Thermochemistry Homework

Specific Heat Capacity Problems: q = Cp · m · ΔT

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)

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

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

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?

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

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

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

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?

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

$$Specific\ Heat\ Capacity = \frac{1251\ J}{(35.2\ \mathrm{grams})(25℃)}$$

$$\boldsymbol{Specific\ Heat\ Capacity\ = 1.42 \frac{\mathrm{J}}{\mathrm{g \cdot ℃}}}$$

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?

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

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

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…

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

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

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

STEP 2: Solve for heat transferred …

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

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

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

STEP 3: Combine to solve for Final Temperature …

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

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

Combine like terms to get…

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

$$\boldsymbol{T_f = 69.52^\circ\mathrm{C}}$$

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?

$$\mathbf{4\ NH_3\ (g) + 5\ O_2(g) \leftrightharpoons 4\ NO(g) + 6\ H_2O\ (g)}$$

$$\Delta \boldsymbol{H}_{\boldsymbol{rxn}} = \sum \Delta \boldsymbol{H}_{\boldsymbol{products}} - \sum \Delta \boldsymbol{H}_{\boldsymbol{reactants}}$$

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

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

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

ΔH is negative so the reaction is EXOTHERMIC

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

$$\mathbf{4\ NH_3\ (g) + 5\ O_2(g) \leftrightharpoons 4\ NO(g) + 6\ H_2O\ (g)}$$

$$\Delta \boldsymbol{S}_{\boldsymbol{rxn}} = \sum \Delta \boldsymbol{S}_{\boldsymbol{products}} - \sum \Delta \boldsymbol{S}_{\boldsymbol{reactants}}$$

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

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

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

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

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

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

a.) $N_2(g) + 3H_2(g) \rightleftharpoons 2\ NH_3(g)$ **Product has fewer moles; entropy decreases**

b.) $H_2O(l) \rightleftharpoons H_2O(g)$ **Product changed to gas so entropy increases.**

c.) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ **Decomposition reactions increase entropy**

d.) $NH_4NO_3(s) + H_2O(l) \rightarrow NH_4^+(aq) + NO_3^-(aq)$ **Dissociated ions increase entropy**

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

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

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

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

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.

$Br_2(l) \rightarrow Br_2(g)$

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

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

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

ΔG is positive so the reaction is NOT spontaneous!