

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)

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

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

- 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?

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

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

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

- 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...

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

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

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

STEP 2: Solve for heat transferred ...

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

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

$$- q_{aluminum} = \left(-849.95 \frac{J}{^\circ C} \right) T_f + 63746.25 J$$

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

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

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

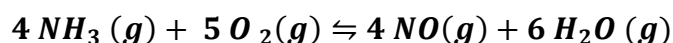
Combine like terms to get...

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

$$T_f = 69.52^\circ C$$

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?



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

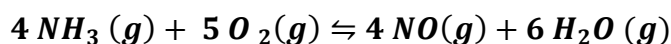
$$\Delta H_{rxn} = \left[\left(4 \text{ mol} \cdot 90.3 \frac{\text{kJ}}{\text{mol}} \right) + \left(6 \text{ mol} \cdot -241.8 \frac{\text{kJ}}{\text{mol}} \right) \right] - \left[\left(4 \text{ mol} \cdot -45.9 \frac{\text{kJ}}{\text{mol}} \right) + \left(5 \text{ mol} \cdot 0 \frac{\text{kJ}}{\text{mol}}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

$$\Delta H_{rxn} = \left[-1089.6 \frac{\text{kJ}}{\text{mol}} \right] - \left[-183.6 \frac{\text{kJ}}{\text{mol}} \right]$$

$$\Delta H_{rxn} = -906 \frac{\text{kJ}}{\text{mol}}$$

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

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



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

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

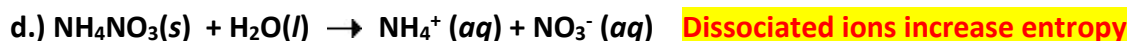
$$\Delta S_{rxn} = \left[1975.4 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right] - \left[1796.2 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right]$$

$$\Delta S_{rxn} = 179.2 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

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.

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

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

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

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.



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

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

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

ΔG is positive so the reaction is NOT spontaneous!