Thermochemistry and Thermodynamics Worksheet 2

1. Given the following: $C_2H_{4(g)} + 3\ O_{2(g)} \text{ -----> } 2\ CO_{2(g)} + 2\ H_2O_{(g)}$ $\Delta H = -1322.9$ kJ
   $\Delta H_f\ C_2H_{4(g)} = +52.3$ kJ/mol
   $\Delta H_f\ H_2O_{(g)} = -241.8$ kJ/mol

   (A) Calculate the heat of formation of $CO_{2(g)}$.

   (B) How much heat will be evolved when 140.0 g $C_2H_{4(g)}$ is consumed?

   (C) How many moles of $C_2H_{4(g)}$ will be required to produce 2,300 kJ of heat?

   (D) If the molar volume of $C_2H_{4(g)}$ is 22.4 L/mol, how many liters of C_2H_4 are required in part (C)?

2. Given the following chemical reaction:
   $2\ Al_{(s)} + Fe_2O_{3(s)} \text{ ------> } Al_2O_{3(s)} + 2\ Fe_{(s)}$
   $\Delta H_f\ Fe_2O_{3(s)} = -822.2$ kJ/mol $\Delta H_f\ Al_2O_{3(s)} = -1669.8$ kJ/mol

   (A) Calculate ΔH for this reaction.

   (B) If all the heat given off by reacting 1 mole of $Fe_2O_{3(s)}$ is absorbed by the products, what would be the change in temperature if the reaction goes to completion. (The specific heat of $Al_2O_{3(s)}$ = 0.19J/g•C and the specific heat of $Fe_{(s)}$ is 0.48J/g•C)

3. The combustion of 1.00 mol of sucrose, $C_{12}H_{22}O_{11}$, evolves 5.65×10^3 kJ of heat. A bomb calorimeter with a calorimeter constant of 1.23 kJ/°C contains 0.600 kg of water. How many grams of sucrose should be burned to raise the temperature of the calorimeter and its contents from 23.0 °C to 50.0 °C? (The calorimeter constant represents the heat capacity of the empty calorimeter.) The specific heat of water is 4.184 J/g ^{o}C.

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4. The heat of reaction for burning 1 mole of a certain compound X is known to be -477.7 kJ. The calorimeter constant of the bomb being used is 2.5×10^3 J/°C and the initial temperature of the water is 23.2 ^oC.

(A) If 96.54 g of compound X (MM = 46) is burned in the bomb calorimeter containing 2000 ml of water (S.H. = 4.184 J/g•°C), what will be the final temperature?

(B) How much water can be warmed from 23.2°C to 56.5°C when 172.0 g of the compound is burned in the bomb?

5. A 50.0-g piece of metal at 60.0°C is placed in 200.0 g of water at 22.0°C contained in a coffee-cup calorimeter. The metal and water come to the same temperature at 32.5°C.

(A) How much heat did the metal give up to the water?

(B) What is the specific heat of the metal?

(C) How many grams of the metal at 80°C would have to be used to heat half as much water (100.0 g) by to the same temperature?

6. Use Hess' Law to calculate ΔH for the following reaction:

$$4\ HCN_{(l)} + 9\ O_{2(g)} \text{ -----> } 4\ CO_{2(g)} + 4\ NO_{2(g)} + 2\ H_2O_{(l)}$$

Given:

$H_{2(g)} + 2\ C_{(s)} + N_{2(g)}$ -----> $2\ HCN_{(l)}$	ΔH = +217.8 kJ
$C_{(s)} + O_{2(g)}$ -----> $CO_{2(g)}$	ΔH = -393.5 kJ
$2\ NO_{2(g)}$ -----> $N_{2(g)} + 2\ O_{2(g)}$	ΔH = -66.40 kJ
$2\ H_{2(g)} + O_{2(g)}$ -----> $2\ H_2O_{(l)}$	ΔH = -571.7 kJ