Energy, Heat and Temperature

Introduction

3 basic types of energy:

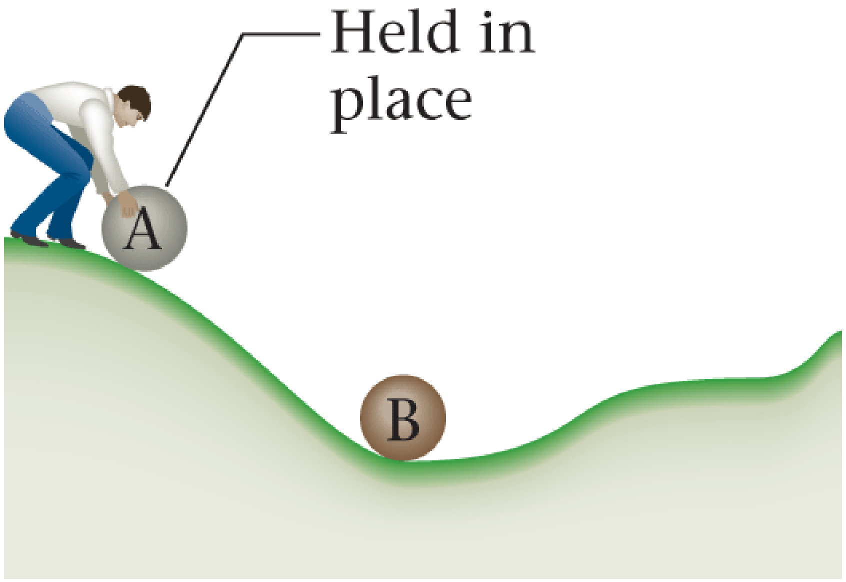
- **Potential** (possibility of doing work because of composition or position)
- **Kinetic** (moving objects doing work)
- Radiant (energy transferred by electromagnetic waves)

A. The Nature of Energy

- Energy is the ability to do work or produce heat.

Potential energy

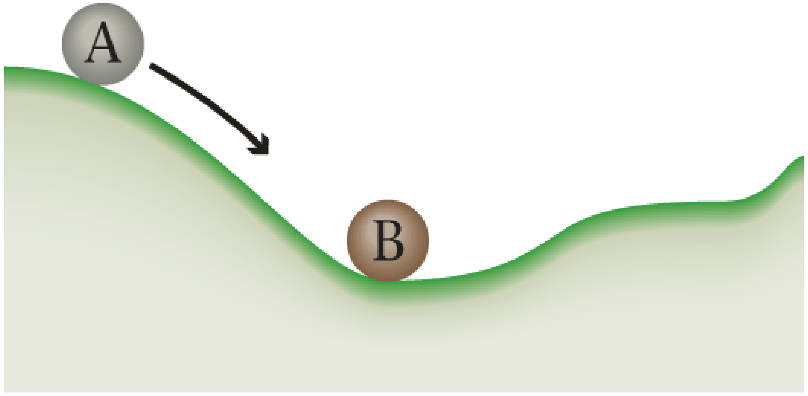
Energy of position



Kinetic energy

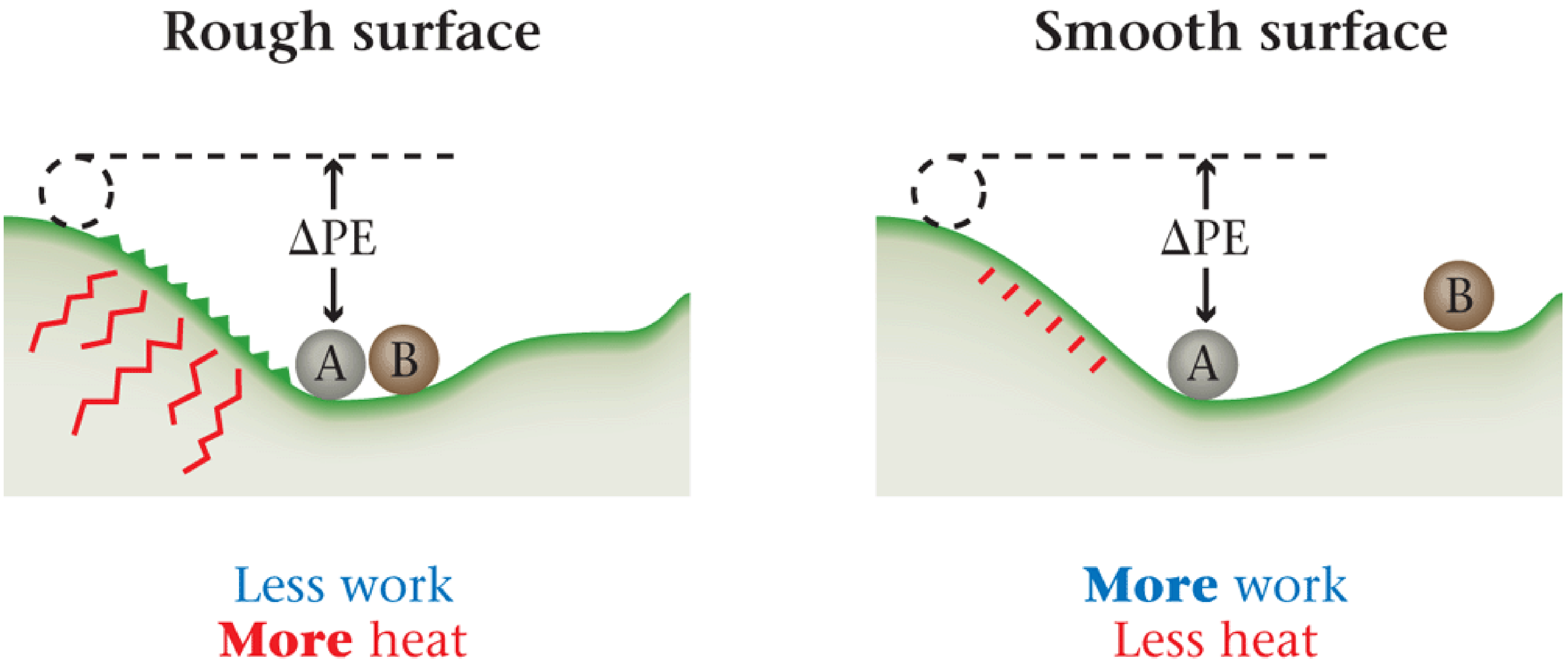
Energy of motion

E = ½ mv2



A. The Nature of Energy

- Law of conservation of energy
  - Energy can be converted from one form to another but can neither be created or destroyed.



The Law of Conservation of Energy

- **States that in any process, *energy cannot be created or destroyed***
  - The SI unit of energy is the **joule** (**J**)
  - You may also encounter the **calorie** (**cal**)
  - The dietary Calorie (note upper case letter), **Cal**, is actually 1 kilocalorie

Thermodynamics

- Internal energy, E – sum of kinetic and potential energies of all the “particles” in a system
  - Internal energy can be changed by two types of energy flow:
    - Heat (q)
    - Work (w)

$$\Delta E = q + w$$

- A **2 kg** object moving at **1 meter per second** has **1 J** of kinetic energy

- 1 calorie is defined as the amount of energy required to raise 1 gram of water by 1 degree Celsius

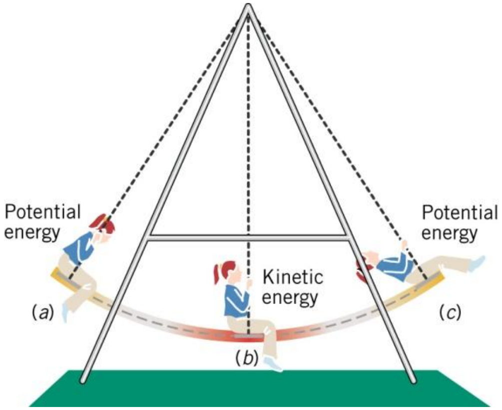
$$1\,\mathrm{J} = \tfrac{1}{2}\,(2\,\mathrm{kg})\left(\frac{1\,\mathrm{m}}{1\,\mathrm{s}}\right)^2 = 1\,\mathrm{kg}\,\mathrm{m}^2\,\mathrm{s}^{-2}$$

$$1\,\mathrm{cal} = 4.184\,\mathrm{J}\ \text{(exactly)}$$

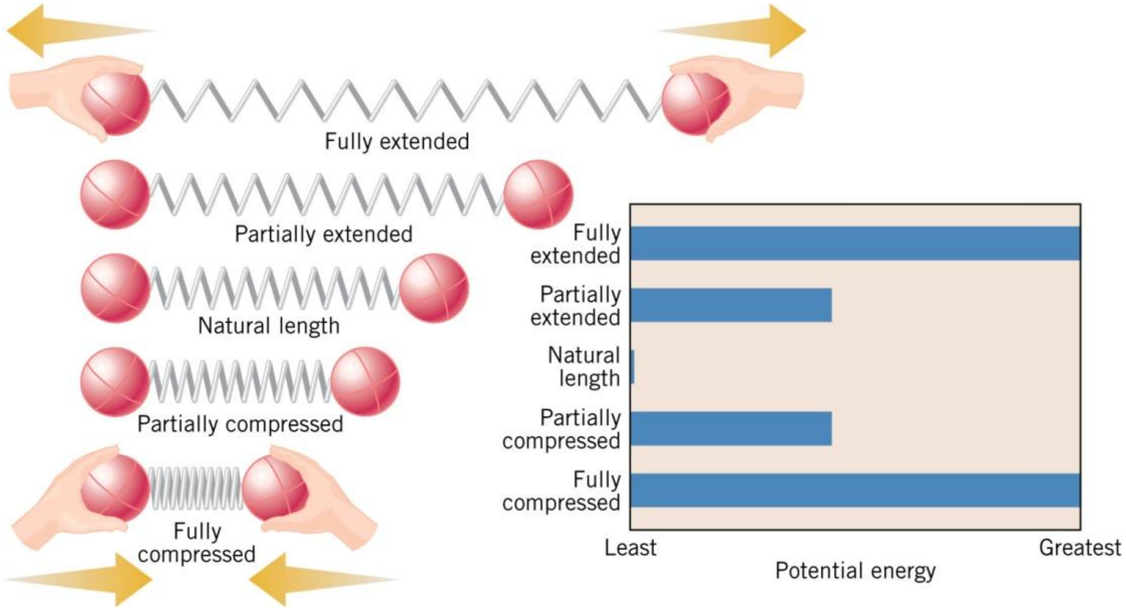
Units of Energy

- SI unit is the Joule, J
  - J = kg·m/s^2
  - If the calculated value is greater than 1000 J, use the kJ
- Another unit is the calorie, cal
  - cal = 4.184 J (exact)
- Nutritional unit is the Calorie (note capital C), which is one kilocalorie
  - 1 Cal = 1 kcal = 4.184 kJ

- KE can be converted into PE and vice versa



When the child is at points (a) and (c) they have only PE; at point (b) only KE. Total energy is conserved (PE + KE = constant).



When fully compressed or extended only PE; at natural length only KE. Total energy is conserved.

Internal Energy is Conserved

- ***1st Law of Thermodynamics***: For an isolated system the internal energy (*E*) is constant:

$$\Delta E = E_f - E_i = 0$$

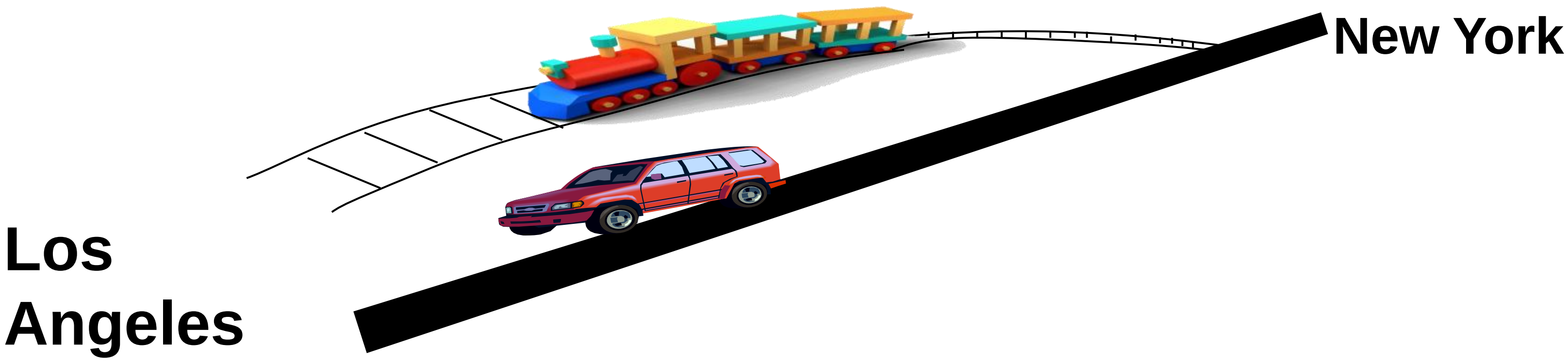
$$\Delta E = E_{\text{product}} - E_{\text{reactant}} = 0$$

- We can’t measure the internal energy of anything, so we measure the changes in energy

- *E* is a state function

State Function

- A property whose value depends only on the present state of the system, not on the method or mechanism used to arrive at that state
- Position is a state function: both train and car travel to the same locations although their paths vary
- The actual distance traveled does vary with path



Heat and Temperature are Not the Same

- The ***temperature*** of an object is proportional to the **average** kinetic energy of its particles—the higher the average kinetic energy, the higher the temperature
- ***Heat*** is energy (also called **thermal energy**) transferred between objects caused by differences in their temperatures until they reach ***thermal equilibrium***

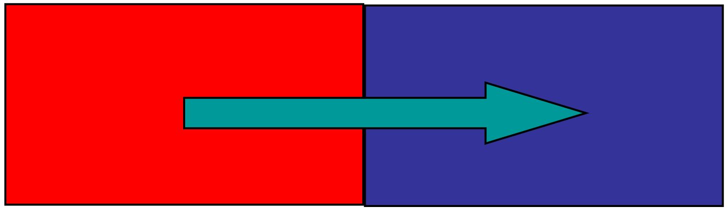
Heat Transfer is a State Function

- Transfer of heat during a reaction is a state function.
- The route taken to arrive at the products does not affect the amount of heat that is transferred.
- The number of steps does not affect the amount of heat that is transferred.

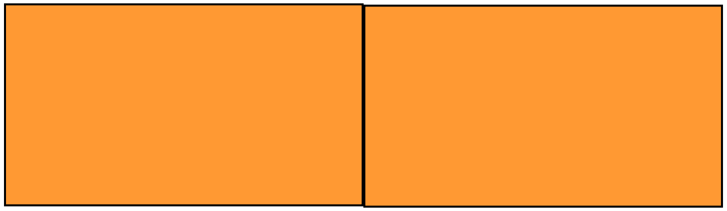
Heat (*q*)

Heat is energy transfer resulting from thermal differences between the system and surroundings

“flows” spontaneously from **higher *T*** → **lower *T***

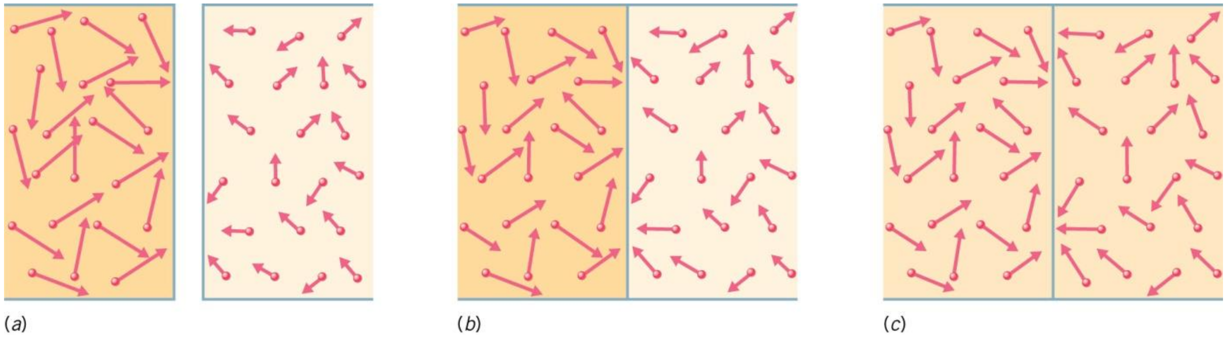


“flow” ceases at *thermal equilibrium*



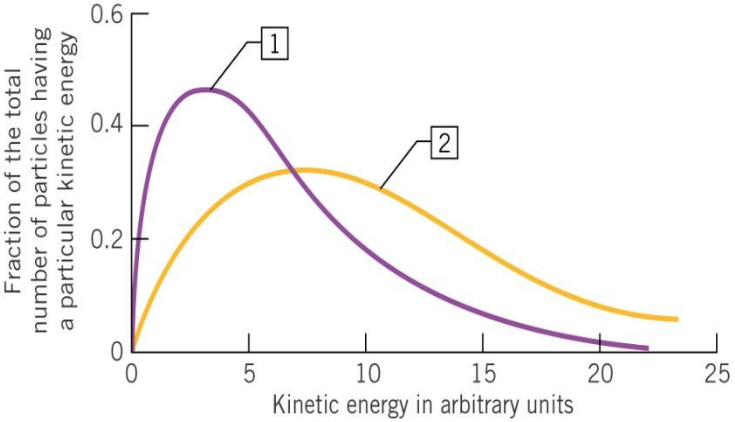
- The energy that is transferred as heat comes from the object’s **internal energy**
- The energy associated with the motion of the object’s molecules is referred to as its **molecular kinetic energy**
- The temperature of an object is related to the average kinetic energy of its atoms and molecules

- Heat is a transfer of energy due to a temperature difference



a) Isolated warm (left) and cold (right) objects
b) Thermal contact is made: thermal energy is transferred from left to right
c) Thermal equilibrium: the same average KE for molecules in both objects

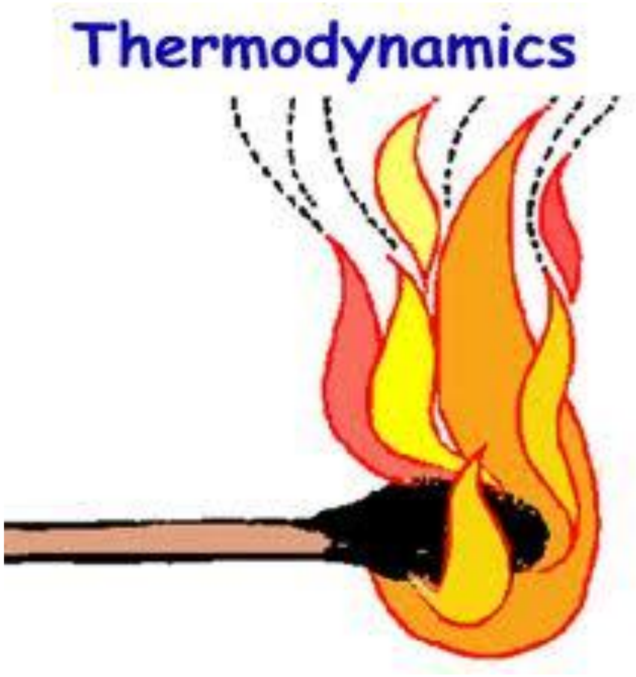
- The temperature of an object is related to the average kinetic energy of its atoms and molecules



The temperature for curve (1) is lower than for curve (2) because the average kinetic energy is lower.

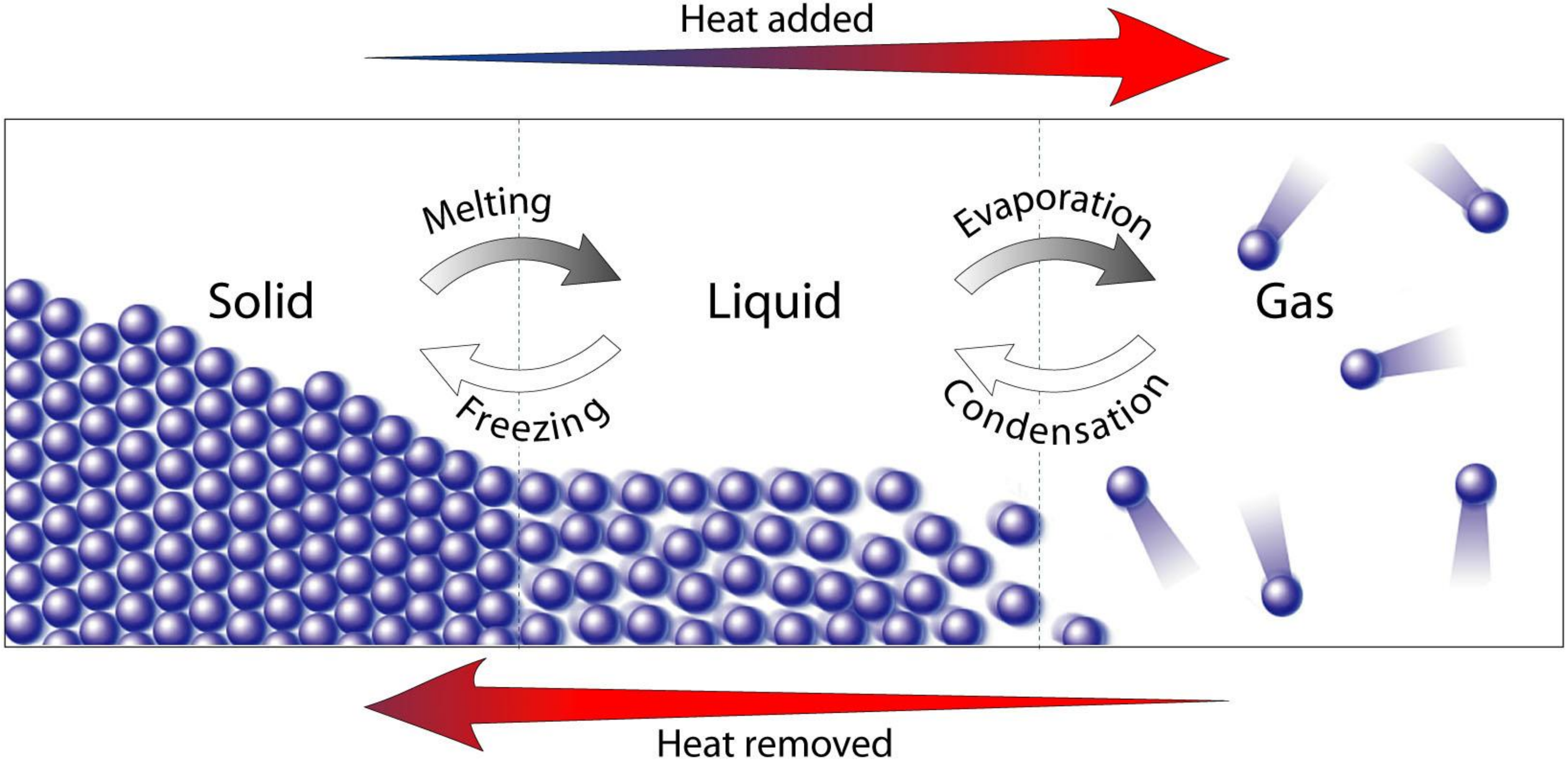
Energy Changes

- Changes in energy can result in chemical and physical changes in matter.



- http://www.childrensbiblestudy.com/nolife/Thermodynamics.htm

Molecules on the move.



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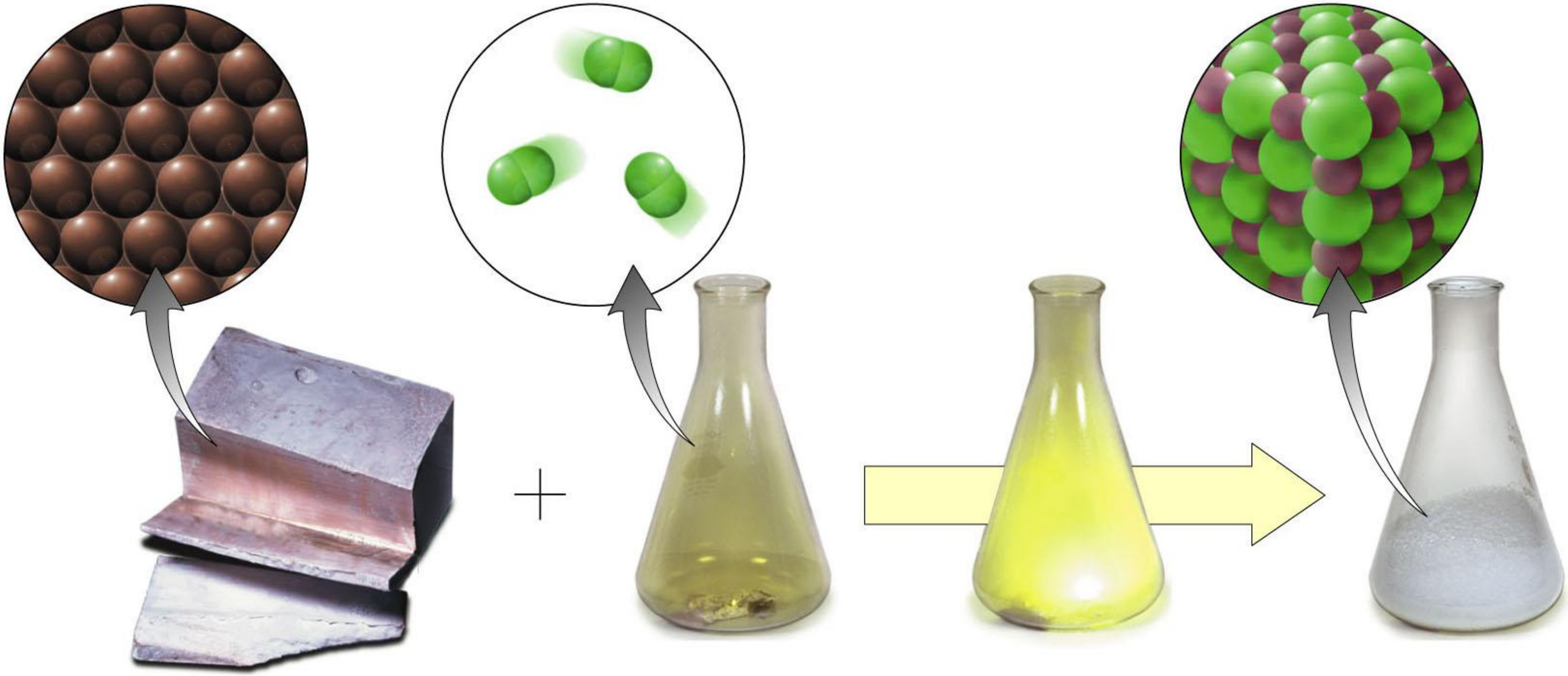

Sodium metal and chlorine gas react to form sodium chloride

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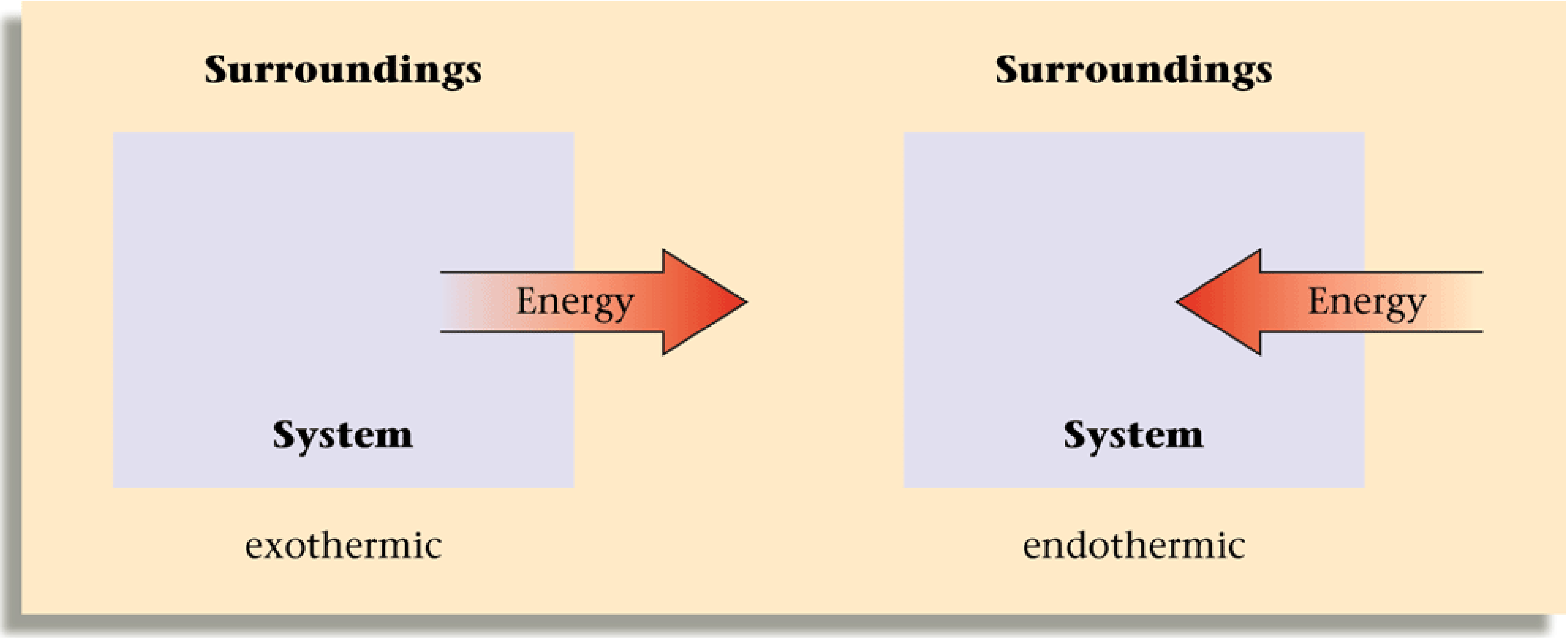
Surroundings / System / Universe

- ***System*** - the reaction or area under study
- ***Surroundings*** - the rest of the universe
- ***Open systems*** can gain or lose mass and energy across their boundaries
  - i.e. the human body
- ***Closed systems*** can absorb or release energy, but not mass, across the boundary
  - i.e. a light bulb
- ***Isolated systems*** (***adiabatic***) cannot exchange matter or energy with their surroundings
  - i.e. a closed Thermos bottle

C. Exothermic and Endothermic Processes

- Exothermic – energy flows out of the system
- Endothermic – energy flows into the system



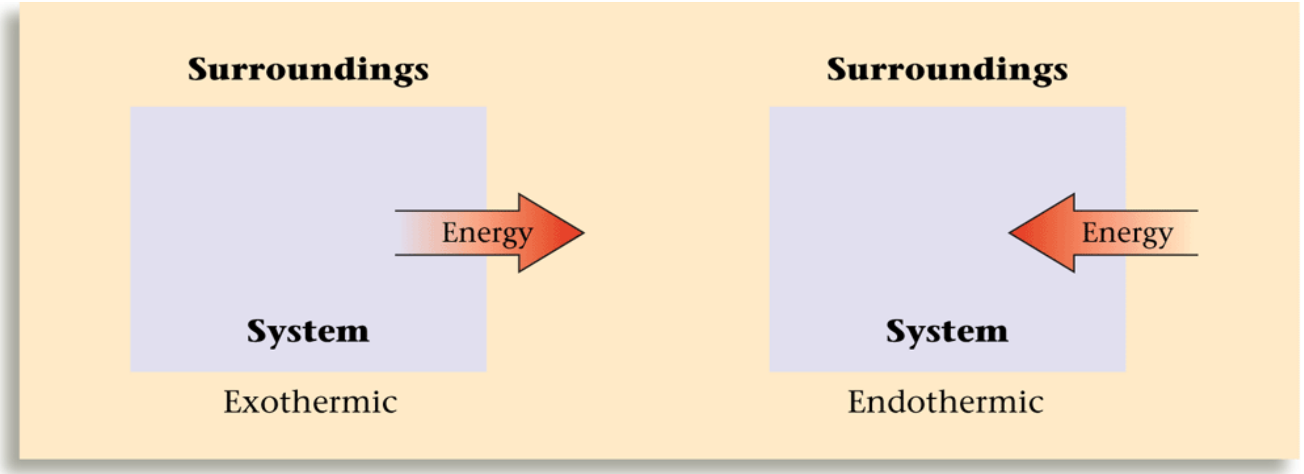
The Sign Convention

- ***Endothermic*** systems require energy to be added to the system, thus the *q* is (+)
- ***Exothermic*** reactions release energy to the surroundings. Their *q* is (-)
- Energy changes are measured from the point of view of the system

- ($\Delta E = q + w$)

Thermodynamics

- Thermodynamic quantities always consist of 2 parts:
  - A number (magnitude of the change)
  - A sign (indicates the direction of flow)
    - Reflects the systems point of view

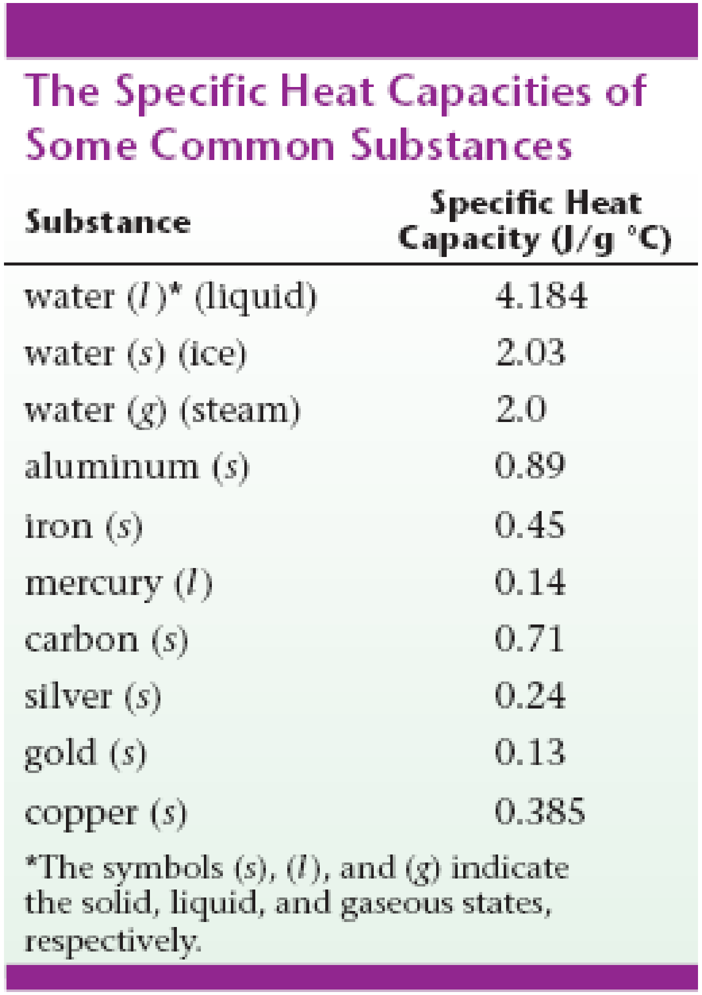


- $q = -x$
- negative $q \rightarrow$ system's energy decreases
- $\Delta E < 0$

- $q = +x$
- positive $q \rightarrow$ system's energy increases
- $\Delta E > 0$

B. Measuring Energy Changes

- Specific heat capacity is the energy required to change the temperature of a mass of one gram of a substance by one Celsius degree.



The Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity (J/g °C)
water (l)* (liquid)	4.184
water (s) (ice)	2.03
water (g) (steam)	2.0
aluminum (s)	0.89
iron (s)	0.45
mercury (l)	0.14
carbon (s)	0.71
silver (s)	0.24
gold (s)	0.13
copper (s)	0.385

*The symbols (s), (l), and (g) indicate the solid, liquid, and gaseous states, respectively.

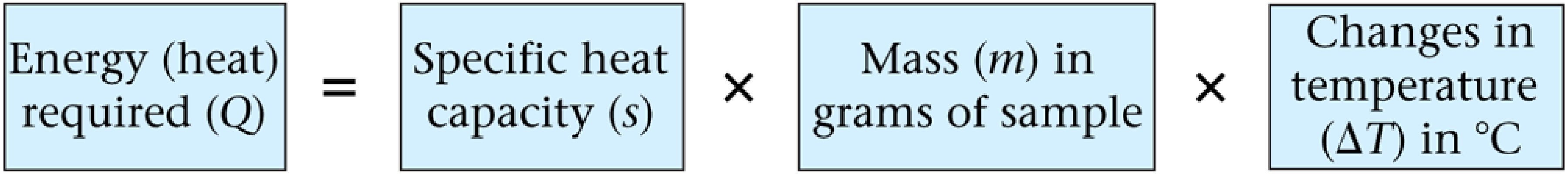
Specific Heats

- Substances with high specific heats resist temperature changes
- Note that water has a very high specific heat
  - This is why coastal temperatures are different from inland temperatures.

Substance	Specific Heat J g^{-1} C^{-1} (25 C)
Carbon (graphite)	0.711
Copper	0.387
Ethyl alcohol	2.45
Gold	0.129
Granite	0.803
Iron	0.4498
Lead	0.128
Olive oil	2.0
Silver	0.235
Water (liquid)	4.18

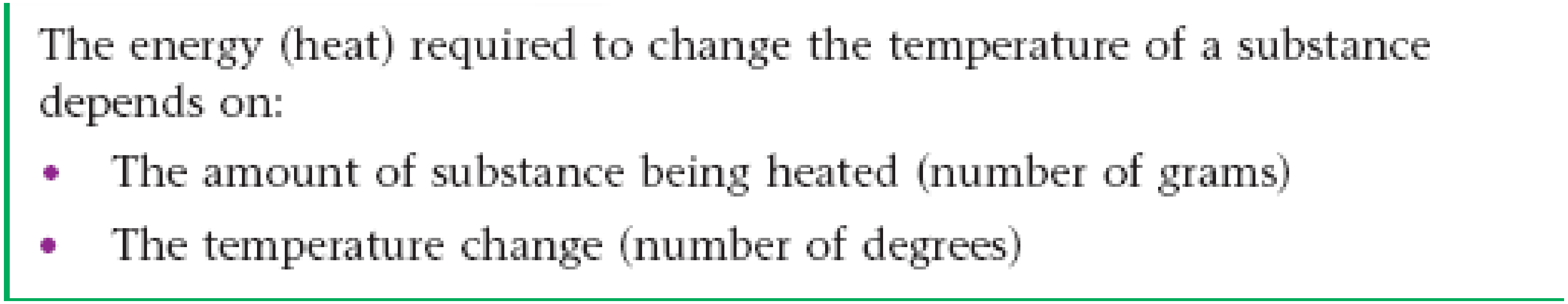
B. Measuring Energy Changes

- To calculate the energy required for a reaction:



Energy (heat) required (Q) = Specific heat capacity (*s*) × Mass (*m*) in grams of sample × Changes in temperature (ΔT) in °C

$$Q = s \times m \times \Delta t$$



The energy (heat) required to change the temperature of a substance depends on:

- The amount of substance being heated (number of grams)
- The temperature change (number of degrees)

Your Turn!

A cast iron skillet is moved from a hot oven to a sink full of water. Which of the following is not true?

A. The water heats
B. The skillet cools
C. The heat transfer for the skillet has a (-) sign
D. The heat transfer for the skillet is the same as the heat transfer for the water
E. None of these are untrue

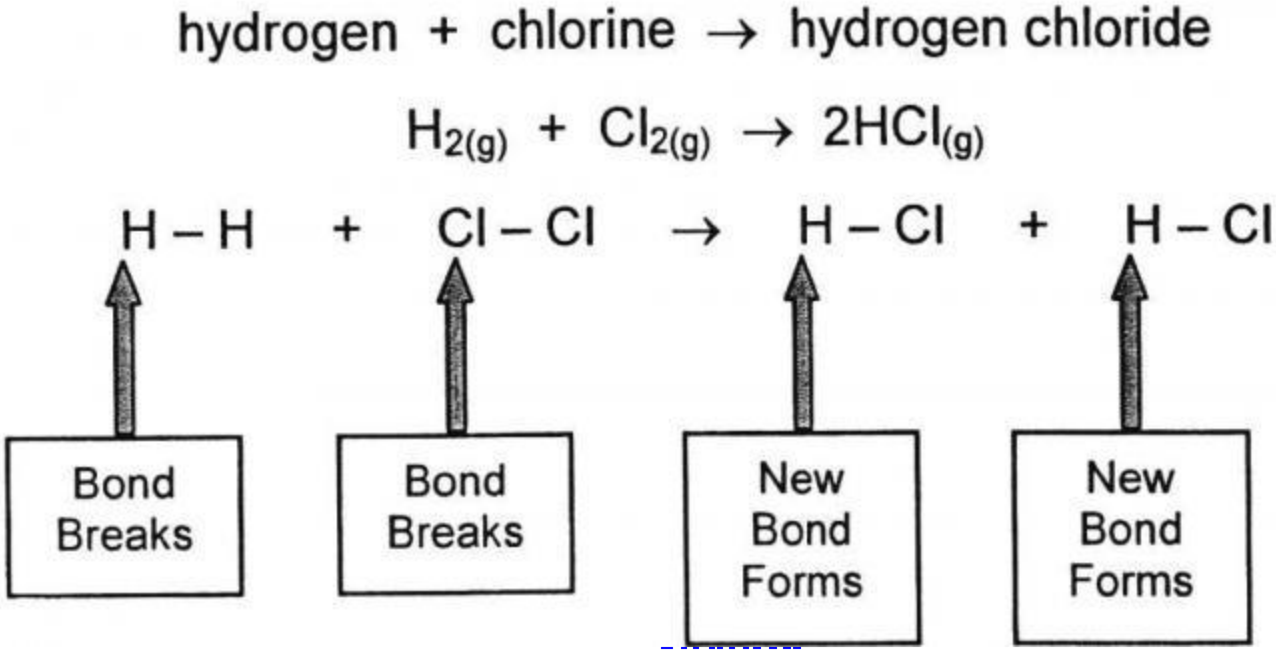
Learning Check

Calculate the specific heat of a metal if it takes 235 J to raise the temperature of a 32.91 g sample by 2.53 °C.

$$q = m \times s \times \Delta t$$

$$s = \frac{q}{m \times \Delta t} = \frac{235\ \mathrm{J}}{32.91\ \mathrm{g} \times 2.53\ ^\circ\mathrm{C}} = 2.82\,\frac{\mathrm{J}}{\mathrm{g}\ ^\circ\mathrm{C}}$$

Graphs of Energy Changes in a Chemical Reaction



hydrogen + chlorine → hydrogen chloride

$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

$H - H \quad + \quad Cl - Cl \quad \rightarrow \quad H - Cl \quad + \quad H - Cl$

Bond Breaks

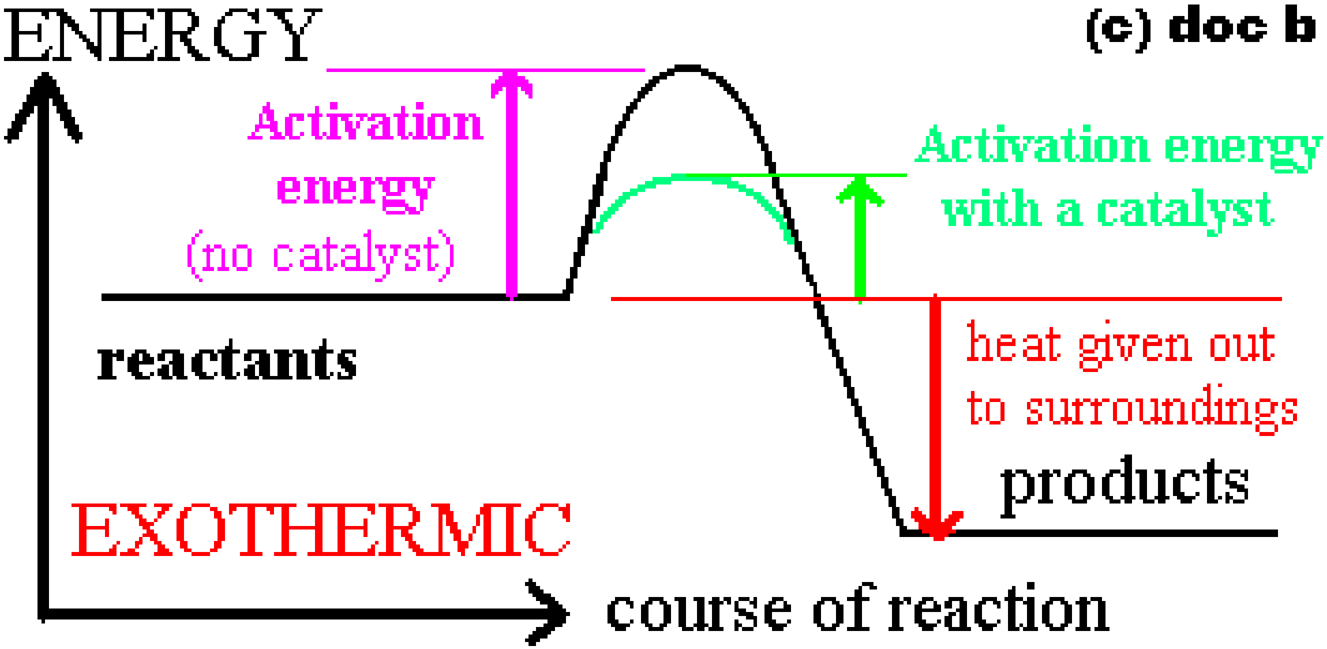
Bond Breaks

New Bond Forms

New Bond Forms

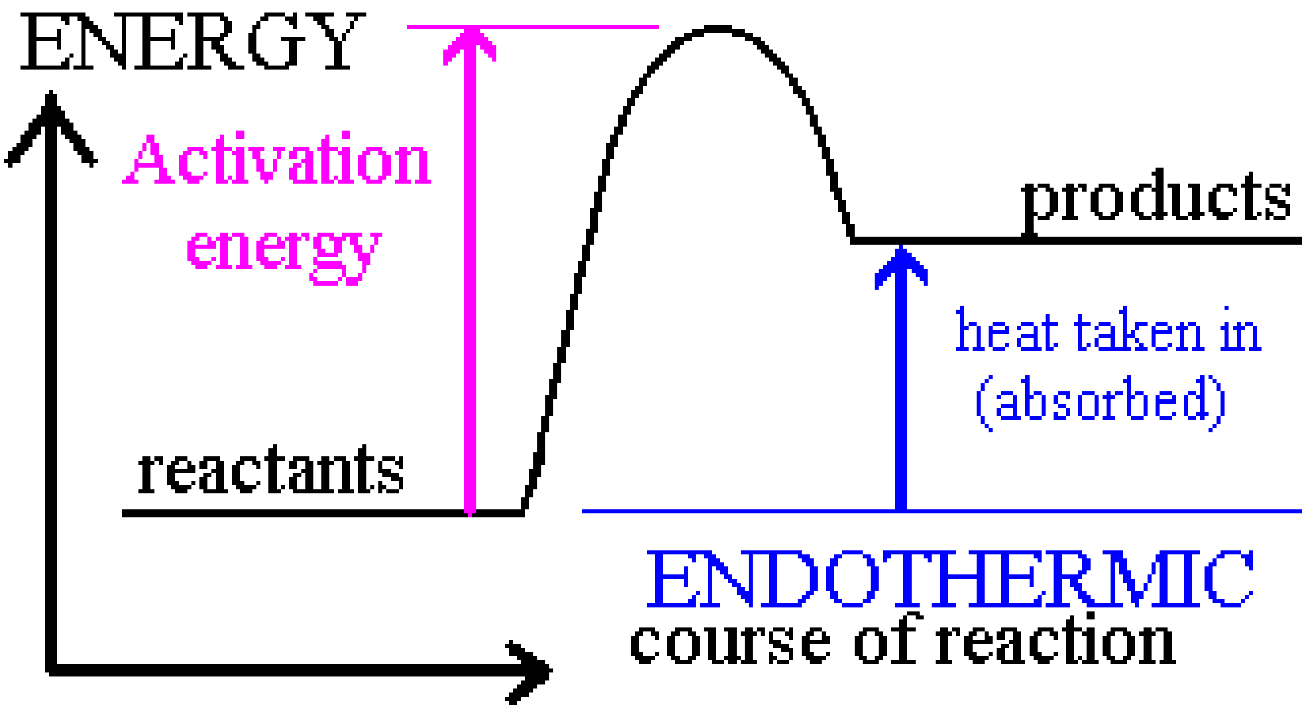
Exothermic Reactions

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + Energy

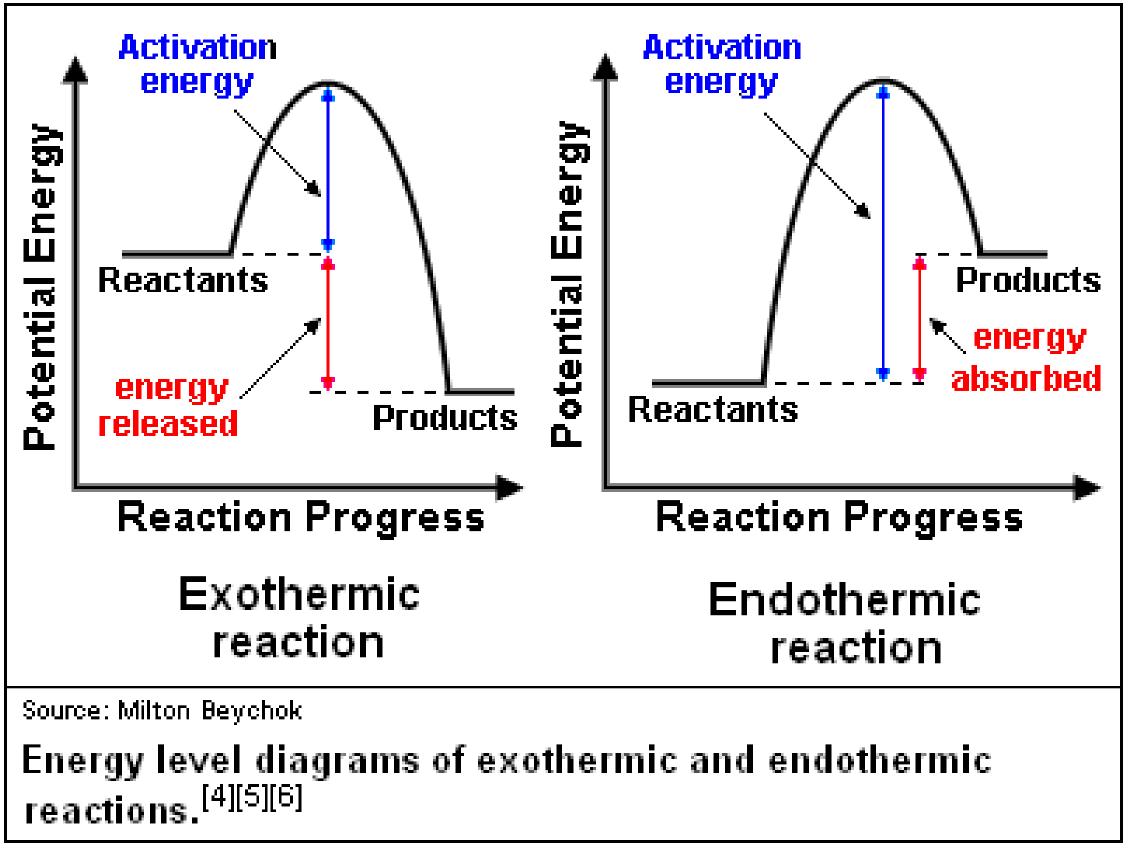


Endothermic Reactions

$6CO_2 + 6H_2O +$ Energy $\rightarrow 6O_2 + C_6H_{12}O_6$



Comparing Exothermic and Endothermic Reaction Energy Graphs



Source: Milton Beychok

Energy level diagrams of exothermic and endothermic reactions.[4][5][6]

Thermochemical Equations

- Relate the energy of a reaction to the quantities involved
- Must be balanced, but may use fractional coefficients
- Quantities are presumed to be in moles
- Example:

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

Enthalpy Diagram

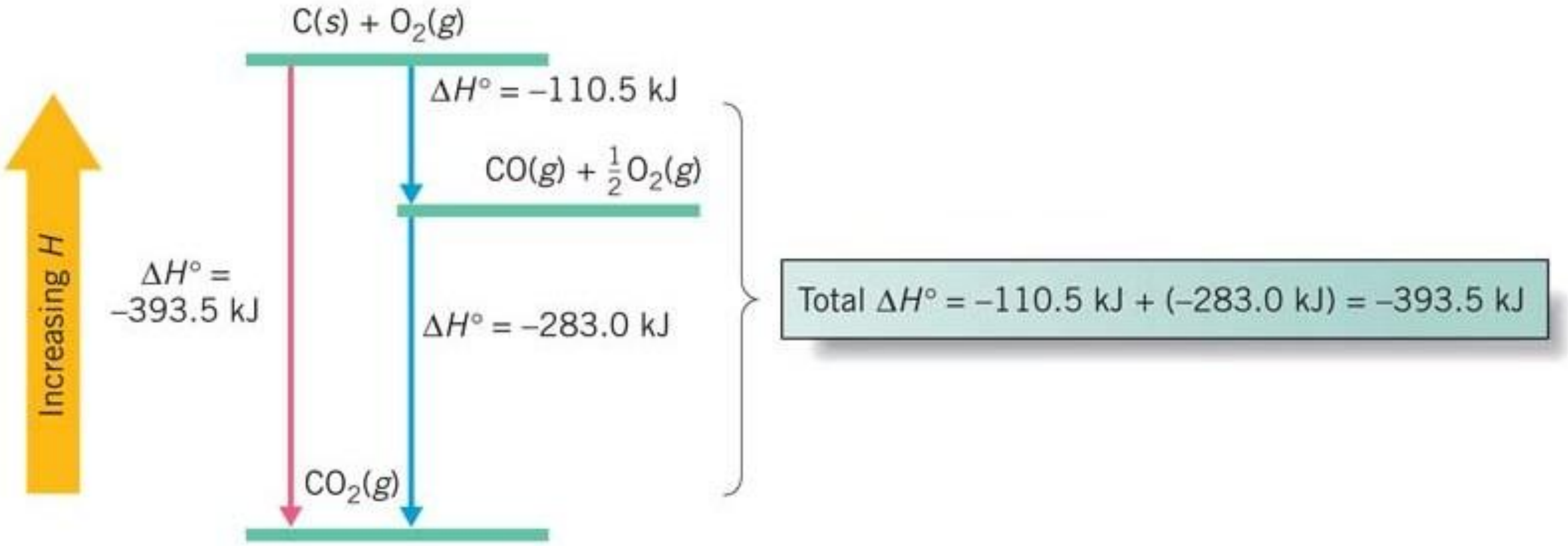


FIG. 6.9 **An enthalpy diagram for the formation of CO_2(*g*) from its elements by two different paths.** On the left is path 1, the direct conversion of C(*s*) and O_2(*g*) to CO_2(*g*). On the right, path 2 shows two shorter, downward pointing arrows. The first step of path 2 takes the elements to CO(*g*), and the second step takes CO(*g*) to CO_2(*g*). The overall enthalpy change is identical for both paths, as it must be, because enthalpy is a state function.

B. Hess’s Law

- For a particular reaction, the change in enthalpy is the same whether the reaction takes place in one step or a series of steps.
- Example:

$$\begin{array}{llr}
N_2(g) + O_2(g) \rightarrow 2NO(g) & \Delta H_2 & = 180\text{ kJ} \\
2NO(g) + O_2(g) \rightarrow 2NO_2(g) & \Delta H_3 & = -112\text{ kJ} \\
\hline
N_2(g) + 2O_2(g) \rightarrow 2NO_2(g) & \Delta H_2 + \Delta H_3 & = 68\text{ kJ}
\end{array}$$

Hess's Law

The overall enthalpy change for a reaction is equal to the sum of the enthalpy changes for individual steps in the reaction

For example:

$2Fe(s) + 3/2O_2(g) \rightarrow Fe_2O_3(s) \qquad \Delta H = -822.2\ kJ$

$Fe_2O_3(s) + 2Al(s) \rightarrow Al_2O_3(s) + 2Fe(s) \qquad \Delta H = -848\ kJ$

$3/2O_2(g) + 2Al(s) \rightarrow Al_2O_3(s) \quad \Delta H = -822.2\ kJ + -848\ kJ$

6.7 Thermochemical equations can be combined because enthalpy is a state function

-1670 kJ

Rules for Adding Thermochemical Reactions

1. When an equation is reversed—written in the opposite direction—the sign of *H* must also be reversed.
2. Formulas canceled from both sides of an equation must be for the substance in identical physical states.
3. If all the coefficients of an equation are multiplied or divided by the same factor, the value of *H* must likewise be multiplied or divided by that factor.

Strategy for Adding Reactions Together

1. Choose the most complex compound in the equation (1)
2. Choose the equation (2 or 3 or…) that contains the compound
3. Write this equation down so that the compound is on the appropriate side of the equation and has an appropriate coefficient for our reaction
4. Look for the next most complex compound

Hess’s Law (Cont.)

5. Choose an equation that allows you to cancel intermediates and multiply by an appropriate coefficient
6. Add the reactions together and cancel like terms
7. Add the energies together, modifying the enthalpy values in the same way that you modified the equation
   - If you reversed an equation, change the sign on the enthalpy
   - If you doubled an equation, double the energy

Learning Check

How can we calculate the enthalpy change for the reaction $2\ H_2(g) + N_2(g) \rightarrow N_2H_4(g)$ using these equations?

$N_2H_4(g) + H_2(g) \rightarrow 2NH_3(g) \qquad \Delta H° = -187.8\ \text{kJ}$

$3H_2(g) + N_2(g) \rightarrow 2NH_3(g) \qquad \Delta H° = -92.4\ \text{kJ}$

Reverse the first reaction (and change sign)

$2NH_3(g) \rightarrow N_2H_4(g) + H_2(g) \qquad \Delta H° = +187.8\ \text{kJ}$

Add the second reaction (and add the enthalpy)

$3H_2(g) + N_2(g) \rightarrow 2NH_3(g) \qquad \Delta H° = -92.4\ \text{kJ}$

$\cancel{2NH_3}(g) + \overset{2}{\cancel{3}}H_2(g) + N_2(g) \rightarrow N_2H_4(g) + \cancel{H_2(g)} + \cancel{2NH_3}(g)$

$2H_2(g) + N_2(g) \rightarrow N_2H_4(g) \qquad (187.8 - 92.4) = +95.4\ \text{kJ}$