Chapter 11 – Thermochemistry – Heat and Chemical Change

Chapter 11: 1 – 35, 57, 60, 61, 71

Section 11.1 – The Flow of Energy - Heat

Practice Problems

1. When 435 J of heat is added to 3.4 g of olive oil at 21°C, the temperature increases to 85°C. What is the specific heat of olive oil?

Knowns: q = 435 J; $m_{olive\ oil}$ = 3.4 g; ΔT = (85℃ – 21℃);

Use $C = q/ m \cdot \Delta T$

$$C = \frac{435\ J}{3.4\ g \cdot 64\ ℃} = 2.0\frac{J}{g \cdot ℃}$$

2. A 1.55-g piece of stainless steel absorbs 141 J of heat when its temperature increases by 178℃. What is the specific heat of the stainless steel?

Knowns: $m_{stainless\ steel}$ = 1.55 g; q = 141 J; ΔT = 178℃

Use $C = q/ m \cdot \Delta T$

$$C = \frac{141\ J}{1.55\ g \cdot 178\ ℃} = 0.511\frac{J}{g \cdot ℃}$$

3. How much heat is required to raise the temperature of 250.0 g of mercury 52℃?

Knowns: m_{Hg} = 250.0 g; ΔT = 52℃; C_{Hg} = 0.14 J/ g · ℃

Use $C = q/ m \cdot \Delta T$; rearrange to solve for q; $q = C \cdot m \cdot \Delta T$

$$q = \left(0.14\frac{J}{g \cdot ℃}\right)(250.0\ g)(52\ ℃) = 1{,}820\ J$$

Section Review 11.1

4. Define *energy* and explain how energy and *heat* are related.

Energy is the capacity for doing work or supplying heat. Heat is energy that is transferred between objects, when their temperatures are different.

5. Explain the difference between *heat capacity* and *specific heat*.

The specific heat of a substance is independent of its mass, whereas the heat capacity of an object is proportional to its mass.

6. Will the specific heat of 50 g of a substance be the same as, or greater than, the specific heat of 10 g of the same substance? Explain.

The specific heat between the two substances will be the same. Specific heat is not dependent upon the mass, whereas *heat capacity* is.

7. On a sunny day, why does the concrete deck around an outdoor swimming pool become hot, while the water stays cool?

The water stays cooler because it has a higher heat capacity than concrete. The sun's heat raises the temperature of the concrete more than that of the water.

8. Using *calories*, calculate how much heat 32.0 g of water absorbs when it is heated from 25.0℃ to 80.0℃. How many joules is this?

Use $C = q / m \cdot \Delta T$; rearrange to solve for q; $q = C \cdot m \cdot \Delta T$

$$q = \left(4.184 \frac{J}{g \cdot ℃}\right)(32.0\ g)(55.0\ ℃) = 7{,}364\ J; (7{,}364\ J)\left(\frac{1\ calorie}{4.184\ J}\right) = 1{,}760\ calories$$

9. A chunk of silver has a heat capacity of 42.8 J/℃. If the silver has a mass of 181 g, calculate the specific heat of silver.

$$C = \left(\frac{42.8 \frac{J}{℃}}{(181\ g)}\right) = 0.236 \frac{J}{g \cdot ℃}$$

10. How many kilojoules of heat are absorbed when 1.00 L of water is heated from 18℃ to 85℃?

Use $C = q / m \cdot \Delta T$; rearrange to solve for q; $q = C \cdot m \cdot \Delta T$

$$q = \left(4.184 \frac{J}{g \cdot ℃}\right)(1000\ g)(67.0\ ℃)\left(\frac{1\ kJ}{1000\ J}\right) = 280\ kJ$$

Section 11.2 – Measuring and Expressing Heat Changes

Practice Problems

11. A student mixed 50.0 mL of an aqueous solution containing 0.50 mol HCl at 22.5℃ with 50.0 mL of another aqueous solution containing 0.50 mol NaOH at 22.5℃ in a foam cup calorimeter. The temperature of the resulting solution increased to 26.0℃. How much heat in kilojoules (kJ) was released by this reaction? Use $\Delta H = mC\Delta T$

Known:

m = 100 g water solution total; C = 4.18 J/ g·℃; ΔT = +3.5℃

Unknown: ΔH in kJ

$$\Delta H = (100\ g)\left(4.18 \frac{J}{g\ \cdot ℃}\right)(3.5\ ℃) = 1{,}463\ J\ or\ 1.5\ kJ$$

12. A small pebble is heated and placed in a foam cup calorimeter containing 25.0 mL of water at 25.0℃. The water reaches a maximum temperature of 26.4℃. How many joules of heat were released by the pebble? Use $\Delta H = mC\Delta T$

Known:

m = 25 g water solution total; C = 4.18 J/ g·℃; ΔT = +1.4℃

Unknown: ΔH in kJ

$$\Delta H = (25\ g)\left(4.18 \frac{J}{g\ \cdot ℃}\right)(1.4\ ℃) = 146.3\ J\ or\ 0.146\ kJ$$

13. When carbon disulfide is formed from its elements, heat is absorbed. Calculate the amount of heat (in kJ) absorbed when 5.66 g of carbon disulfide is formed.

$C(s) + 2\ S(s) \rightarrow CS_2(l)$ $\quad\quad$ ΔH = 89.3 kJ

$$(5.66\ g\ CS_2)\left(\frac{1\ mol\ CS_2}{76.12\ g\ CS_2}\right)\left(\frac{89.3\ kJ}{1\ mol\ CS_2}\right) = 6.64\ kJ$$

14. The production of iron and carbon dioxide from iron(III) oxide and carbon monoxide is an exothermic reaction. How many kilojoules of heat are produced when 3.40 mol Fe_2O_3 reacts with an excess of CO?

$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g) + 26.3\ kJ$

$$(3.40\ mol\ Fe_2O_3)\left(\frac{26.3\ kJ}{1\ mol\ Fe_2O_3}\right) = 89.4\ kJ$$

Section 11.2

15. When 2 mol of solid magnesium (Mg) combines with 1 mole of oxygen gas (O_2), 2 mol of solid magnesium oxide (MgO) is formed and 1204 kJ of heat is released. Write the thermochemical equation for this combustion reaction.

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s) + 1204\ kJ$ or, $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)\ \Delta H = -1204\ kJ$

16. Gasohol contains ethanol (C_2H_5OH) (l), which when burned reacts with oxygen to produce CO_2(g) and H_2O(g). How much heat is released when 12.5 g of ethanol burns?

$C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$ $\quad \Delta H = -1235\ kJ$

$$(12.5\ g\ C_2H_5OH\ )\left(\frac{1\ mol\ C_2H_5OH}{46.1\ g\ C_2H_5OH}\right)\left(\frac{1{,}235\ kJ}{1\ mol\ C_2H_5OH}\right) = 335\ kJ$$

17. Explain the term *heat of reaction*.

Heat of reaction refers to the heat released or absorbed in a chemical change.

18. Hydrogen gas and fluorine gas react to produce hydrogen fluoride. Calculate the heat change (in kJ) for the conversion of 15.0 g of hydrogen gas to hydrogen fluoride gas at constant pressure.

$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$ $\quad \Delta H = -536\ kJ$

$$(15.0\ g\ H_2)\left(\frac{1\ mol\ H_2}{2.02\ g\ H_2}\right)\left(\frac{-536\ kJ}{1\ mol\ H_2}\right) = -4020\ kJ$$

19. Why is it important to give the physical state of a substance in thermochemical reaction?

Phase changes always involve an energy change.

Section 11.3 – Heat in Changes of State

Practice Problems

20. How many grams of ice at 0℃ and 101.3 kPa could be melted by the addition of 0.400 kJ of heat?

Look to Table 11.5 to first find the ΔH_{fus} for water; it is 6.01 kJ/ mol.

$$g_{ice} = (0.400\ kJ)\left(\frac{1\ mol\ ice}{6.01\ kJ}\right)\left(\frac{18.0\ g\ ice}{1\ mol\ ice}\right) = 1.20\ g\ ice$$

21. How many kilojoules of heat are required to melt a 10.0 g popsicle at 0℃ and 101.3 kPa? Assume the popsicle has the same molar mass and heat capacity as water.

$$(10.0\ g\ ice)\left(\frac{1\ mol\ ice}{18.0\ g\ ice}\right)\left(\frac{6.01\ kJ}{1\ mol\ ice}\right) = 3.34\ kJ$$

22. How much heat (in kJ) is absorbed when 63.7 g $H_2O(l)$ at 100℃ is converted to steam at 100℃?

Look to Table 11.5 to first find the ΔH_{vap} for water; it is 40.7 kJ/ mol.

$$(63.7\ g\ H_2O)\left(\frac{1\ mol\ H_2O}{18.0\ g\ H_2O}\right)\left(\frac{40.7\ kJ}{1\ mol\ H_2O}\right) = 144\ kJ$$

23. How many kilojoules of heat are absorbed when 0.46 g of chloroethane (C_2H_5Cl, bp 12.3℃) vaporizes at its boiling point? The molar heat of vaporization of chloroethane is 26.4 kJ/ mol.

$$(0.46\ g\ C_2H_5Cl)\left(\frac{1\ mol\ C_2H_5Cl}{64.51\ g\ C_2H_5Cl}\right)\left(\frac{26.4\ kJ}{1\ mol\ C_2H_5Cl}\right) = 0.188\ kJ$$

24. How much heat (in kJ) is released when 0.677 mol NaOH(s) is dissolved in water?

$$(0.677\ mol\ NaOH)\left(\frac{-445.1\ kJ}{1\ mol\ NaOH}\right) = 301\ kJ$$

25. How many moles of $NH_4NO_3(s)$ must be dissolved in water so that 88.0 kJ of heat is released from the water?

$$(88.0\ kJ)\left(\frac{1\ mol\ NH_4NO_3}{25.7\ kJ}\right) = 3.42\ mol\ NH_4NO_3$$

Section Review 11.3

26. Identify each heat change by name and classify each change as exothermic or endothermic.

a. 1 mol $C_3H_8(l) \rightarrow$ 1 mol $C_3H_8(g)$ molar heat of vaporization; endothermic

b. 1 mol NaCl(s) + 3.88 kJ/ mol $\rightarrow$ 1 mol NaCl(aq) molar heat of solution; endothermic

c. 1 mol NaCl(s) $\rightarrow$ 1 mol NaCl(l) molar heat of fusion; endothermic

d. 1 mol $NH_3(g) \rightarrow$ 1 mol $NH_3(l)$ molar heat of condensation; exothermic

e. 1 mol Hg(l) $\rightarrow$ 1 mol Hg(s) molar heat of solidification; exothermic

27. Heavy water, in which the hydrogens are hydrogen-2 instead of the more common hydrogen-1, is called deuterium oxide (D_2O). Solid D_2O melts at 3.78℃. The molar heat of fusion of $D_2O(s)$ is 6.34 kJ/ mol. How much heat is released when 8.46 g $D_2O(l)$ solidifies at its melting point?

$$(8.46\ g\ D_2O)\left(\frac{1\ mol\ D_2O}{20.0\ g\ D_2O}\right)\left(\frac{6.34\ kJ}{1\ mol\ D_2O}\right) = 2.68\ kJ$$

28. Why is a burn from steam potentially far more serious than a burn from very hot water?

When a single mole of steam condenses, it releases a substantial amount of heat! There is much more kinetic energy in steam particles as well.

29. Why does an ice cube melt at room temperature?

The ice absorbs sufficient heat from the surroundings enabling it to change from the solid to the liquid state.

Section 11.4 – Calculating Heat Changes

Practice Problems

30. Use the standard heats of formation to calculate the standards heats of reaction (ΔH^0) for these reactions.

Look at Table 11.6 in text to determine the ΔH_f^0 for bromine in it's liquid form, and in its gaseous form. We know $\Delta H^0 = \Delta H_f^0(\text{prod}) - \Delta H_f^0(\text{react})$

a. $Br_2(g) \rightarrow Br_2(l)$

$\Delta H^0 = \Delta H_f^0$(0 kJ/ mol)-$\Delta H_f^0$(30.91 kJ/ mol) = -30.91 kJ energy released per mole of Br_2 gas converted to liquid form.

b. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$\Delta H^0 = \Delta H_f^0$(-635.1 kJ/ mol + -393.5 kJ/ mol)-ΔH_f^0(-1207 kJ/ mol) = 178.4 kJ energy absorbed per mole of $CaCO_3$ decomposed.

c. $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

$\Delta H^0 = \Delta H_f^0$[(2)(33.85 kJ/mol)] -$\Delta H_f^0$[(2)(90.37 kJ/mol) + (1)(0 kJ/mol)] = -113.0 kJ energy released per mole of NO_2 synthesized.

31. With one exception, the standard heats of formation of Na(s), $O_2(g)$, $Br_2(l)$, CO(g), Fe(s), and He(g) are identical. What is the exception? Explain.

All the substances are in their elemental states, thus their standard heats of formation is zero, with the exception of CO, which is a compound.

Section Review 11.4

32. Calculate the enthalpy change (ΔH) in kJ for the following reaction.

$2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

Use the enthalpy changes for the combustion of aluminum and iron:

$2\ Al(s) + \cancel{1\tfrac{1}{2}\ O_2(g)} \rightarrow Al_2O_3(s)$ ΔH = -1669.8 kJ

~~$2\ Fe(s) + 1\tfrac{1}{2}\ O_2(g) \rightarrow Fe_2O_3(s)$ ΔH = -824.2 kJ~~ Reverse this rxn, rewrite.

$Fe_2O_3(s) \rightarrow 2\ Fe(s) + \cancel{1\tfrac{1}{2}\ O_2(g)}$ ΔH = +824.2 kJ Combine the ΔH

$2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$ ΔH = -845.6 kJ

33. What is the standard heat of reaction (ΔH^0) for the decomposition of hydrogen peroxide?

$2\ H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g)$

Use Table 11.6 to solve:

$\Delta H^0 = \Delta H_f^0$ [(2)(-285.8 kJ/ mol)+ (0.0 kJ/ mol)] – [(2)(-187.8 kJ/mol)]

= -196.0 kJ energy

34. State Hess's law of heat summation in your own words. Explain its usefulness.

When solving for the enthalpy change of a theoretical reaction, other chemical equations can be added algebraically, along with their enthalpies.

35. What happens to the sign of ΔH when the reverse of a chemical reaction is written? Why?

The sign of ΔH must be changed, since we are rewriting an equation in its reverse.

Chapter 11 Review

57. Calculate the heat change for the formation of lead(IV) chloride by the reaction of lead(II) chloride with chlorine. 11.4

$PbCl_2(s) + Cl_2(g) \rightarrow PbCl_4(l)$ $\Delta H = ?$

Use the following thermochemical equations.

$\cancel{Pb(s)} + \cancel{2}\ 1\ Cl_2(g) \rightarrow PbCl_4(l)$ $\Delta H = -329.2$ kJ

~~$Pb(s) + Cl_2(g) \rightarrow PbCl_2(s)$ $\Delta H = -359.4$ kJ~~

$PbCl_2(s) \rightarrow \cancel{Pb(s)} + \cancel{Cl_2}(g)$ $\Delta H = 359.4$ kJ

$PbCl_2(s) + Cl_2(g) \rightarrow PbCl_4(l)$ $\Delta H = 30.2$ kJ heat

Reverse the second reaction, rewrite it and change the ΔH sign from -359.4 kJ to +359.4 kJ.

Cancel out terms that appear on both sides, and add.

60. What is the standard heat of formation of a free element in its standard state? 11.4

Zero

61. Consider the statement, "the more negative the value of ΔH_f^0, the more stable the compound." Is this statement true or false? Explain. 11.4

This statement is true, because stability implies lower energy. The greater the release of heat in any reaction, the more stable is the compound, relative to its elements (all of which have ΔH_f^0 equal to zero).

71. The molar heat of vaporization of ethanol (C_2H_5OH) (l) is 43.5 kJ/ mol. Calculate the heat required to vaporize 25.0 g of ethanol at its boiling point.

$$(25.0\ g\ C_2H_5OH)\left(\frac{1\ mol\ C_2H_5OH}{46.07\ g\ C_2H_5OH}\right)\left(\frac{43.5\ kJ}{1\ mol\ C_2H_5OH}\right) = 23.6\ kJ$$