Chapter 17.1-The Flow of Energy

Chapter 17.2-Thermochemical Equations

Chapter 17.3-Heat in Changes of State

Chapter 17.4-Calculating Heats of Reaction

Text Reference- p. 505-532
Study Guide- p. 534

Terms

- thermochemistry-the study of energy changes that occur during chemical reactions or changes of state.
- heat (represented by 'q') is energy that transfers from one object to another because of a temperature difference between them.

Heat ALWAYS flows from a warmer object to a cooler one until the temperature is equalized.

- exothermic-the 'system' loses heat as the surroundings heat up
- endothermic-the 'system' gains heat as the surroundings cool down

Units of Heat Measurement

Heat flow is measured in two common units, calorie or joule.

- calorie (cal) is defined as the quantity of heat (q) needed to raise the temperature of 1g of pure water 1° C.
- 1 dietary Calorie, is equivalent to 1 kilocalorie, 1000 calories
- the joule is the SI unit
- 1 J of heat raises the temperature of water 0.239 ° C

∴ 1 J = 0.239 cal 4.184 J = 1 cal

Heat Capacity and Specific Heat

Heat Capacity = the amount of heat required to raise the temperature of an object exactly 1℃

Specific Heat Capacity (represented by C)
-also called simply 'specific heat'
-the amount of heat required to raise 1g of a substance exactly 1℃

Table 17.1

Specific Heats of Some Common Substances

Substance	Specific Heat	
	J/(g•°C)	cal/(g•°C)
Water	4.18	1.00
Grain alcohol	2.4	0.58
Ice	2.1	0.50
Steam	1.7	0.40
Chloroform	0.96	0.23
Aluminum	0.90	0.21
Iron	0.46	0.11
Silver	0.24	0.057
Mercury	0.14	0.033

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Calculation of Specific Heat

$$C = \frac{q}{m \times \Delta T} = \frac{\text{heat (joules or calories)}}{\text{mass (g)} \times \text{change in temperature (°C)}}$$

Note!

$$\Delta T = T_f - T_i$$

Units are either

$\frac{J}{g\,℃}$ $\frac{cal}{g\,℃}$

SAMPLE PROBLEM 17.1

Calculating the Specific Heat of a Metal

The temperature of a 95.4-g piece of copper increases from 25.0°C to 48.0°C when the copper absorbs 849 J of heat. What is the specific heat of copper?

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#3,4
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Practice Problems

3. 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 the olive oil?

4. How much heat is required to raise the temperature of 250.0 g of mercury 52°C?

17.2-Thermochemical Equations

- a thermochemical equation is simply a regular chemical equation that identifies the reactants and products, but includes the enthalpy change that occurs as a result of the reaction
- the terms 'heat' and 'enthalpy change' are used interchangeably
- this means that $q = \Delta H$

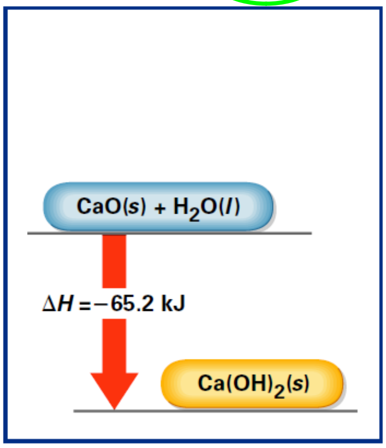
Example 1:

$CaO + H_2O(s) \longrightarrow Ca(OH)_2(l) + 65.2kJ$

ΔH

Heat of Reaction

ΔH_{rxn}



a Exothermic Reaction

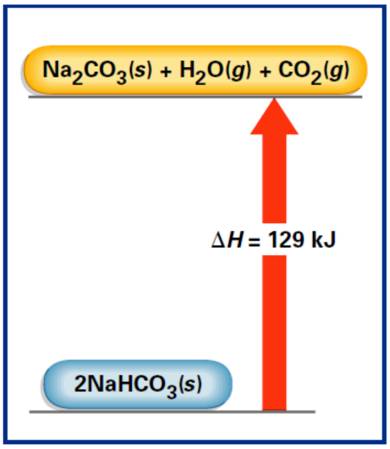
Example 2:

$2NaHCO_3(s) + 129kJ \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

ΔH

Heat of Reaction

ΔH_{rxn}



b Endothermic Reaction

ΔH for endothermic reactions is (+)

ΔH for exothermic reactions is (-)

SAMPLE PROBLEM 17.3

Using the Heat of Reaction to Calculate Enthalpy Change

Using the thermochemical equation in Figure 17.7b on page 515, calculate the amount of heat (in kJ) required to decompose 2.24 mol $NaHCO_3(s)$.

$$2NaHCO_3(s) + 129kJ \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

2 moles of $NaHCO_3$
requires 129kJ of heat

2 moles NaHCO3 : 129kJ

$\therefore$ 2.24 moles NaHCO3 : x kJ

Cross Multiply

$(2.24)(129) = 2x$

$\therefore \quad x = \dfrac{(2.24)(129)}{2} = 144kJ$

Practice Problems

14. 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.

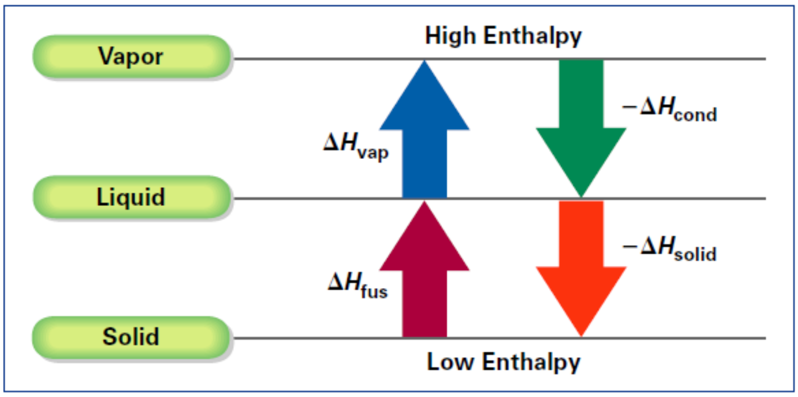
$$C(s) + 2S(s) \longrightarrow CS_2(l) \quad \Delta H = 89.3 \text{ kJ}$$

15. 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?

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g) + 26.3 \text{ kJ}$$

17.3-Heat In Changes of State

- All matter is made of tiny particles which are in constant motion.
- the temperature of a substance is a reflection of 'average' speed of the particles.
- heating the substance increases the average speed and therefore increases the temperature.
- Heat energy added to the system does NOT always increase the temperature, it simply increases the seperation between the particles, thereby changing them from solids to liquids or liquids to gases.
- Melting and evaporation are endothermic, since the system absorbds heat energy from the surroundings.
- the reverse, condensation or solidification, release heat to the surroundings, and therefore are exothermic.



Terms

Molar Heat of Vaporization - the amount of heat needed for 1 mole of a substance to change from a liquid to a gas

$$\Delta H_{vap}$$

Molar Heat of Fusion - the amount of heat released when 1 mole of a substance changes from a gas to a liquid.

$$\Delta H_{fus}$$

similar terms for:

heat of condensation ΔH_{cond} and heat of solidification ΔH_{solid}

- each reverse process has the same heat value
- the heat of fusion (melting) is the same as the heat of solidification (freezing)
- the heat of vaporization (evaporation) is the same as the heat of condensation (condensing)

$$\Delta H_{fus} \text{ of water} = 6.01 \text{kJ/mol}$$

means that 6.01kJ of heat energy is needed to melt 1 mole of solid water to make 1 mol of liquid water **AT THE SAME TEMPERATURE**

SAMPLE PROBLEM 17.4

Using the Heat of Fusion in Phase-Change Calculations

How many grams of ice at 0°C will melt if 2.25 kJ of heat are added?

Knowns
- Initial and final temperatures are 0°C
- $\Delta H_{fus} = 6.01$ kJ/mol
- $\Delta H = 2.25$ kJ

Unknown
- $m_{ice} = ?$ g

Use the thermochemical equation

$$H_2O(s) + 6.01\text{ kJ} \longrightarrow H_2O(l)$$

1 mole : 6.01kJ
x mole : 2.25kJ $\therefore$ (6.01) x = 2.25 $x = \frac{2.25}{6.01} = 0.374\text{mol}$

1 mole : 18g
0.37 mole : x g $\therefore$ x = (0.374)(18) = 6.73g

Heat of Vaporization and Condensation

- similar events occur when liquids change to gases and vice versa

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Heats of Physical Change

Substance	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)
Ammonia (NH_3)	5.65	23.4
Ethanol (C_2H_5OH)	4.60	43.5
Hydrogen (H_2)	0.12	0.90
Methanol (CH_3OH)	3.16	35.3
Oxygen (O_2)	0.44	6.82
Water (H_2O)	6.01	40.7

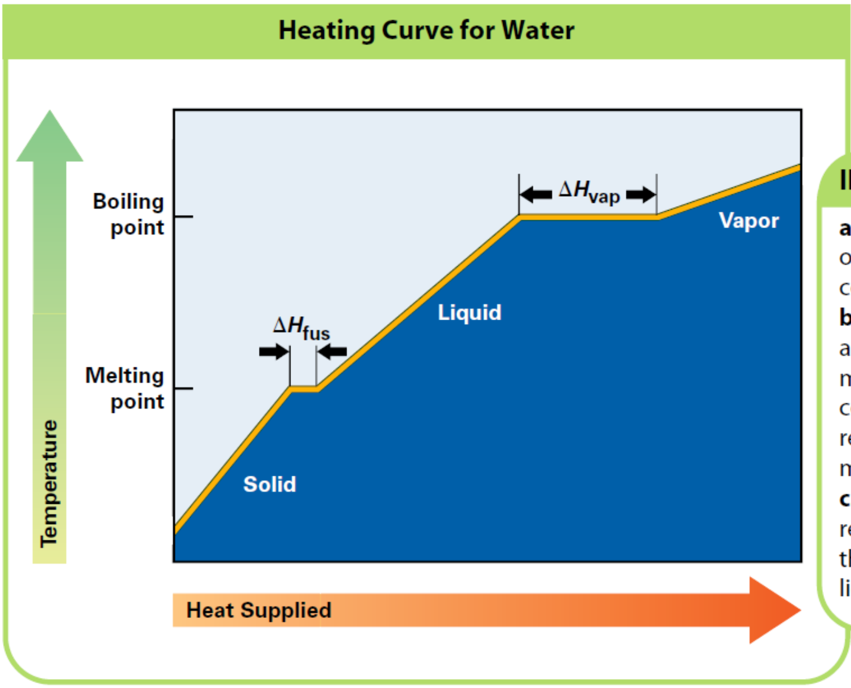

Figure 17.10 A heating curve graphically describes the enthalpy changes that take place during phase changes.

INTERPRETING GRAPHS

a. Identify In which region(s) of the graph is temperature constant?

b. Describe How does the amount of energy required to melt a given mass of ice compare to the energy required to vaporize the same mass of water? Explain.

c. Apply Concepts Which region of the graph represents the coexistence of solid and liquid? Liquid and vapor?

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21. How many kilojoules of heat are required to melt a 10.0-g popsicle at 0°C? Assume the popsicle has the same molar mass and heat of fusion as water.

22. How many grams of ice at 0°C could be melted by the addition of 0.400 kJ of heat?

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23. How much heat is absorbed when 63.7g $H_2O(l)$ at 100°C and 101.3 kPa is converted to steam at 100°C? Express your answer in kJ.

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

17.4-Calculating Heats of Reaction