Chapter 17
“Thermochemistry”

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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 capacity* *(also called simply 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

- u Gasoline contains a significant amount of chemical potential energy
- u 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:
  - **release** of heat, or
  - **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*** (gets warmer) 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*** (gets cooler) 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* (q is positive)
- On the right, the system (the body) cools as perspiration evaporates, and heat flows to the surroundings
  - this is *exothermic* (q is negative)

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

- 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

- The calorie is also related to the **Joule**, the SI unit of heat and energy
  - named after James Prescott Joule
  - 4.184 J = 1 cal
- **Heat Capacity** - the amount of heat needed to increase the temperature of an object exactly 1 °C
  - Depends on both the object’s mass and its chemical composition

Heat Capacity and Specific Heat

- **Specific Heat Capacity** (abbreviated "C") - the amount of heat it takes to raise the temperature of *1 gram* of the substance by 1 °C
  - often called simply "Specific Heat"
  - Note Table 17.1, page 508 (next slide)
- Water has a **HUGE** value, when it is compared to other chemicals

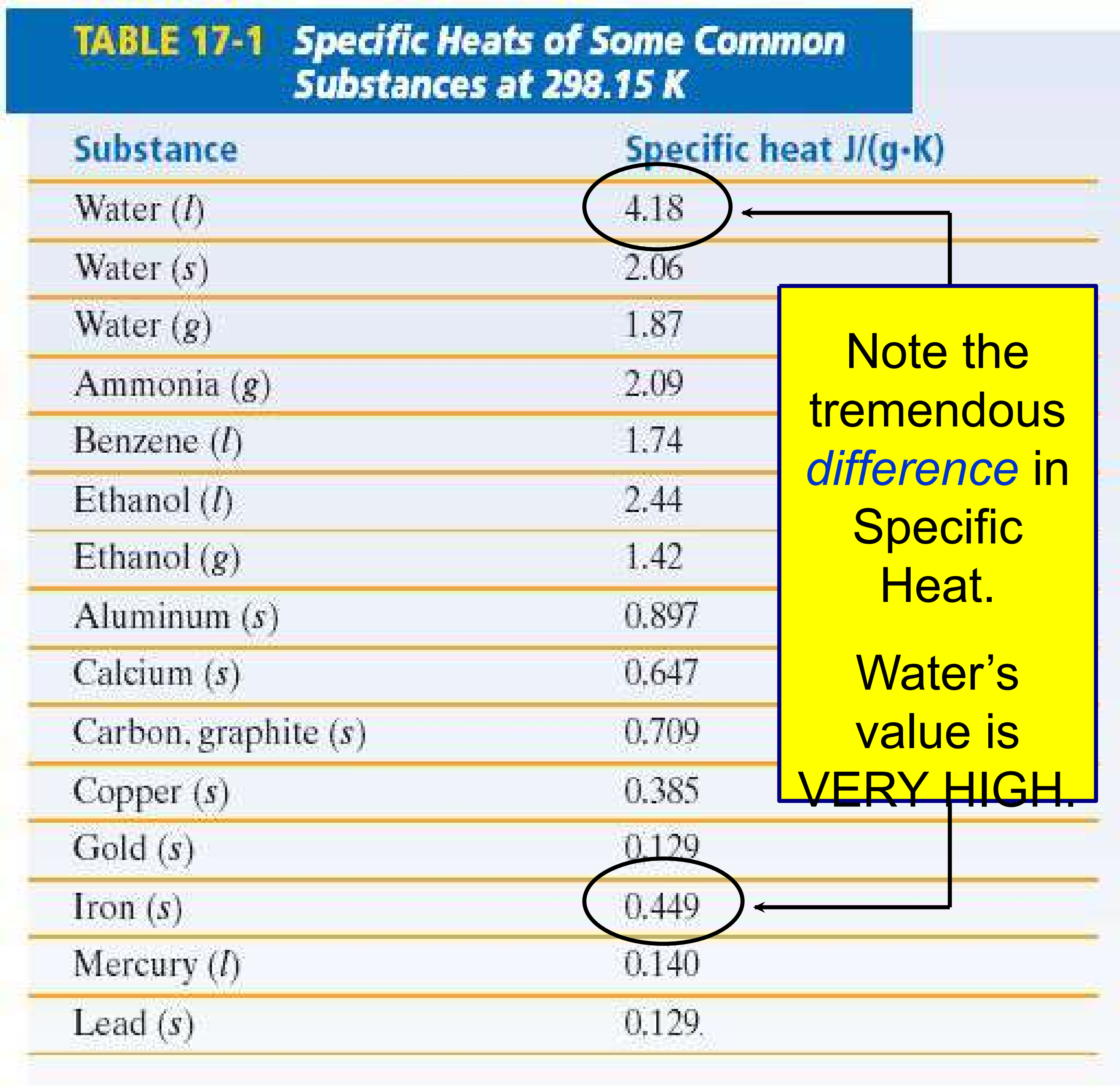


TABLE 17-1 ***Specific Heats of Some Common Substances at 298.15 K***

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

Note the tremendous *difference* in Specific Heat.

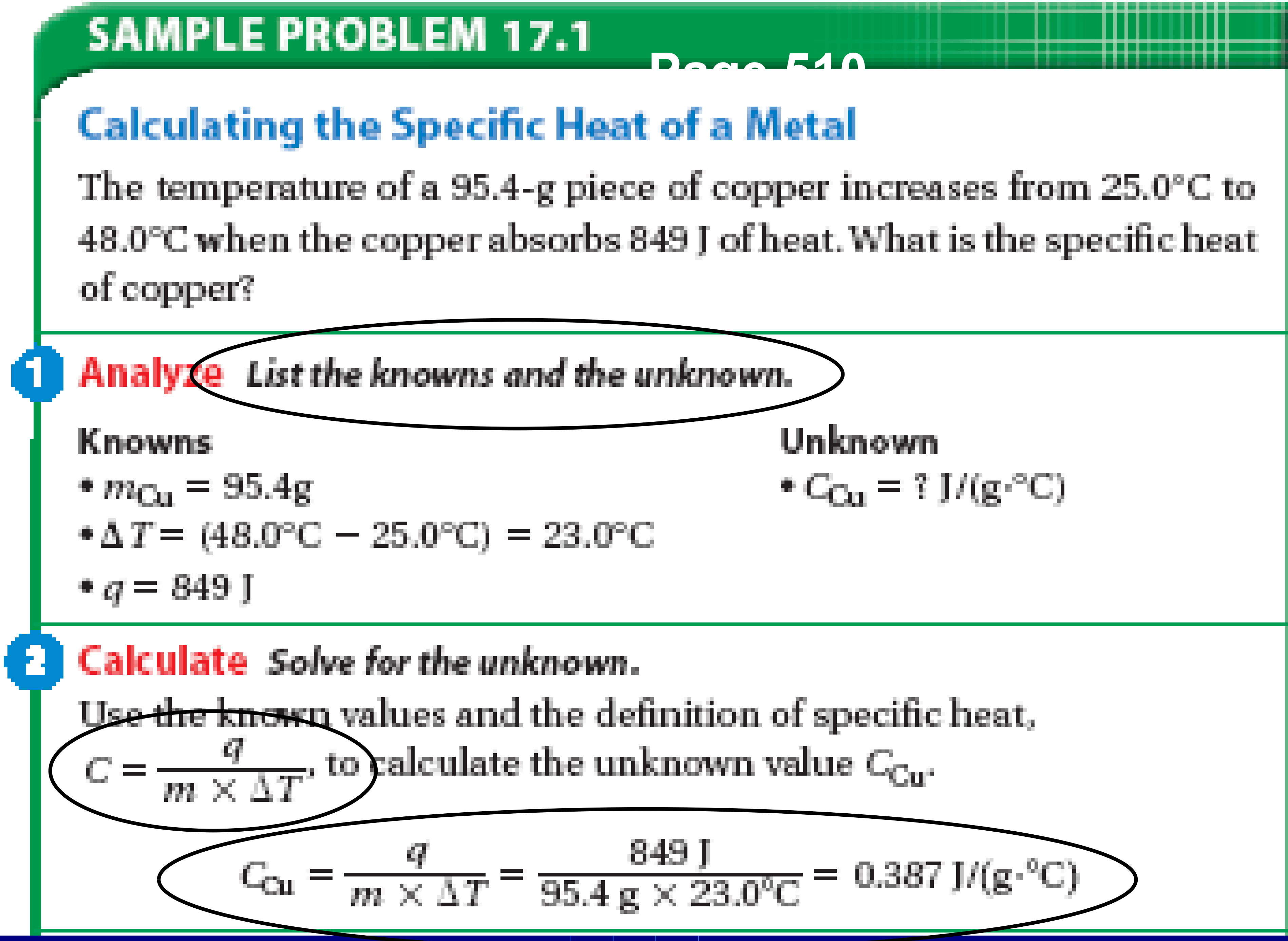
Water's value is VERY HIGH.

Heat Capacity and Specific Heat

- For water, C = 4.18 J/(g °C) in Joules, and C = 1.00 cal/(g °C) in calories.
- 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 = mass (in grams) x ΔT x C
- heat is abbreviated as “q”
- ΔT = change in temperature
- C = Specific Heat
  - Units are either: J/(g °C) or cal/(g °C)



SAMPLE PROBLEM 17.1

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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	**Unknown**
$m_{Cu} = 95.4\,g$	$C_{Cu} = ?\ J/(g \cdot °C)$
$\Delta T = (48.0°C - 25.0°C) = 23.0°C$	
$q = 849\ J$	

2 Calculate *Solve for the unknown.*

Use the known values and the definition of specific heat, $C = \dfrac{q}{m \times \Delta T}$, to calculate the unknown value C_{Cu}.

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

Section 17.2
Measuring and Expressing Enthalpy Changes

- **OBJECTIVES:**
  - **Describe** how *calorimeters* are used to measure heat flow.

Section 17.2

Measuring and Expressing Enthalpy Changes

- **OBJECTIVES:**
  - **Construct** thermochemical *equations*.

Section 17.2
Measuring and Expressing Enthalpy Changes

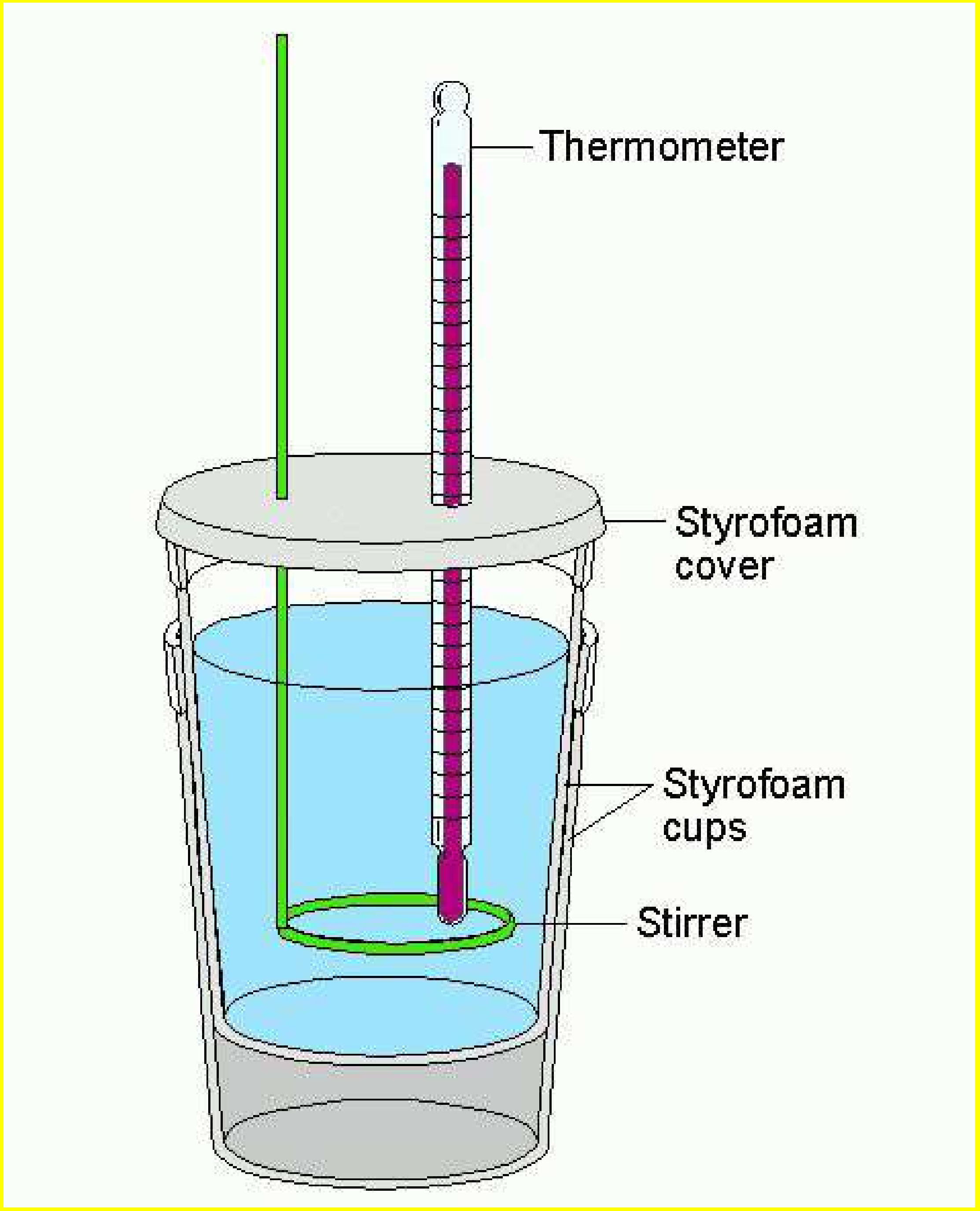
- **OBJECTIVES:**
  - **Solve** for enthalpy changes in chemical reactions by using *heats of reaction*.

Calorimetry

- Calorimetry - the measurement of the heat *into* or *out of* a system for chemical and physical processes.
  - Based on the fact that the *heat released = the heat absorbed*
- 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 under *constant pressure*. (They are good because they are well-insulated.)
  - Fig. 17.5, page 511
  - What about a Dewar’s flask?
- 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

u Changes in enthalpy = ΔH

u $q = \Delta H$ These terms will be used interchangeably in this textbook

u Thus, $q = \Delta H = m \times C \times \Delta T$

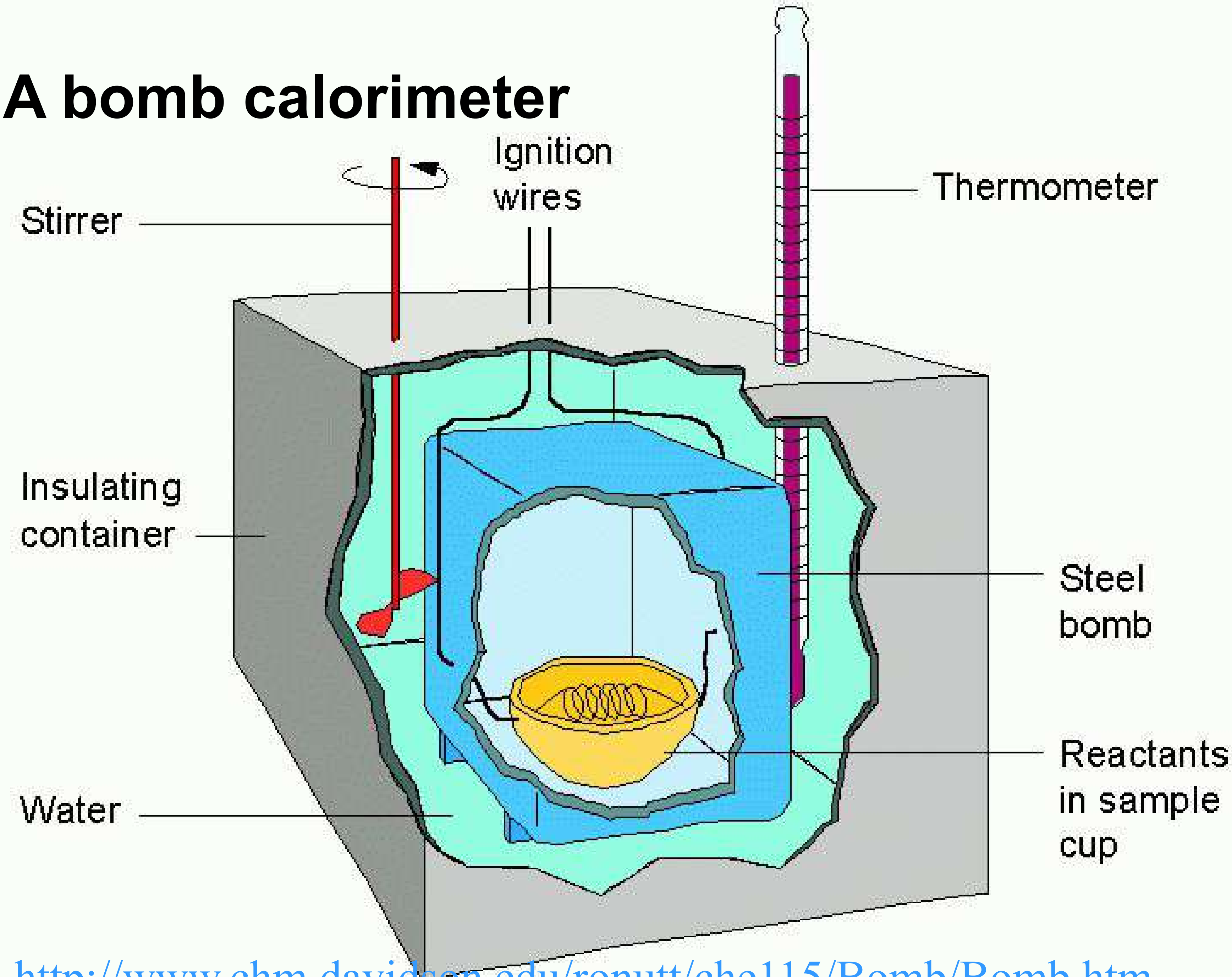
4 ΔH is **negative** for an exothermic reaction

4 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



http://www.chm.davidson.edu/ronutt/che115/Bomb/Bomb.htm

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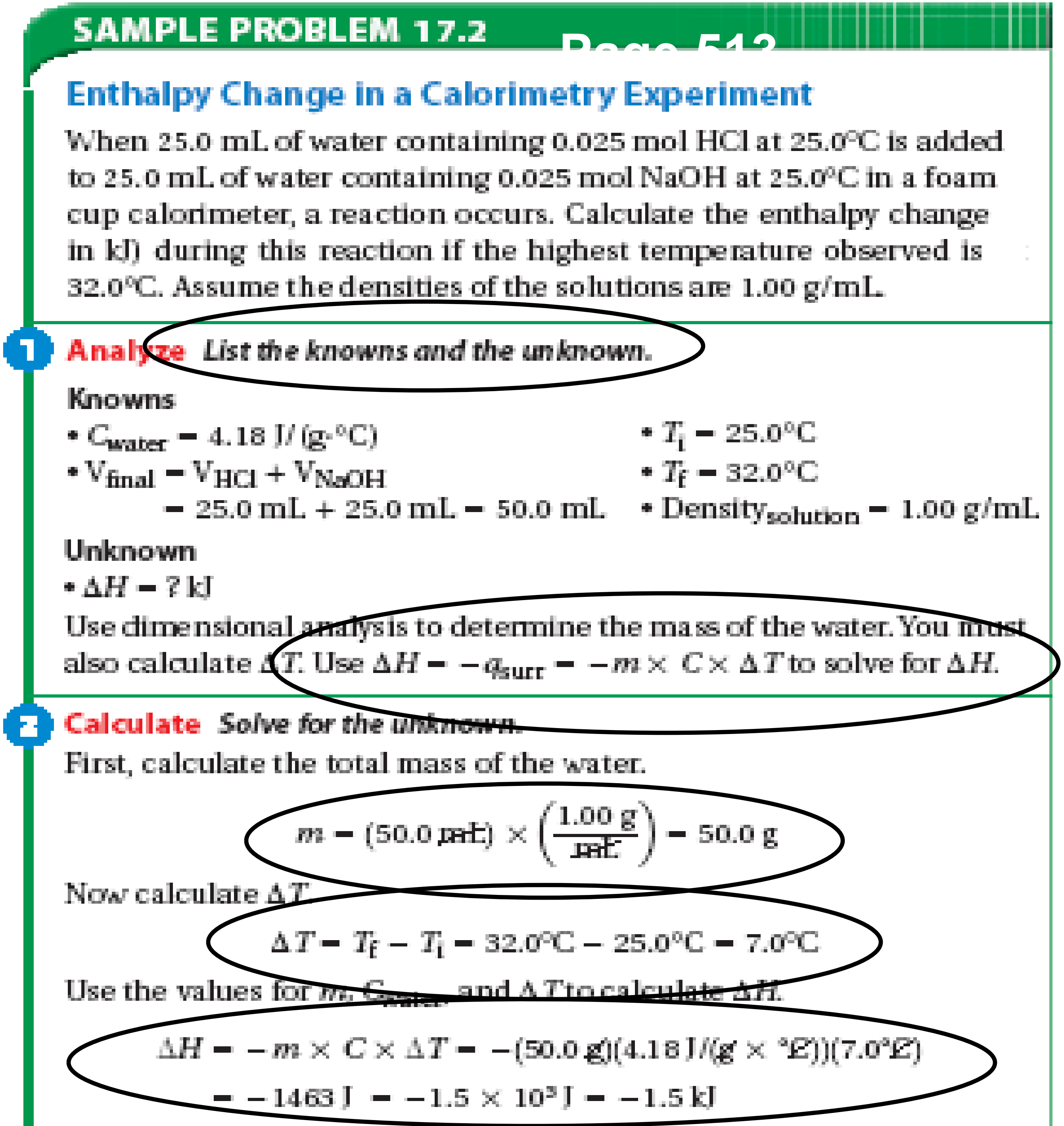


SAMPLE PROBLEM 17.2

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

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

Unknown

- $\Delta H = ?\ kJ$

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

2 **Calculate** ***Solve for the unknown.***

First, calculate the total mass of the water.

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

Now calculate ΔT.

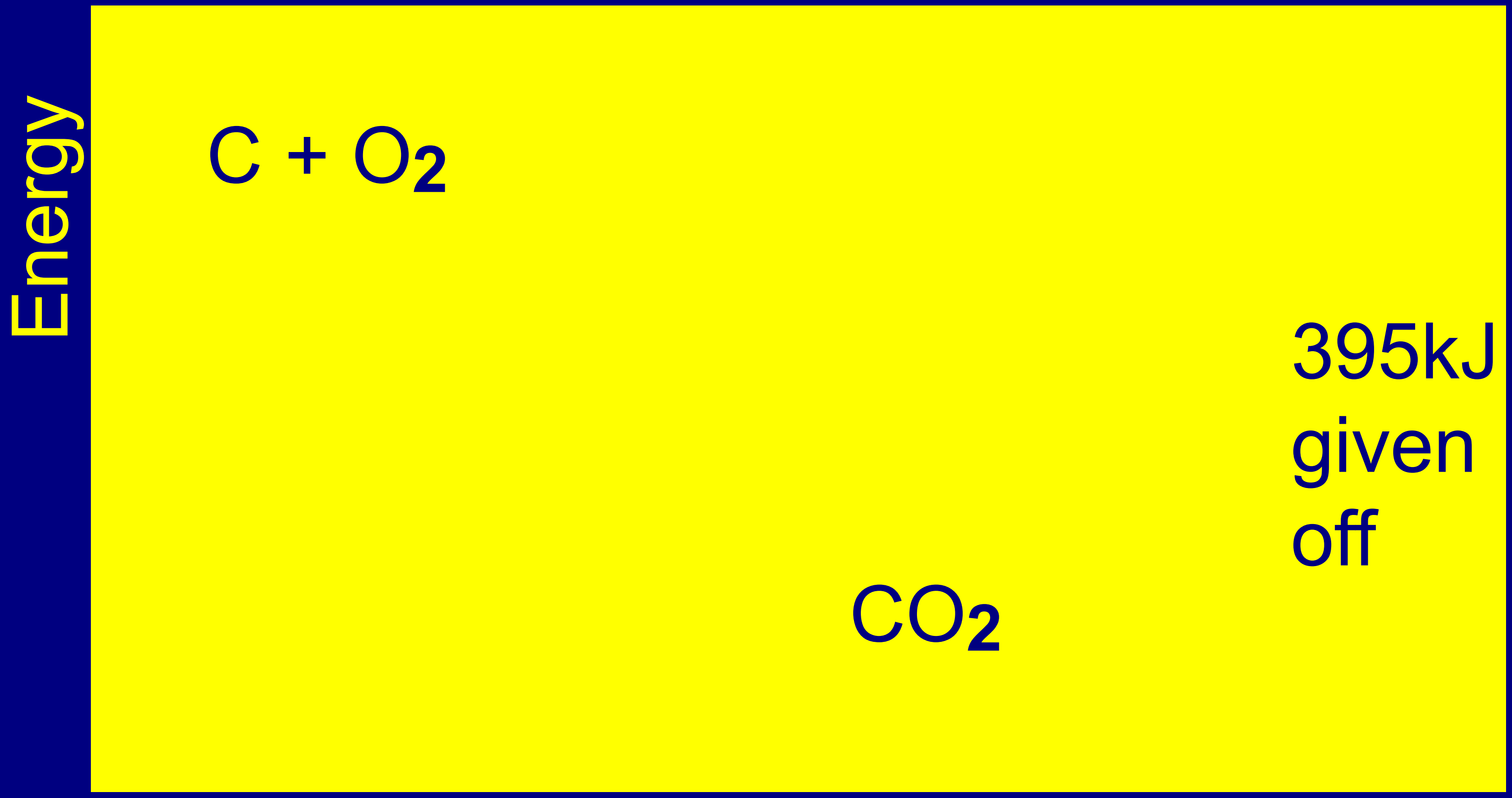
$$\Delta T = T_f - T_i = 32.0^{\circ}C - 25.0^{\circ}C = 7.0^{\circ}C$$

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

$$\Delta H = -m \times C \times \Delta T = -(50.0\ \cancel{g})(4.18\ J/(\cancel{g} \times \cancel{{}^{\circ}C}))(7.0\cancel{{}^{\circ}C})$$

$$= -1463\ J = -1.5 \times 10^3\ J = -1.5\ kJ$$

$C + O_2 \rightarrow CO_2 + 395\ kJ$



Energy

$C + O_2$

395kJ
given
off

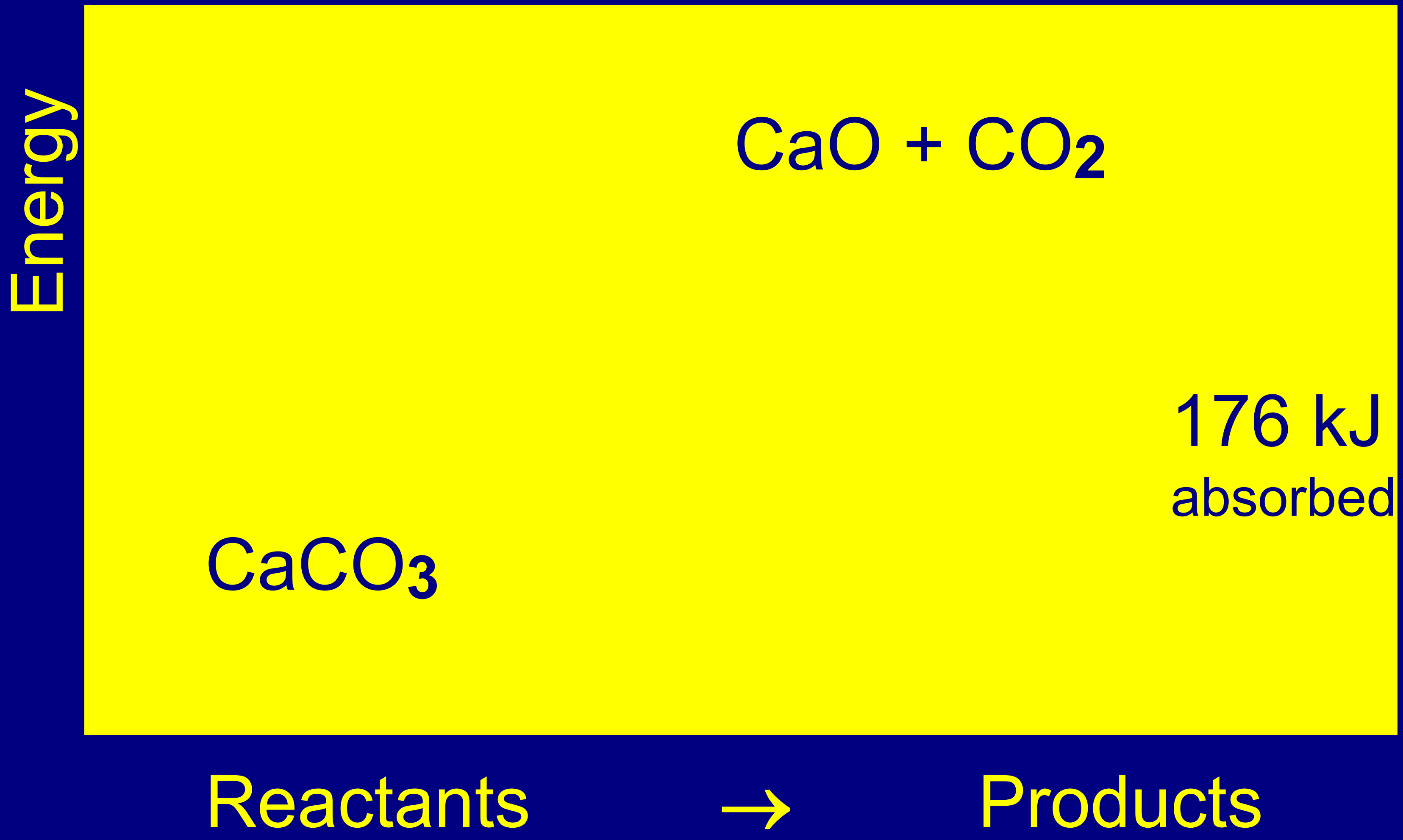
CO_2

Reactants → Products

Exothermic

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

$CaCO_3$ + 176 kJ → $CaO + CO_2$



$CaO + CO_2$

176 kJ
absorbed

$CaCO_3$

Energy

Reactants → Products

Endothermic

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

Chemistry Happens in

MOLES

- An equation that *includes energy* is called a **thermochemical equation**
- $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 802.2\ kJ$
  - 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

- The 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* (SC) for the reaction is 101.3 kPa (1 atm.) and 25 °C (different from STP)

$$1\ CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)} + 802.2\ kJ$$

- If 10. 3 grams of CH_4 are burned completely, how much heat will be produced?

Start with known value

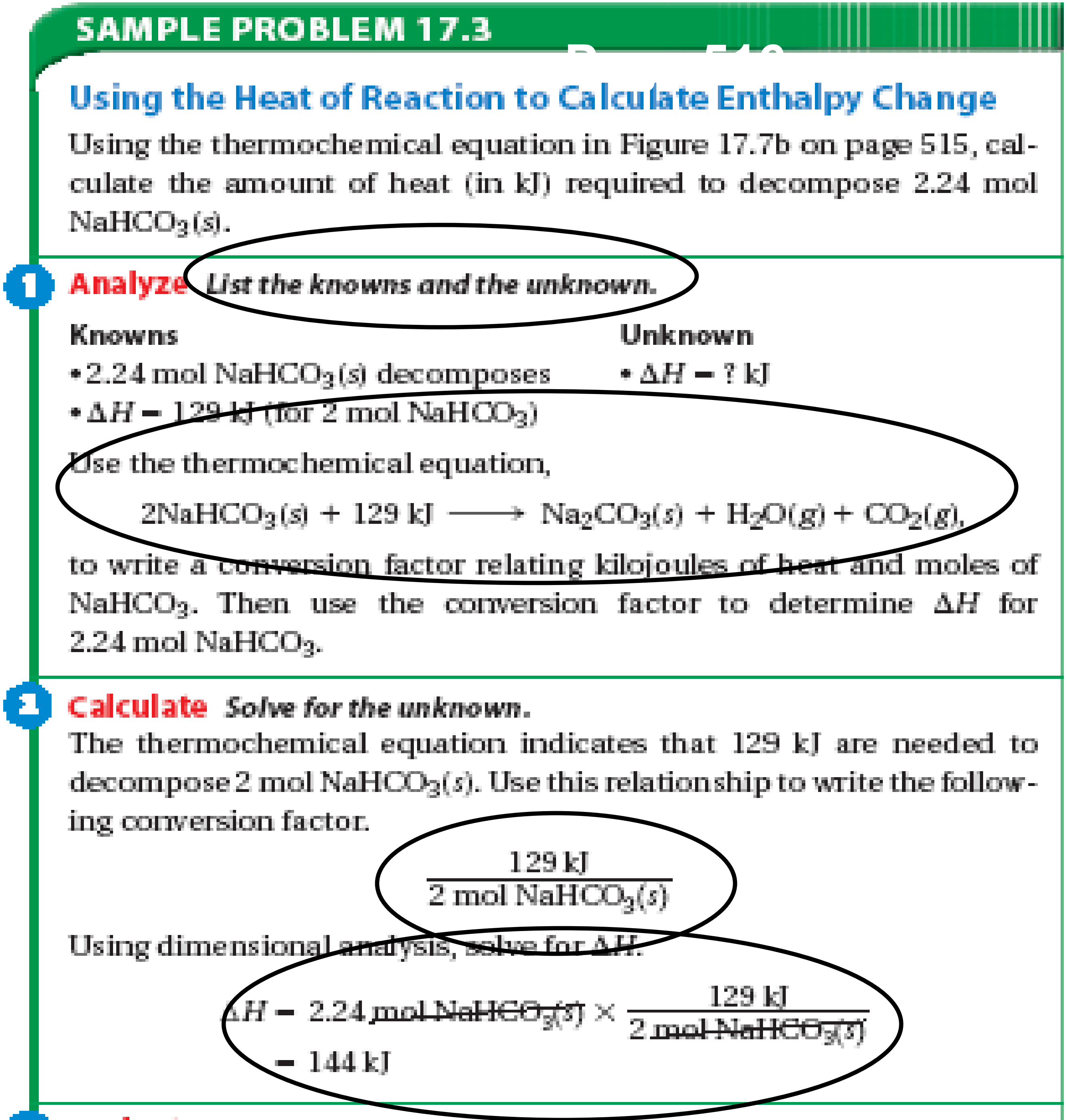
Convert to moles

Convert moles to desired unit

$$10.\ 3\ g\ CH_4 \quad \frac{1\ mol\ CH_4}{16.05\ g\ CH_4} \quad \frac{802.2\ kJ}{1\ mol\ CH_4} = 514\ kJ$$

Ratio from balanced equation

ΔH = -514 kJ, which means the heat is *released* for the reaction of 10.3 grams CH_4



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 $NaHCO_3(s)$.

1 **Analyze** *List the knowns and the unknown.*

Knowns	**Unknown**
• 2.24 mol $NaHCO_3(s)$ decomposes	• $\Delta H = ?$ kJ
• $\Delta H = 129$ kJ (for 2 mol $NaHCO_3$)	

Use the thermochemical equation,

$$2NaHCO_3(s) + 129\text{ kJ} \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g),$$

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 $NaHCO_3(s)$. 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.

$$\Delta H = 2.24\ \cancel{\text{mol }NaHCO_3(s)} \times \frac{129\text{ kJ}}{2\ \cancel{\text{mol }NaHCO_3(s)}}$$
$$= 144\text{ kJ}$$

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

Change is down

ΔH is <0

= Exothermic

(heat is given off)

Reactants → Products

Energy

Change is up

ΔH is > 0

= Endothermic

(heat is absorbed)

Reactants → Products

Heat of Reaction

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

Heat of Combustion

- The heat from the reaction that completely burns **1 mole** of a substance:
- $C + O_2(g) \rightarrow CO_2(g) + 393.5\ kJ$
- $C + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$
- Note Table 17.2, page 517
- DVD: The Thermite Reaction

Section 17.3

Heat in Changes of State

- **OBJECTIVES:**
  - **Classify** the *enthalpy change* that occurs when a substance: a) melts, b) freezes, c) boils, d) condenses, or e) dissolves.

Section 17.3
Heat in Changes of State

- **OBJECTIVES:**
  - **Solve** for the *enthalpy change* that occurs when a substance: a) melts, b) freezes, c) boils, d) condenses, or e) dissolves.

Heat in Changes of State

1. Molar Heat of Fusion ($\Delta H_{fus.}$) = the heat absorbed by one mole of a substance in *melting* from a solid to a liquid

$q = mol \times \Delta H_{fus.}$ (**no** temperature change)

Values given in Table 17.3, page 522

2. Molar Heat of Solidification ($\Delta H_{solid.}$) = the heat lost when one mole of liquid *solidifies (or freezes)* to a solid

$q = mol \times \Delta H_{solid.}$ (**no** temperature change)

Heat in Changes of State

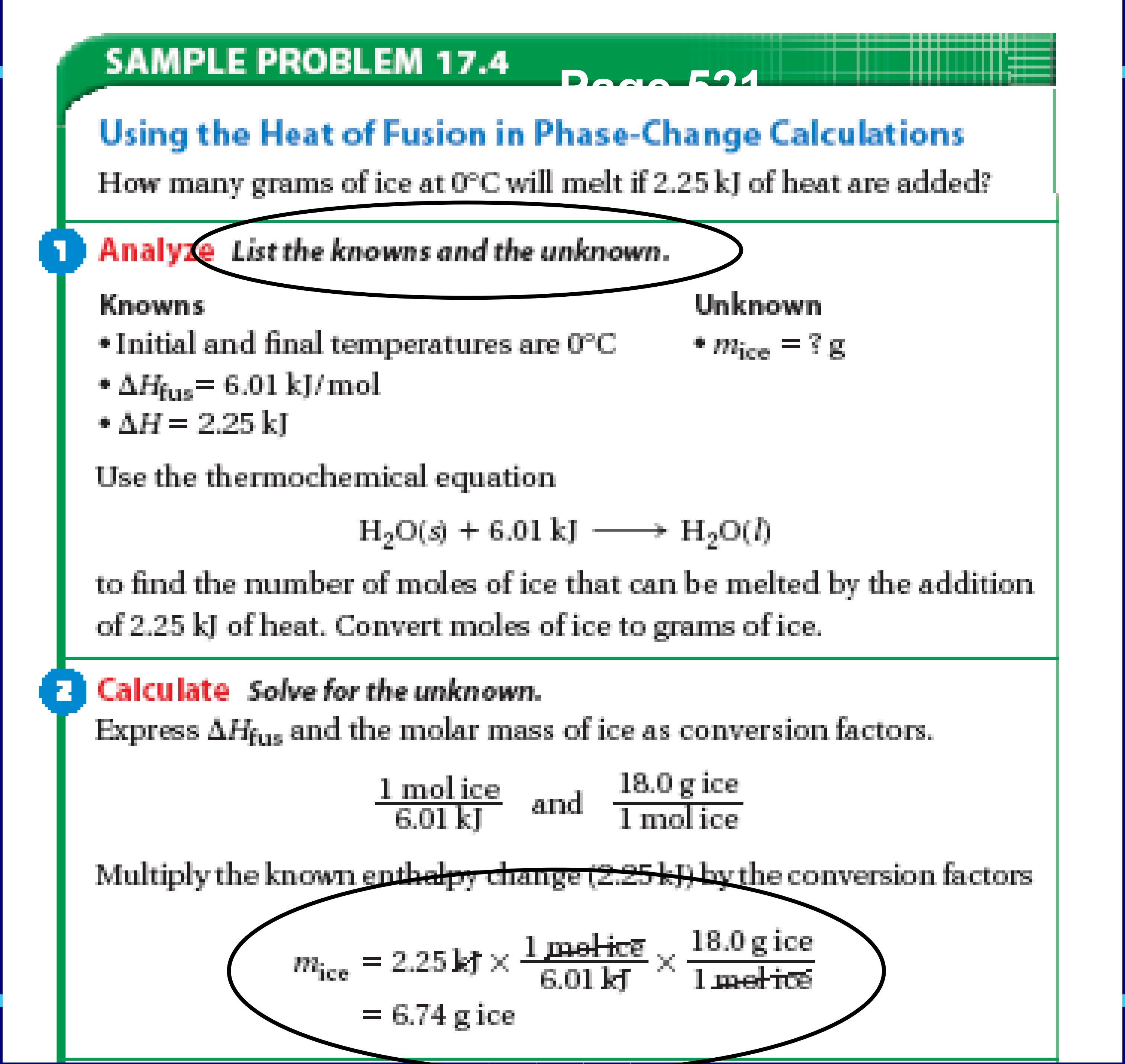
- Note: You may also have the value of these equations as:
  $q = \text{mass x } \Delta H$

This is because some textbooks give the value of ΔH as kJ/*gram*, instead of kJ/*mol*

Heat in Changes of State

- Heat absorbed by a melting solid ***is equal*** to heat lost when a liquid solidifies
  - Thus, $\Delta H_{fus.} = -\Delta H_{solid.}$
- Note Table 17.3, page 522 for the number values. Why is there ***no value*** listed for the molar heat of solidification?

Page 521



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?

1 **Analyze** *List the knowns and the unknown.*

Knowns	**Unknown**
• Initial and final temperatures are 0°C	• m_{ice} = ? g
• ΔH_{fus} = 6.01 kJ/mol	
• ΔH = 2.25 kJ	

Use the thermochemical equation

$$H_2O(s) + 6.01\ kJ \longrightarrow H_2O(l)$$

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.

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

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

$$m_{ice} = 2.25\ \cancel{kJ} \times \frac{1\ \cancel{mol\ ice}}{6.01\ \cancel{kJ}} \times \frac{18.0\ g\ ice}{1\ \cancel{mol\ ice}}$$
$$= 6.74\ g\ ice$$

Heats of Vaporization and Condensation

- When liquids absorb heat at their boiling points, they become **vapors**.

3. Molar Heat of Vaporization ($\Delta H_{vap.}$) = the amount of heat necessary to vaporize one mole of a given liquid.

$q = mol \times \Delta H_{vap.}$ (**no** temperature change)

- Table 17.3, page 522

Heats of Vaporization and Condensation

u Condensation is the opposite of vaporization.

4. Molar Heat of Condensation ($\Delta H_{cond.}$) = amount of heat released when one mole of vapor condenses to a liquid

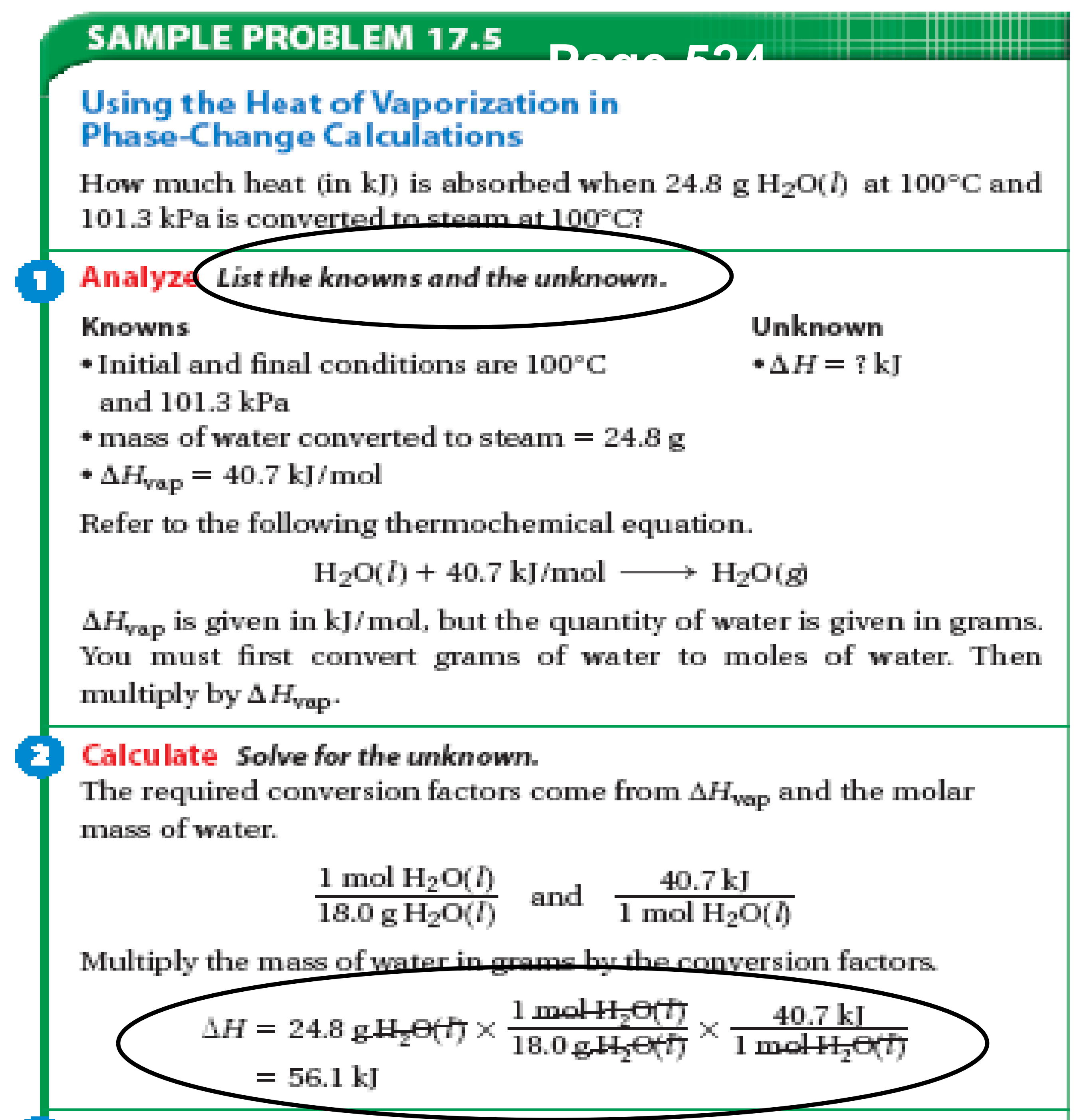
$q = mol \times \Delta H_{cond.}$ (**no** temperature change)

u $\Delta H_{vap.} = - \Delta H_{cond.}$

Heats of Vaporization and Condensation

- Lets look at Table 17.3, page 522…
- The **large values** for water $\Delta H_{vap.}$ and $\Delta H_{cond.}$ is the reason hot vapors such as steam are very dangerous!
  - You can receive a scalding burn from steam when the heat of condensation is released!

$H_20_{(g)} \rightarrow H_20_{(l)} \quad \Delta H_{cond.} = -40.7 kJ/mol$



SAMPLE PROBLEM 17.5

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Using the Heat of Vaporization in Phase-Change Calculations

How much heat (in kJ) is absorbed when 24.8 g $H_2O(l)$ 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
- $\Delta H_{vap} = 40.7\ kJ/mol$

Unknown

- $\Delta H = ?\ kJ$

Refer to the following thermochemical equation.

$$H_2O(l) + 40.7\ kJ/mol \longrightarrow H_2O(g)$$

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.

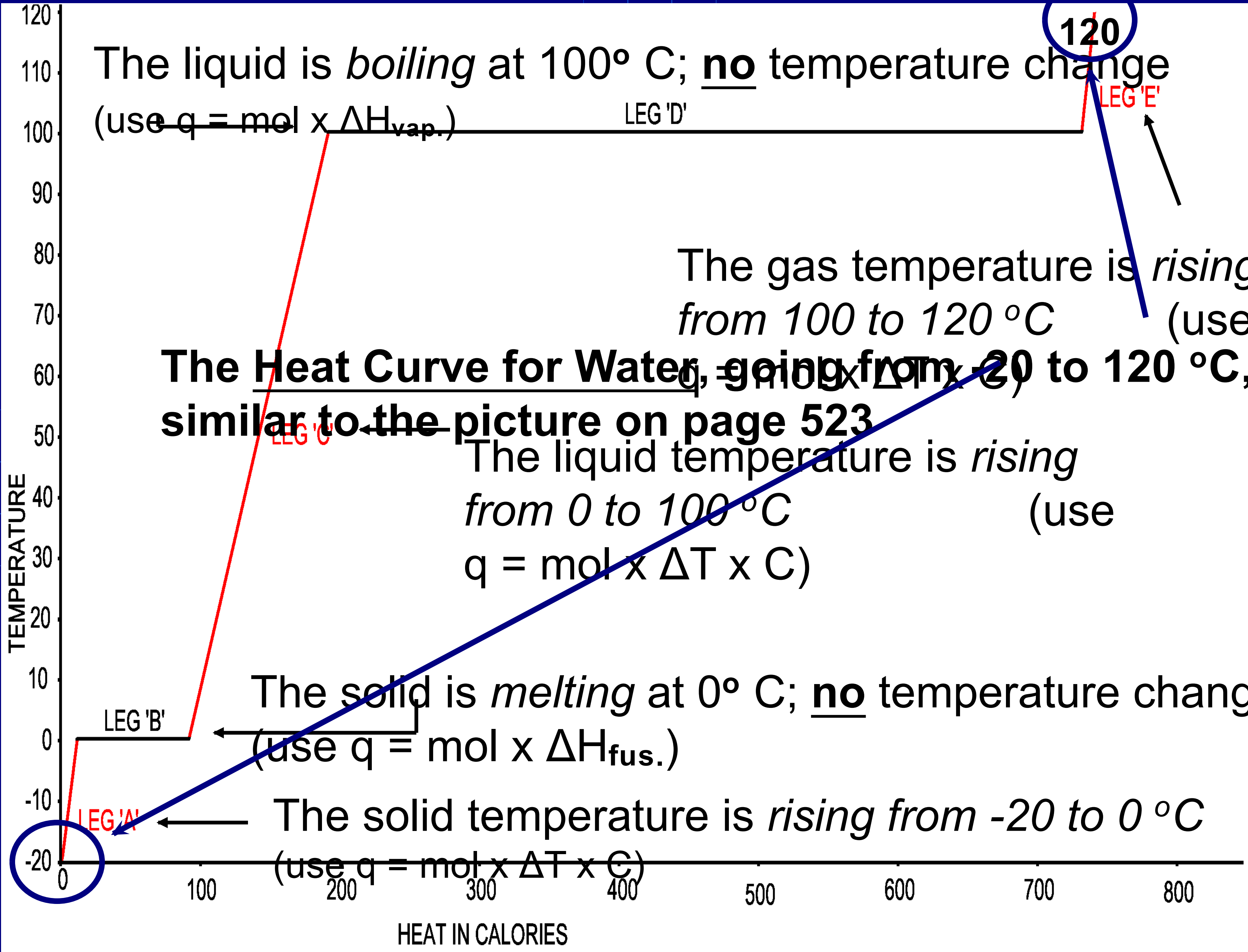
$$\frac{1\ mol\ H_2O(l)}{18.0\ g\ H_2O(l)} \quad \text{and} \quad \frac{40.7\ kJ}{1\ mol\ H_2O(l)}$$

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

$$\Delta H = 24.8\ g\ \cancel{H_2O(l)} \times \frac{1\ \cancel{mol\ H_2O(l)}}{18.0\ g\ \cancel{H_2O(l)}} \times \frac{40.7\ kJ}{1\ \cancel{mol\ H_2O(l)}}$$

$$= 56.1\ kJ$$

The Heat Curve for Water, going from -20 to 120 °C, similar to the picture on page 523



The liquid is *boiling* at 100° C; **no** temperature change (use q = mol x $\Delta H_{vap.}$)

The gas temperature is *rising from 100 to 120 °C* (use q = mol x ΔT x C)

The liquid temperature is *rising from 0 to 100 °C* (use q = mol x ΔT x C)

The solid is *melting* at 0° C; **no** temperature change (use q = mol x $\Delta H_{fus.}$)

The solid temperature is *rising from -20 to 0 °C* (use q = mol x ΔT x C)

Heat of Solution

- Heat changes can also occur when a solute *dissolves* in a solvent.

5. Molar Heat of Solution **($\Delta H_{soln.}$)** = heat change caused by dissolution of one mole of substance

   $q = \text{mol} \times \Delta H_{soln.}$ (**no** temperature change)

- Sodium hydroxide provides a good example of an exothermic molar heat of solution (next slide)

Heat of Solution

$$NaOH_{(s)} \xrightarrow{H_2O_{(l)}} Na^{1+}_{(aq)} + OH^{1-}_{(aq)}$$

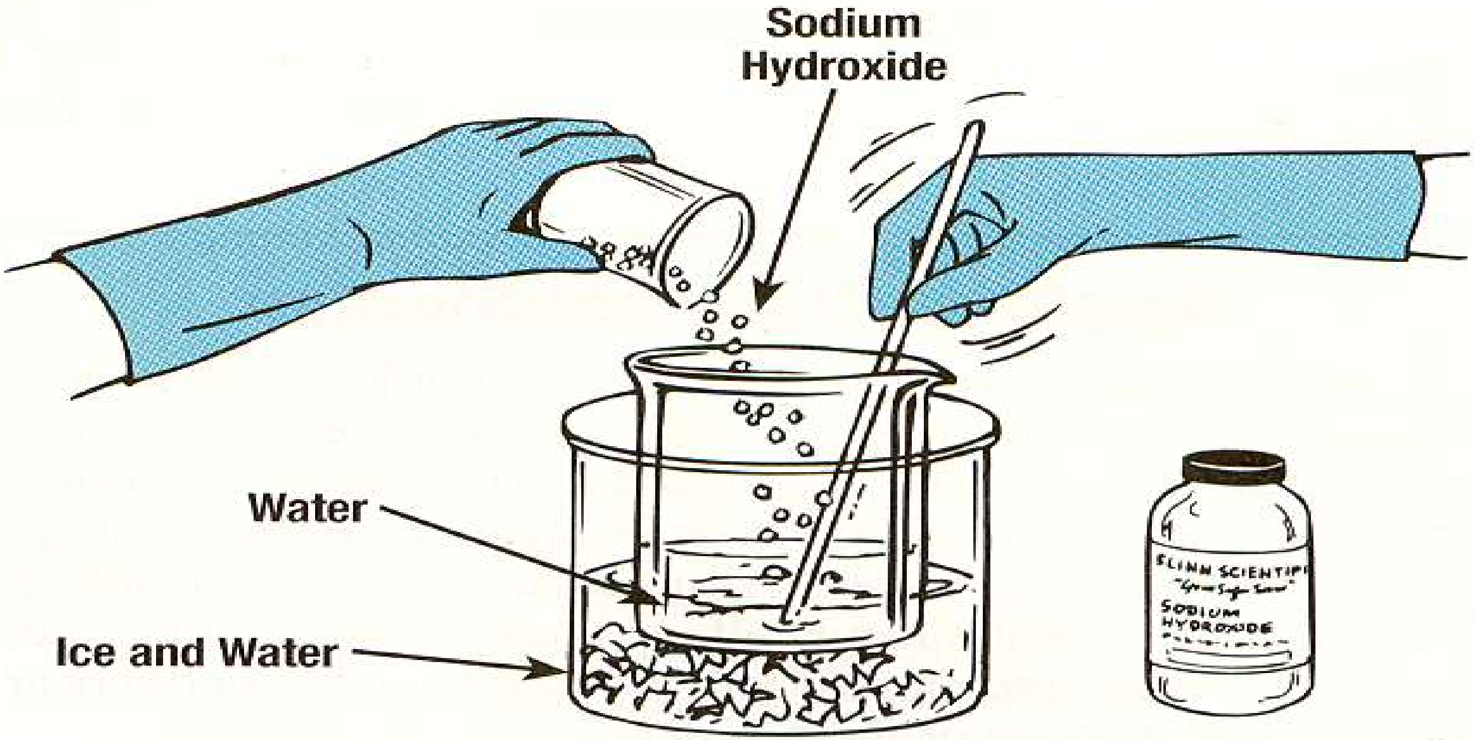
$$\Delta H_{soln.} = -\ 445.1 \text{ kJ/mol}$$

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

Sodium Hydroxide and Heat

Preparing Sodium Hydroxide Solution?

A great amount of heat is released when sodium hydroxide dissolves in water. The temperature of the solution may increase very rapidly. In fact, the temperature may rise so fast that the solution will boil and possibly spatter a hot, caustic solution. Immerse the beaker or flask in an ice–water bath to control the solution temperature. Add ingredients slowly while stirring continuously. Use only borosilicate glassware and check to make sure that there are no scratches, chips or breaks.



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

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

$$\text{NaOH}(s) \xrightarrow{\text{H}_2\text{O}(l)} \text{Na}^+(aq) + \text{OH}^-(aq) + 445.1\ \text{kJ/mol}$$

2 **Calculate** *Solve for the unknown.*

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

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

Section 17.4
Calculating Heats of Reaction

- **OBJECTIVES:**
  - State *Hess's Law of Heat Summation*, and describe how (or why) it is used in chemistry.

Section 17.4

Calculating Heats of Reaction

u **OBJECTIVES:**

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

Hess's Law

(developed in 1840)

Germain Henri Hess
(1802-1850)

- 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

How Does It Work?

- If you <u>turn an equation around</u>, you change the sign:

If $H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(g)\ \Delta H = -285.5\ kJ$

then the reverse is: $H_2O(g) \rightarrow H_2(g) + 1/2\ O_2(g)\ \ \Delta H = +285.5\ kJ$

- If you <u>multiply the equation by a number</u>, you multiply the heat by that number:

$2\ H_2O(g) \rightarrow 2\ H_2(g) + O_2(g)\ \ \Delta H = +571.0\ kJ$

- Or, you can just <u>leave the equation</u> “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)

u Note samples from pages 528 and 529

Standard Heats of Formation

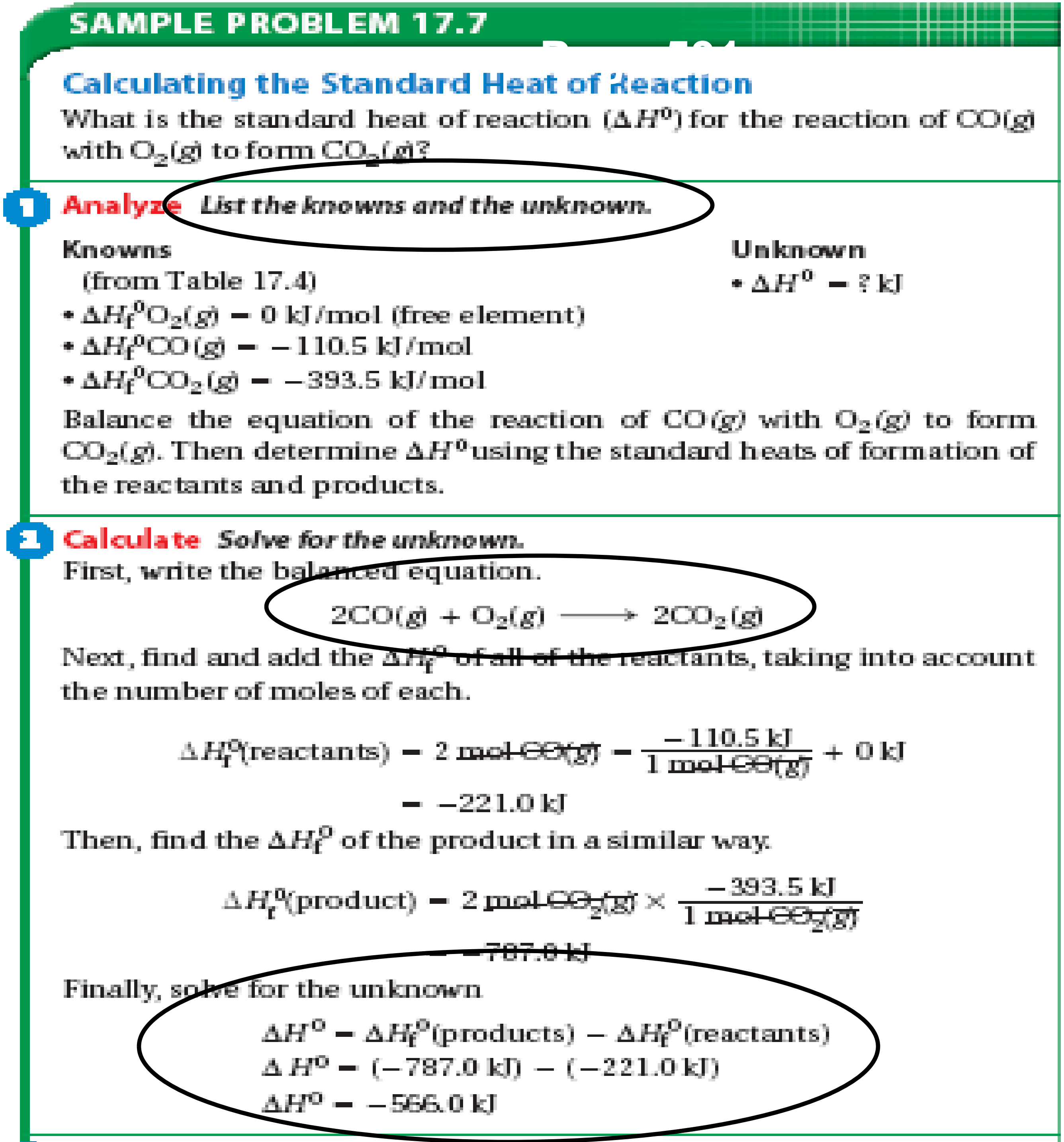
- The ΔH for a reaction that produces (or forms) 1 mol of a compound from its elements at standard conditions
- Standard conditions: 25°C and 1 atm.
- Symbol is: ΔH_f^0
- **The standard heat of formation of an <u>element</u> in it's <u>*standard state*</u> is arbitrarily set at "0"**
- **This includes the diatomic elements**

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

$$\boxed{\Delta H^{o} = (\Delta H_{f}^{0}\ \text{Products}) - (\Delta H_{f}^{0}\ \text{Reactants})}$$

Remember, from balanced equation: **Products - Reactants**



SAMPLE PROBLEM 17.7

Calculating the Standard Heat of Reaction

What is the standard heat of reaction (ΔH^0) for the reaction of $CO(g)$ with $O_2(g)$ to form $CO_2(g)$?

1 Analyze *List the knowns and the unknown.*

Knowns	Unknown
(from Table 17.4)	• $\Delta H^0 = ?$ kJ
• $\Delta H_f^0 O_2(g) = 0$ kJ/mol (free element)	
• $\Delta H_f^0 CO(g) = -110.5$ kJ/mol	
• $\Delta H_f^0 CO_2(g) = -393.5$ kJ/mol	

Balance the equation of the reaction of $CO(g)$ with $O_2(g)$ to form $CO_2(g)$. Then determine ΔH^0 using the standard heats of formation of the reactants and products.

2 Calculate *Solve for the unknown.*

First, write the balanced equation.

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

Next, find and add the ΔH_f^0 of all of the reactants, taking into account the number of moles of each.

$$\Delta H_f^0(\text{reactants}) = 2\ \cancel{\text{mol CO}(g)} \times \frac{-110.5\ \text{kJ}}{1\ \cancel{\text{mol CO}(g)}} + 0\ \text{kJ}$$

$$= -221.0\ \text{kJ}$$

Then, find the ΔH_f^0 of the product in a similar way.

$$\Delta H_f^0(\text{product}) = 2\ \cancel{\text{mol CO}_2(g)} \times \frac{-393.5\ \text{kJ}}{1\ \cancel{\text{mol CO}_2(g)}}$$

$$= -787.0\ \text{kJ}$$

Finally, solve for the unknown

$$\Delta H^0 = \Delta H_f^0(\text{products}) - \Delta H_f^0(\text{reactants})$$

$$\Delta H^0 = (-787.0\ \text{kJ}) - (-221.0\ \text{kJ})$$

$$\Delta H^0 = -566.0\ \text{kJ}$$

Another Example

- $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

ΔH_f^0 CH_4 (g) = - 74.86 kJ/mol

ΔH_f^0 O_2(g) = 0 kJ/mol (Because it is an element)

ΔH_f^0 CO_2(g) = - 393.5 kJ/mol

ΔH_f^0 H_2O(g) = - 241.8 kJ/mol

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

$\Delta H = -\ 802.24$ kJ (endothermic or exothermic?)

End of Chapter 17