AP CHEMISTRY

TOPIC 9: THERMODYNAMICS, PART D,

Day 111:

- Gibbs Free Energy (continued)

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1. Acetylene gas, C_2H_2, is used in gas welding procedures and is a very important commercial gas. Use the data below to answer the following questions about the combustion of acetylene gas.

Substance	ΔH^0_f (kJ mol^{-1})	S^0 (J mol^{-1} K^{-1})
$C_2H_{2\,(g)}$	227	200.9
$O_{2\,(g)}$	0	205.1
$CO_{2\,(g)}$	-393.5	213.7
$H_2O_{\,(l)}$	-285.8	188.8

Use the data above to answer the questions that follow. Assume all reactions take place at 25 ^{0}C.

a) Write a complete balanced chemical equation for the combustion of acetylene. Assume that carbon dioxide and water are the only products.

$$2\ C_2H_{2\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ CO_{2\,(g)} + 2\ H_2O_{\,(l)}$$

b) Calculate the standard enthalpy change, ΔH^0, for the combustion of one mole of acetylene gas.

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

$$\Delta H^0{}_{comb} = \left[4\left(\frac{-393.5\ kJ}{mol}\right)+2\left(\frac{-285.8\ kJ}{mol}\right)\right]-\left[2\left(\frac{227.0\ kJ}{mol}\right)+5\left(\frac{0\ kJ}{mol}\right)\right] = -2600\ \frac{kJ}{mol}$$

Typically the question wants to know the enthalpy change for one mole of acetylene: Therefore,

$$\Delta H^0{}_{comb} = -2600\ kJ \quad \text{for 2 moles of } C_2H_{2\,(g)}$$

$$C_2H_{2\,(g)} + \frac{5}{2}\ O_{2\,(g)} \rightarrow 2\ CO_{2\,(g)} + H_2O_{\,(l)}$$

Simply drop the unit mole, since we know that this is for one mole: $\dfrac{-2600\ kJ}{2} = -1300\ kJ$

c) Calculate the standard entropy change, ΔS^0, for the combustion of **one mole of acetylene gas.**

$$\Delta S^0{}_{rxn} = \Sigma\ \Delta S^0\ \text{(products)} - \Sigma\ \Delta S^0\ \text{(reactants)}$$

$$\Delta S^0{}_{rxn} = \left[2\left(\frac{213.7\ J}{mol\cdot K}\right)+1\left(\frac{188.8\ J}{mol\cdot K}\right)\right]-\left[1\left(\frac{200.9\ J}{mol\cdot K}\right)+\frac{5}{2}\left(\frac{205.0\ J}{mol\cdot K}\right)\right] = -97.2\ \frac{J}{mol\cdot K}$$

d) Determine the value of ΔG^0 for the reaction.

$$\Delta G = \Delta H - T\Delta S$$

$$\text{Convert (J to kJ):}\quad \Delta S = \left(\frac{-97.2\ J}{mol\cdot K}\right)\left(\frac{1\ kJ}{1000\ J}\right) = -0.0972\frac{kJ}{mol\cdot K}$$

$$\Delta G = -1300\ \frac{kJ}{mol} - (\ 298\ K\ )\left(-0.0972\frac{kJ}{mol\cdot K}\right) = -1271\ \frac{kJ}{mol}$$

The question does not ask, but the ΔG is negative, therefore, the reaction is spontaneous in the forward direction.

e) If 1 mole of acetylene gas is burned and all of the evolved heat is used to heat a 6.00 kg sample of pure water, what will the temperature change of the water be? The specific heat for water is 4.18 J g^{-1} K^{-1}.

The heat of combustion (from part b) is equal to the amount of heat the acetylene gas will generate. $\Delta H = q$.

Also, since this is a calorimeter problem, the energy absorbed by the water is POSITIVE (change the sign).

$$q = m\ c\ \Delta T$$

$$1.30 \times 10^6 \text{ J} = (6000 \text{ g})(4.18 \text{ J g}^{-1} \text{ K}^{-1})\,\Delta T$$

$$\Delta T = \frac{1.30\times10^6\ J\ (g\cdot K)}{(6000\ g)(4.18\ J)} = 51.8\ K\ \ or\ \ 51.8\ ^0C$$

2. An ice cube is placed in a flask filled with water at room temperature and allowed to sit until the ice cube has completely melted. Describe the changes in enthalpy, entropy, free energy, and temperature that occur during the time interval.

	Ice Cube	*Water*
Enthalpy = ΔH	*Endothermic*	*Exothermic*
Entropy = ΔS	*Increase (+)*	*Decrease (-)*
Free Energy = ΔG	*Negative*	*Negative*
Temperature	*increases*	*decreases*

- *The enthalpy for the ice cube is increasing (positive value - Endothermic) since the water (surrounding the ice cube) has more energy than the ice cube – energy always moves from areas of high energy to areas of low energy until the two items (the system) are at the same temperature. Since the water is losing (giving) energy to the ice cube the water is exothermic.*

- *The entropy of the ice cube is increasing the disorder (positive value): changing from a solid to a liquid. Where the entropy of the water is decreasing the disorder (negative value) since the temperature of the water is decreasing and allowing the molecules to become more ordered – decreasing the disorder (getting closer to becoming a solid).*

- *This process occurs spontaneously, so ΔG must be negative.*

3. 40.0 g of sodium hydroxide pellets are added to 500 mL of water, and most of it dissolves very quickly. The temperature of the system increases. Describe the changes in enthalpy, entropy, free energy.

	NaOH	*Water*
Enthalpy = ΔH	*Exothermic*	*Endothermic*
Entropy = ΔS	*Increase (+)*	*Increase (+)*
Free Energy = ΔG	*Negative*	*Negative*
Temperature	*decreases*	*increases*

- *The enthalpy for the NaOH is decreasing since the water (surrounding the NaOH) has less energy than the NaOH – The breaking of the ionic bonds in sodium hydroxide is a very exothermic event. We can reason that the breaking of the bonds with NaOH (in this case) is exothermic since the water (solution) heats up as the compound dissociates. Since water is gaining (absorbing) energy from the NaOH, the water is endothermic.*

- *The entropy of the NaOH is increasing the disorder (positive value): changing from a solid to an aqueous (ions surrounded by water) solution. Where the entropy of the water is increasing in disorder (positive value) since the temperature of the water is increasing and the molecules are becoming less ordered (getting closer to becoming a gas).*

- *This process (dissolving) occurs spontaneously, so ΔG must be negative.*

4. Commercial instant ice packs are available that contains a mixture of ammonium nitrate and water separated by a barrier. When the ice pack is twisted, the barrier breaks and the two substances mix. The temperature rapidly decreases as the ammonium nitrate dissolves in the water. Describe the changes in enthalpy, entropy, free energy.

	NH_4NO_3	*Water*
Enthalpy = ΔH	*Endothermic*	*Exothermic*
Entropy = ΔS	*Increase (+)*	*Decrease (-)*
Free Energy = ΔG	*Negative*	*Negative*
Temperature	*increases*	*decreases*

- *The enthalpy for the* NH_4NO_3 *is increasing since the water (surrounding the* NH_4NO_3*) has more energy than the* NH_4NO_3 *– The* NH_4NO_3 *absorbs energy (Endothermic) to break the ionic bonds. Since the water is losing energy to the* NH_4NO_3 *the water is exothermic (as a result the solution becomes cooler).*

- *The entropy of the* NH_4NO_3 *is increasing the disorder (positive value): changing from a solid to an aqueous (ions surrounded by water)solution. Where the entropy of the water is decreasing the disorder (negative value) since the temperature of the water is decreasing and becoming more ordered (getting closer to becoming a solid).*

- *This process occurs spontaneously, so ΔG must be negative.*