

Chapter 5

“Thermochemistry”



Section 17.1

The Flow of Energy – Heat and Work

◆ OBJECTIVES:

- Explain how energy, heat, and work are related.

Section 17.1

The Flow of Energy – Heat and Work

◆ OBJECTIVES:

- Classify processes as either exothermic or endothermic.

Section 17.1

The Flow of Energy – Heat and Work

◆ OBJECTIVES:

- Identify the units used to measure heat transfer.

Section 17.1

The Flow of Energy – Heat and Work

◆ OBJECTIVES:

- Distinguish between heat capacity and specific heat.

Energy Transformations

- ◆ Thermochemistry - concerned with heat changes that occur during chemical reactions
- ◆ Energy - capacity for doing work or supplying heat
 - weightless, odorless, tasteless
 - if within the chemical substances- called chemical potential energy

Energy Transformations

- ◆ Gasoline contains a significant amount of chemical potential energy
- ◆ Heat - represented by “ q ”, is energy that transfers from one object to another, because of a temperature difference between them.
 - *only changes* can be detected!
 - flows from warmer → cooler object

Exothermic and Endothermic Processes

- ◆ Essentially all chemical reactions, and changes in physical state, involve either:
 - a) release of heat, or
 - b) absorption of heat

Exothermic and Endothermic Processes

- ◆ In studying heat changes, think of defining these two parts:
 - the system - the part of the universe on which you focus your attention
 - the surroundings - includes everything else in the universe

Exothermic and Endothermic Processes

- ◆ Together, the system and it's surroundings constitute the *universe*
- ◆ Thermochemistry is concerned with the flow of heat from the system to it's surroundings, and vice-versa.

Exothermic and Endothermic Processes

- ◆ **The Law of Conservation of Energy states that in any chemical or physical process, energy is neither created nor destroyed.**
 - All the energy is accounted for as work, stored energy, or heat.

Exothermic and Endothermic Processes

- ◆ Heat flowing into a system from it's surroundings:
 - defined as positive
 - q has a positive value
 - called endothermic
 - system *gains heat* as the surroundings cool down

Exothermic and Endothermic Processes

- ◆ Heat flowing out of a system into it's surroundings:
 - defined as negative
 - q has a negative value
 - called exothermic
 - system *loses heat* as the surroundings heat up

Exothermic and Endothermic

- ◆ Fig. 17.2, page 506 - on the left, the system (the people) gain heat from it's surroundings (the fire)
 - this is *endothermic*
- ◆ On the right, the system (the body) cools as perspiration evaporates, and heat flows to the surroundings
 - this is *exothermic*

Exothermic and Endothermic

- ◆ Every reaction has an energy change associated with it
- ◆ **** Gummy Bear Sacrifice ****
- ◆ Exothermic reactions release energy, usually in the form of heat.
- ◆ Endothermic reactions absorb energy
- ◆ Energy is stored in bonds between atoms

Units for Measuring Heat Flow

- 1) A calorie is defined as the quantity of heat needed to raise the temperature of 1 g of pure water 1 °C.
 - Used except when referring to food
 - a Calorie, (written with a capital C), always refers to the energy in food
 - 1 Calorie = 1 kilocalorie = 1000 cal.

Units for Measuring Heat Flow

2) The calorie is also related to the Joule, the SI unit of heat and energy

- named after James Prescott Joule

- $4.184 \text{ J} = 1 \text{ cal}$

- ◆ Heat Capacity - the amount of heat needed to increase the temperature of an object exactly $1 \text{ }^{\circ}\text{C}$

- Depends on both the object's mass and its chemical composition

Heat Capacity and Specific Heat

- ◆ Specific Heat Capacity - the amount of heat it takes to raise the temperature of *1 gram* of the substance by $1\text{ }^{\circ}\text{C}$ (abbreviated “C”)
 - often called simply “Specific Heat”
 - Note Table 17.1, page 508
- ◆ Water has a **HUGE** value, when it is compared to other chemicals

TABLE 17-1 Specific Heats of Some Common Substances at 298.15 K

Substance	Specific heat J/(g·K)
Water (<i>l</i>)	4.18 ←
Water (<i>s</i>)	2.06
Water (<i>g</i>)	1.87
Ammonia (<i>g</i>)	2.09
Benzene (<i>l</i>)	1.74
Ethanol (<i>l</i>)	2.44
Ethanol (<i>g</i>)	1.42
Aluminum (<i>s</i>)	0.897
Calcium (<i>s</i>)	0.647
Carbon, graphite (<i>s</i>)	0.709
Copper (<i>s</i>)	0.385
Gold (<i>s</i>)	0.129
Iron (<i>s</i>)	0.449 ←
Mercury (<i>l</i>)	0.140
Lead (<i>s</i>)	0.129

Heat Capacity and Specific Heat

- ◆ For water, $C = 4.18 \text{ J}/(\text{g } ^\circ\text{C})$, and also $C = 1.00 \text{ cal}/(\text{g } ^\circ\text{C})$
- ◆ Thus, for water:
 - it takes a long time to heat up, and
 - it takes a long time to cool off!
- ◆ Water is used as a coolant!
 - Note Figure 17.4, page 509

Heat Capacity and Specific Heat

- ◆ To calculate, use the formula:

$$q = \text{mass (g)} \times \Delta T \times C$$

- ◆ heat is abbreviated as “q”
- ◆ ΔT = change in temperature
- ◆ C = Specific Heat
- ◆ Units are either $\text{J}/(\text{g } ^\circ\text{C})$ or $\text{cal}/(\text{g } ^\circ\text{C})$

SAMPLE PROBLEM 17.1 - Page 510

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?

1 Analyze *List the knowns and the unknown.*

Knowns

- $m_{\text{Cu}} = 95.4\text{g}$
- $\Delta T = (48.0^\circ\text{C} - 25.0^\circ\text{C}) = 23.0^\circ\text{C}$
- $q = 849\text{ J}$

Unknown

- $C_{\text{Cu}} = ?\text{ J}/(\text{g}\cdot^\circ\text{C})$

2 Calculate *Solve for the unknown.*

Use the known values and the definition of specific heat,

$C = \frac{q}{m \times \Delta T}$, to calculate the unknown value C_{Cu} .

$$C_{\text{Cu}} = \frac{q}{m \times \Delta T} = \frac{849\text{ J}}{95.4\text{ g} \times 23.0^\circ\text{C}} = 0.387\text{ J}/(\text{g}\cdot^\circ\text{C})$$

Section 11.2

Measuring and Expressing Enthalpy Changes

◆ OBJECTIVES:

- Describe how calorimeters are used to measure heat flow.

Section 11.2

Measuring and Expressing Enthalpy Changes

◆ OBJECTIVES:

- Construct
thermochemical
equations.

Section 11.2

Measuring and Expressing Enthalpy Changes

◆ OBJECTIVES:

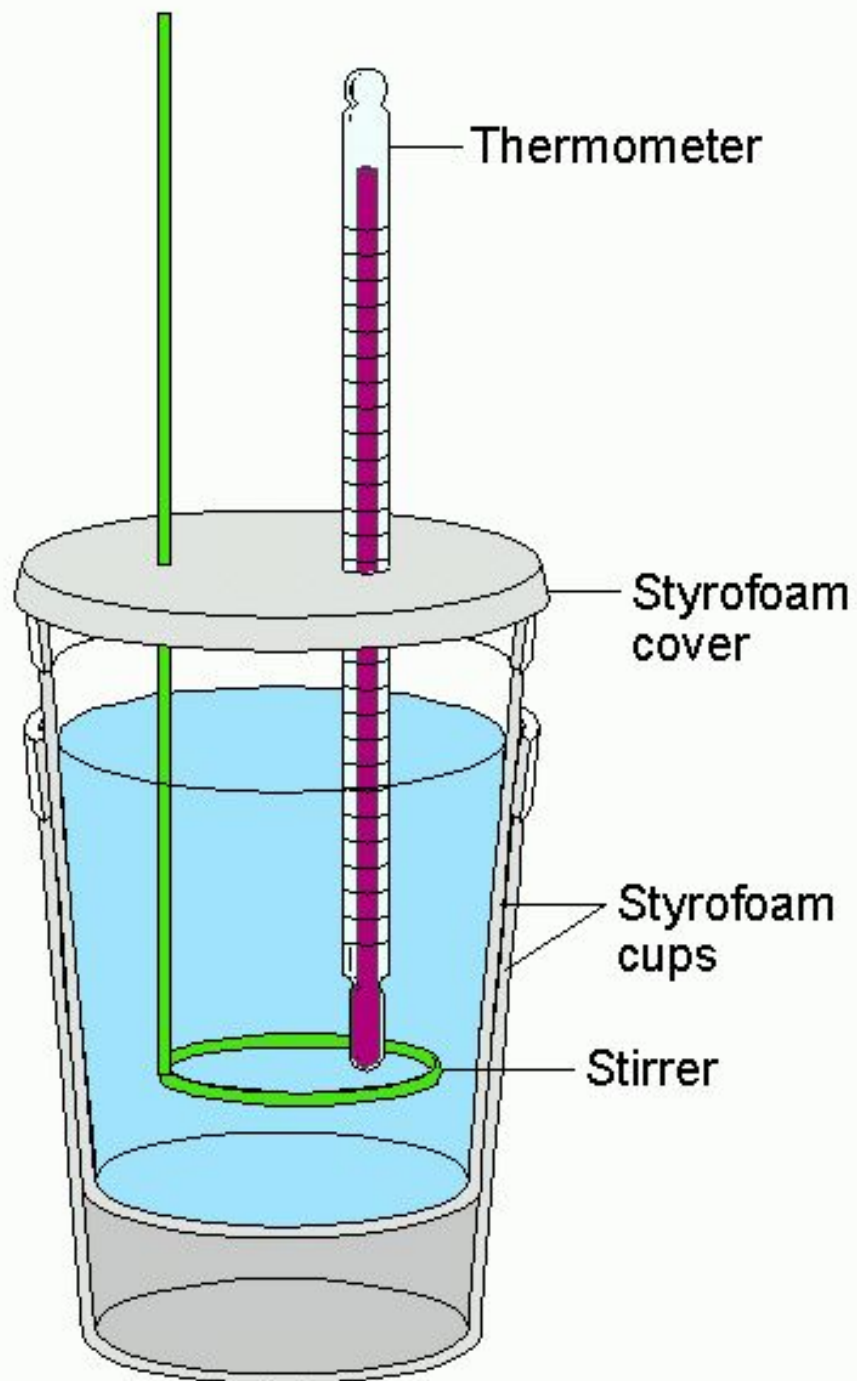
- Solve for enthalpy changes in chemical reactions by using heats of reaction.

Calorimetry

- ◆ Calorimetry - the precise measurement of the heat into or out of a system for chemical and physical processes.
- ◆ The device used to *measure* the absorption or release of heat in chemical or physical processes is called a Calorimeter

Calorimetry

- ◆ Foam cups are excellent heat insulators, and are commonly used as simple calorimeters
 - Fig. 17.5, page 511
- ◆ For systems at constant pressure, the heat content is the same as a property called Enthalpy (H) of the system



A foam cup calorimeter – here, two cups are nestled together for better insulation

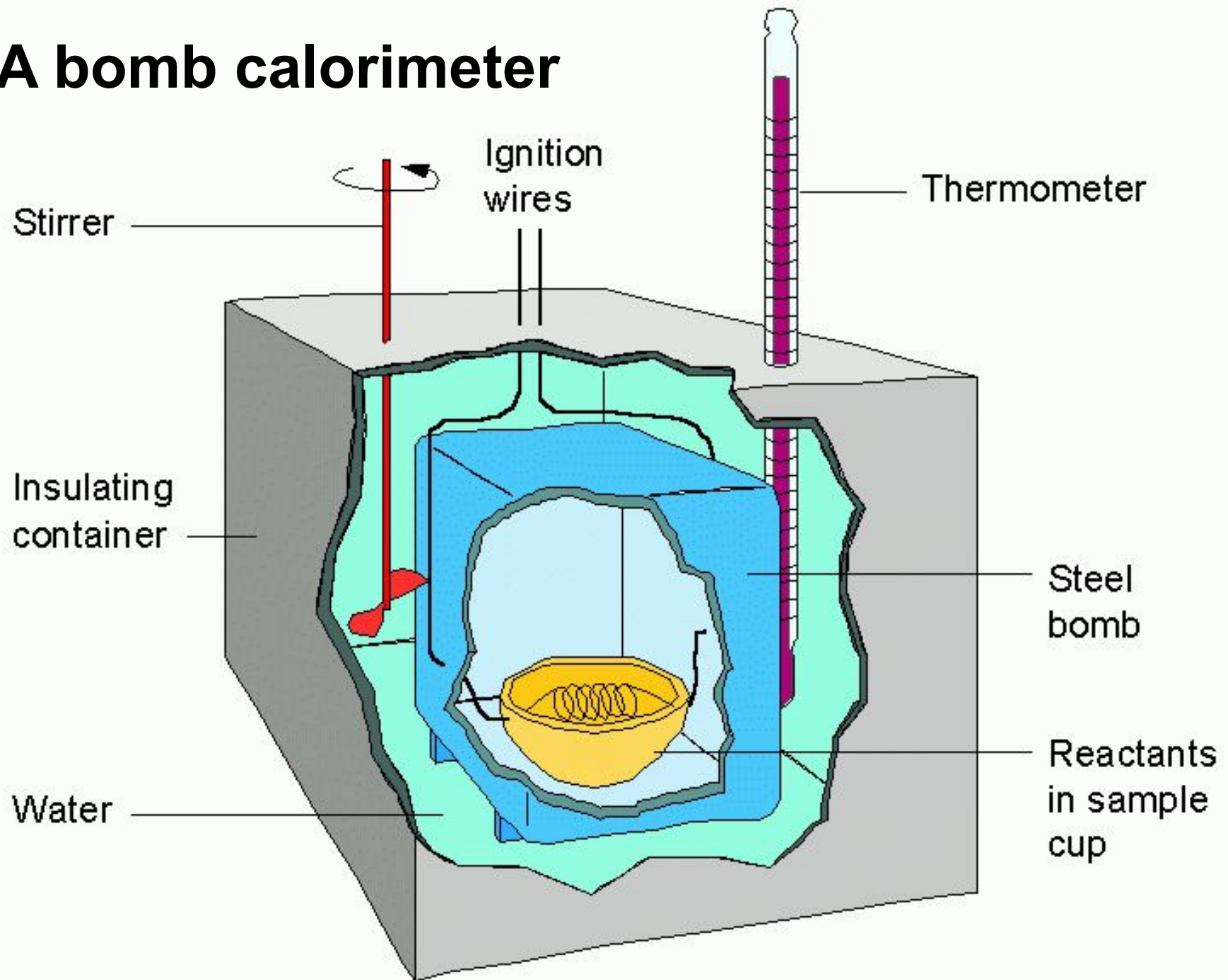
Calorimetry

- ◆ Changes in enthalpy = ΔH
- ◆ $q = \Delta H$ These terms will be used interchangeably in this textbook
- ◆ Thus, $q = \Delta H = m \times C \times \Delta T$
- ◆ ΔH is negative for an exothermic reaction
- ◆ ΔH is positive for an endothermic reaction

Calorimetry

- ◆ Calorimetry experiments can be performed at a constant volume using a device called a “bomb calorimeter” - a closed system
 - Used by nutritionists to measure energy content of food

A bomb calorimeter



SAMPLE PROBLEM 17.2 - Page 513

Enthalpy Change in a Calorimetry Experiment

When 25.0 mL of water containing 0.025 mol HCl at 25.0°C is added to 25.0 mL of water containing 0.025 mol NaOH at 25.0°C in a foam cup calorimeter, a reaction occurs. Calculate the enthalpy change in kJ) during this reaction if the highest temperature observed is 32.0°C. Assume the densities of the solutions are 1.00 g/mL.

1 Analyze List the knowns and the unknown.

Knowns

- $C_{\text{water}} = 4.18 \text{ J}/(\text{g} \cdot ^\circ\text{C})$
- $T_i = 25.0^\circ\text{C}$
- $V_{\text{final}} = V_{\text{HCl}} + V_{\text{NaOH}}$
- $T_f = 32.0^\circ\text{C}$
- $-25.0 \text{ mL} + 25.0 \text{ mL} = 50.0 \text{ mL}$
- $\text{Density}_{\text{solution}} = 1.00 \text{ g/mL}$

Unknown

- $\Delta H = ? \text{ kJ}$

Use dimensional analysis to determine the mass of the water. You must also calculate ΔT . Use $\Delta H = -q_{\text{surr}} = -m \times C \times \Delta T$ to solve for ΔH .

2 Calculate Solve for the unknown.

First, calculate the total mass of the water.

$$m = (50.0 \text{ mL}) \times \left(\frac{1.00 \text{ g}}{\text{mL}} \right) = 50.0 \text{ g}$$

Now calculate ΔT .

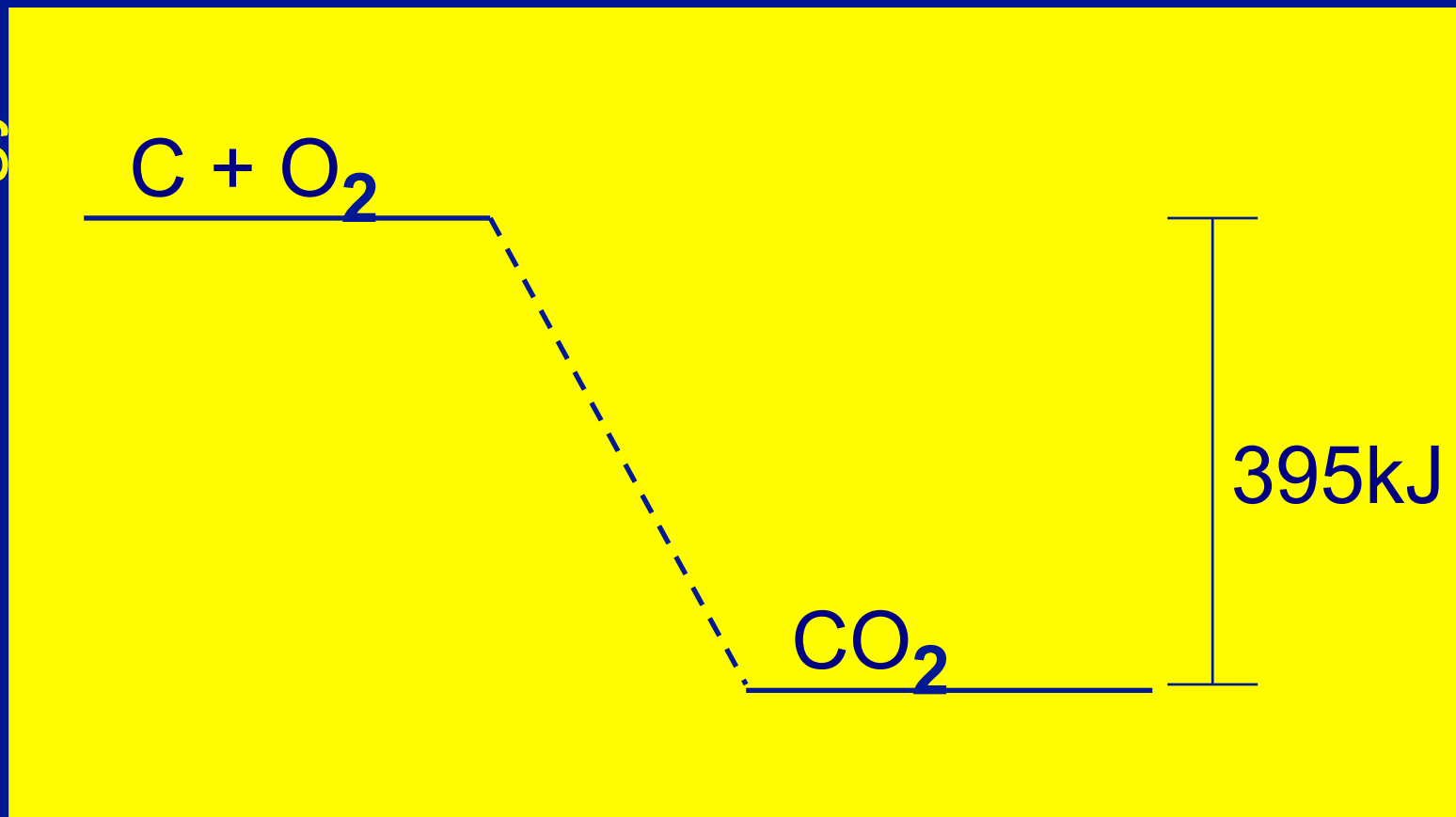
$$\Delta T = T_f - T_i = 32.0^\circ\text{C} - 25.0^\circ\text{C} = 7.0^\circ\text{C}$$

Use the values for m , C_{water} , and ΔT to calculate ΔH .

$$\begin{aligned} \Delta H &= -m \times C \times \Delta T = -(50.0 \text{ g})(4.18 \text{ J}/(\text{g} \times ^\circ\text{C}))(7.0^\circ\text{C}) \\ &= -1463 \text{ J} = -1.5 \times 10^3 \text{ J} = -1.5 \text{ kJ} \end{aligned}$$



Energy



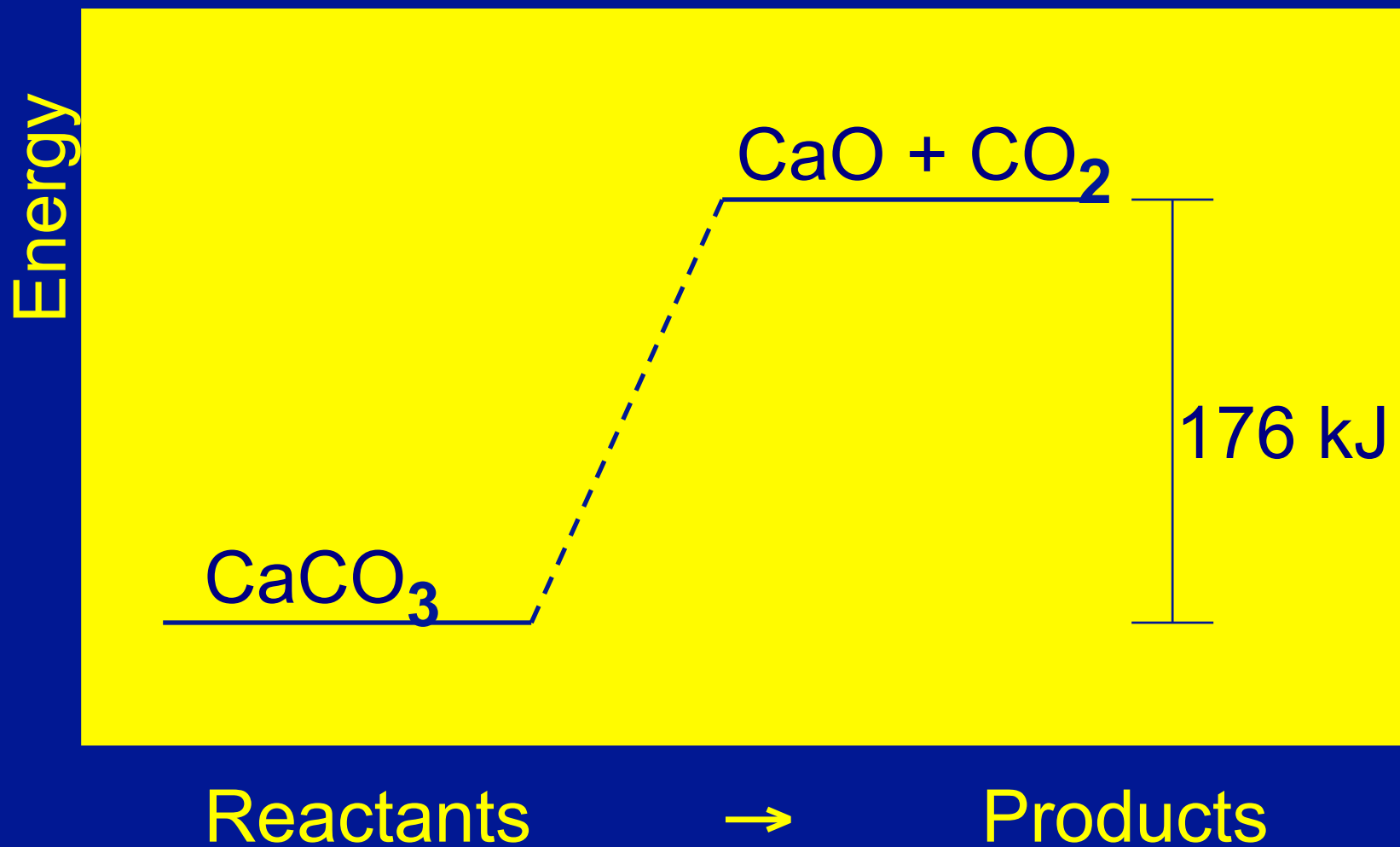
Reactants



Products

Exothermic

- ◆ The products are lower in energy than the reactants
- ◆ Thus, energy is released.
- ◆ $\Delta H = -395 \text{ kJ}$
 - The negative sign does not mean negative energy, but instead that energy is lost.



Endothermic

- ◆ The products are higher in energy than the reactants
- ◆ Thus, energy is absorbed.
- ◆ $\Delta H = +176 \text{ kJ}$
 - The positive sign means energy is absorbed

Chemistry Happens in **MOLES**

- ◆ An equation that includes energy is called a thermochemical equation
- ◆ $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 802.2 \text{ kJ}$
 - 1 mole of CH_4 releases 802.2 kJ of energy.
 - When you make 802.2 kJ you also make 2 moles of water

Thermochemical Equations

- ◆ A heat of reaction is the heat change for the equation, exactly as written
 - The physical state of reactants and products must also be given.
 - Standard conditions for the reaction is 101.3 kPa (1 atm.) and 25 °C (different from STP)



- ◆ If 10.3 grams of CH_4 are burned completely, how much heat will be produced?

$$10.3 \text{ g CH}_4 \left(\frac{1 \text{ mol CH}_4}{16.05 \text{ g CH}_4} \right) \left(\frac{802.2 \text{ kJ}}{1 \text{ mol CH}_4} \right) = 514 \text{ kJ}$$

Thus, $\Delta H = -514 \text{ kJ}$ for this reaction

SAMPLE PROBLEM 17.3 - Page 516

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 $\text{NaHCO}_3(s)$.

1 Analyze List the knowns and the unknown.

Knowns

- 2.24 mol $\text{NaHCO}_3(s)$ decomposes
- $\Delta H = 129 \text{ kJ}$ (for 2 mol NaHCO_3)

Unknown

- $\Delta H = ? \text{ kJ}$

Use the thermochemical equation,



to write a conversion factor relating kilojoules of heat and moles of NaHCO_3 . Then use the conversion factor to determine ΔH for 2.24 mol NaHCO_3 .

2 Calculate Solve for the unknown.

The thermochemical equation indicates that 129 kJ are needed to decompose 2 mol $\text{NaHCO}_3(s)$. Use this relationship to write the following conversion factor.

$$\frac{129 \text{ kJ}}{2 \text{ mol NaHCO}_3(s)}$$

Using dimensional analysis, solve for ΔH .

$$\begin{aligned} \Delta H &= 2.24 \text{ mol NaHCO}_3(s) \times \frac{129 \text{ kJ}}{2 \text{ mol NaHCO}_3(s)} \\ &= 144 \text{ kJ} \end{aligned}$$



Summary, so far...

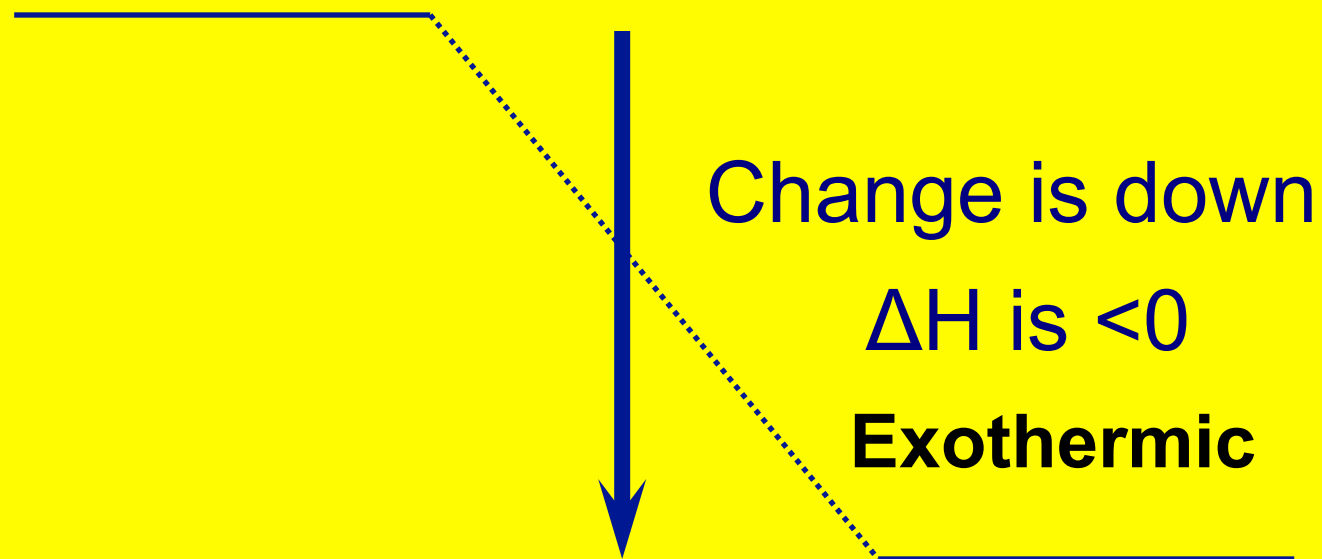
Enthalpy

- ◆ The heat content a substance has at a given temperature and pressure
- ◆ Can't be measured directly because there is no set starting point
- ◆ The reactants *start* with a heat content
- ◆ The products *end up* with a heat content
- ◆ So we can measure how much **enthalpy changes**

Enthalpy

- ◆ Symbol is H
- ◆ Change in enthalpy is ΔH (delta H)
- ◆ If heat is *released*, the heat content of the products is *lower*
 ΔH is negative (exothermic)
- ◆ If heat is *absorbed*, the heat content of the products is *higher*
 ΔH is positive (endothermic)

Energy



Reactants → Products

Energy



Reactants →

Products

Heat of Reaction

- ◆ The heat that is released or absorbed in a chemical reaction
- ◆ Equivalent to ΔH
- ◆ $C + O_2(g) \rightarrow CO_2(g) + 393.5 \text{ kJ}$
- ◆ $C + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -393.5 \text{ kJ}$
- ◆ In thermochemical equation, it is important to indicate the physical state
- ◆ $H_2(g) + 1/2O_2(g) \rightarrow H_2O(g) \quad \Delta H = -241.8 \text{ kJ}$
- ◆ $H_2(g) + 1/2O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8 \text{ kJ}$

Heat of Combustion

- ◆ The heat from the reaction that completely burns 1 mole of a substance:
- ◆ $\text{C} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 393.5 \text{ kJ}$
- ◆ $\text{C} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \quad \Delta\text{H} = -393.5 \text{ kJ}$
- ◆ Note Table 17.2, page 517

Section 17.3

Heat in Changes of State

◆ OBJECTIVES:

- Classify the enthalpy change that occurs when a substance melts, freezes, boils, condenses, or dissolves.

Section 17.3

Heat in Changes of State

◆ OBJECTIVES:

- Solve for the enthalpy change that occurs when a substance melts, freezes, boils, condenses, or dissolves.

Heat in Changes of State

- ◆ Molar Heat of Fusion (ΔH_{fus}) – the heat absorbed by one mole of a substance in *melting* from a solid to a liquid
- ◆ $q = \text{mass} \times \Delta H_{\text{fus}}$ (there is no temperature change)
- ◆ Molar Heat of Solidification (ΔH_{solid}) = heat lost when one mole of liquid *solidifies (or freezes)* to a solid
- ◆ $q = \text{mass} \times \Delta H_{\text{solid}}$ (no temperature change)

Heat in Changes of State

- ◆ Heat absorbed by a melting solid is equal to heat lost when a liquid solidifies
 - Thus, $\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$
- ◆ Note Table 17.3, page 522

SAMPLE PROBLEM 17.4 - Page 521

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?

1 Analyze List the knowns and the unknown.

Knowns

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

Unknown

- $m_{\text{ice}} = ? \text{ g}$

Use the thermochemical equation



to find the number of moles of ice that can be melted by the addition of 2.25 kJ of heat. Convert moles of ice to grams of ice.

2 Calculate Solve for the unknown.

Express ΔH_{fus} and the molar mass of ice as conversion factors.

$$\frac{1 \text{ mol ice}}{6.01 \text{ kJ}} \quad \text{and} \quad \frac{18.0 \text{ g ice}}{1 \text{ mol ice}}$$

Multiply the known enthalpy change (2.25 kJ) by the conversion factors

$$\begin{aligned} m_{\text{ice}} &= 2.25 \text{ kJ} \times \frac{1 \text{ mol ice}}{6.01 \text{ kJ}} \times \frac{18.0 \text{ g ice}}{1 \text{ mol ice}} \\ &= 6.74 \text{ g ice} \end{aligned}$$

Heats of Vaporization and Condensation

- ◆ When liquids absorb heat at their boiling points, they become vapors.
- ◆ Molar Heat of Vaporization (ΔH_{vap}) = the amount of heat necessary to vaporize one mole of a given liquid.
- ◆ $q = \text{mass} \times \Delta H_{\text{vap}}$ (no temperature change)
- ◆ Table 17.3, page 522

Heats of Vaporization and Condensation

- ◆ Condensation is the opposite of vaporization.
- ◆ Molar Heat of Condensation (ΔH_{cond}) = amount of heat released when one mole of vapor condenses to a liquid
- ◆ $\Delta H_{\text{vap}} = - \Delta H_{\text{cond}}$

Heats of Vaporization and Condensation

- ◆ Note Figure 17.10, page 523
- ◆ The large values for water ΔH_{vap} and ΔH_{cond} are the reason hot vapors such as steam is very dangerous
 - You can receive a scalding burn from steam when the heat of condensation is released!



SAMPLE PROBLEM 17.5 -Page 524

Using the Heat of Vaporization in Phase-Change Calculations

How much heat (in kJ) is absorbed when 24.8 g $\text{H}_2\text{O}(l)$ at 100°C and 101.3 kPa is converted to steam at 100°C ?

1 Analyze List the knowns and the unknown.

Knowns

- Initial and final conditions are 100°C and 101.3 kPa
- mass of water converted to steam = 24.8 g
- $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$

Unknown

- $\Delta H = ? \text{ kJ}$

Refer to the following thermochemical equation.



ΔH_{vap} is given in kJ/mol, but the quantity of water is given in grams. You must first convert grams of water to moles of water. Then multiply by ΔH_{vap} .

2 Calculate Solve for the unknown.

The required conversion factors come from ΔH_{vap} and the molar mass of water.

$$\frac{1 \text{ mol H}_2\text{O}(l)}{18.0 \text{ g H}_2\text{O}(l)} \quad \text{and} \quad \frac{40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}(l)}$$

Multiply the mass of water in grams by the conversion factors.

$$\begin{aligned} \Delta H &= 24.8 \text{ g H}_2\text{O}(l) \times \frac{1 \text{ mol H}_2\text{O}(l)}{18.0 \text{ g H}_2\text{O}(l)} \times \frac{40.7 \text{ kJ}}{1 \text{ mol H}_2\text{O}(l)} \\ &= 56.1 \text{ kJ} \end{aligned}$$

Heat of Solution

- ◆ Heat changes can also occur when a solute *dissolves* in a solvent.
- ◆ Molar Heat of Solution (ΔH_{soln}) = heat change caused by dissolution of one mole of substance
- ◆ Sodium hydroxide provides a good example of an exothermic molar heat of solution:

Heat of Solution



$$\Delta H_{\text{soln}} = -445.1 \text{ kJ/mol}$$

- ◆ The heat is released as the ions separate (by dissolving) and interact with water, releasing 445.1 kJ of heat as ΔH_{soln} -thus becoming so hot it steams!

SAMPLE PROBLEM 17.6 - Page 526

Calculating the Heat Change in Solution Formation

How much heat (in kJ) is released when 2.500 mol NaOH(*s*) is dissolved in water?

1 Analyze List the knowns and the unknown.

Knowns

- $\Delta H_{\text{soln}} = -445.1 \text{ kJ/mol}$
- amount of NaOH(*s*) dissolved = 2.500 mol

Unknown

- $\Delta H = ? \text{ kJ}$

Use the heat of solution from the following chemical equation to solve for the amount of heat released (ΔH).



2 Calculate Solve for the unknown.

Multiply the number of moles of NaOH by ΔH_{soln} .

$$\Delta H = 2.500 \text{ mol NaOH}(s) \times \frac{-445.1 \text{ kJ}}{1 \text{ mol NaOH}(aq)} = -1113 \text{ kJ}$$

Section 17.4

Calculating Heats of Reaction

◆ OBJECTIVES:

- State Hess' s law of heat summation and describe how it is used in chemistry.

Section 17.4

Calculating Heats of Reaction

◆ OBJECTIVES:

- Solve for enthalpy changes by using Hess's law or standard heats of formation.

Hess' s Law



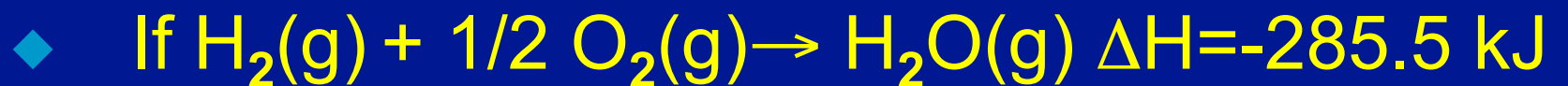
- ◆ If you add two or more thermochemical equations to give a final equation, then you can also add the heats of reaction to give the final heat of reaction.

Called Hess' s law of heat summation

- ◆ Example shown on page 528 for graphite and diamonds

How Does It Work?

1) If you turn an equation around, you change the sign:

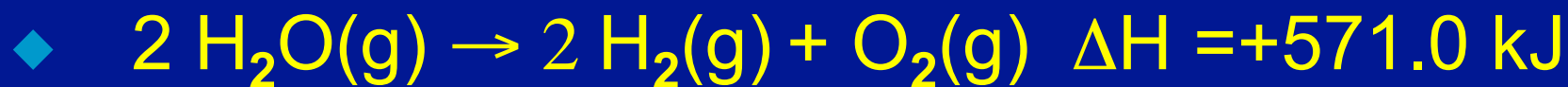


◆ then,



◆ also,

2) If you multiply the equation by a number, you multiply the heat by that number:



3) Or, you can just leave the equation “as is”

Hess' s Law - Procedure Options:

- ◆ 1. Use the equation as written
- ◆ 2. Reverse the equation (and change heat sign + to -, etc.)
- ◆ 3. Increase the coefficients in the equation (and increase heat by same amount)
- ◆ Note sample from page 529

Standard Heats of Formation

- ◆ The ΔH for a reaction that produces 1 mol of a compound from its elements at standard conditions
- ◆ Standard conditions: 25°C and 1 atm.
- ◆ Symbol is ΔH_f^0
 - ◆ The standard heat of formation of an element in its standard state is arbitrarily set at “0”
- ◆ This includes the diatomics

Standard Heats of Formation

- ◆ Table 17.4, page 530 has standard heats of formation
- ◆ The heat of a reaction can be calculated by:
 - subtracting the heats of formation of the reactants from the products

$$\Delta H^{\circ} = (\Delta H_{\text{f}}^{\circ} \text{ Products}) - (\Delta H_{\text{f}}^{\circ} \text{ Reactants})$$

SAMPLE PROBLEM 17.7 - Page 531

Calculating the Standard Heat of Reaction

What is the standard heat of reaction (ΔH°) for the reaction of $\text{CO}(g)$ with $\text{O}_2(g)$ to form $\text{CO}_2(g)$?

1 Analyze List the knowns and the unknown.

Knowns

(from Table 17.4)

- $\Delta H_f^\circ \text{O}_2(g) = 0 \text{ kJ/mol}$ (free element)
- $\Delta H_f^\circ \text{CO}(g) = -110.5 \text{ kJ/mol}$
- $\Delta H_f^\circ \text{CO}_2(g) = -393.5 \text{ kJ/mol}$

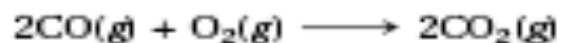
Unknown

- $\Delta H^\circ = ? \text{ kJ}$

Balance the equation of the reaction of $\text{CO}(g)$ with $\text{O}_2(g)$ to form $\text{CO}_2(g)$. Then determine ΔH° using the standard heats of formation of the reactants and products.

2 Calculate Solve for the unknown.

First, write the balanced equation.



Next, find and add the ΔH_f° of all of the reactants, taking into account the number of moles of each.

$$\begin{aligned}\Delta H_f^\circ(\text{reactants}) &= 2 \text{ mol CO}(g) \times \frac{-110.5 \text{ kJ}}{1 \text{ mol CO}(g)} + 0 \text{ kJ} \\ &= -221.0 \text{ kJ}\end{aligned}$$

Then, find the ΔH_f° of the product in a similar way:

$$\begin{aligned}\Delta H_f^\circ(\text{product}) &= 2 \text{ mol CO}_2(g) \times \frac{-393.5 \text{ kJ}}{1 \text{ mol CO}_2(g)} \\ &= -787.0 \text{ kJ}\end{aligned}$$

Finally, solve for the unknown

$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactants}) \\ \Delta H^\circ &= (-787.0 \text{ kJ}) - (-221.0 \text{ kJ}) \\ \Delta H^\circ &= -566.0 \text{ kJ}\end{aligned}$$

Another Example



$$\Delta H_f^0 \text{ CH}_4(\text{g}) = -74.86 \text{ kJ/mol}$$

$$\Delta H_f^0 \text{ O}_2(\text{g}) = 0 \text{ kJ/mol}$$

$$\Delta H_f^0 \text{ CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^0 \text{ H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$$

$$\blacklozenge \Delta H = [-393.5 + 2(-241.8)] - [-74.86 + 2(0)]$$

$$\Delta H = -802.4 \text{ kJ}$$

End of Chapter 17