

ASSIGNMENT SHEET #11 APQ ANSWERS

#1

a. The unit for q must be an energy unit, typically Joules or calories.

The unit for mass is the gram.

The unit for specific heat is J per gram-degree or calorie per gram-degree.

The unit for ΔT is degree (either Celsius or K).

b. Measurements include:

Initial temperature of each solution, final temperature of the mixed solution, and the volumes of each solution. For kicks, get the height and weight of your lab partner. Maybe mass a bagel or two.

c i. Since water is made in a one-to-one ratio, the moles of water formed will equal the moles of H^+ (and, in this case, moles of OH^-) added initially. The moles can be calculated by multiplying the volume of the acid by the molarity of the acid (or work similarly with the base solution).

c ii. Once q is calculated, divide q by the number of moles of acid added (or moles of water produced). This will give you energy per mole of water produced, ΔH .

d i. If the concentration of the reactants is increased, the amount of heat released will increase, so q should be larger.

d ii. The molar heat of neutralization will not change because you will divide a larger q value with a proportionately larger number of moles of acid used to determine ΔH (as done in c ii above).

e. If heat is lost to the air, ΔT will be too small; hence the calculated q would be too small. Dividing a smaller q value by the number of moles of acid will result in a smaller ΔH .

#5 c

5.00 grams of $C_5H_{12} = 0.0694$ moles of C_5H_{12} ; 243 kJ per 0.0694 mole = 3,500 kJ per mole.

#7 ab

a. ΔS is (+) for this reaction due to the solid-to-gas phase change going from reactants to products.

b. When the temperature is higher, the percent of CO (which is the product of the forward reaction) is greater, which will only happen if the forward reaction is endothermic according to Le Chatelier's principle. Therefore, ΔH is (+).

#8

a. 73.1 grams of NO = 2.44 moles of NO $2.44 \text{ moles} \times \frac{114.1 \text{ kJ}}{2 \text{ mole NO}} = 139 \text{ kJ}$

bi. $\Delta G^\circ = -RT \ln K_{\text{eq}}$

$R = 8.314 \text{ J/mole-K}$ $T = 25 + 273 = 298 \text{ K}$ $\Delta G = -70,400 \text{ J}$ $K_{\text{eq}} = 2.19 \times 10^{12}$

b ii.

According to the equation $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, as T increases, given that both ΔH° and ΔS° are (-) values, ΔG° should increase; i.e., become less negative.

OR . . . say that as T increases, ΔG° doesn't change, but that ΔG does change. Recall that ΔG° is a standard measurement, pretty much a constant.

c. $\Delta S_{\text{rxn}} = \sum S_{\text{products}} - \sum S_{\text{reactants}} \therefore -146.5 \text{ J} = 2(240.1 \text{ J/mol}) - [2(210.8 \text{ J/mol}) + x]$

$x = 205.1 \text{ J} = S^\circ$ for O_2 gas

d. $\Delta H_{\text{rxn}} = (+) \text{ BE of reactants} + (-) \text{ BE of products}$ (or, a la Chang: $\Delta H_{\text{rxn}} = \sum \text{BE}_{\text{reactants}} - \sum \text{BE}_{\text{products}}$)

$\therefore -114.1 \text{ kJ} = 2(607 \text{ kJ}) + 495 \text{ kJ} + 2[(-x \text{ kJ})]$ $x = \text{BE of NO bond in NO}_2 = 455.8 \text{ kJ}$

#17 abc

a. $\Delta H_{\text{rxn}} = \sum \Delta H_{\text{f products}} - \sum \Delta H_{\text{f reactants}} = -200 \text{ kJ}$

b. ΔS is zero. There are no phase changes and the number of gaseous reactants equals the number of gaseous products.

c. ΔG is (-) since $\Delta H = (-)$, according to $\Delta G = \Delta H - T\Delta S$.

#23

a i. ΔS is (-) due to the gas-to-solid/liquid phase change.

a ii. Since the reaction is spontaneous, ΔG is (-). Therefore, since ΔS is (-), which is unfavorable in terms of achieving a (-) ΔG according to the equation $\Delta G = \Delta H - T\Delta S$, the principal driving force behind this reaction has to be the (-) ΔH .

a iii. Here we go again!

According to the equation $\Delta G = \Delta H - T\Delta S$, as T increases, and given that both ΔH and ΔS are (-) values, then ΔG should become less (-) (it should increase).

b i. ΔG° tells us nothing about activation energy or the kinetics of the reaction. It simply allows us to predict spontaneity (actually, ΔG does, not ΔG° .)

b ii. None. As stated above, the thermodynamics and kinetics of a reaction are unrelated. The catalyst decreases the activation energies of both the forward and reverse reactions, which affects both rate

constants, k_F and k_R , proportionately. The equilibrium constant, K_{eq} , is the ratio of k_F and k_R , and is not changed by the catalyst. If K_{eq} doesn't change, ΔG° doesn't change according to $\Delta G^\circ = -RT \ln K_{eq}$.

#31

a. 2.000 grams of C_6H_5OH (phenol) = 0.0213 moles of phenol

\therefore 64.98 kJ per 0.0213 moles = 3,050.7 kJ per mole

$\therefore \Delta H_{comb} = -3,050.7$ kJ/mol

b. Since $\Delta H_{rxn} = \sum \Delta H_f \text{ products} - \sum \Delta H_f \text{ reactants}$, ΔH_f phenol = - 167.85 kJ/mole

c. $\Delta S_{rxn} = \sum S_{products} - \sum S_{reactants} = -87.67$ J

$\therefore \Delta G = \Delta H - T\Delta S = -3,050,700 \text{ J} - (298)(-87.67 \text{ J}) = -3,024,574 \text{ J} = -3,024.6$ kJ

d. Gas law review!

Moles of gaseous products (based on 0.0213 moles of phenol) = 0.1278 mol CO_2 + 0.0639 mol H_2O

$$P = \frac{nRT}{V} = \frac{(0.1917 \text{ mole})(0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mole} \cdot \text{K}})(383 \text{ K})}{10 \text{ L}} = 0.60 \text{ atm}$$

#44 ab

a. ΔS is (+) because the number of gaseous particles increases over the course of the reaction.

b. Since the amount of products increases as temperature increases, according to Le Chatelier, this reaction must be endothermic (the forward reaction is favored with increased T). Therefore, $\Delta H = (+)$. Since ΔH and ΔS are (+), according to the equation $\Delta G = \Delta H - T\Delta S$, if the temperature is increased, ΔG must decrease; i.e. become more (-).

#48

a. Since $\Delta S^\circ_{rxn} = \sum S^\circ_{products} - \sum S^\circ_{reactants}$, S° for $C_2H_6 = 229.6$ J

b. $\Delta H^\circ_{rxn} = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants} = -311.4$ kJ

\therefore since $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -311,400 \text{ J/mole} - (298 \text{ K})(-232.7 \text{ J/mole} \cdot \text{K}) = -242,055 \text{ J} = -242.1$ kJ

Since the sign of ΔG° is negative, the forward reaction occurs spontaneously.

c. At equilibrium, $\Delta G^\circ = -RT \ln K_{eq} \quad \therefore -242,055 \text{ J} = -8.314 \text{ J/mole} \cdot \text{K} (298 \text{ K})(\ln K_{eq})$

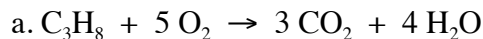
$$K_{eq} = 2.69 \times 10^{42}$$

d. $\Delta H_{rxn} = (+) BE_{reactants} + (-) BE_{products}$ let x = the C-to-C triple bond energy

$\therefore -311.4 \text{ kJ} = x + 2(414 \text{ kJ/mol}) + 2(436 \text{ kJ/mol}) + (-347 \text{ kJ/mol}) + 6(-414 \text{ kJ/mol})$

x = 819.6 kJ/mol

#51



b. Gas law review!

10 g of propane = 0.227 mole, which needs 5 times as much O_2 in order to react; $n \text{O}_2 = 1.14$ mole.

$$\therefore V \text{ of } \text{O}_2 = \frac{1.14 \text{ mole} (0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mole} \cdot \text{K}}) (30 + 273 \text{K})}{1 \text{ atm}} = 28.3 \text{ L} = 21\% \text{ of air needed}$$

$$\therefore V \text{ of air needed} = \frac{28.3 \text{ L}}{0.21} = 134.6 \text{ L}$$

c. $\Delta H_f^\circ \text{C}_3\text{H}_8 = -101.8 \text{ kJ/mol}$ Note: This value is per 1 mole of propane combusted.

d. Hooray! Calorimetry! $\Delta T = \sim 45^\circ\text{C}$

#59

a. ΔS is (+) due to the creation of an aqueous solution from a pure liquid and a pure solid.

b. Since $\Delta H = (+)$, according to Le Chatelier, a decrease in temperature will favor the reverse reaction, reducing the number of ions in solution. This will decrease the value of K_{sp} .

c. Adding additional solid does not change the concentration of either ion. Adding solid to an equilibrium system that already has some solid does not change the concentration of the solid, so it is not a Le Chatelier stress.

d. Recall $\Delta G = \Delta G^\circ + RT \ln Q$

The additional Pb^{+2} ion will make Q_{sp} greater than K_{sp} and will favor the reverse reaction according to Le Chatelier.

Quantitatively, according to the equation above, you will be adding a larger value to ΔG° , so the value of ΔG , which had been zero, will become (+).

Qualitatively, if you favor the reverse reaction to proceed spontaneously, we expect ΔG to be (+).

#63

a. $\Delta S^\circ = (-)$ because of the gas-to-solid-and-gas phase change.

b. As temperature increases, according to the equation $\Delta G = \Delta H - T\Delta S$, the value of ΔG will increase since $\Delta H = (-)$.

c. Increasing the temperature favors the endothermic reaction according to Le Chatelier, which is the reverse reaction in this case. Therefore, there will be less product and more reactant, which means that the value of K_{eq} will be smaller.

d. We would use $\Delta G = \Delta H - T\Delta S$, or, rearranged for when $\Delta G = 0$, $T = \frac{\Delta H}{\Delta S}$. This will only give an

approximate temperature value because we would be at a non-standard temperature but would be using standard values of ΔH and ΔS . Now there's a sin if ever there was one 😊.