

Thermochemistry Homework

Specific Heat Capacity Problems: $q = C_p \cdot m \cdot \Delta T$

- 1) Calculate the amount of energy required to raise the temperature of 145 grams of water from 22.3°C to 75°C. (Specific Heat of Water = 4.184 J/g·C)

$$q = (4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}) (145 \text{ g})(75^\circ\text{C} - 22.3^\circ\text{C})$$

$$q = 31972 \text{ Joules}$$

- 2) The specific heat capacity of iron is 0.45 J/g·C. If 47 Joules of energy is required to raise the temperature of a sample of iron from 25°C to 90°C, what is the mass of the sample?

$$m = \frac{q}{C_p \cdot \Delta T}$$

$$\text{mass} = \frac{47 \text{ J}}{(0.45 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}) (90^\circ\text{C} - 25^\circ\text{C})}$$

$$\text{mass} = 1.6 \text{ grams}$$

- 3) A 35.2 gram sample requires 1251 Joules of energy to heat the sample by 25°C. What is the specific heat capacity of the sample?

$$C_p = \frac{q}{m \cdot \Delta T}$$

$$\text{Specific Heat Capacity} = \frac{1251 \text{ J}}{(35.2 \text{ grams})(25^\circ\text{C})}$$

$$\text{Specific Heat Capacity} = 1.42 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}}$$

- 4) A 955 g bar of aluminum at 75°C with a heat capacity of 0.89 J/g·°C is dropped into a calorimeter containing 25.0 g of liquid water at 25°C. What will be the final temperature of the mixture?

$$+q_{\text{water}} = -q_{\text{aluminum}}$$

$$+[C_p \cdot m \cdot (T_f - T_i)] = -[C_p \cdot m \cdot (T_f - T_i)]$$

Use the distributive property to separate temperatures. Watch for sign changes!!!

$$[(Cp \cdot m \cdot T_f) - (Cp \cdot m \cdot T_i)] = [- (Cp \cdot m \cdot T_f) + (Cp \cdot m \cdot T_i)]$$

STEP 1: Solve for heat absorbed...

$$+ q_{water} = [(Cp \cdot m \cdot T_f) - (Cp \cdot m \cdot T_i)]$$

$$+ q_{water} = [(4.184 \text{ J/g}^\circ\text{C} \cdot 25 \text{ g} \cdot T_f) - (4.184 \text{ J/g}^\circ\text{C} \cdot 25 \text{ g} \cdot 25^\circ\text{C})]$$

$$+ q_{water} = (104.6 \frac{\text{J}}{^\circ\text{C}}) T_f - 2615 \text{ J}$$

STEP 2: Solve for heat transferred ...

$$- q_{aluminum} = [-(Cp \cdot m \cdot T_f) + (Cp \cdot m \cdot T_i)]$$

$$- q_{aluminum} = [-(0.89 \text{ J/g} \cdot ^\circ\text{C}) \cdot (955 \text{ g}) \cdot T_f + ((0.89 \text{ J/g} \cdot ^\circ\text{C}) \cdot (955 \text{ g}) \cdot (75^\circ\text{C}))]$$

$$- q_{aluminum} = (-849.95 \text{ J/g}^\circ\text{C}) \cdot T_f + 63746.25 \text{ J/g}^\circ\text{C}$$

STEP 3: Combine to solve for Final Temperature ...

$$+ q_{water} = - q_{aluminum}$$

$$(104.6 \frac{\text{J}}{^\circ\text{C}}) T_f - 2615 \text{ J} = (-849.95 \frac{\text{J}}{^\circ\text{C}}) T_f + 63746.25 \text{ J}$$

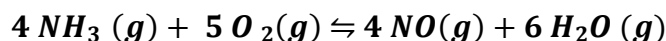
Combine like terms to get...

$$(954.55 \frac{\text{J}}{^\circ\text{C}}) T_f = 66361.25 \text{ J}$$

$$T_f = 69.52^\circ\text{C}$$

Enthalpy, Entropy, and Gibb's Free Energy Problems:

- 1.) Calculate the Enthalpy (ΔH) required to react ammonia with oxygen to form nitric oxide gas and water based on the reaction below. Is the reaction exothermic or endothermic?

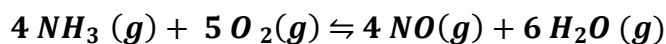


$$\Delta H_{rxn} = \sum \Delta H_{products} - \sum \Delta H_{reactants}$$

$$\begin{aligned} \Delta H_{rxn} &= \left[(4 \text{ mol} \cdot 90.3 \frac{\text{kJ}}{\text{mol}}) + (6 \text{ mol} \cdot -241.8 \frac{\text{kJ}}{\text{mol}}) \right] - \left[(4 \text{ mol} \cdot -45.9 \frac{\text{kJ}}{\text{mol}}) + (5 \text{ mol} \cdot 0 \frac{\text{kJ}}{\text{mol}}) \right] \\ \Delta H_{rxn} &= [-1089.6 \frac{\text{kJ}}{\text{mol}}] - [-183.6 \frac{\text{kJ}}{\text{mol}}] \\ \Delta H_{rxn} &= -906 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

ΔH is negative so the reaction is **EXOTHERMIC**

- 2.) Calculate the Entropy (ΔS) change for the reaction above.



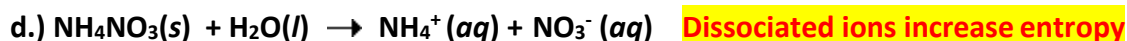
$$\Delta S_{rxn} = \sum \Delta S_{products} - \sum \Delta S_{reactants}$$

$$\begin{aligned} \Delta S_{rxn} &= \left[(4 \text{ mol} \cdot 210.8 \frac{\text{J}}{\text{mol} \cdot \text{K}}) + (6 \text{ mol} \cdot 188.7 \frac{\text{J}}{\text{mol} \cdot \text{K}}) \right] - \left[(4 \text{ mol} \cdot 192.8 \frac{\text{J}}{\text{mol} \cdot \text{K}}) + (5 \text{ mol} \cdot 205 \frac{\text{J}}{\text{mol} \cdot \text{K}}) \right] \\ \Delta S_{rxn} &= [1975.4 \frac{\text{J}}{\text{mol} \cdot \text{K}}] - [1796.2 \frac{\text{J}}{\text{mol} \cdot \text{K}}] \\ \Delta S_{rxn} &= 179.2 \frac{\text{J}}{\text{mol} \cdot \text{K}} \end{aligned}$$

NOTE: Gibb's Free Energy Requires units to be in kilojoules!

$$\Delta S_{rxn} = 0.1792 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}$$

3.) Which of the following processes will lead to an increase in the entropy of the system?



4.) Use the Gibb's Free Energy equation to determine if the reaction in Problem #1 will be spontaneous at STP.

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

$$\Delta G = (-906 \frac{\text{kJ}}{\text{mol}}) - (273 \text{ K})(0.1792 \frac{\text{kJ}}{\text{mol} \cdot \text{K}})$$

$$\Delta G = -954.9 \frac{\text{kJ}}{\text{mol}}$$

ΔG is negative so the reaction is Spontaneous!

5.) Use the Gibb's Free Energy equation to determine the reaction below will be spontaneous at STP. Assume Enthalpy (ΔH) = 30.91 kJ/mol and Entropy (ΔS) = 93.2 J/mol·K.



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

$$\Delta G = (30.91 \frac{\text{kJ}}{\text{mol}}) - (273 \text{ K})(0.0932 \frac{\text{kJ}}{\text{mol} \cdot \text{K}})$$

$$\Delta G = +5.4664 \frac{\text{kJ}}{\text{mol}}$$

ΔG is positive so the reaction is NOT spontaneous!