CH1020 Exercises (Worksheet 18)
(Hess's law, standard enthalpy of reactions)

1. State Hess's law. Why is it important to thermochemistry?
   If a reaction can be described as a series of steps, ΔH for the reaction is the sum of the enthalpy changes for each step. As long as we can describe a route where ΔH for each step is known, ΔH for any process can be calculated.
2. What is the connection between Hess's law and the fact that H is a state function?
   Hess's law is a consequence of the fact that enthalpy is a state function. Since ΔH is independent of path, we can describe a process by any series of steps that add up to the overall process and ΔH for the process is the sum of the ΔH values for the steps.
3. Consider the following hypothetical reactions:
   $A \rightarrow B \qquad \Delta H = +30\ kJ$
   $B \rightarrow C \qquad \Delta H = +60\ kJ$
   Use Hess's law to calculate the enthalpy change for the reaction $A \rightarrow C$.

$$\Delta H_{rxn} = \Delta H_1 + \Delta H_2 = 30\,kJ + 60\,kJ = 90\,kJ$$

4. Suppose you are given the following hypothetical reactions:

   $X \rightarrow Y \qquad \Delta H = -40\ kJ$
   $X \rightarrow Z \qquad \Delta H = -95\ kJ$
   Use Hess's law to calculate the enthalpy change for the reaction $Y \rightarrow Z$
   -55 kJ

5. Two gaseous pollutants that form in auto exhaust are CO and NO. An environmental chemist is studying ways to convert them to less harmful gases through the following equation:
   $CO(g) + NO(g) \rightarrow CO_2(g) + \frac{1}{2} N_2(g) \quad \Delta H = ?$
   Given the following information, calculate the unknown ΔH:
   Equation A: $CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g) \qquad \Delta H_A = -283.0\ kJ$
   Equation B: $N_2(g) + O_2(g) \rightarrow 2NO(g) \qquad \Delta H_B = 180.6\ kJ$
   -373.3 kJ

6. Given:
   $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5\ kJ$
   $2CO(g) + O_2(g) \rightarrow 2CO_2(g) \qquad \Delta H = -566.0\ kJ$
   calculate ΔH for the reaction:
   $C(s) + \frac{1}{2} O_2(g) \rightarrow CO(g) \quad \Delta H = ?$

   Solution:
   $C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$
   $CO_2 \rightarrow CO(g) + \frac{1}{2}O_2(g) \qquad \Delta H = \frac{1}{2}(+566.0\ kJ) = +283\ kJ$

   $C(s) + \frac{1}{2} O_2(g) \rightarrow CO(g) \qquad \Delta H = -110.5\ kJ$

7. Given the following thermochemical equations

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -571.6$ kJ

$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l)$ $\Delta H = -73.7$ kJ

$\frac{1}{2} N_2(g) + 3/2\ O_2(g) + \frac{1}{2} H_2(g) \rightarrow HNO_3(l)$ $\Delta H = -174.1$ kJ

calculate ΔH for the formation of one mole of dinitrogen pentoxide from its elements.

+11.3 kJ

8. Given the following data:

$NH_3(g) \rightarrow \frac{1}{2} N_2(g) + 3/2\ H_2(g)$ $\Delta H = 46$ kJ

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\Delta H = -484$ kJ

calculate ΔH for the reaction

$2N_2(g) + 6H_2O(g) \rightarrow 3O_2(g) + 4NH_3(g)$

On the basis of the enthalpy change, is this a useful reaction for the synthesis of ammonia?

+1268 kJ

No. Since the reaction is very endothermic (requires a lot of heat), it would not be a practical way of making ammonia due to the high energy costs required.

9. From the following enthalpies of reaction:

$H_2(g) + F_2(g) \rightarrow 2HF(g)$ $\Delta H = -537$ kJ

$C(s) + 2F_2(g) \rightarrow CF_4(g)$ $\Delta H = -680$ kJ

$2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$ $\Delta H = +52.3$ kJ

calculate ΔH for the reaction of ethylene with F_2:

$C_2H_4(g) + 6F_2(g) \rightarrow 2CF_4(g) + 4HF(g)$

-2486.3 kJ

10. The bombardier beetle uses an explosive discharge as a defensive measure. The chemical reaction involved is the oxidation of hydroquinone by hydrogen peroxide to produce quinone and water:

$C_6H_4(OH)_2\ (aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l)$

Calculate ΔH for this reaction from the following data:

$C_6H_4(OH)_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g)$ $\Delta H = +177.4$ kJ

$H_2(g) + O_2(g) \rightarrow H_2O_2(aq)$ $\Delta H = -191.2$ kJ

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ kJ

$H_2O(g) \rightarrow H_2O(l)$ $\Delta H = -43.8$ kJ

$C_6H_4(OH)_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g)$ $\Delta H = +177.4$ kJ

$H_2O_2(aq) \rightarrow H_2(g) + O_2(g)$ $\Delta H = +191.2$ kJ

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\Delta H = -483.6$ kJ

$2H_2O(g) \rightarrow 2H_2O(l)$ $\Delta H = -87.6$ kJ

$C_6H_4(OH)_2\ (aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l)$

$\Delta H_{rxn} = -202.6$ kJ

11. Given the following data:

$P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(g)$ $\Delta H = -1225.6$ kJ

$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$ $\Delta H = -2967.3$ kJ

$PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$ $\Delta H = -84.2$ kJ

$PCl_3(g) + \frac{1}{2} O_2(g) \rightarrow Cl_3PO(g)$ $\Delta H = -285.7$ kJ

Calculate ΔH for the reaction

$P_4O_{10}(s) + 6PCl_5(g) \rightarrow 10Cl_3PO(g)$

-610.1 kJ

12. Why are tables of standard enthalpies of formation so useful?

Tables of $\Delta H^o{}_f$ are useful because, according to Hess's law, the standard enthalpy of any reaction can be calculated from the standard enthalpies of formation for the reactants and products :

$\Delta H^o{}_{reaction} = \Sigma\, \Delta H^o{}_f\,(products) - \Sigma\, \Delta H^o{}_f\,(reactants)$

13. What is the value of the standard enthalpy of formation of an element in its most stable form?

The standard enthalpy of formation for any element in its standard state is zero. Elements in their standard states are the reference point for the enthalpy of formation scale.

14. Calculate $\Delta H^0{}_f$ for the following reactions:

a. $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

-1235 kJ

b. $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(aq)$

-320. kJ

c. $MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(s)$

-37 kJ

15. Calculate ΔH^0 for each of the following reactions:

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

-832 kJ

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

-368 kJ

$2Na(s) + CO_2(g) \rightarrow Na_2O(s) + CO(g)$

-133kJ

Explain why a water or carbon dioxide fire extinguisher might not be effective in putting out a sodium fire.

In both cases, Na metal reacts with the 'extinguishing agent'. Both reactions are exothermic and each reaction produces a flammable gas, H_2 and CO, respectively.

16. The space shuttle orbiter utilizes the oxidation of methyl hydrazine by dinitrogen tetroxide for propulsion:

$4N_2H_3CH_3(l) + 5N_2O_4(l) \rightarrow 12H_2O(g) + 9N_2(g) + 4CO(g)$

Calculate ΔH^0 for the reaction.

-3460 kJ

17. The standard enthalpy of combustion of ethane gas, $C_2H_4(g)$, is –1411.1 kJ/mol at 298 K. Given the following enthalpies of formation, calculate the enthalpy of formation for $C_2H_4(g)$:

$CO_2(g)$	-393.5 kJ/mol	$H_2O(l)$	-285.8 kJ/mol

52.5 kJ/mol