,(g)} \rightarrow H_2O_{(l)} \qquad \Delta H^0 = -285.8\ kJ$$

$$N_2O_{5\,(g)} + H_2O_{(l)} \rightarrow 2\,HNO_{3\,(l)} \qquad \Delta H^0 = -76.6\ kJ$$

$$\frac{1}{2}\,N_{2\,(g)} + \frac{3}{2}\,O_{2\,(l)} + \frac{1}{2}\,H_{2\,(g)} \rightarrow HNO_{3\,(l)} \qquad \Delta H^0 = -174.1\ kJ$$

Answer

$$2\ (\ H_2O_{(l)} \rightarrow H_{2\,(g)} + \frac{1}{2}O_{2\,(g)}) \quad \Delta H^0 = 2\,(+285.8\ kJ) = +571.6\ kJ$$

$$2\ (\,2\,HNO_{3\,(l)} \rightarrow N_2O_{5\,(g)} + H_2O_{(l)}\,) \quad \Delta H^0 = 2\,(+76.6\ kJ) = +153.3\ kJ$$

$$4\ (\frac{1}{2}\,N_{2\,(g)} + \frac{3}{2}\,O_{2\,(l)} + \frac{1}{2}\,H_{2\,(g)} \rightarrow HNO_{3\,(l)}\,) \quad \Delta H^0 = 4\,(-174.1\ kJ) = -696.4\ kJ$$

$$2\,N_{2\,(g)} + 5\,O_{2\,(g)} \rightarrow 2\,N_2O_{5\,(g)} \quad \Delta H^0 = +28.5\ kJ$$