### Ch 16 Study Guide: Thermochemistry KEY

**Thermochemistry**
- exothermic vs. endothermic reactions
- measuring heat flow
- calorimetry (coffee cup calorimeters; bomb calorimeters)
- specific heat; heat capacity
- enthalpy, entropy, free energy
- heat in balanced chemical equations
- calculating ΔH (Hess' Law)
- enthalpies of formation
- spontaneity of reactions

**Equations:**
- \( q = mc \Delta t \) (applies to constant state only)
- \( q = (C_{cal}) (\Delta t) \)
- \( q_{rxn} = -q_{cal} \)
- \( \Delta H = (\Delta H_{products}) - (\Delta H_{reactants}) \) (eqn also applies to S and G)
- \( \Delta G = \Delta H - T \Delta S \)

**Phase Change Equations:**
- \( q = \text{(mass)} (\text{heat of fusion}) \)
- \( q = \text{(mass)} (\text{heat of vaporization}) \)

**Constants:**
- Specific heat of ice = 2.09 J/g°C
- Specific heat of water = 4.18 J/g°C
- Specific heat of steam = 2.03 J/g°C
- Heat of fusion of water = 334 J/g
- Heat of vaporization of water = 2260 J/g

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1. A copper pot with a mass of 772 grams absorbs 22.7 kJ of heat. Its final temperature is 137.0°C. What was its initial temperature? (the specific heat of copper is 0.385 J/g°C)
   
   \[
   q = mc \Delta t \\
   22,700 \text{ J} = (772 \text{ g})(0.385 \text{ J/g°C}) \Delta T \\
   \Delta T = 76.4 ^\circ \text{C} \\
   \Delta T = T_f - T_i \\
   76.4 ^\circ \text{C} = 137.0 ^\circ \text{C} - T_i \\
   T_i = 60.6 ^\circ \text{C}
   \]

2. How much heat is absorbed by a 15.5 g piece of gold as it is heated from 4.5°C to 177.4°C? (the specific heat of gold is 0.129 J/g°C)
   
   \[
   q = mc \Delta t \\
   q = (15.5 \text{ g})(0.129 \text{ J/g°C})(172.9 \text{ °C}) \\
   q = 346 \text{ J}
   \]

3. A container full of water absorbs 64.4 kJ of heat and its temperature rises from 22.0°C to 73.4°C. What is the volume of water in mL? (the density of water = 1 g/mL)
   
   \[
   q = mc \Delta t \\
   64,400 \text{ J} = (m)(4.184 \text{ J/g°C})(51.4°C) \\
   m = 299.45 \text{ g} \\
   v = \frac{m}{d} = \frac{299.45 \text{ g}}{(1.00 \text{ g/mL})} \\
   v = 299 \text{ mL}
   \]

4. A sample of fructose, \( C_6H_{12}O_6 \), weighing 7.55 g is burned in a bomb calorimeter. The heat capacity of the calorimeter is 2.155 x 10^4 J/°C. The temperature in the calorimeter rises from 22.54°C to 29.56°C.
   (a) What is \( q \) when the 7.55 g of fructose is burned?

   \[
   \begin{align*}
   q_{cal} &= (C_{cal})(\Delta t) \\
   q_{cal} &= (2.155 \times 10^4 \text{ J/°C})(7.02°C) \\
   q_{cal} &= 150930 \text{ J} = 151 \text{ kJ} \\
   q_{rxn} &= -151 \text{ kJ}
   \end{align*}
   \]

   (b) What is \( q \) for the combustion of 1 mole of fructose?

   \[
   \frac{-151 \text{ kJ}}{7.55 \text{ g} \ C_6H_{12}O_6} \times \frac{180.0 \text{ g}}{1 \text{ mol}} = -3600 \text{ kJ/mol} = -3.60 \times 10^3 \text{ kJ/mol}
   \]

5. Naphthalene, \( C_{10}H_8 \), is the compound present in moth balls. When one mole of naphthalene is burned, 5.15 x 10^3 kJ of heat is evolved. A sample of naphthalene burned in a bomb calorimeter (heat capacity = 9832 J/°C) increases the temperature in the calorimeter from 25.1°C to 28.4°C. How many milligrams of naphthalene were burned?

   \[
   \begin{align*}
   q_{cal} &= (C_{cal})(\Delta t) \\
   q_{cal} &= (9832 \text{ J/°C})(3.3°C) \\
   q_{cal} &= 32445.6 \text{ J} = 32.4 \text{ kJ} \\
   q_{rxn} &= -32.4 \text{ kJ}
   \end{align*}
   \]
1 mol C_{10}H_{8} = 128.0 g = 128,000 mg

\[
\frac{-32,445.6 \text{ kJ}}{\text{mg}} = \frac{-5.15 \times 10^{3} \text{ kJ}}{128,000 \text{ mg}}
\]

\[x = 806 \text{ mg}\]

6. Nitrogen monoxide (NO) has been found to react with oxygen gas (O_{2}) to produce the brown gas nitrogen dioxide (NO_{2}). When one mole of NO reacts with oxygen, 57.0 kJ of heat is evolved.

(a) Write the thermochemical equation for the reaction between one mole of nitrogen monoxide and oxygen to produce nitrogen dioxide.

\[
\text{NO} + \frac{1}{2} \text{O}_{2} \rightarrow \text{NO}_{2} \quad \Delta H = -57.0 \text{ kJ}
\]

(b) Is the reaction exothermic or endothermic?

exothermic

(c) What is \(\Delta H\) when 5.00 g of nitrogen monoxide reacts?

\[
-57.0 \text{ kJ} \times \frac{1 \text{ mol NO}}{30.0 \text{ g}} \times 5.00 \text{ g} = -9.50 \text{ kJ}
\]

(d) How many grams of nitrogen monoxide must react with an excess of oxygen to produce 10.0 kJ of heat?

\[
\frac{-9.50 \text{ kJ}}{5.00 \text{ g}} = \frac{x}{10.0 \text{ kJ}}
\]

\[x = 5.26 \text{ g}\]

7. Strontium metal (Sr) combines with graphite (C) and oxygen gas (O_{2}) to produce strontium carbonate (SrCO_{3}). The formation of one mole of SrCO_{3} releases 1.220 \times 10^{3} \text{ kJ} of heat.

(a) Write a balanced thermochemical equation for the reaction resulting in the formation of one mole of SrCO_{3}.

\[
\text{Sr} + \text{C} + \frac{1}{2} \text{O}_{2} \rightarrow \text{SrCO}_{3} \quad \Delta H = -1.220 \times 10^{3} \text{ kJ}
\]

(b) What is \(\Delta H\) when 10.00 g of strontium reacts with excess graphite and oxygen?

\[
-1.220 \times 10^{3} \text{ kJ} \times \frac{1 \text{ mol Sr}}{87.6 \text{ g}} \times 10.00 \text{ g} = -139.3 \text{ kJ}
\]

(c) What mass of SrCO_{3} forms when 2355 kJ of heat are also formed?

\[
\frac{-1220 \text{ kJ}}{147.6 \text{ g SrCO}_{3}} = \frac{-2355 \text{ kJ}}{x}
\]

\[x = 284.9 \text{ g}\]

8. Given: 2CuO(s) \rightarrow 2Cu(s) + O_{2}(g) \quad \Delta H = 314.6 \text{ kJ}

(a) Determine the heat of formation of CuO(s).

\[
2\text{Cu(s)} + \frac{1}{2} \text{O}_{2}(g) \rightarrow 2\text{CuO(s)} \quad \Delta H = -314.6 \text{ kJ}
\]

\[
\text{Cu(s)} + \frac{1}{2} \text{O}_{2}(g) \rightarrow \text{CuO(s)} \quad \Delta H = -157.3 \text{ kJ}
\]

(Since \(\Delta H = (\Delta H_{\text{products}}) - (\Delta H_{\text{reactants}})\) and the heats of formation of reactants are both zero, heat of formation of one mole of CuO = -157.3 kJ)

(b) Calculate \(\Delta H\) for the formation of 13.58 g of CuO.

\[
-157.3 \text{ kJ} \times \frac{1 \text{ mol CuO}}{79.5 \text{ g}} \times 13.58 \text{ g} = -26.87 \text{ kJ}
\]

9. Limestone, CaCO_{3}, when subjected to a temperature of 900°C in a kiln, decomposes to solid calcium oxide and carbon dioxide gas.

(a) Write a balanced chemical equation for this reaction.

\[
\text{CaCO}_{3} \rightarrow \text{CaO} + \text{CO}_{2}
\]
(b) Determine $\Delta H$ for the reaction using the handout of standard heats of formation.

$$\Delta H = \left[ (-635.1) + (-393.5) \right] - (-1206.9) - (-1206.9) = 178.3 \text{ kJ/mol}$$

(c) How much heat is evolved or absorbed when one gram of limestone decomposes?

$$\frac{178.3 \text{ kJ}}{1 \text{ mol CaCO}_3} \times \frac{1 \text{ mol CaCO}_3}{100.1 \text{ g}} \times 1.000 \text{ g} = 1.781 \text{ kJ} \text{ (absorbed)}$$

10. How much energy is released when 52.3 g of steam at 136.5°C cools and condenses to form water at 93.2°C?

$$q_{136.5\to100} = (52.3 \text{ g})(-2.03 \text{ J/g°C})(36.5°C) = -3875 \text{ J}$$

$$q_{\text{vap}} = (52.3 \text{ g})(-2260 \text{ J/g}) = -118,198 \text{ J}$$

$$q_{100\to93.2} = (52.3 \text{ g})(-4.184 \text{ J/g°C})(6.8°C) = -1488 \text{ J}$$

$$q_{\text{total}} = -123,561 \text{ J} = -124 \text{ kJ}$$

11. How much energy is needed to heat a 42.3 g sample of ice at -35.7°C to steam at 112.0°C?

$$q_{35.7\to0} = (42.3 \text{ g})(2.09 \text{ J/g°C})(35.7°C) = 3156 \text{ J}$$

$$q_{\text{fus}} = (42.3 \text{ g})(334 \text{ J/g}) = 14,128 \text{ J}$$

$$q_{0\to100} = (42.3 \text{ g})(4.184 \text{ J/g°C})(100.0°C) = 17,698 \text{ J}$$

$$q_{\text{vap}} = (42.3 \text{ g})(2260 \text{ J/g}) = 95,598 \text{ J}$$

$$q_{100\to112} = (42.3 \text{ g})(2.03 \text{ J/g°C})(12.0°C) = 1030 \text{ J}$$

$$q_{\text{total}} = 131,610 \text{ J} = 132 \text{ kJ}$$

12. Hess's Law and $H$:

(a) Explain why Hess's law is used in the chemistry laboratory?

This law can be used to allow us to calculate heats of reaction for reactions that cannot be carried out in a simple calorimeter. Instead, we can add together the enthalpies of reaction for reactions we can measure, then add them in such a way that we find the enthalpy of another reaction.

(b) How can $H$ be calculated for an equation in which the coefficients have been multiplied by a factor of two?

Multiply $H$ by 2

(c) What happens to the sign of $H$ if a reaction is run in the reverse direction from the way it is written?

Reverse the sign of $H$

13. What is meant by the terms of heat of fusion and heat of vaporization?

The amount of energy released or absorbed to cause one gram of a substance to change state. Heat of fusion relates to melting and freezing, while heat of vaporization relates to evaporating and condensing.

14. From the following enthalpy changes,

$$2\text{PbS(s)} + 3\text{O}_2(g) \rightarrow 2\text{PbO(s)} + 2\text{SO}_2(g) \quad H^\circ = -827.0 \text{ kJ}$$

$$\text{PbO(s)} + \text{C(s)} \rightarrow \text{Pb(s)} + \text{CO(g)} \quad H^\circ = +106.8 \text{ kJ}$$

(a) Calculate the value of $H^\circ$ in the following reaction:

$$2\text{PbS(s)} + 3\text{O}_2(g) + 2\text{C(s)} \rightarrow 2\text{Pb(s)} + 2\text{CO(g)} + 2\text{SO}_2(g).$$

Multiply second equation by two

Add the two equations together

$$H = (-827.0 \text{ kJ}) + (213.6) = -613.4 \text{ kJ}$$

(b) Is the reaction endothermic or exothermic?

exothermic
15. Determine the change in enthalpy for the following reaction: \( \text{C (graphite) + 2H}_2(g) \rightarrow \text{CH}_4(g) \)

Use these reaction equations:

\[
\begin{align*}
\text{C (graphite) + O}_2(g) & \rightarrow \text{CO}_2(g) \quad H^\circ = -394 \text{ kJ} \\
\text{H}_2(g) + \frac{1}{2} \text{O}_2(g) & \rightarrow \text{H}_2\text{O}(l) \quad H^\circ = -286 \text{ kJ} \\
\text{CO}_2(g) + 2\text{H}_2\text{O}(l) & \rightarrow \text{CH}_4(g) + 2\text{O}_2(g) \quad H^\circ = +890.3 \text{ kJ}
\end{align*}
\]

Multiply middle equation by 2
Add the three equations together
\[ H = (-394 \text{ kJ}) + (2)(-286) + (+890.3) = -75.7 \text{ kJ} \]

16. A reaction at 45 °C has the following enthalpy and entropy: \( \Delta H^\circ = -86.6 \text{ kJ} \) and \( \Delta S^\circ = -382 \text{ J/K} \).

(a) Calculate \( \Delta G^\circ \)

\[ \Delta G = \Delta H - T\Delta S \\
\Delta G = (-86.6 \text{ kJ}) - (318\text{K})(-0.312 \text{ kJ/K}) \\
\Delta G = 12.6 \text{ kJ} \]

(b) Is the reaction spontaneous at this temperature?

NO