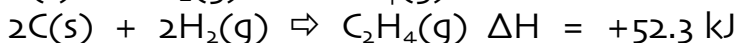
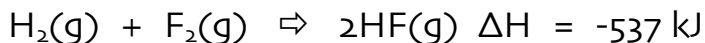


PRACTICE#2/ADDITIONAL PRACTICE PROBLEMS/THERMOCHEMISTRY/AP CHEMISTRY

1. A 1.800-g sample of phenol ( $C_6H_5OH$ ) was burned in a bomb calorimeter whose total heat capacity is  $11.66 \text{ kJ/}^\circ\text{C}$ . The temperature of the calorimeter plus contents increased from  $21.36^\circ\text{C}$  to  $26.37^\circ\text{C}$ .

- A) write a balanced chemical eq. for the bomb calorimeter reaction.  
 B) What is the heat of combustion per gram of phenol? Per mole of phenol?

2. From the enthalpies of reaction:



calculate  $\Delta H$  for the reaction of ethylene with  $F_2$ :

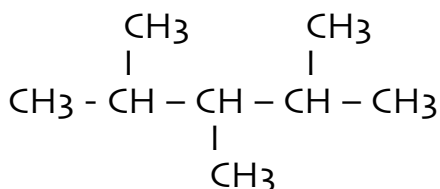


3. The following is known as the thermite reaction:



This highly exothermic reaction is used for welding massive units, such as propellers for large ships. Using enthalpies of formation in Appendix C, calculate the  $\Delta H^\circ$  for this reaction.

4. Gasoline is composed primarily of hydrocarbons, including many with eight carbon atoms, called octanes. One of the cleanest burning octanes is a compound called 2,3,4-trimethylpentane, which has the following structural formula:



The complete combustion of one mole of this compound to  $CO_2(g)$  and  $H_2O(g)$  leads to  $\Delta H^\circ = -5069 \text{ kJ}$

- A) Write a balanced equation for the combustion of 1 mol of  $C_8H_{18}(l)$ .  
 B) Write a balanced equation for the formation of  $C_8H_{18}(l)$  from its elements.  
 C) By using the information in this problem and data in table 5.3, calculate the  $\Delta H^\circ_f$  for 2,3,4-trimethylpentane.

5. When a mole of Dry Ice<sup>®</sup>,  $CO_2(s)$ , is converted to  $CO_2(g)$  at atmospheric pressure and  $-78^\circ\text{C}$ , the heat absorbed by the system exceeds the increase in internal energy of the  $CO_2$ . Why is this so? What happens to the remaining energy?

6. An aluminum can of a soft drink is placed in a freezer. Later the can is found to be split open and its contents frozen. Work was done on the can in splitting it open. Where did the energy to accomplish this work come from?

7. A coffee-cup calorimeter contains 150.0-g of water at 25.1°C. A 121.0-g block of copper metal is heated to 100.4°C by putting it into a beaker of boiling water. The specific heat of Cu(s) is 0.385 J/gK. The copper is added to the calorimeter and the temperatures allowed to equilibrate to 30.1°C.

A)

Determine the amount of heat, in J, lost by the copper block.

B)

Determine the amount of heat gained by the water. The specific heat of water is 4.18 J/gK

C)

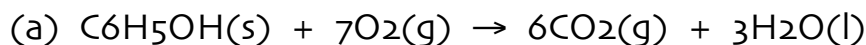
The difference between your answers for parts a and b is due to the heat loss through the Styrofoam® cups and the heat necessary to raise the temperature of the inner wall of the apparatus. The heat capacity of the calorimeter is the amount of heat necessary to raise the temperature of the apparatus (cups, lid, thermometer) by 1 K. Calculate the heat capacity of the calorimeter in J/K.

D)

What would the final temperature of the system be if all of the heat lost by the copper block was absorbed by the water in the calorimeter?

ANSWERS TO ADDITIONAL PRACTICE PROBLEMS:

1.



$$(b) q_{\text{bomb}} = -q_{\text{rxn}}; \Delta T = 26.37^\circ\text{C} - 21.36^\circ\text{C} = 5.01^\circ\text{C}$$

$$q_{\text{bomb}} = 11.66 \text{ kJ}/1^\circ\text{C} \times 5.01^\circ\text{C} = 58.417 = 58.4 \text{ kJ}$$

At constant volume,  $q_v = \Delta E$ ,  $\Delta E$  and  $\Delta H$  are very similar.

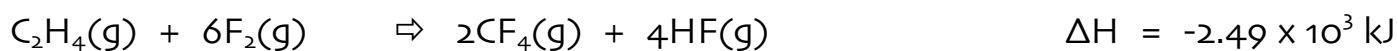
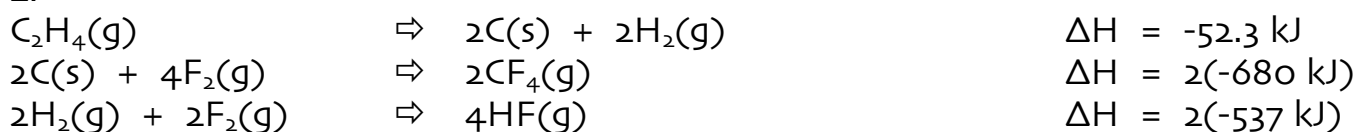
$$\Delta H_{\text{rxn}} = \Delta E_{\text{rxn}} = q_{\text{rxn}} = -q_{\text{bomb}} = -58.417 \text{ kJ}/1.800 \text{ g C}_6\text{H}_5\text{OH} = -32.454 =$$

$$-32.5 \text{ kJ/g C}_6\text{H}_5\text{OH}$$

$$\Delta H_{\text{rxn}} = -32.454 \text{ kJ}/1 \text{ g C}_6\text{H}_5\text{OH} \times 94.11 \text{ g C}_6\text{H}_5\text{OH}/1 \text{ mol C}_6\text{H}_5\text{OH} =$$

$$-3.054 \times 10^3 \text{ kJ/mol C}_6\text{H}_5\text{OH} = -3.05 \times 10^3 \text{ kJ/mol C}_6\text{H}_5\text{OH}$$

2.



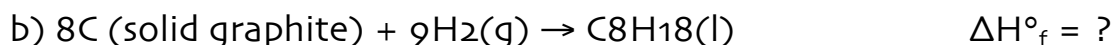
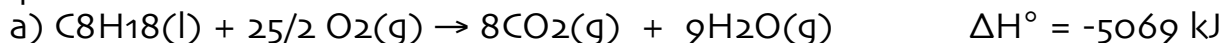
3.

Plan:  $\Delta H_{\text{rxn}} = \sum n\Delta H_f^\circ(\text{products}) - \sum n\Delta H_f^\circ(\text{reactants})$

$$\Delta H_{\text{rxn}} = \Delta H_f^\circ \text{Al}_2\text{O}_3(s) + 2\Delta H_f^\circ \text{Fe}(s) - \Delta H_f^\circ \text{Fe}_2\text{O}_3 - 2\Delta H_f^\circ \text{Al}(s)$$

$$\Delta H_{\text{rxn}} = (-1669.8 \text{ kJ}) + 2(0) - (-822.16 \text{ kJ}) - 2(0) = -847.6 \text{ kJ}$$

4.



$$c) \Delta H_{\text{rxn}}^\circ = 8\Delta H_f^\circ \text{CO}_2(g) + 9\Delta H_f^\circ \text{H}_2\text{O}(g) - \Delta H_f^\circ \text{C}_8\text{H}_{18}(l) - 25/2 \Delta H_f^\circ \text{O}_2(g)$$

$$\Delta H_f^\circ \text{C}_8\text{H}_{18}(l) = 8(-393.5 \text{ kJ}) + 9(-241.82 \text{ kJ}) + 5069 \text{ kJ} = -255 \text{ kJ}$$

5.

In the process described, one mole of solid  $\text{CO}_2$  is converted to one mole of gaseous  $\text{CO}_2$ . The volume of the gas is much greater than the volume of the solid. Thus, the system (the  $\text{CO}_2$ ) must work against atmospheric pressure when it expands. To accomplish this work while maintaining a constant temperature requires the absorption of additional heat beyond that required to increase the internal energy of the  $\text{CO}_2$ . The remaining energy is turned into work.

6.

Freezing is an exothermic process (the opposite of melting – which is clearly endothermic). When the system (the soft drink) freezes, it releases energy to the surroundings (the can). Some of this energy is used to work the can apart (splitting).

7.

$$\text{a) } q_{\text{Cu}} = 0.385 \text{ J/gK} \times 121.0 \text{ g Cu} \times (30.1^\circ\text{C} - 100.4^\circ\text{C}) = -3274.9 = -3.27 \times 10^3 \text{ J}$$

the negative sign indicates the  $3.27 \times 10^3 \text{ J}$  are lost by the Cu block.

$$\text{b) } q_{\text{H}_2\text{O}} = 4.184 \text{ J/gK} \times 150.0 \text{ g H}_2\text{O} \times (30.1^\circ\text{C} - 25.1^\circ\text{C}) = 3138 = +3.1 \times 10^3 \text{ J}$$

the positive sign indicates that  $3.1 \times 10^3 \text{ J}$  are gained by the H<sub>2</sub>O.

c) The difference between the heat lost by the Cu and the heat gained by the water is

$$3.275 \times 10^3 \text{ J} - 3.3138 \times 10^3 = 0.137 \times 10^3 = 1 \times 10^2 \text{ J. The temperature change of the}$$

calorimeter is  $5.0^\circ\text{C}$ . The heat capacity of the calorimeter in J/K is

$$0.137 \times 10^3 \text{ J} \times 1/5.0^\circ\text{C} = 27.4 = 3 \times 10 \text{ J/K} = 30 \text{ J/K}$$

Since  $q_{\text{H}_2\text{O}}$  is known to 1 decimal place, the difference has 1 decimal place and the result has 1 sig. fig.

If the rounded results from a) and b) are used,  $C_{\text{calorimeter}} = 0.2 \times 10^3 \text{ J}/5.0^\circ\text{C} = 4 \times 10 \text{ J/K}$

$$\text{d) } q_{\text{H}_2\text{O}} = 3.275 \times 10^3 \text{ J} = 4.184 \text{ J/gK} \times 150.0 \text{ g} \times (\Delta T)$$

$$\Delta T = 5.22^\circ\text{C}; T_f = 25.1^\circ\text{C} + 5.22^\circ\text{C} = 30.3^\circ\text{C}$$