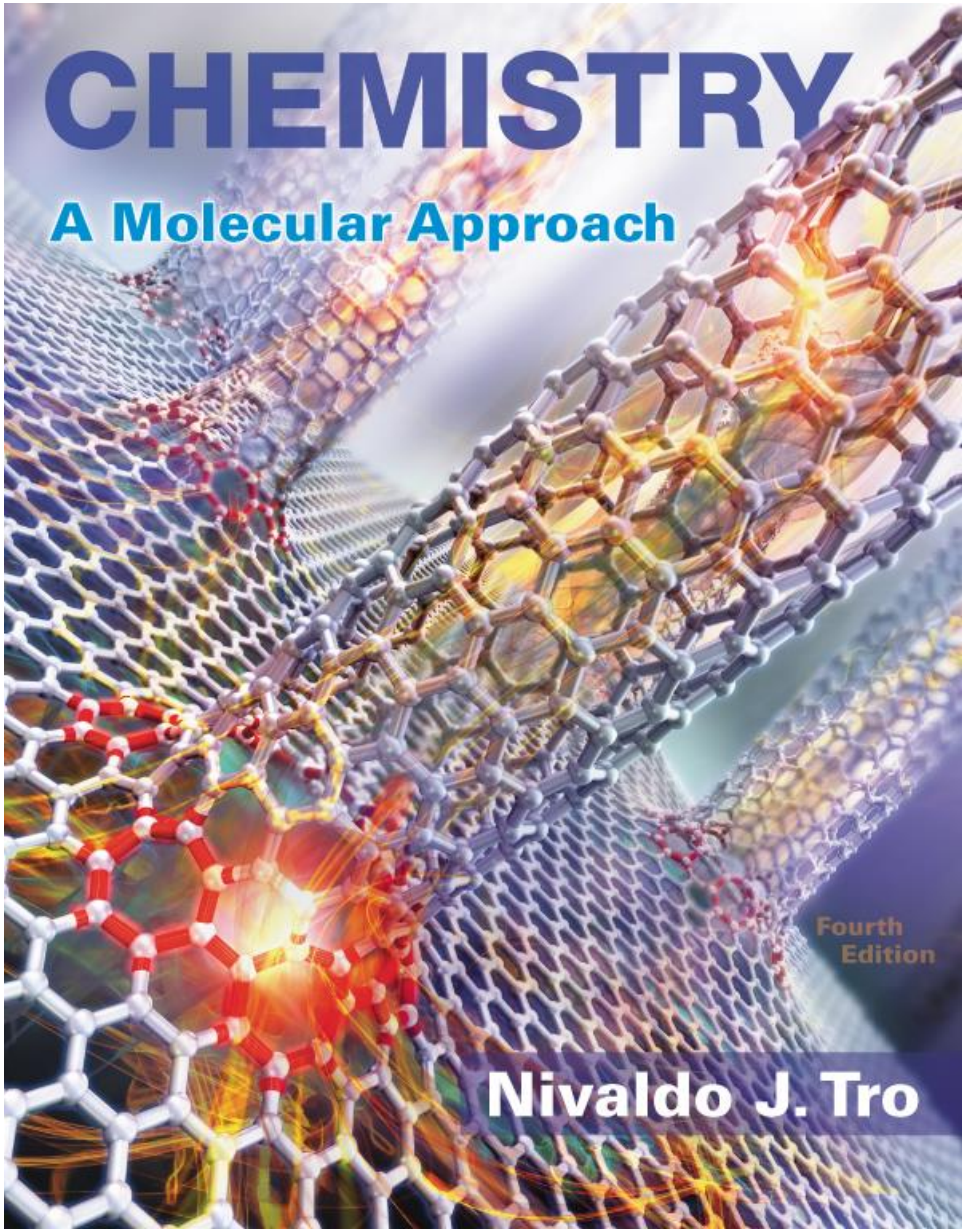


Lecture Presentation

Chapter 6

Thermochemistry

Chemical Hand Warmers

- Most hand warmers work by using the heat released from the slow oxidation of iron:

$$4\ \mathrm{Fe}(s) + 3\ \mathrm{O_2}(g) \rightarrow 2\ \mathrm{Fe_2O_3}(s)$$

- Exothermic reaction
- The amount your hand temperature rises depends on several factors:
  - The size of the hand warmer
  - The size of your glove, etc.

  Mainly, the amount of heat released by the reaction

Nature of Energy

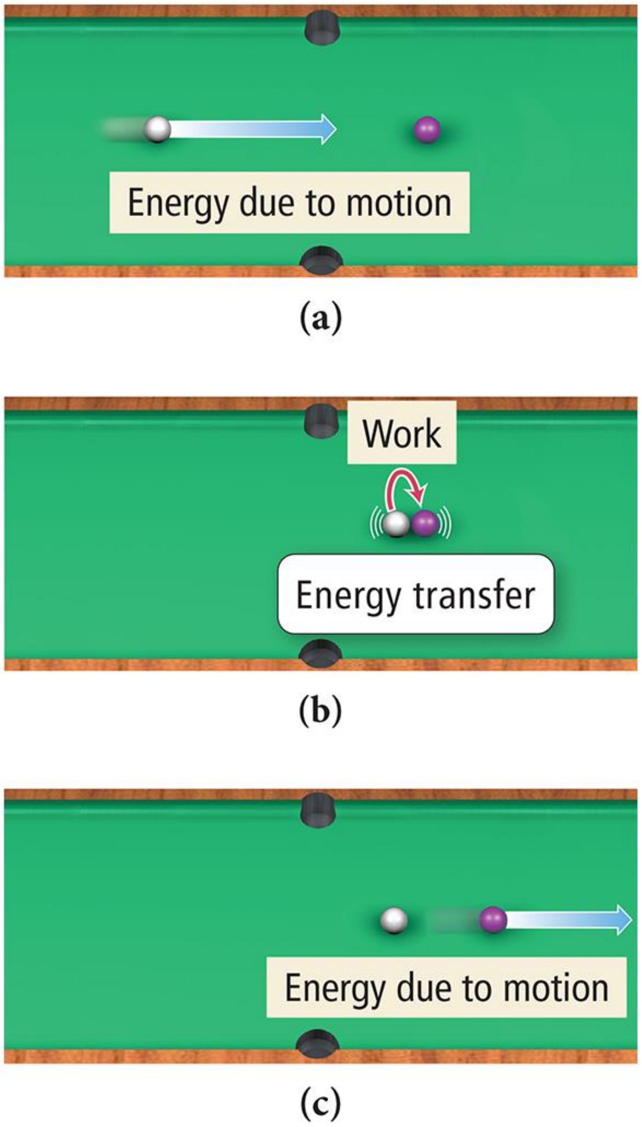
- Even though chemistry is the study of matter, energy affects matter.
- **Energy** is anything that has the capacity to do work.
- **Work** is a force acting over a distance.
  - Energy = work = force × distance
- **Heat** is the flow of energy caused by a difference in temperature.
- Energy can be exchanged between objects through contact.
  - For example, through collisions

Energy, Heat, and Work

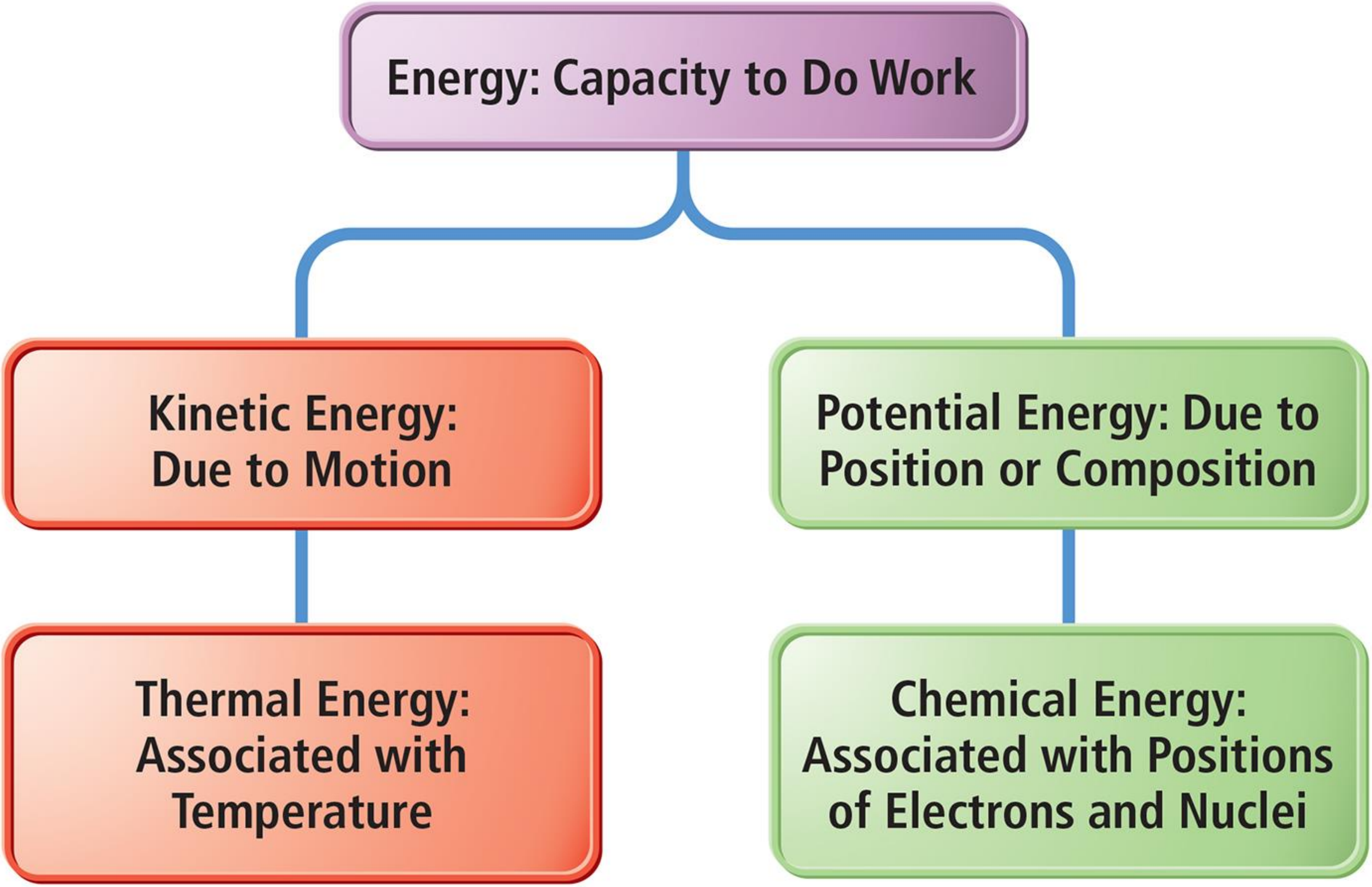
- You can think of **energy** as a quantity an object can possess or as a collection of objects.
- You can think of **heat and work** as the two different ways that an object can exchange energy with other objects.
  - Either out of it or into it

Classification of Energy

- **Kinetic energy** is energy of motion, or energy that is being transferred.
- **Thermal energy** is the energy associated with temperature.
    - Thermal energy is a form of kinetic energy.



Manifestations of Energy



Energy: Capacity to Do Work

Kinetic Energy: Due to Motion

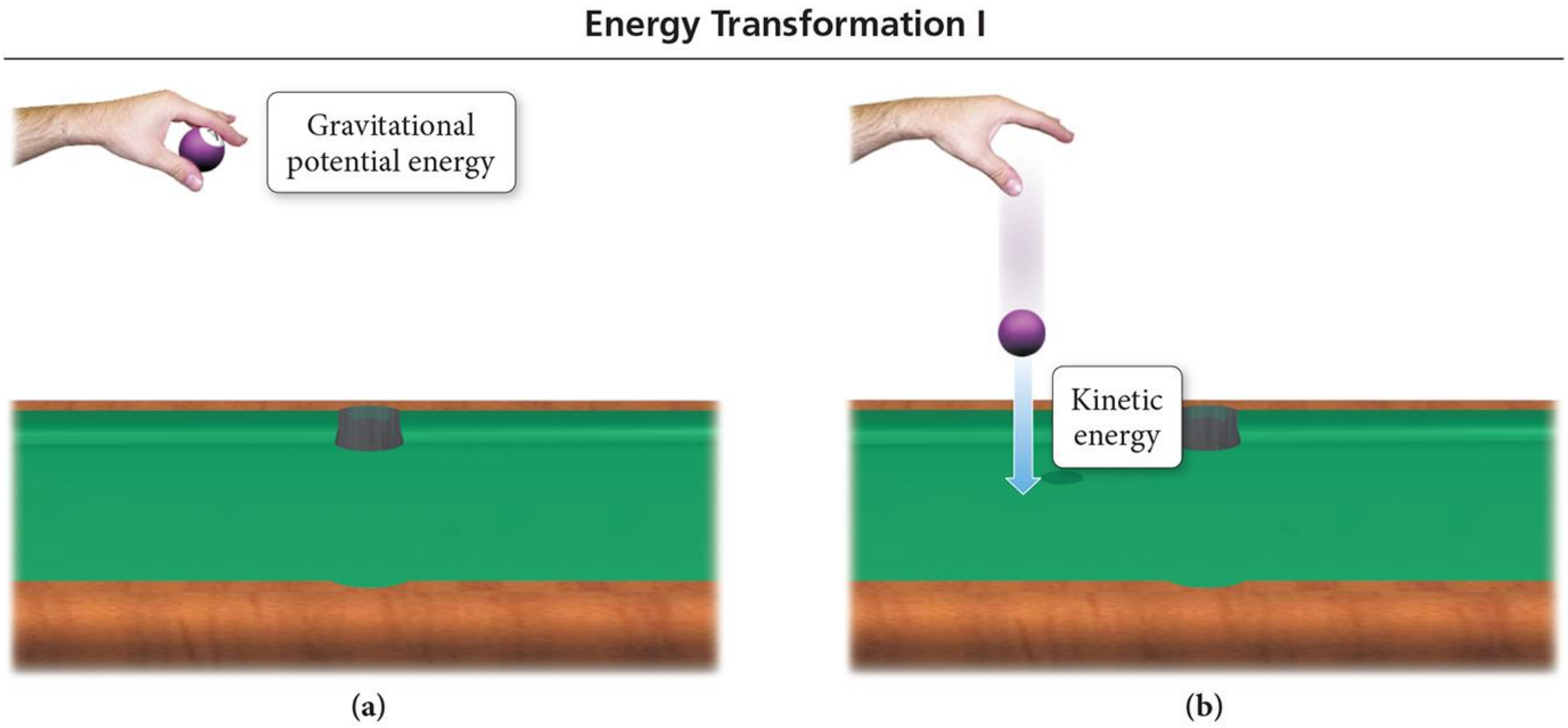
Potential Energy: Due to Position or Composition

Thermal Energy: Associated with Temperature

Chemical Energy: Associated with Positions of Electrons and Nuclei

Classification of Energy

- **Potential energy** is energy that is stored in an object, or energy associated with the composition and position of the object.
  - Energy stored in the structure of a compound is potential energy.



Some Forms of Energy

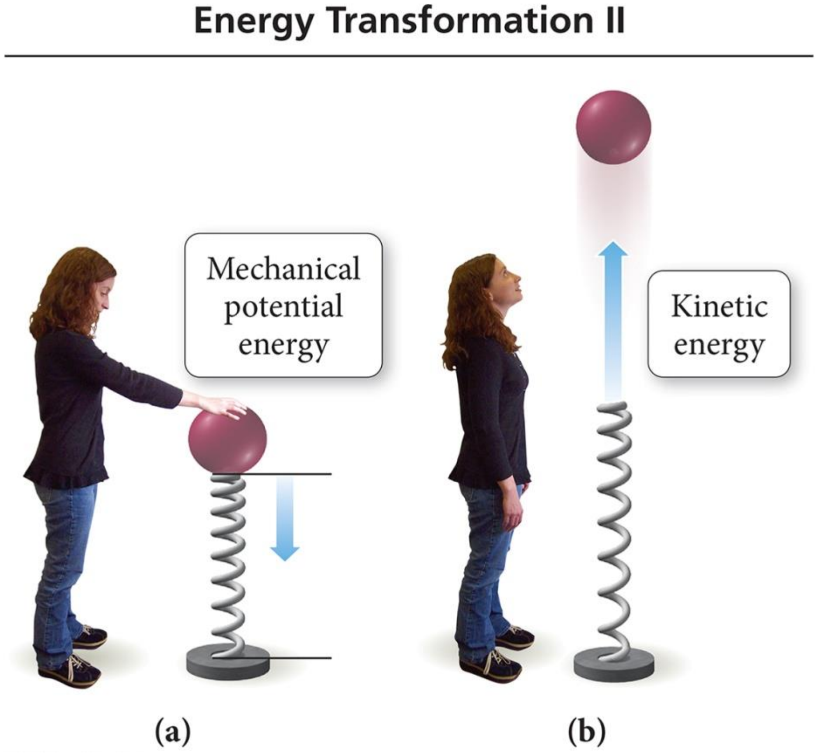
- **Chemical energy**:

  Potential energy due to the structure of the atoms, the attachment between atoms, the atoms' positions relative to each other in the molecule, or the molecules' relative positions in the structure

Conservation of Energy

- The **law of conservation of energy** states that energy cannot be created or destroyed.

- When energy is transferred between objects, or converted from one form to another, the total amount of energy present at the beginning must be present at the end.

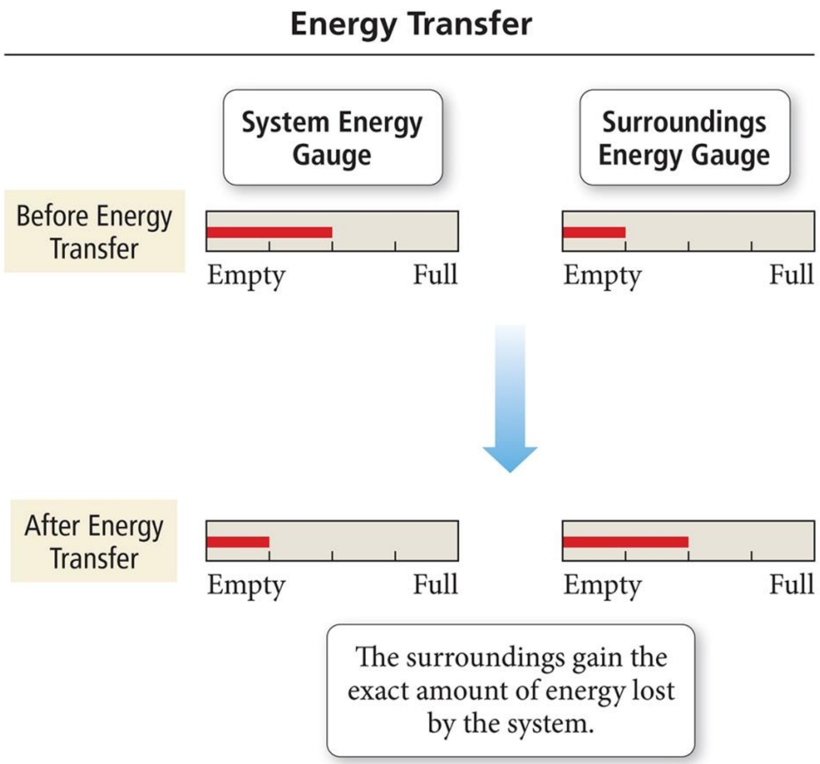


System and Surroundings

- We define the **system** as the material or process within which we are studying the energy changes within.
- We define the **surroundings** as everything else with which the system can exchange energy.
- What we study is the exchange of energy between the system and the surroundings.

Comparing the Amount of Energy in the System and Surroundings during Transfer

- Conservation of energy means that the amount of energy gained or lost by the system has to be equal to the amount of energy lost or gained by the surroundings.



The surroundings gain the exact amount of energy lost by the system.

Units of Energy

- The amount of kinetic energy an object has is directly proportional to its mass and velocity.
  - KE = $\frac{1}{2}mv^2$
- When the mass is in kg and velocity is in m/s, the unit for kinetic energy is $\frac{\text{kg}\bullet\text{m}^2}{\text{s}^2}$
- 1 joule of energy is the amount of energy needed to move a 1 kg mass at a speed of 1 m/s.
  - $1\text{ J} = 1\ \frac{\text{kg}\bullet\text{m}^2}{\text{s}^2}$

3.6×10^5 J or 0.10 kWh
used in 1 hour

Units of Energy

- A **joule (J)** is the amount of energy needed to move a 1 kg mass a distance of 1 meter.
  - **$1\ J = 1\ N \cdot m = 1\ kg \cdot m^2/s^2$**
- A **calorie (cal)** is the amount of energy needed to raise the temperature of one gram of water 1 °C.
  - kcal = energy needed to raise 1000 g of water 1 °C
  - food Calories = kcals

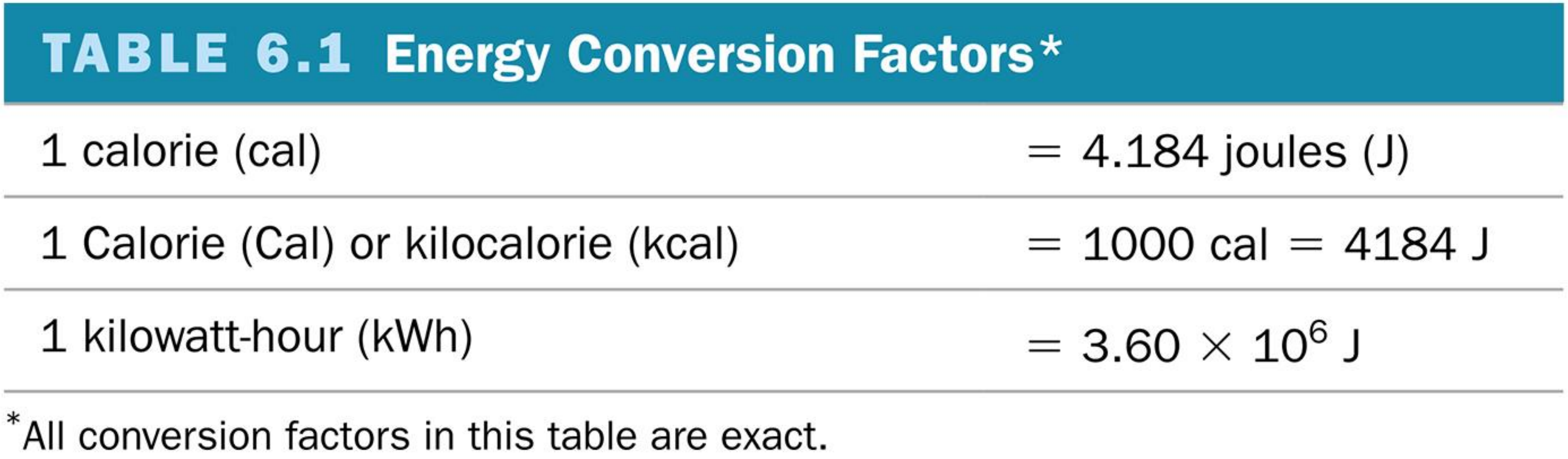


TABLE 6.1 Energy Conversion Factors*

1 calorie (cal)	= 4.184 joules (J)
1 Calorie (Cal) or kilocalorie (kcal)	= 1000 cal = 4184 J
1 kilowatt-hour (kWh)	= 3.60×10^6 J

*All conversion factors in this table are exact.

Energy Use

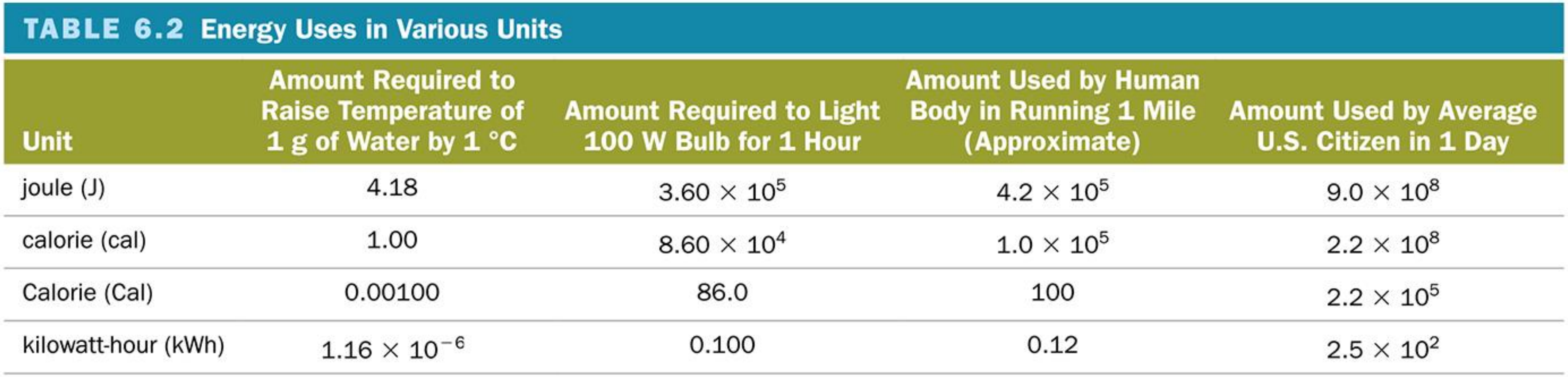


TABLE 6.2 Energy Uses in Various Units

Unit	Amount Required to Raise Temperature of 1 g of Water by 1 °C	Amount Required to Light 100 W Bulb for 1 Hour	Amount Used by Human Body in Running 1 Mile (Approximate)	Amount Used by Average U.S. Citizen in 1 Day
joule (J)	4.18	3.60×10^5	4.2×10^5	9.0×10^8
calorie (cal)	1.00	8.60×10^4	1.0×10^5	2.2×10^8
Calorie (Cal)	0.00100	86.0	100	2.2×10^5
kilowatt-hour (kWh)	1.16×10^{-6}	0.100	0.12	2.5×10^2

The First Law of Thermodynamics: Law of Conservation of Energy

- **Thermodynamics** is the study of energy and its interconversions.

- The first law of thermodynamics is the law of conservation of energy.
  - This means that the total amount of energy in the universe is constant.

- Therefore, you can never design a system that will continue to produce energy without some source of energy.

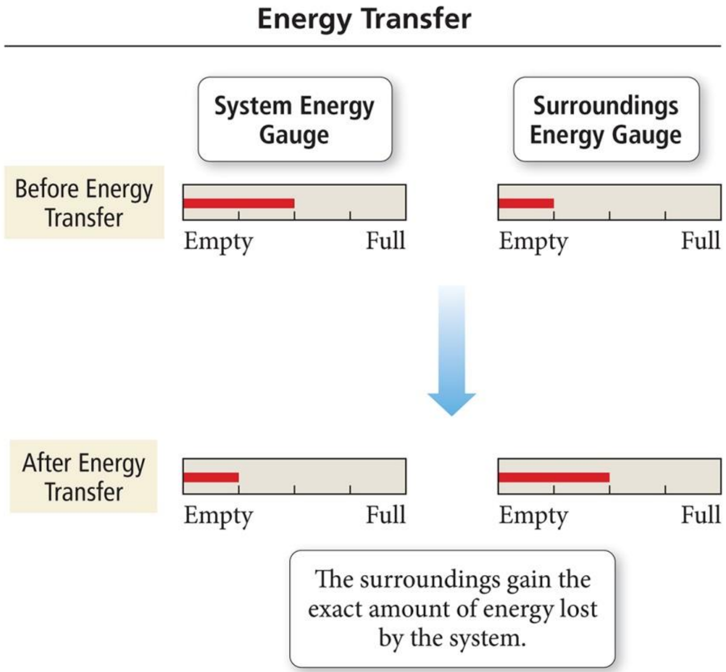
Energy Flow and Conservation of Energy

- Conservation of energy requires that the sum of the energy changes in the system and the surroundings must be zero.

  $\Delta\text{Energy}_{\text{universe}} = 0 = \Delta\text{Energy}_{\text{system}} + \Delta\text{Energy}_{\text{surroundings}}$

  Δ is the symbol that is used to mean change.

  - Final amount–initial amount



Internal Energy

- The **internal energy** is the sum of the kinetic and potential energies of all of the particles that compose the system.
- The change in the internal energy of a system depends only on the amount of energy in the system at the beginning and end.
  - A **state function** is a mathematical function whose result depends only on the initial and final conditions, not on the process used.

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

$$\Delta E_{\text{reaction}} = E_{\text{products}} - E_{\text{reactants}}$$

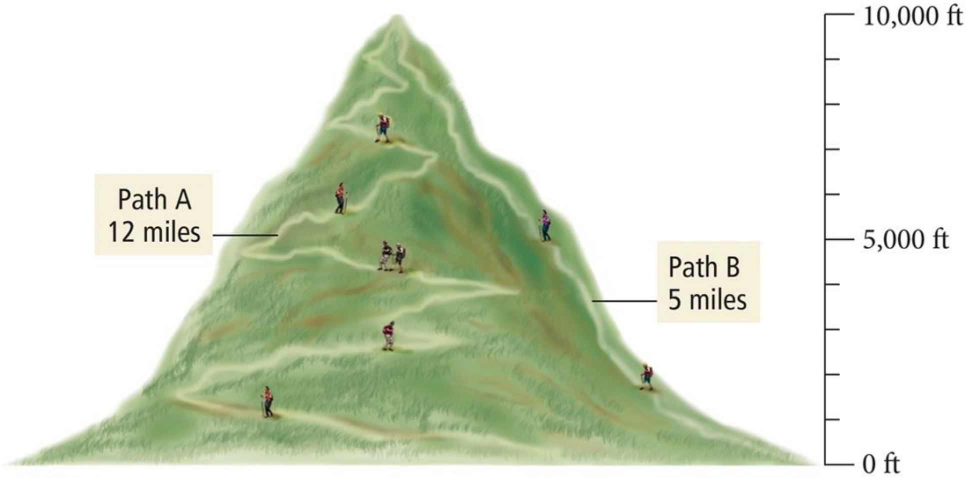
State Function

To reach the top of the mountain there are two trails:

1. Long and winding
2. Short but steep

Regardless of the trail, when you reach the top, you will be 10,000 ft above the base.

A State Function
Change in altitude depends only on the difference between the initial and final values, not on the path taken.



The distance from the base to the peak of the mountain is a state function. It depends only on the difference in elevation between the base and the peak, not on how you arrive there!

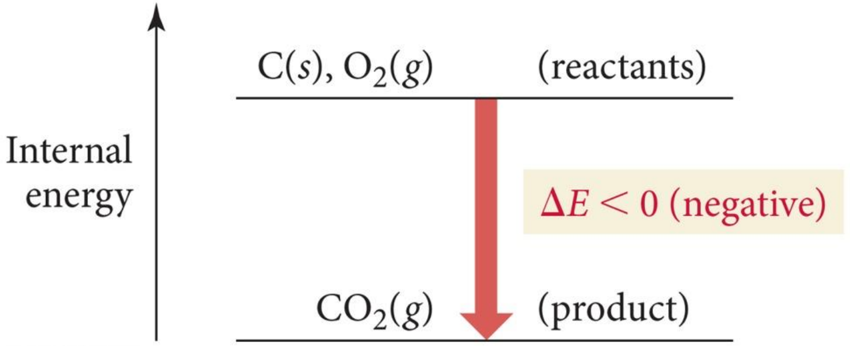
Energy Flow in a Chemical Reaction

- The total amount of internal energy in 1 mole of C(*s*) and 1 mole of $O_2(g)$ is greater than the internal energy in 1 mole of $CO_2(g)$.
  - At the same temperature and pressure

- In the reaction $C(s) + O_2(g) \rightarrow CO_2(g)$, there will be a net release of energy into the surroundings.
  - $-\Delta E_{\text{reaction}} = \Delta E_{\text{surroundings}}$

- In the reaction $CO_2(g) \rightarrow C(s) + O_2(g)$, there will be an absorption of energy from the surroundings into the reaction.

$$\Delta E_{\text{reaction}} = -\,\Delta E_{\text{surroundings}}$$



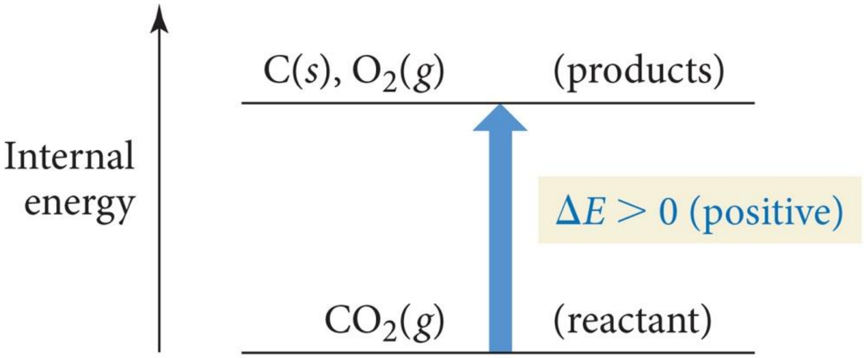
Energy Flow in a Chemical Reaction

- The total amount of internal energy in 1 mole of C(*s*) and 1 mole of $O_2(g)$ is greater than the internal energy in 1 mole of $CO_2(g)$.
  - At the same temperature and pressure

- In the reaction $C(s) + O_2(g) \rightarrow CO_2(g)$, there will be a net release of energy into the surroundings.
  - $-\Delta E_{\text{reaction}} = \Delta E_{\text{surroundings}}$

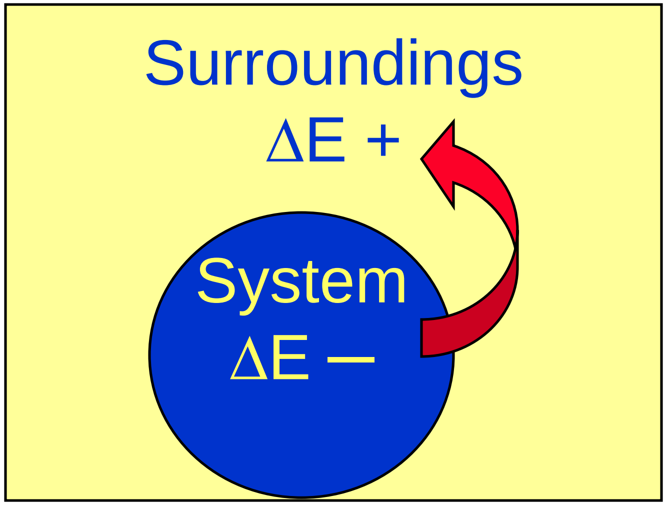
- In the reaction $CO_2(g) \rightarrow C(s) + O_2(g)$, there will be an absorption of energy from the surroundings into the reaction.

$$\Delta E_{\text{reaction}} = -\Delta E_{\text{surroundings}}$$



Energy Flow

- When energy flows out of a system, it must all flow into the surroundings.
- When energy flows out of a system, ΔE_{system} is negative.
- When energy flows into the surroundings, $\Delta E_{\text{surroundings}}$ is positive.
- Therefore,
  - $-\Delta E_{\text{system}} = \Delta E_{\text{surroundings}}$.

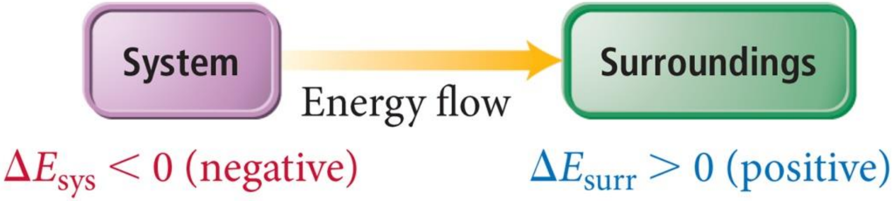


Surroundings

ΔE +

System

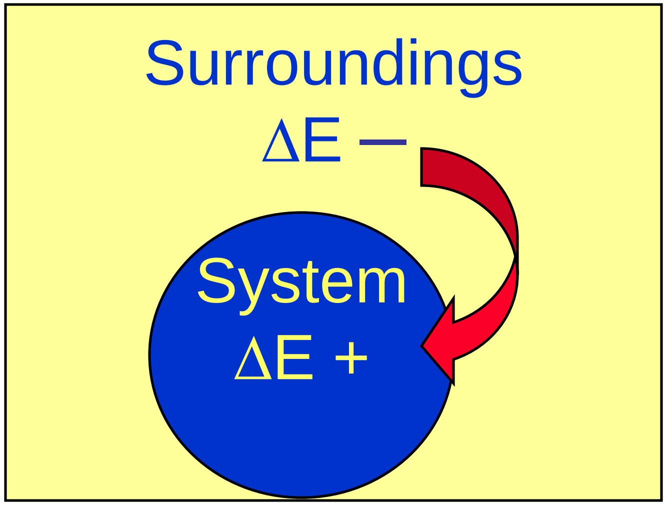
ΔE –

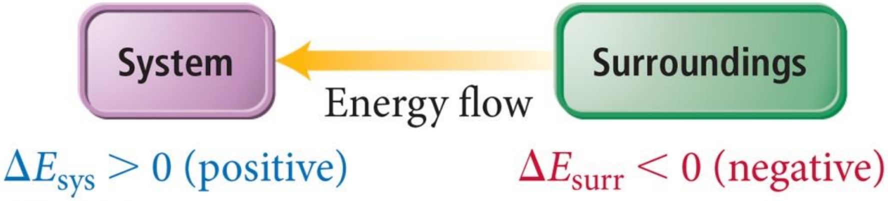


Energy Flow

- When energy flows into a system, it must all come from the surroundings.
- When energy flows into a system, ΔE_{system} is positive.
- When energy flows out of the surroundings, $\Delta E_{\text{surroundings}}$ is negative.
- Therefore,

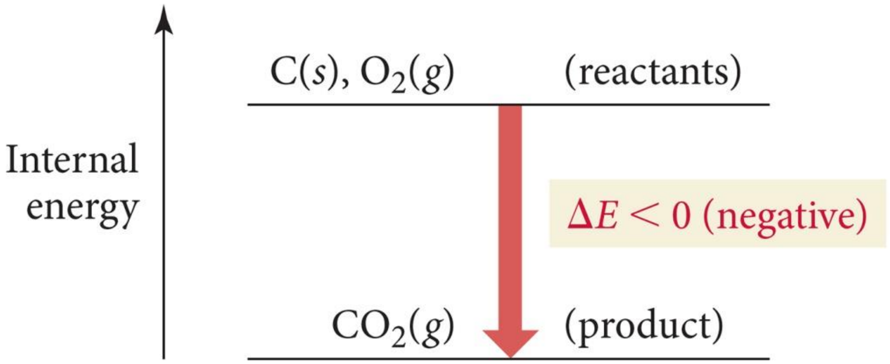
$$\Delta E_{\text{system}} = -\,\Delta E_{\text{surroundings}}.$$



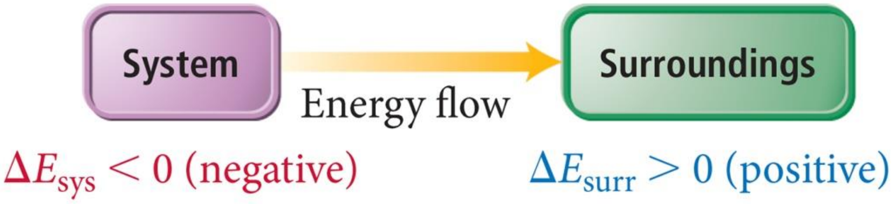


Energy Diagrams

- Energy diagrams are a “graphical” way of showing the direction of energy flow during a process.

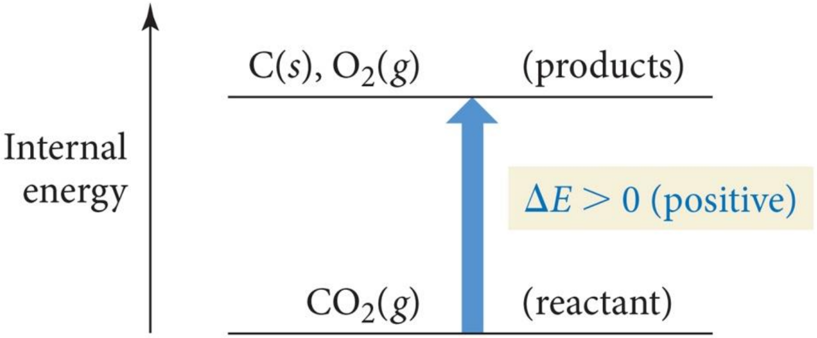


- If the reactants have a lower internal energy than the products, the change in energy will be positive.



Energy Diagrams

- Energy diagrams are a “graphical” way of showing the direction of energy flow during a process.



- If the reactants have a higher internal energy than the products, the change in energy will be negative.

Energy Exchange

- Energy is exchanged between the system and surroundings through heat and work.
  - q = heat (thermal) energy
  - w = work energy
  - q and w are NOT state functions; their value depends on the process.

$$\Delta E = q + w$$

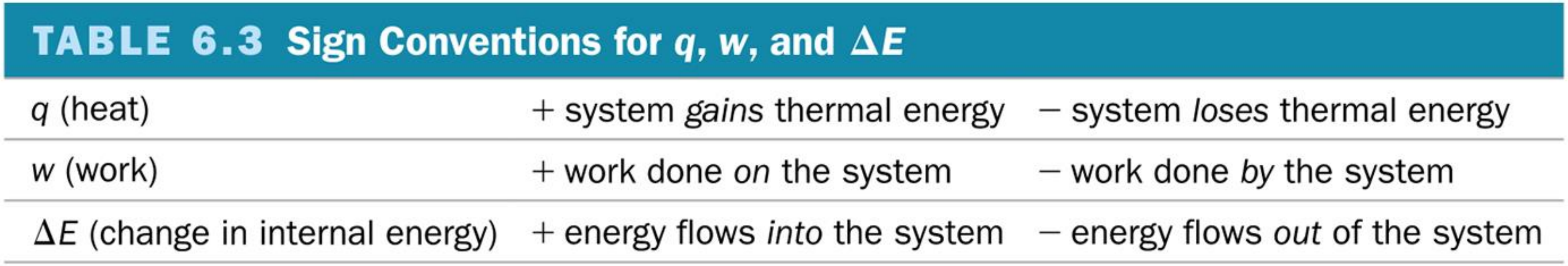
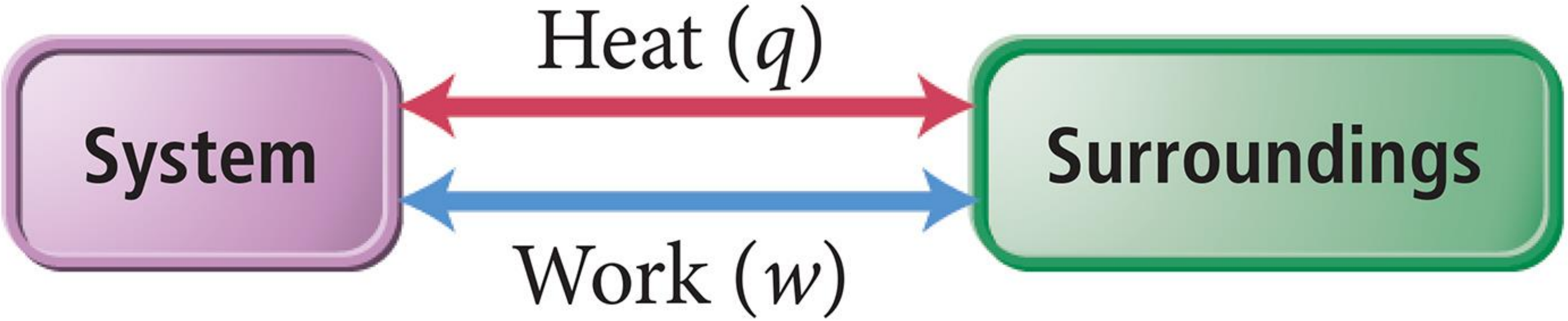


TABLE 6.3 Sign Conventions for q, w, and ΔE

q (heat)	+ system *gains* thermal energy	− system *loses* thermal energy
w (work)	+ work done *on* the system	− work done *by* the system
ΔE (change in internal energy)	+ energy flows *into* the system	− energy flows *out* of the system

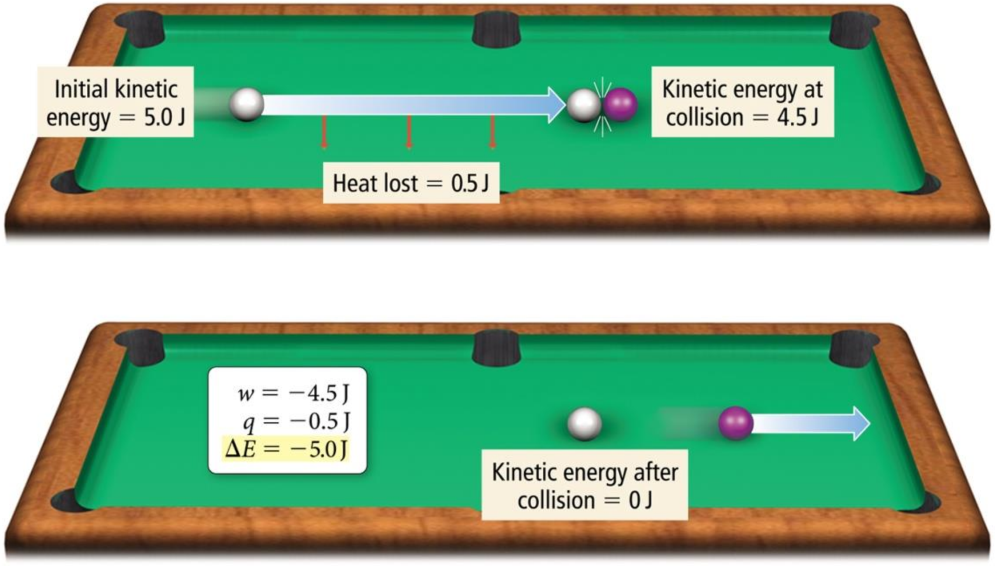
Energy Exchange



- Energy is exchanged between the system and surroundings through either heat exchange or work being done.

Heat and Work

- The white ball has an initial amount of 5.0 J of kinetic energy.
- As it rolls on the table, some of the energy is converted to heat by friction.
- The rest of the kinetic energy is transferred to the purple ball by collision.



(a) Smooth table

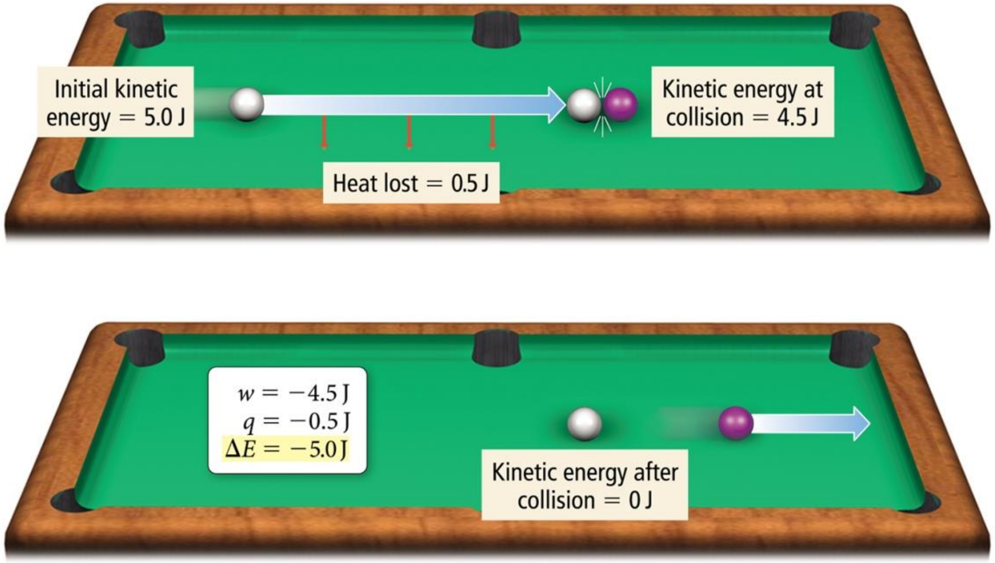
Heat and Work

On a smooth table, most of the kinetic energy is transferred from the white ball to the purple ball, with a small amount lost through friction.

Energy change for the white ball is as follows:

$\Delta E = KE_{\text{final}} - KE_{\text{initial}}$

$= 0\text{ J} - 5.0\text{ J} = -5.0\text{ J}$



(a) Smooth table

Kinetic energy transferred to the purple ball is $w = -4.5$ J.

Kinetic energy lost as heat is $q = -0.5$ J.

$q + w = (-0.5\text{ J}) + (-4.5\text{ J})$

$= -5.0\text{ J} = \Delta E$

Heat and Work

On a rough table, most of the kinetic energy of the white ball is lost through friction—less than half is transferred to the purple ball.

Energy change for the white ball is as follows:

$\Delta E = KE_{\text{final}} - KE_{\text{initial}}$

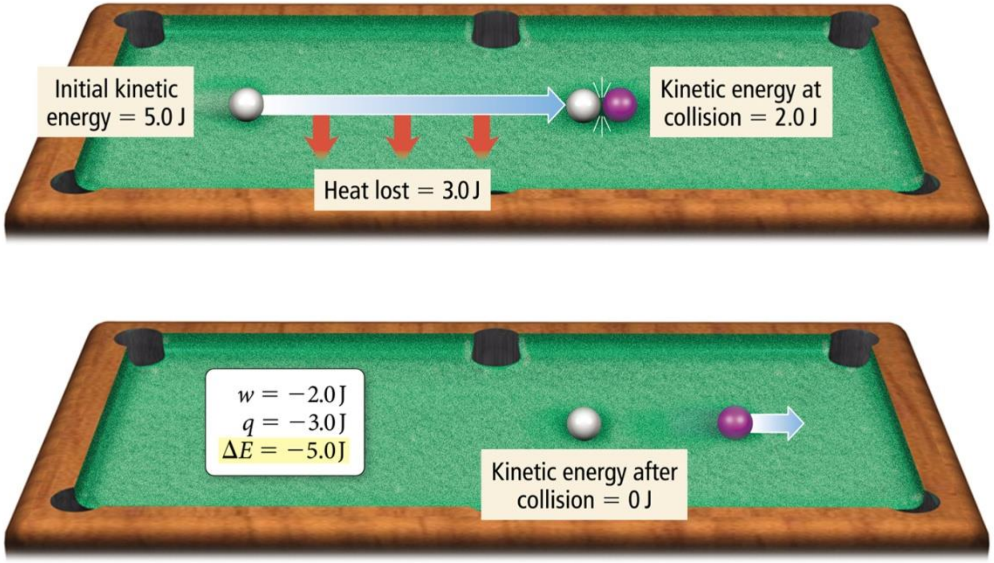
$= 0\ \text{J} - 5.0\ \text{J} = -5.0\ \text{J}$

Kinetic energy transferred to the purple ball is $w = -3.0$ J.

Kinetic energy lost as heat is $q = -2.0$ J.

$q + w = (-2.0\ \text{J}) + (-3.0\ \text{J})$

$= -5.0\ \text{J} = \Delta E$



(b) Rough table

Heat, Work, and Internal Energy

- In the previous billiard ball example, the ΔE of the white ball is the same for both cases, but q and w are not.
- On the rougher table, the heat loss, q, is greater.
  - q is a more negative number.
- But on the rougher table, less kinetic energy is transferred to the purple ball, so the work done by the white ball, w, is less.
  - w is a less negative number.
- The ΔE is a state function and depends only on the velocity of the white ball before and after the collision.
  - In both cases it started with 5.0 kJ of kinetic energy and ended with 0 kJ because it stopped.
  - $q + w$ is the same for both tables, even though the values of q and w are different.

Heat Exchange

- **Heat** is the *exchange* of thermal energy between a system and surroundings.
- Heat exchange occurs when system and surroundings have a difference in temperature.
- **Temperature** is the *measure* of the thermal energy within a sample of matter.
- Heat flows from matter with high temperature to matter with low temperature until both objects reach the same temperature.
  - Thermal equilibrium

Quantity of Heat Energy Absorbed: Heat Capacity

- When a system absorbs heat, its temperature increases.

- The increase in temperature is directly proportional to the amount of heat absorbed.

- The proportionality constant is called the **heat capacity, C**.
  - Units of C are J/°C or J/K.

$$\boldsymbol{q = C \times \Delta T}$$

- The larger the heat capacity of the object being studied, the smaller the temperature rise will be for a given amount of heat.

Factors Affecting Heat Capacity

- The heat capacity of an object depends on its amount of matter.
  - It is usually measured by its mass.
  - 200 g of water requires twice as much heat to raise its temperature by 1 °C as does 100 g of water.
- The heat capacity of an object depends on the type of material.
  - 1000 J of heat energy will raise the temperature of 100 g of sand 12 °C but raise the temperature of 100 g of water by only 2.4 °C.

Specific Heat Capacity

- Measure of a substance’s *intrinsic* ability to absorb heat.
- The **specific heat capacity** is the amount of heat energy required to raise the temperature of one gram of a substance 1 °C.
  - **C_s**
  - Units J/(g · °C)
- The **molar heat capacity** is the amount of heat energy required to raise the temperature of one mole of a substance 1 °C.

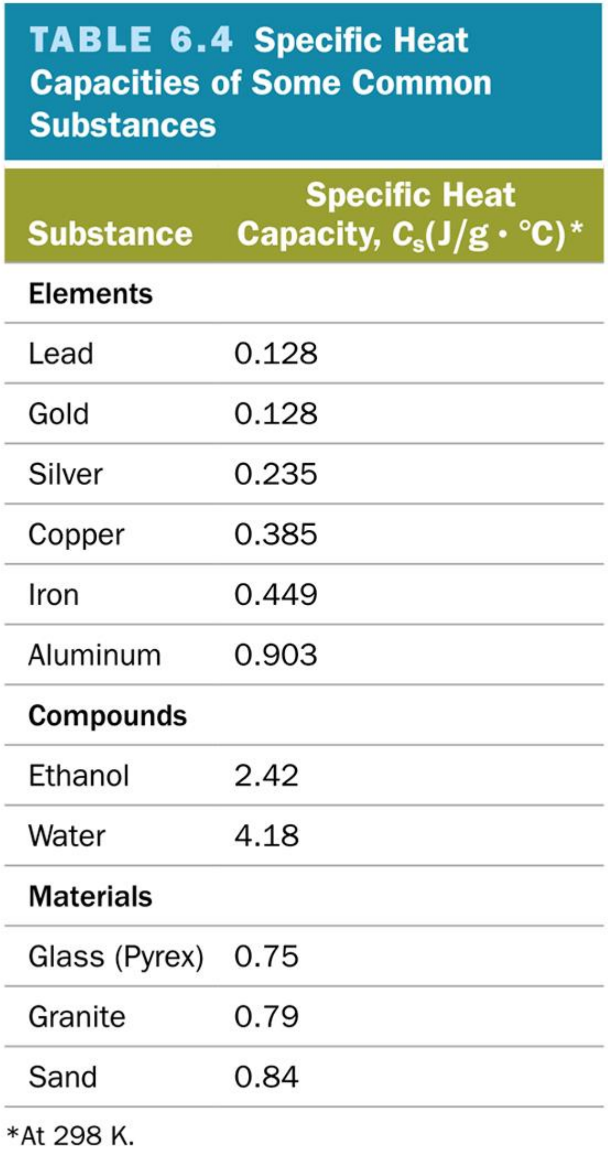


TABLE 6.4 Specific Heat Capacities of Some Common Substances

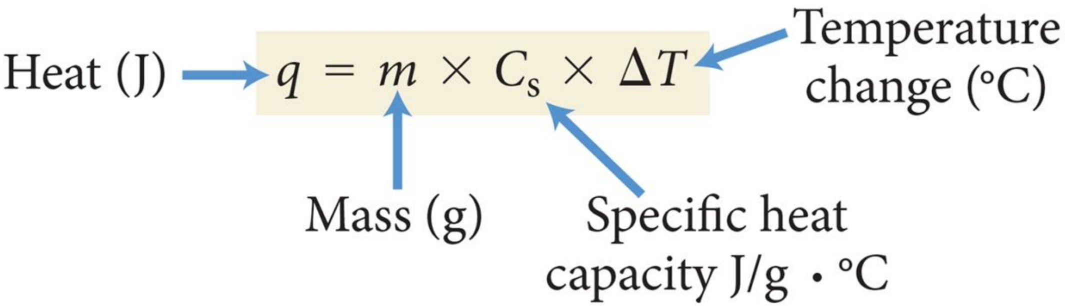
Substance	Specific Heat Capacity, C_s(J/g · °C)*
Elements	
Lead	0.128
Gold	0.128
Silver	0.235
Copper	0.385
Iron	0.449
Aluminum	0.903
Compounds	
Ethanol	2.42
Water	4.18
Materials	
Glass (Pyrex)	0.75
Granite	0.79
Sand	0.84

*At 298 K.

Quantifying Heat Energy

- The heat capacity of an object is proportional to the following:
  - Its mass
  - The specific heat of the material

- So, we can calculate the quantity of heat absorbed by an object if we know the mass, the specific heat, and the temperature change of the object.



Heat (J)

$$q = m \times C_s \times \Delta T$$

Temperature change (°C)

Mass (g)

Specific heat capacity J/g • °C

Specific Heat of Water

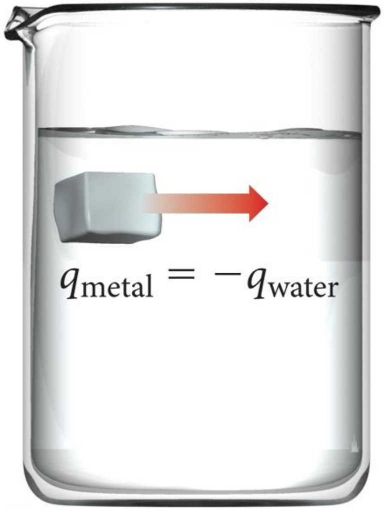
- Water can absorb a lot of heat energy without a large increase in its temperature due to its high specific heat capacity.
- The large amount of water absorbing heat from the air keeps beaches cool in the summer.
  - Temperature difference in summer between Sacramento (an inland city) and San Francisco (a coastal city) may be as much as 18 °C (30 °F).
- Water is commonly used as a coolant because it can absorb a lot of heat and remove it from important mechanical parts to keep them from overheating.
  - Water can even prevent melting.
  - It can also be used to transfer the heat to something else because it is a fluid.

Thermal Energy Transfer

- When two objects at different temperatures are placed in contact, heat flows from the material at the higher temperature to the material at the lower temperature.

- The amount of heat energy lost by the hot material equals the amount of heat gained by the cold material.

$$q_{sys} = -q_{surr}$$



Thermal Energy Transfer

- A block of metal at 55 °C is submerged in water initially at 25 °C.
- Thermal energy transfers heat from the metal to the water.
- The exact temperature change depends on the following:
  - The mass of the metal
  - The mass of water
  - Specific heat capacities of the metal and of water

$$q_{\text{metal}} = -q_{\text{water}}$$

$$m_{\text{metal}} \times C_{\text{s, metal}} \times \Delta T_{\text{metal}} = -m_{\text{water}} \times C_{\text{s, water}} \times \Delta T_{\text{water}}$$

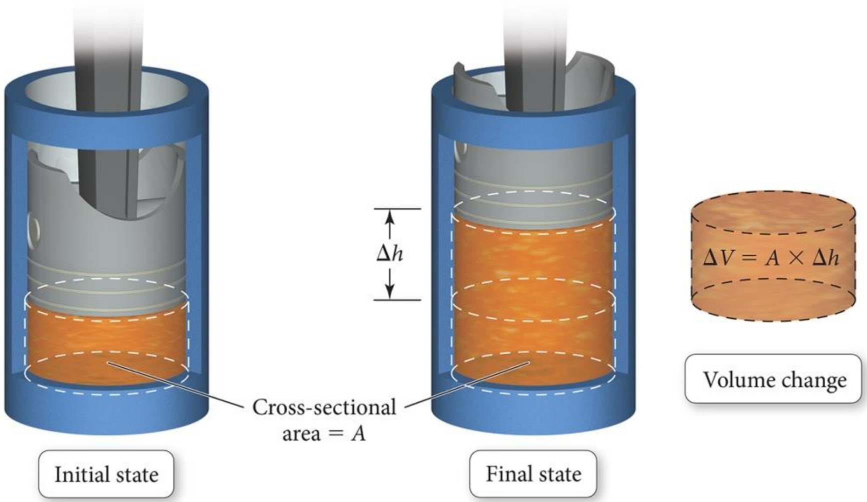
Pressure–Volume Work

- PV work is work caused by a volume change against an external pressure.
- When gases expand, ΔV is positive, but the system is doing work on the surroundings, so w_{gas} is negative.
- As long as the external pressure is kept constant,

$$Work_{gas} = \text{External Pressure} \times \text{Change in } Volume_{gas}$$

$$w = -P\Delta V.$$

  - To convert the units to joules, use 101.3 J = 1 atm · L.



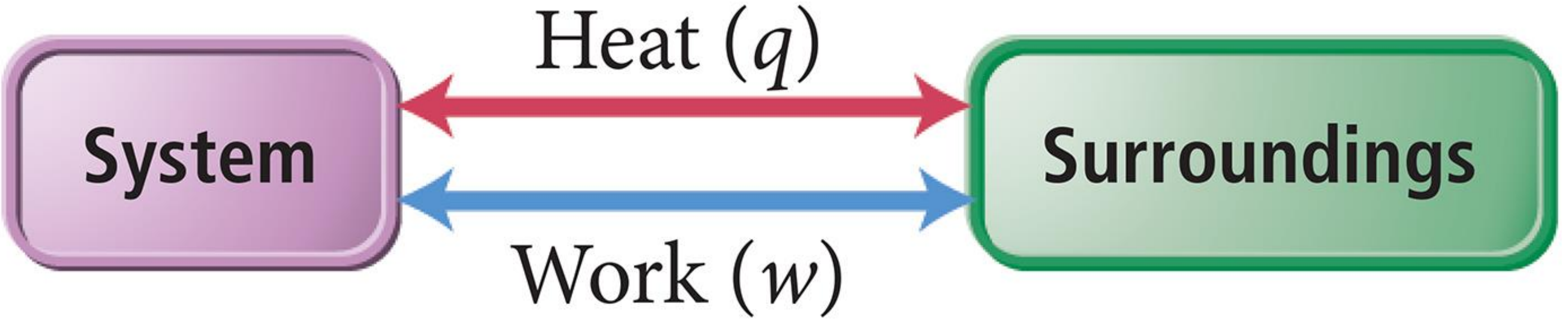
Exchanging Energy between System and Surroundings

- Exchange of heat energy

$$q = \text{mass} \times \text{specific heat} \times \Delta\text{Temperature}$$

- Exchange of work

$$w = -\text{Pressure} \times \Delta\text{Volume}$$



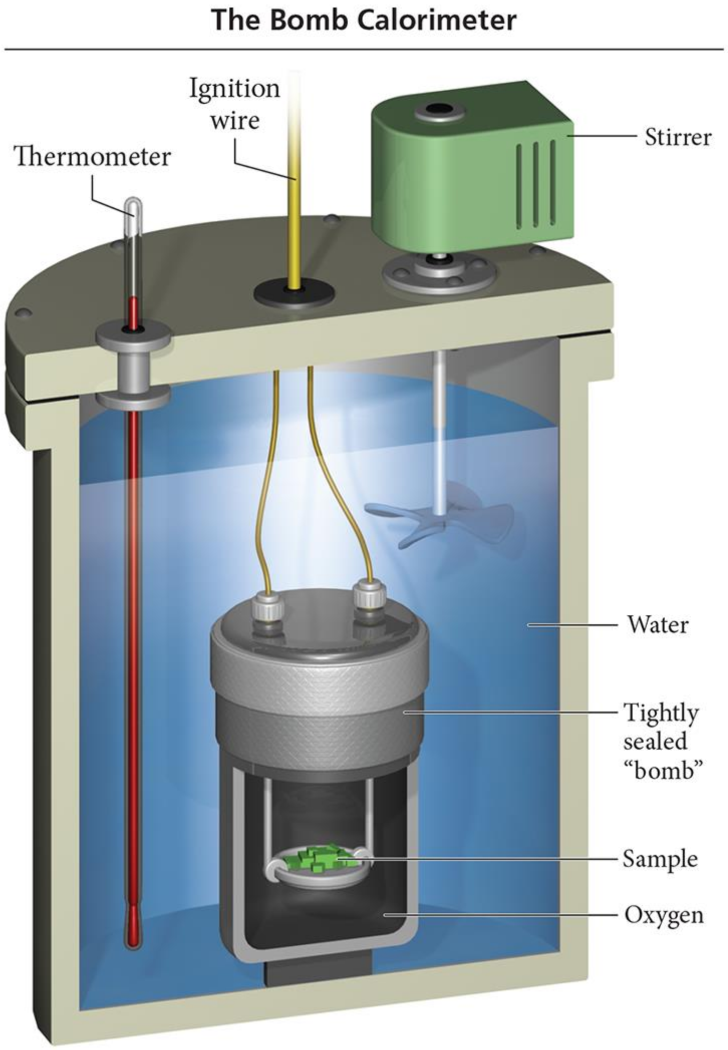
Measuring ΔE: Calorimetry at Constant Volume

- Because $\Delta E = q + w$, we can determine ΔE by measuring q and w.
- In practice, it is easiest to do a process in such a way that there is no change in volume, so $w = 0$.
  - At constant volume, $\Delta E_{\text{system}} = q_{\text{system}}$.
- In practice, we cannot observe the temperature changes of the individual chemicals involved in a reaction, so instead we measure the temperature change in the surroundings.
  - Use insulated, controlled surroundings.
  - $q_{\text{system}} = -q_{\text{surroundings}}$
- The surrounding area is called a **bomb calorimeter** and is usually made of a sealed, insulated container filled with water.

$$\boldsymbol{q_{\text{surroundings}} = q_{\text{calorimeter}} = -q_{\text{system}}}$$

Calorimetry

- **Calorimetry** is used to measure thermal energy exchanged between reaction and surroundings.
- A **bomb calorimeter** has a constant volume and is used to measure ΔE for combustion reactions.
- The heat capacity of the calorimeter is the amount of heat absorbed by the calorimeter for each degree rise in temperature and is called the **calorimeter constant.**
  - C_{cal}, kJ/°C



Enthalpy

- The **enthalpy**, ***H***, of a system is the sum of the internal energy of the system and the product of pressure and volume.
  - *H* is a state function.

$$H = E + PV$$

- The **enthalpy change**, **Δ*H***, of a reaction is the heat involved in a reaction at constant pressure.

$$\Delta H_{\text{reaction}} = q_{\text{reaction at constant pressure}}$$

- Usually ΔH and ΔE are similar in value; the difference is largest for reactions that produce or use large quantities of gas.

Endothermic and Exothermic Reactions

- When ΔH is negative, heat is being released by the system.
  - – This is called an **exothermic reaction**.

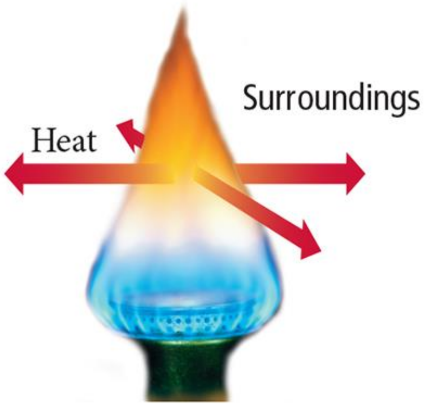
- When ΔH is positive, heat is being absorbed by the system.
  - – This is called an **endothermic reaction**.

Endothermic and Exothermic Reactions

- Chemical heat packs contain iron filings that are oxidized in an exothermic reaction: Your hands get warm because the released heat of the reaction is transferred to your hands.

- Chemical cold packs contain NH_4NO_3 that dissolves in water in an endothermic process: Your hands get cold because the pack is absorbing your heat.

Endothermic



Exothermic

Molecular View of Exothermic Reactions

- For an exothermic reaction, the surrounding’s temperature rises due to a release of thermal energy by the reaction.
- This extrathermal energy comes from the conversion of some of *the chemical potential energy* in the reactants into kinetic energy in the form of heat.
- During the course of a reaction, existing bonds are broken and new bonds are made.
- The products of the reaction have *less* chemical potential energy than the reactants.
- The difference in energy is released as heat.

Molecular View of Endothermic Reactions

- In an endothermic reaction, the surrounding’s temperature drops due to absorption of some of its thermal energy by the reaction.
- During the course of a reaction, existing bonds are broken and new bonds are made.
- The products of the reaction have *more* chemical potential energy than the reactants.
- To acquire this extra energy, some of the thermal energy of the surroundings is converted into chemical potential energy stored in the products.

Enthalpy of Reaction

- The enthalpy change in a chemical reaction is an extensive property.
  - The more reactants you use, the larger the enthalpy change.

- By convention, we calculate the enthalpy change for the number of moles of reactants in the reaction as written.

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g) \qquad \Delta H = -2044\ \text{kJ}$

$1\ \text{mol}\ C_3H_8(g) = -2044\ \text{kJ}$ or $5\ \text{mol}\ O_2(g) = -2044\ \text{kJ}$

Measuring ΔH Calorimetry at Constant Pressure

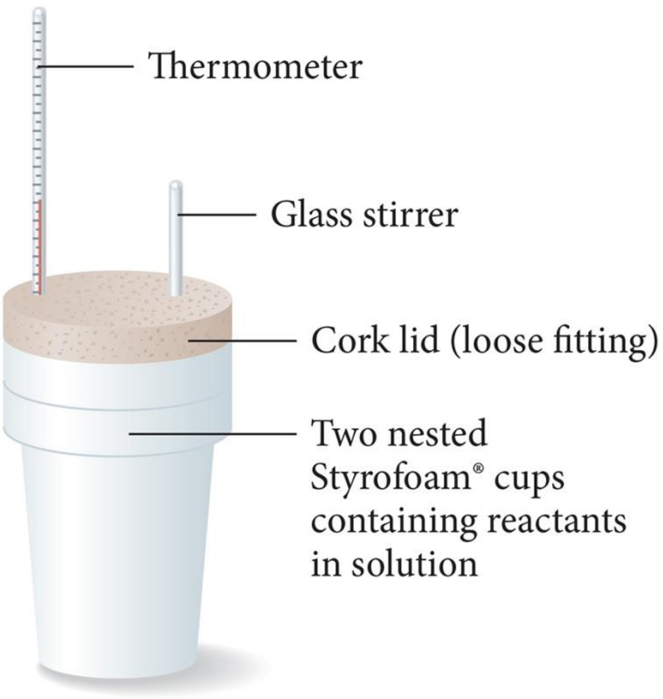
- Reactions done in aqueous solution are at constant pressure.
  - Open to the atmosphere
- The calorimeter is often nested foam cups containing the solution.

$$q_{\text{reaction}} = -q_{\text{solution}}$$

$$= -(\text{mass}_{\text{solution}} \times C_{\text{s, solution}} \times \Delta T)$$

- $\Delta H_{\text{reaction}} = q_{\text{constant pressure}} = q_{\text{reaction}}$
  - To get $\Delta H_{\text{reaction}}$ per mol, divide by the number of moles.

The Coffee-Cup Calorimeter



Relationships Involving ΔH_{rxn}

- When a reaction is multiplied by a factor, ΔH_{rxn} is multiplied by that factor.
  - Because ΔH_{rxn} is extensive,

$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ \text{kJ}$

$2\ C(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g) \quad \Delta H = 2(-393.5\ \text{kJ}) = -787.0\ \text{kJ}.$

- If a reaction is reversed, then the sign of ΔH is changed.

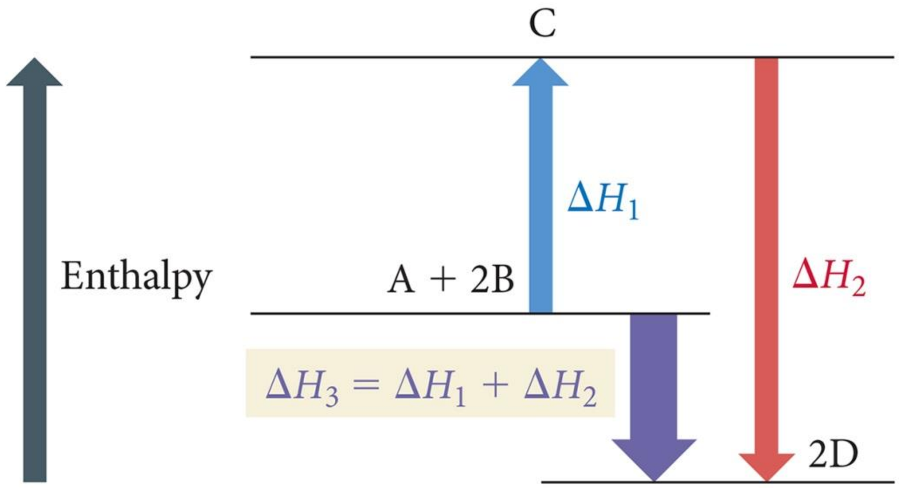
$CO_2(g) \rightarrow C(s) + O_2(g) \qquad \Delta H = +393.5\ \text{kJ}$

Relationships Involving ΔH_{rxn}: Hess's Law

- If a reaction can be expressed as a series of steps, then the ΔH_{rxn} for the overall reaction is the sum of the heats of reaction for each step.

Hess's Law

The change in enthalpy for a stepwise process is the sum of the enthalpy changes of the steps.



Standard Conditions

- The **standard state** is the state of a material at a defined set of conditions.
  - Pure gas at exactly 1 atm pressure
  - Pure solid or liquid in its most stable form at exactly 1 atm pressure and temperature of interest
    - Usually 25 °C
  - Substance in a solution with concentration 1 M
- The **standard enthalpy change**, **$\Delta H°$**, is the enthalpy change when all reactants and products are in their standard states.
- The **standard enthalpy of formation**, **$\Delta H_f°$**, is the enthalpy change for the reaction forming 1 mole of a pure compound from its constituent elements.
  - The elements must be in their standard states.
  - The $\Delta H_f°$ for a pure element in its standard state = 0 kJ/mol.

Standard Enthalpies of Formation

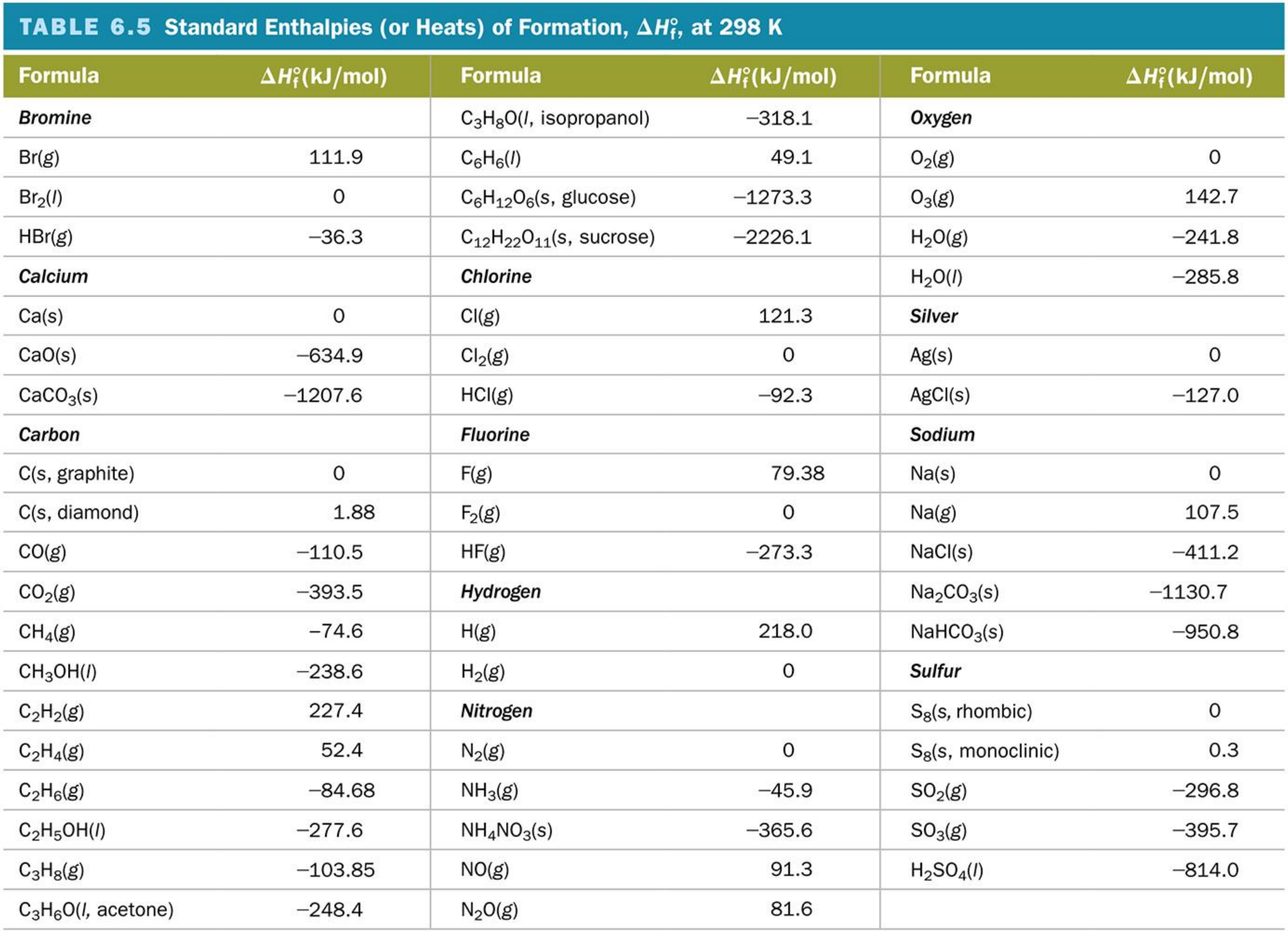


TABLE 6.5 Standard Enthalpies (or Heats) of Formation, ΔH_f°, at 298 K

Formula	ΔH_f°(kJ/mol)	Formula	ΔH_f°(kJ/mol)	Formula	ΔH_f°(kJ/mol)
Bromine		C_3H_8O(*l*, isopropanol)	−318.1	***Oxygen***	
$Br(g)$	111.9	$C_6H_6(l)$	49.1	$O_2(g)$	0
$Br_2(l)$	0	$C_6H_{12}O_6$(s, glucose)	−1273.3	$O_3(g)$	142.7
$HBr(g)$	−36.3	$C_{12}H_{22}O_{11}$(s, sucrose)	−2226.1	$H_2O(g)$	−241.8
Calcium		***Chlorine***		$H_2O(l)$	−285.8
$Ca(s)$	0	$Cl(g)$	121.3	***Silver***	
$CaO(s)$	−634.9	$Cl_2(g)$	0	$Ag(s)$	0
$CaCO_3(s)$	−1207.6	$HCl(g)$	−92.3	$AgCl(s)$	−127.0
Carbon		***Fluorine***		***Sodium***	
C(s, graphite)	0	$F(g)$	79.38	$Na(s)$	0
C(s, diamond)	1.88	$F_2(g)$	0	$Na(g)$	107.5
$CO(g)$	−110.5	$HF(g)$	−273.3	$NaCl(s)$	−411.2
$CO_2(g)$	−393.5	***Hydrogen***		$Na_2CO_3(s)$	−1130.7
$CH_4(g)$	−74.6	$H(g)$	218.0	$NaHCO_3(s)$	−950.8
$CH_3OH(l)$	−238.6	$H_2(g)$	0	***Sulfur***	
$C_2H_2(g)$	227.4	***Nitrogen***		S_8(s, rhombic)	0
$C_2H_4(g)$	52.4	$N_2(g)$	0	S_8(s, monoclinic)	0.3
$C_2H_6(g)$	−84.68	$NH_3(g)$	−45.9	$SO_2(g)$	−296.8
$C_2H_5OH(l)$	−277.6	$NH_4NO_3(s)$	−365.6	$SO_3(g)$	−395.7
$C_3H_8(g)$	−103.85	$NO(g)$	91.3	$H_2SO_4(l)$	−814.0
C_3H_6O(*l*, acetone)	−248.4	$N_2O(g)$	81.6		

Formation Reactions

- Reactions of elements in their standard state to form 1 mole of a pure compound
  - If you are not sure what the standard state of an element is, find the form in Appendix IIB that has $\Delta H_f^\circ = 0$.
  - Because the definition requires 1 mole of compound be made, the coefficients of the reactants may be fractions.

Writing Formation Reactions: Write the Formation Reaction for CO(*g*)

- The formation reaction is the reaction between the elements in the compound, which are C and O.

$$C + O \rightarrow CO(g)$$

- The elements must be in their standard state.
  - There are several forms of solid C, but the one with $\Delta H_f° = 0$ is graphite.
  - Oxygen's standard state is the diatomic gas.

$$C(s, \text{graphite}) + O_2(g) \rightarrow CO(g)$$

- The equation must be balanced, but the coefficient of the product compound must be 1.
  - Use whatever coefficient in front of the reactants is necessary to make the atoms on both sides equal without changing the product coefficient.

$$C(s, \text{graphite}) + \tfrac{1}{2}\, O_2(g) \rightarrow CO(g)$$

Calculating Standard Enthalpy Change for a Reaction

- Any reaction can be written as the sum of formation reactions (or the reverse of formation reactions) for the reactants and products.

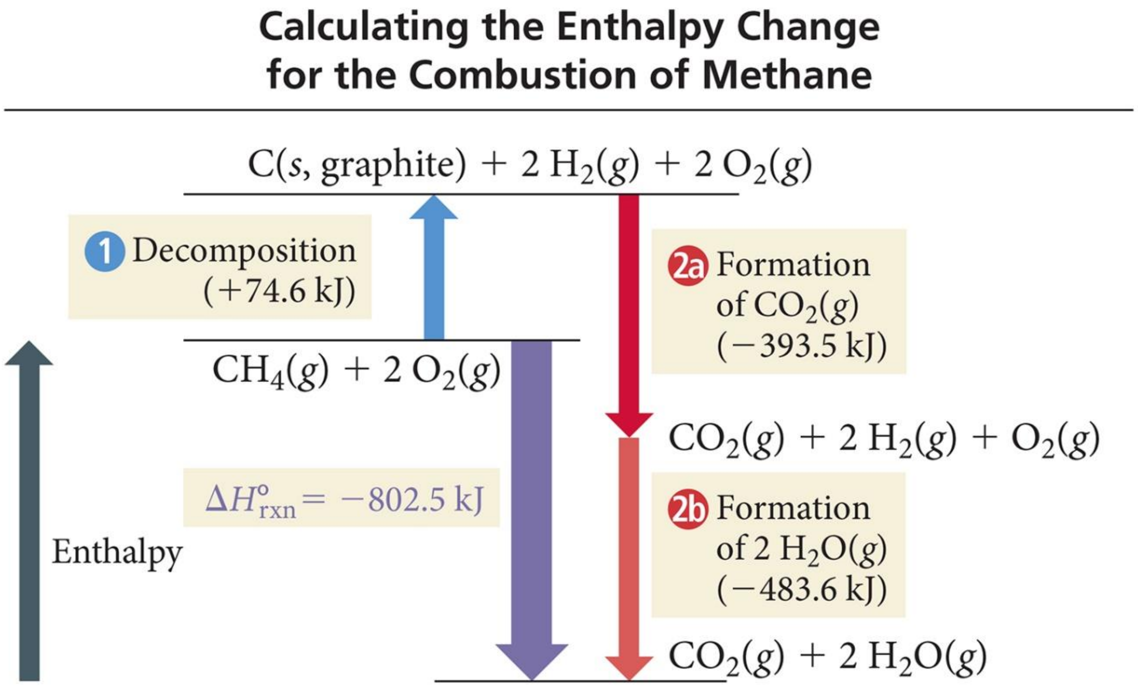
- The $\Delta H°$ for the reaction is then the sum of the $\Delta H_f°$ for the component reactions.

$$\boldsymbol{\Delta H°_{\text{reaction}} = \Sigma\, n\, \Delta H_f° \text{ (products)} - \Sigma\, n\, \Delta H_f° \text{ (reactants)}}$$

Σ means sum.

n is the coefficient of the reaction.

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

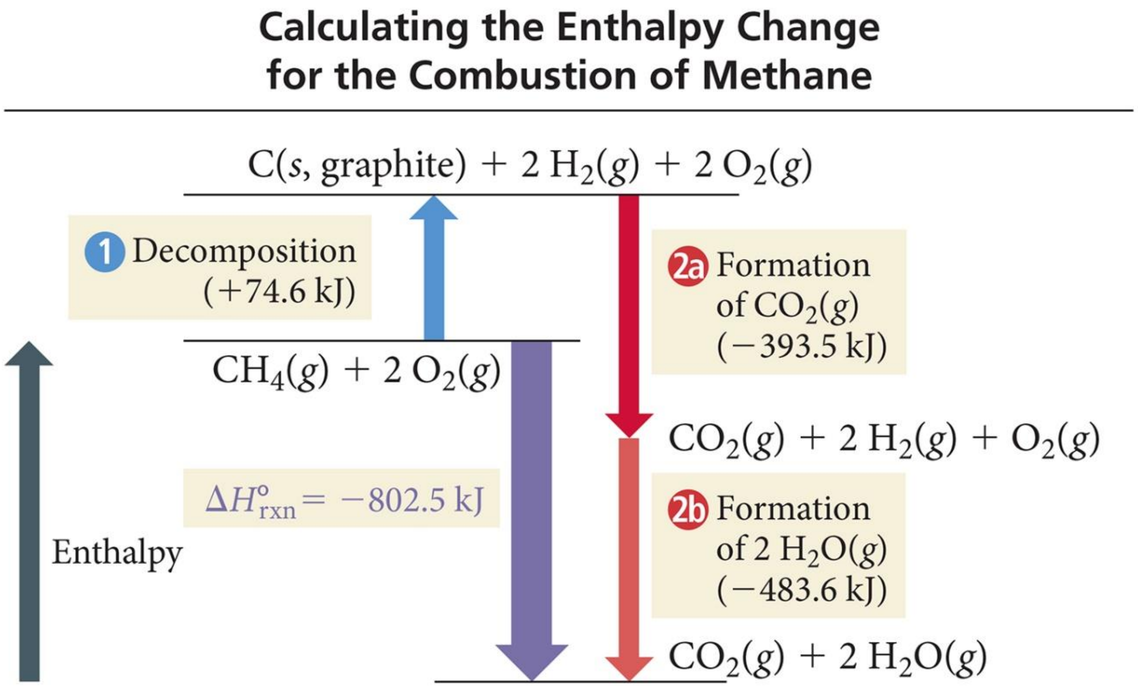


$C(s, \text{graphite}) + 2\ H_2(g) \rightarrow CH_4(g)$ $\Delta H_f° = -\ 74.6\ \text{kJ/mol}\ CH_4$

$C(s, \text{graphite}) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_f° = -393.5\ \text{kJ/mol}\ CO_2$

$H_2(g) + \tfrac{1}{2}\ O_2(g) \rightarrow H_2O(g)$ $\Delta H_f° = -241.8\ \text{kJ/mol}\ H_2O$

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


Calculating the Enthalpy Change for the Combustion of Methane

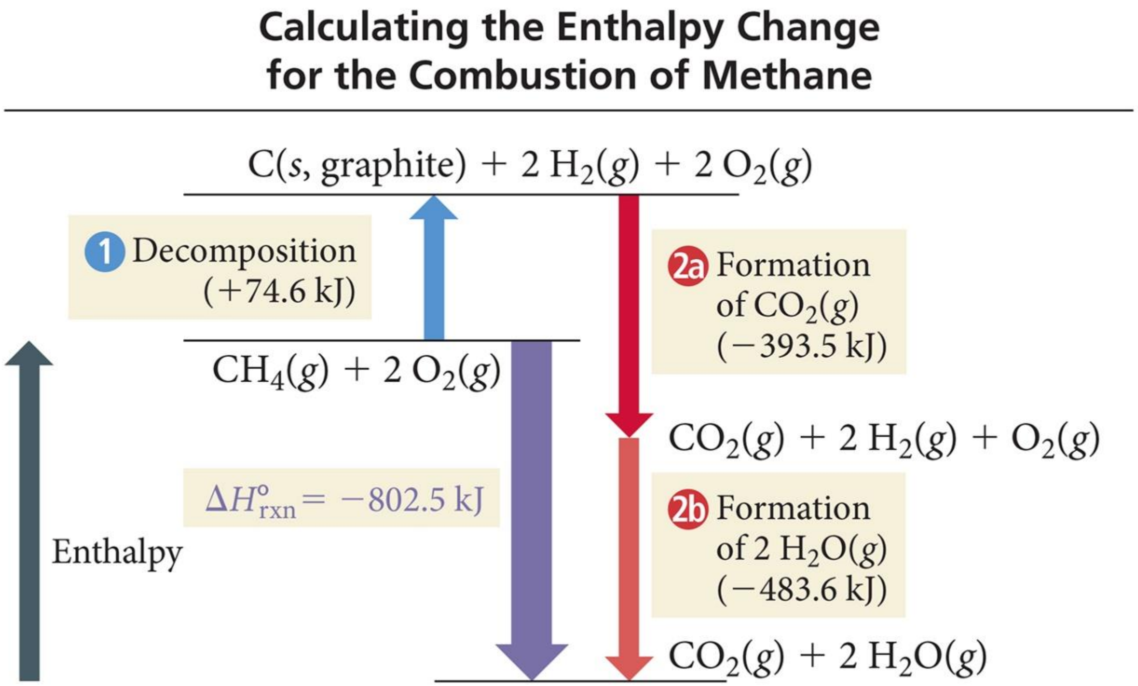
$CH_4(g) \rightarrow \cancel{C(s,\ graphite)} + \cancel{2\ H_2(g)}$ $\Delta H° = +74.6\ \text{kJ}$

$\cancel{C(s,\ graphite)} + O_2(g) \rightarrow CO_2(g)$ $\Delta H_f° = -393.5\ \text{kJ/mol}\ CO_2$

$\cancel{2\ H_2(g)} + O_2(g) \rightarrow 2\ H_2O(g)$ $\Delta H° = -483.6\ \text{kJ}$

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H° = -802.5\ \text{kJ}$

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


Calculating the Enthalpy Change for the Combustion of Methane

$$\Delta H^\circ = [((-393.5\text{ kJ}) + 2(-241.8\text{ kJ})) - ((-74.6\text{ kJ}) + 2(0\text{ kJ}))] = -802.5\text{ kJ}$$

$$\Delta H^\circ = [(\Delta H_f^\circ\ CO_2(g) + 2 \cdot \Delta H_f^\circ\ H_2O(g)) - (\Delta H_f^\circ\ CH_4(g) + 2 \cdot \Delta H_f^\circ\ O_2(g))]$$

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) \Delta H^\circ = -802.5\text{ kJ}$$

Energy Use and the Environment

- In the United States, each person uses over 10^5 kWh of energy per year.

- Most energy comes from the combustion of fossil fuels.
  - Combustible materials that originate from ancient life

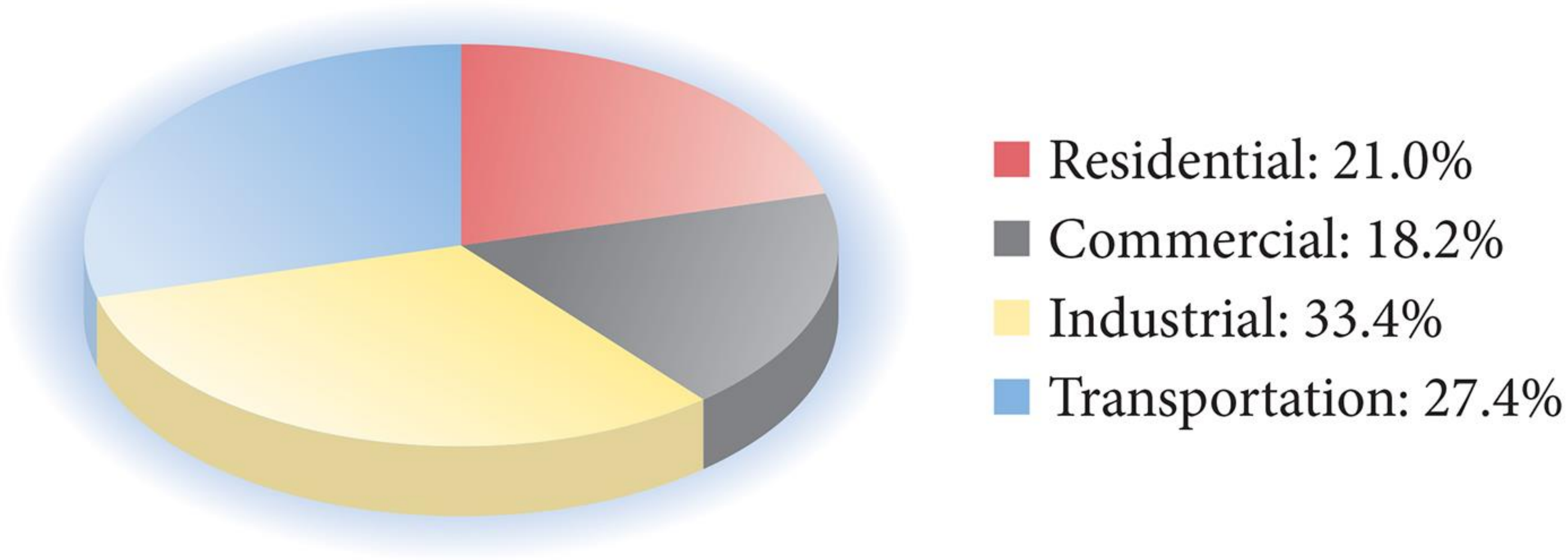
$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H°_{rxn} = -393.5$ kJ

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H°_{rxn} = -802.3$ kJ

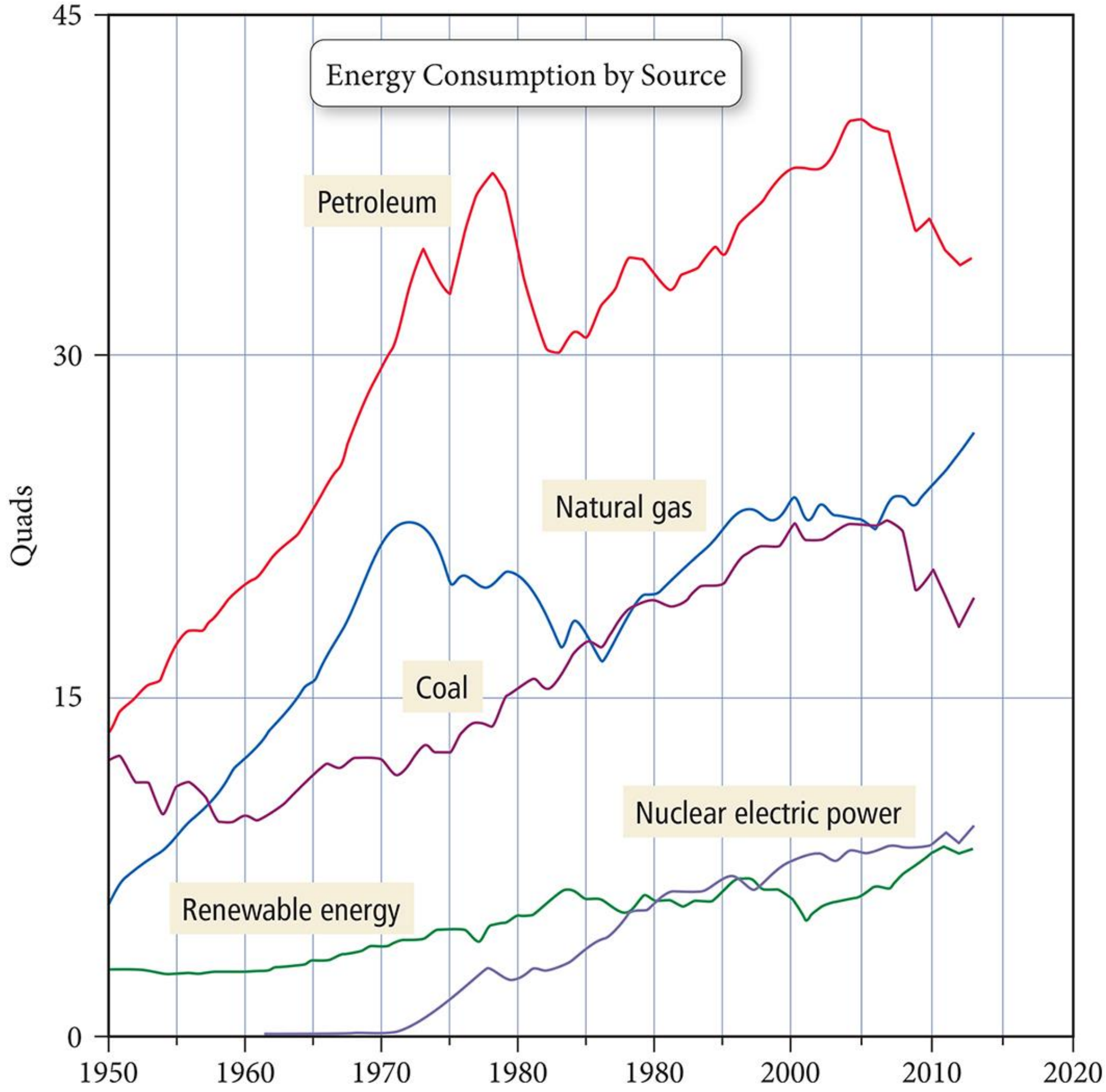
$C_8H_{18}(g) + 12.5\ O_2(g) \rightarrow 8\ CO_2(g) + 9\ H_2O(g)$ $\Delta H°_{rxn} = -5074.1$ kJ

Energy Use and the Environment

- Fossil fuels cannot be replenished.
- At current rates of consumption, oil and natural gas supplies will be depleted in 50–100 years.



Energy Consumption



The Effect of Combustion Products on Our Environment

- Because of additives and impurities in the fossil fuel, incomplete combustion, and side reactions, harmful materials are added to the atmosphere when fossil fuels are burned for energy.
- Therefore, fossil fuel emissions contribute to air pollution, acid rain, and global warming.

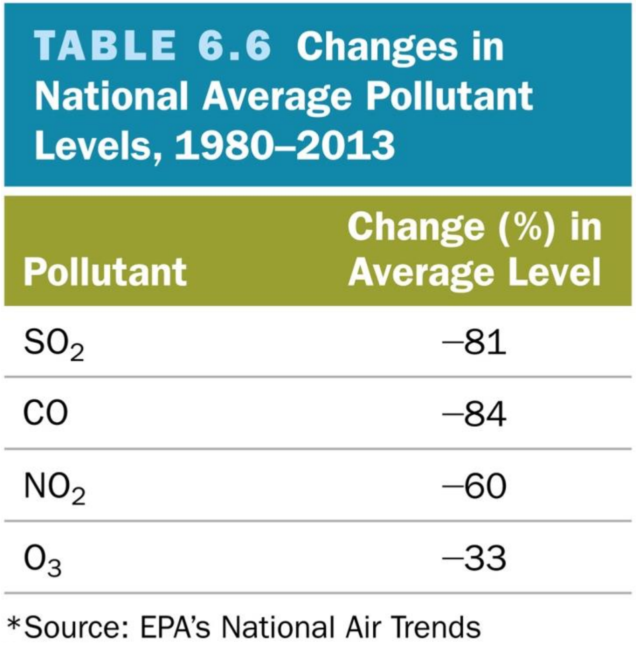


TABLE 6.6 Changes in National Average Pollutant Levels, 1980–2013

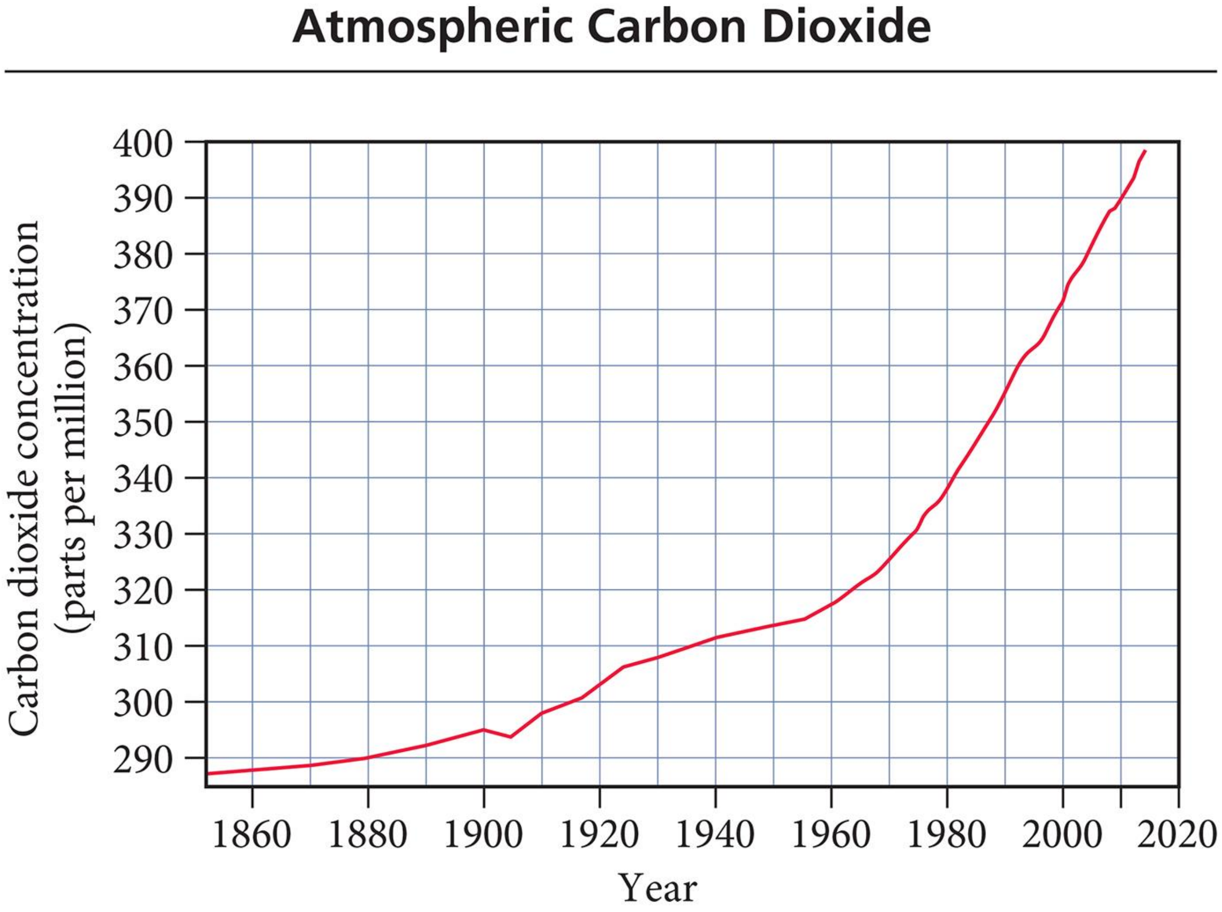
Pollutant	Change (%) in Average Level
SO_2	−81
CO	−84
NO_2	−60
O_3	−33

*Source: EPA's National Air Trends

Global Warming

- CO_2 is a greenhouse gas.
  - It allows light from the sun to reach Earth but does not allow the heat (infrared light) reflected off Earth to escape into outer space.
    - It acts like a blanket.
- CO_2 levels in the atmosphere have been steadily increasing.
- Current observations suggest that the average global air temperature has risen 0.6 °C in the past 100 years.
- Atmospheric models suggest that the warming effect could worsen if CO_2 levels are not curbed.
- Some models predict that the result will be more severe storms, more floods and droughts, shifts in agricultural zones, rising sea levels, and changes in habitats.

Global Warming



Renewable Energy

- Our greatest unlimited supply of energy is the sun.
- New technologies are being developed to capture the energy of sunlight.
  - Parabolic troughs, solar power towers, and dish engines concentrate the sun’s light to generate electricity.
  - Solar energy is used to decompose water into $H_2(g)$ and $O_2(g)$; the H_2 can then be used by fuel cells to generate electricity.

$$H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow H_2O(l) \quad \Delta H^\circ_{\text{rxn}} = -285.8\ \text{kJ}$$

Parabolic troughs

Solar power tower

Dish/engine

Renewable Energy

- Hydroelectric power
- Wind power