CH 222 Practice Problem Set #3

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit. Answers to each problem can be found at the end of this assignment.*

Covering: **Chapter Nine** *and* **Chapter Guide Three**
Important Tables/Constants: **$C(H_2O)$ = 4.184 J g^{-1} K^{-1}** *and the* ***Thermodynamic Values*** *found in problem set #3 and here:* **http://mhchem.org/thermo**

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1. `The specific heat capacity of copper is 0.385 J/g·K. What quantity of heat is required to heat 168 g of copper from -12.2 °C to +25.6 °C?
2. The initial temperature of a 344 g sample of iron is 18.2 °C. If the sample absorbs 2.25 kJ of heat, what is its final temperature? C_{Fe} = 0.449 J/g·K
3. One beaker contains 156 g of water at 22 °C and a second beaker contains 85.2 g of water at 95 °C. The water in the two beakers is mixed. What is the final water temperature?
4. A 237 g piece of molybdenum, initially at 100.0 °C, is dropped into 244 g of water at 10.0 °C. When the system comes to thermal equilibrium, the temperature is 15.3 °C. What is the specific heat capacity of molybdenum?
5. What quantity of heat is required to vaporize 125 g of benzene, C_6H_6, at its boiling point, 80.1 °C? The heat of vaporization of benzene is 30.8 kJ/mol.
6. Isooctane (2,2,4-trimethylpentane), one of the many hydrocarbons that make up gasoline, burns in air to give water and carbon dioxide.

   $2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(l)\quad \Delta H°_{rxn} = -10{,}922\ kJ$

   If you burn 1.00 L of isooctane (density = 0.69 g/mL), what quantity of heat is evolved?
7. The enthalpy changes for the following reactions can be measured:

   $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)\qquad \Delta H° = -802.4\ kJ$

   $CH_3OH(g) + {}^3/_2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)\qquad \Delta H° = -676\ kJ$

   Use these values and Hess's law to determine the enthalpy change for the reaction:

   $CH_4(g) + {}^1/_2\ O_2(g) \rightarrow CH_3OH(g)\qquad \Delta H° = ?$
8. Enthalpy changes for the following reactions can be determined experimentally:

   $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)\qquad \Delta H° = -91.8\ kJ$

   $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)\qquad \Delta H° = -906.2\ kJ$

   $H_2(g) + {}^1/_2\ O_2\ (g) \rightarrow H_2O(g)\qquad \Delta H° = -241.8\ kJ$

   Use these values to determine the enthalpy change for the formation of NO(g) from the elements (an enthalpy change that cannot be measured directly because the reaction is reactant-favored) of

   ${}^1/_2\ N_2(g) + {}^1/_2\ O_2(g) \rightarrow NO(g)\qquad \Delta H° = ?$
9. Write a balanced chemical equation for the formation of Li_2CO_3(s) from the elements in their standard states. Find the value of $\Delta H_f°$ for Li_2CO_3(s) in the appendix of your textbook.
10. Use standard heats of formation in the appendix of your textbook to calculate standard enthalpy changes for the following:
    a. 1.0 g of white phosphorus burns, forming P_4O_{10}(s)
    b. 0.20 mol of NO(g) decomposes to N_2(g) and O_2(g)
    c. 2.40 g of NaCl is formed from Na(s) and excess Cl_2(g)
    d. 250 g of iron is oxidized with oxygen to Fe_2O_3(s)

11. The Romans used calcium oxide, CaO, to produce a strong mortar to build stone structures. The CaO was mixed with water to give $Ca(OH)_2$, which reacted slowly with CO_2 in the air to give $CaCO_3$. **$Ca(OH)_2(s) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(g)$**
   a. Calculate the standard enthalpy change for this reaction.
   b. What quantity of heat is evolved or absorbed if 1.00 kg of $Ca(OH)_2$ reacts with a stoichiometric amount of CO_2?
12. The compound oxygen difluoride is quite reactive, giving oxygen and HF when treated with water:
   $OF_2(g) + H_2O(g) \rightarrow O_2(g) + 2\ HF(g) \quad \Delta H°_{rxn} = -318\ kJ$
   Using bond energies, calculate the bond dissociation energy of the O-F bond in OF_2.
13. Which compound in each of the following pairs should require the higher temperature to melt?
   a. NaCl or RbCl
   b. BaO or MgO
   c. NaCl or MgS
14. Calcium carbide, CaC_2, is manufactured by the reaction of CaO with carbon at high temperatures. Calcium carbide can then be used to make acetylene. Using the reaction below, is this reaction endothermic or exothermic? If 10.0 g of CaO is allowed to react with an excess of carbon, what quantity of heat is absorbed or evolved by the reaction?

$$CaO(s) + 3\ C(s) \rightarrow CaC_2(s) + CO(g) \quad \Delta H^\circ_{rxn} = +464.8\ kJ$$

Answers to the Practice Problem Set:

1. 2440 J
2. 32.8 °C
3. 48 °C
4. 0.27 J/g·K
5. 49.3 kJ
6. 3.3×10^4 kJ heat evolved
7. -126 kJ
8. 90.3 kJ
9. $2\,Li(s) + C(s) + {}^{3}/_{2}\,O_2(g) \rightarrow Li_2CO_3(s)$ $\Delta H_f^\circ = -1216.04$ kJ *(OpenStax)*
10. a. -24 kJ b. -18 kJ c. -16.9 kJ d. -1800 kJ
11. a. -83.1 kJ b. -1120 kJ evolved
12. 195 kJ/mol
13. a. NaCl b. MgO c. MgS
14. endothermic, 82.9 kJ heat required