

## CHAPTER FIVE

### Bond Energy and Standard Enthalpy

**5.11** What is meant by the standard-state condition?

**5.12** How are the standard enthalpies of an element and of a compound determined?

**5.13** What is meant by the standard enthalpy of a reaction?

**5.14** Write the equation for calculating the enthalpy of a reaction. Define all the terms.

**5.15** State Hess's law. Explain, with one example, the usefulness of Hess's law in thermochemistry.

**5.16** Describe how chemists use Hess's law to determine the  $\Delta H^\circ_f$  of a compound by measuring its heat (enthalpy) of combustion.

**5.17** Which of the following standard enthalpy of formation values is not zero at 25°C? Na(s), Ne(g), CH<sub>4</sub>(g), S<sub>8</sub>(s), Hg(l), H(g).

**5.18** The  $\Delta H^\circ_f$  values of the two allotropes of oxygen, O<sub>2</sub> and O<sub>3</sub>, are 0 and 142.2 kJ/mol, respectively, at 25°C. Which is the more stable form at this temperature?

**5.19** Which is the more negative quantity at 25°C:  $\Delta H^\circ_f$  for H<sub>2</sub>O(l) or for  $\Delta H^\circ_f$  H<sub>2</sub>O(g)?

**5.20** Predict the value of  $\Delta H^\circ_f$  (greater than, less than, or equal to zero) for these elements at 25°C: (a) Br<sub>2</sub>(g); Br<sub>2</sub>(l), (b) I<sub>2</sub>(g); I<sub>2</sub>(s).

**5.21** In general, compounds with negative  $\Delta H^\circ_f$  values are more stable than those with positive  $\Delta H^\circ_f$  values. H<sub>2</sub>O<sub>2</sub>(l) has a negative  $\Delta H^\circ_f$  (see Table 6.3). Why, then, does H<sub>2</sub>O<sub>2</sub>(l) have a tendency to decompose to H<sub>2</sub>O(l) and O<sub>2</sub>(g)?

**5.22** Suggest ways (with appropriate equations) that would allow you to measure the  $\Delta H^\circ_f$  values of Ag<sub>2</sub>O(s) and CaCl<sub>2</sub>(s) from their elements. No calculations are necessary.

**5.23** Calculate the heat of decomposition for this process at constant pressure and 25°C:



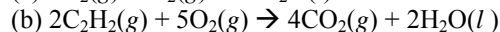
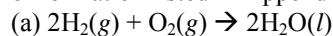
(Look up the standard enthalpy of formation of the reactant and products in Table 6.3.)

**5.24** The standard enthalpies of formation of ions in aqueous solutions are obtained by arbitrarily assigning a value of zero to H<sup>+</sup> ions; that is,  $\Delta H^\circ_f$  [H<sup>+</sup>(aq)] = 0. (a) For the following reaction

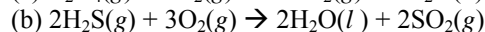
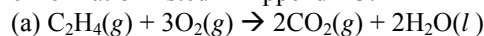


calculate  $\Delta H^\circ_f$  for the Cl<sup>-</sup> ions. (b) Given that  $\Delta H^\circ_f$  for OH<sup>-</sup> ions is -229.6 kJ/mol, calculate the enthalpy of neutralization when 1 mole of a strong monoprotic acid (such as HCl) is titrated by 1 mole of a strong base (such as KOH) at 25°C.

**5.25** Calculate the heats of combustion for the following reactions from the standard enthalpies of formation listed in Appendix 3:

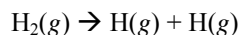


**5.26** Calculate the heats of combustion for the following reactions from the standard enthalpies of formation listed in Appendix 3:



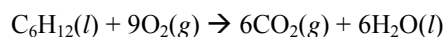
**5.27** Methanol, ethanol, and *n*-propanol are three common alcohols. When 1.00 g of each of these alcohols is burned in air, heat is liberated as follows: (a) methanol (CH<sub>3</sub>OH), -22.6 kJ; (b) ethanol (C<sub>2</sub>H<sub>5</sub>OH), -29.7 kJ; (c) *n*-propanol (C<sub>3</sub>H<sub>7</sub>OH), -33.4 kJ. Calculate the heats of combustion of these alcohols in kJ/mol.

**5.28** The standard enthalpy change for the following reaction is 436.4 kJ:



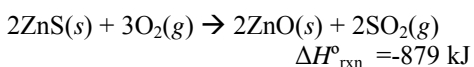
Calculate the standard enthalpy of formation of atomic hydrogen (H).

**5.29** From the standard enthalpies of formation, calculate  $\Delta H^\circ_{\text{rxn}}$  for the reaction



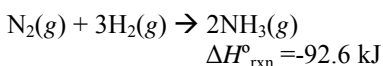
For C<sub>6</sub>H<sub>12</sub>(l),  $\Delta H^\circ_f$  = -151.9 kJ/mol.

**5.30** The first step in the industrial recovery of zinc from the zinc sulfide ore is roasting, that is, the conversion of ZnS to ZnO by heating:



Calculate the heat evolved (in kJ) per gram of ZnS roasted.

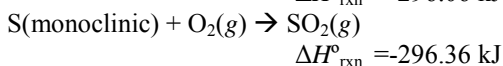
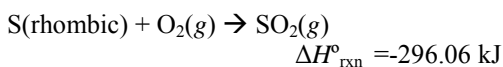
**5.31** Determine the amount of heat (in kJ) given off when  $1.26 \times 10^4$  g of ammonia are produced according to the equation



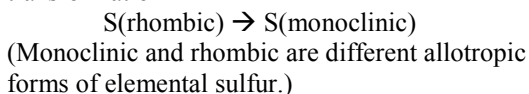
Assume that the reaction takes place under standardstate conditions at 25°C.

**5.32** At 850°C, CaCO<sub>3</sub> undergoes substantial decomposition to yield CaO and CO<sub>2</sub>. Assuming that the  $\Delta H^\circ_f$  values of the reactant and products are the same at 850°C as they are at 25°C, calculate the enthalpy change (in kJ) if 66.8 g of CO<sub>2</sub> are produced in one reaction.

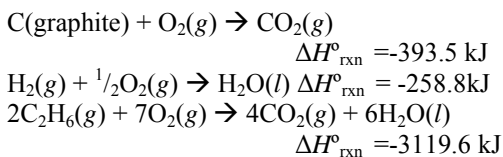
**5.33** From these data,



calculate the enthalpy change for the transformation



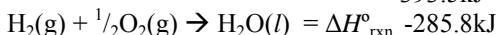
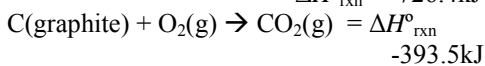
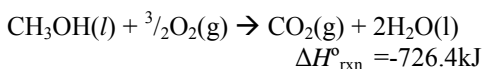
**5.34** From the following data,



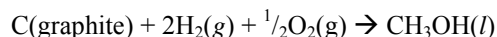
calculate the enthalpy change for the reaction

$$2\text{C}(\text{graphite}) + 3\text{H}_2(g) \rightarrow \text{C}_2\text{H}_6(g)$$

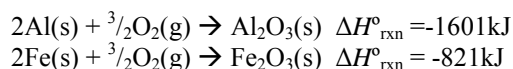
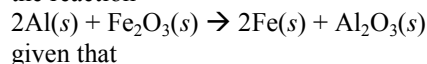
**5.35** From the following heats of combustion,



calculate the enthalpy of formation of methanol (CH<sub>3</sub>OH) from its elements:



**5.36** Calculate the standard enthalpy change for the reaction



**5.37** Consider the following two reactions:



Determine the enthalpy change for the process

$$2\text{B} \rightarrow \text{C}$$

**5.38** The standard enthalpy change  $\Delta H^\circ$  for the thermal decomposition of silver nitrate according to the following equation is +78.67 kJ:



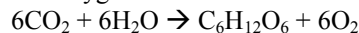
The standard enthalpy of formation of AgNO<sub>3</sub>(s) is -123.02 kJ/mol. Calculate the standard enthalpy of formation of AgNO<sub>2</sub>(s).

**5.39** Hydrazine, N<sub>2</sub>H<sub>4</sub>, decomposes according to the following reaction:



(a) Given that the standard enthalpy of formation of hydrazine is 50.42 kJ/mol, calculate  $\Delta H^\circ$  for its decomposition. (b) Both hydrazine and ammonia burn in oxygen to produce H<sub>2</sub>O(l) and N<sub>2</sub>(g). Write balanced equations for each of these processes and calculate  $\Delta H^\circ$  for each of them. On a mass basis (per kg), would hydrazine or ammonia be the better fuel?

**5.40** Photosynthesis produces glucose, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>, and oxygen from carbon dioxide and water:

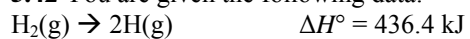


(a) How would you determine experimentally the  $\Delta H^\circ_{\text{rxn}}$  value for this reaction? (b) Solar radiation produces about  $7.0 \times 10^{14}$  kg glucose a year on Earth. What is the corresponding  $\Delta H^\circ$  change?

**5.41** 2.10-mole sample of crystalline acetic acid, initially at 17.0°C, is allowed to melt at 17.0°C

and is then heated to 118.1°C (its normal boiling point) at 1.00 atm. The sample is allowed to vaporize at 118.1°C and is then rapidly quenched to 17.0°C, so that it recrystallizes. Calculate  $\Delta H^\circ$  for the total process as described.

**5.42** You are given the following data:



Calculate  $\Delta H^\circ$  for the reaction

