CH1020 Exercises (Worksheet 14)
(Enthalpy, Thermochemical equations)

1. Under what condition will the enthalpy change of a process equal the amount of heat transferred into or out of the system?
   When a process occurs under constant pressure, the enthalpy change (ΔH) equals the amount of heat transferred. $\Delta H = q_P$

2. Does the enthalpy of a system in a particular state depend on how the system reached that state?   Explain.
   No. Enthalpy is a state function, so it is totally defined by the current conditions (state) of the system, not the history of how the system arrived at its current state.

3. During a constant-pressure process the system absorbs heat from the surroundings. Does the enthalpy of the system increase or decrease in the process?
   $\Delta H = q_P$ If the system absorbs heat, q and ΔH are positive and the enthalpy of the system increases.

4. Why is the change in enthalpy a meaningful for many chemical processes?
   For the many laboratory and real world processes that occur at constant atmospheric pressure, the enthalpy change is a meaningful measure of the energy change associated with the process. At constant pressure, most of the energy change is transferred as heat ($\Delta H = q_P$), even if gases are involved in the process.

5. H is a state function but q is not a state function. Explain.
   Only under conditions of constant pressure is ΔH for a process equal to the heat transferred during the process.

6. For a given process at constant pressure, ΔH is negative. Is the process endothermic or exothermic?
   ΔH is negative, the enthalpy of the system decreases and the process is exothermic

7. The complete combustion of $CH_3OH(l)$ to form $H_2O(l)$ and $CO_2(g)$ at constant pressure produces 726.7 kJ of heat per mole of CH_3OH. Write a balanced thermochemical equation for this reaction.
   $2CH_3OH(l) + 3O_2(g)$ $\rightarrow 4H_2O(l) + 2CO_2(g)$ $\Delta H = -1435.4 \ kJ$

8. The decomposition reaction of $NH_3(g)$ to form $N_2(g)$ and $H_2(g)$ at constant pressure requires 46.19 kJ of heat per mole of $NH_3(g)$. Write a balanced thermochemical equation for the reaction.
   $2NH_3(g) \rightarrow N_2(g) + 3H_2(g) \ \Delta H = 92.38 \ kJ$

9. You are given ΔH for a process that occurs at constant pressure. What additional information is needed to determine ΔE for the process?
   At constant pressure $\Delta E = \Delta H - P\Delta V$. In order to calculate ΔE, more information about the conditions of the reaction must be known. For an ideal gas at constant pressure and temperature, $P\Delta V = RT\Delta n$. The values of either P and ΔV or T and Δn must be known to calculate ΔE from ΔH

10. What is the advantage of using enthalpy rather than internal energy to describe energy changes in reactions?
ΔH= q_P, q_P is something we can measure or readily calculate. Since many chemical and physical changes of interest occur at constant pressure, enthalpy is a more useful function than internal energy.

11. The reaction: $SO_3(g)$ + $H_2O(l)$ → $H_2SO_4(aq)$
is the last step in the commercial production of sulfuric acid. The enthalpy change for this reaction is –227 kJ. In designing a sulfuric acid plant, is it necessary to provide for heating or cooling of the reaction mixture? Explain.
One should try to cool the reaction mixture or provide some means of removing heat since the reaction is very exothermic (heat is released). The sulfuric acid will get very hot and possibly boil unless cooling is provided.

12. Are the following processes exothermic or endothermic?
   a. when solid KBr is dissolved in water, the solution gets colder.
            endothermic
   b. Natural gas (CH_4) is burned in a furnace   *exothermic*
   c. When concentrated H_2SO_4 is added to water, the solution gets very hot
       exothermic
   d. Water is boiled in a teakettle *endothermic*
   e. the combustion of gasoline in a car engine *exothermic*
   f. water condensing on a cold pipe *exothermic*
   g. CO_2 (s) → CO_2 (g) *endothermic*
   h. F_2 (g) → 2F (g)   *endothermic*

13. Which of the following has the highest enthalpy at a given temperature and pressure: $H_2O(s)$; $H_2O(l)$; are $H_2O(g)$? Which has the lowest enthalpy?
            Heat must be added to convert solid → liquid → gas
            $H_2O(s)$ < H_2O (l) < $H_2O(g)$

14. Consider the following reaction: $3O_2(g)$ → $2O_3(g)$ ΔH = +284.6 kJ
      Under the conditions of this reaction, does $O_2(g)$ or $O_3(g)$ have the higher enthalpy?
      O_3 has higher enthalpy (ΔH is positive)

15. The overall reaction in commercial heat packs can be represented as

        4Fe(s) + 3O_2(g) → 2Fe_2O_3(s)   ΔH = -1652 kJ
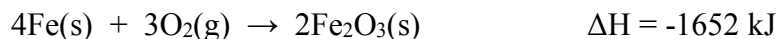
   a. How much heat is released when 1.00 mol Fe_2O_3 is produced?
            -826kJ
   b. How much heat is released when 1.00 g of iron is reacted with excess O_2?
            -7.4 kJ
   c. How much heat is released when 4.00 moles of Fe and 4.00 moles of O_2 are reacted? *-1652 kJ*

     d. How much heat is released when 0.540 moles of Fe and 15.0 g of O_2 react?   *-224 kJ*

   e. How many grams of Fe is required to react to release 1200. kJ of heat?
      162.3 g

16. Consider the following reaction:

$$2H_2(g) \,+\, O_2(g) \;\rightarrow\; 2H_2O(l) \;\Delta H = -572 \text{ kJ}$$
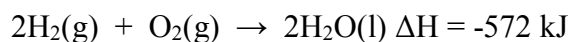
    a. How much heat is evolved for the production of 1.00 mol of $H_2O(l)$?
        -286 kJ
    b. How many grams of hydrogen is required to react, so that it produces 1400. kJ of heat?
        9.869 g
    c. How much heat is evolved when 4.03 g of hydrogen is reacted with excess oxygen?
        -572 kJ

17. Consider the combustion reaction of ethane gas, $C_2H_6(g)$:

$$C_2H_6(g) \,+\, 7/2O_2(g) \;\rightarrow\; 2CO_2(g) \,+\, 3H_2O(g) \qquad \Delta H = -1430 \text{ kJ}$$
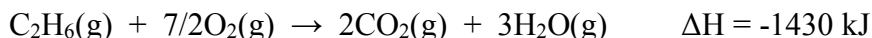
    a. What is the enthalpy change for the reverse reaction?
        +1430 kJ
    b. Balance the forward reaction with whole-number coefficients. What is the ΔH for the reaction represented by this equation?
        -2860 kJ
    c. Which is more likely to be thermodynamically favored, the forward reaction or the reverse reaction?
    The exothermic forward reaction is more likely to be thermodynamically favored.

    d. If the reaction were written to produce $H_2O(l)$ instead of $H_2O(g)$, would you expect the magnitude of ΔH to increase, decrease, or stay the same? Explain.
    Vaporization (liquid to gas) is endothermic, so the reverse process, condensation (gas to liquid) is exothermic. If the product was $H_2O(l)$, the reaction would be more exothermic and ΔH would have a larger negative value.

18. Consider the decomposition of liquid benzene, $C_6H_6(l)$ to gaseous acetylene, $C_2H_2(g)$:

$$1/3C_6H_6(l) \;\rightarrow\; C_2H_2(g) \qquad\qquad \Delta H = +210 \text{ kJ}$$

a. What is the enthalpy change for the reverse reaction?

           -210 kJ

b. What is ΔH for decomposition of 78.1g of benzene to acetylene?

           630 kJ

c. Which is more likely to be thermodynamically favored, the forward reaction or the reverse reaction?

Exothermic reverse reaction is more likely to be thermodynamically favored

d. If $C_6H_6(g)$ were consumed instead of $C_6H_6(l)$, would you expect the magnitude of ΔH to increase, decrease, or stay the same? Explain.

If the reactant is in the higher enthalpy gas phase, the overall ΔH for the reaction has a smaller positive value.