Introductory Chemistry

Nivaldo J. Tro

Chapter 3

Matter and Energy

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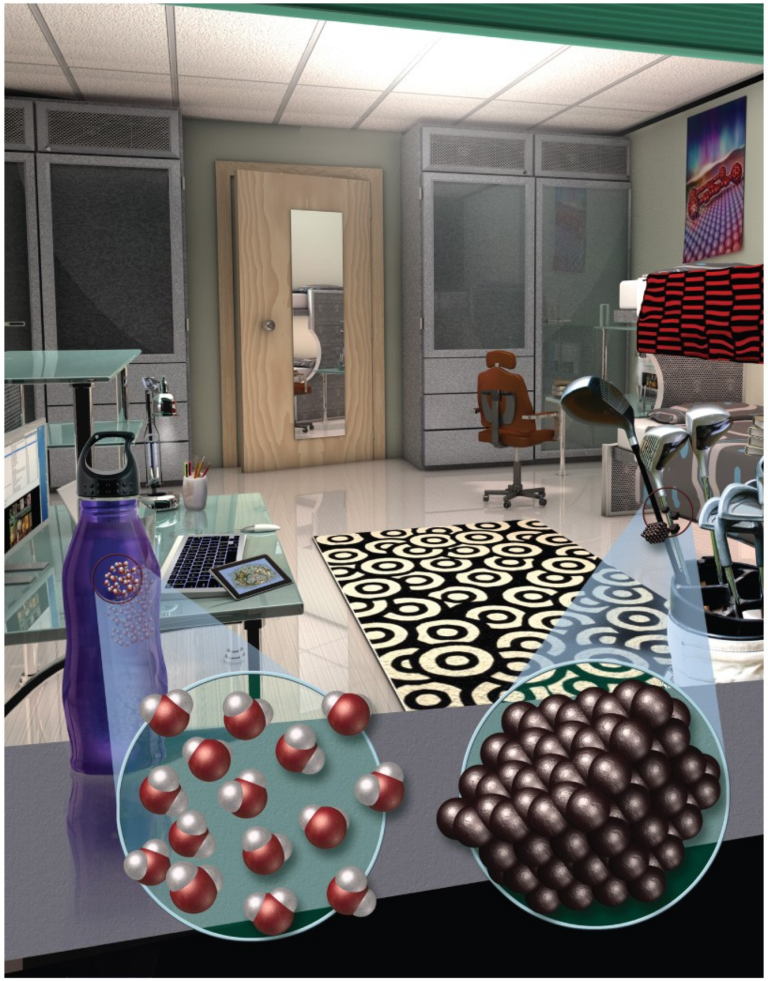
Updated January 25, 2025)

Lecture material with self-check tutorial videos !

3.1 In Your Room

- Everything that you can see in this room is made of **matter.**
- Different kinds of matter are related to the **differences between the molecules and atoms** that compose the matter.
- **The molecular structures**:
- Water molecules on the left
- Carbon atoms in graphite on the right.



3.2 What Is Matter?

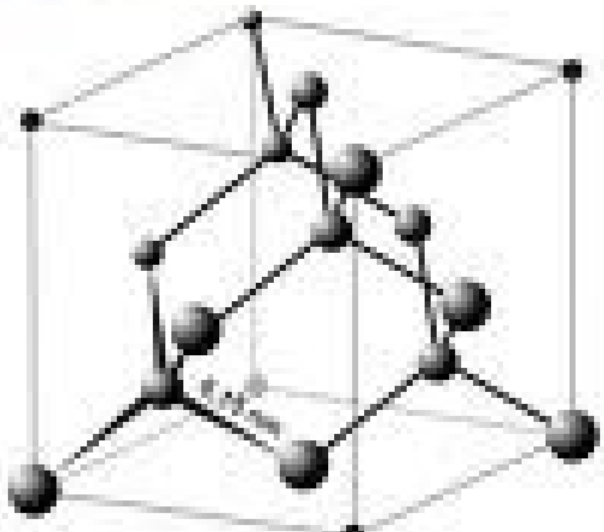
- **Matter** is defined as anything that **occupies space and has mass**.
- **Some types of matter**—such as steel, water, wood, and plastic—are easily visible to our eyes.
- **Other types of matter**—such as air or microscopic dust—are impossible to see without magnification.

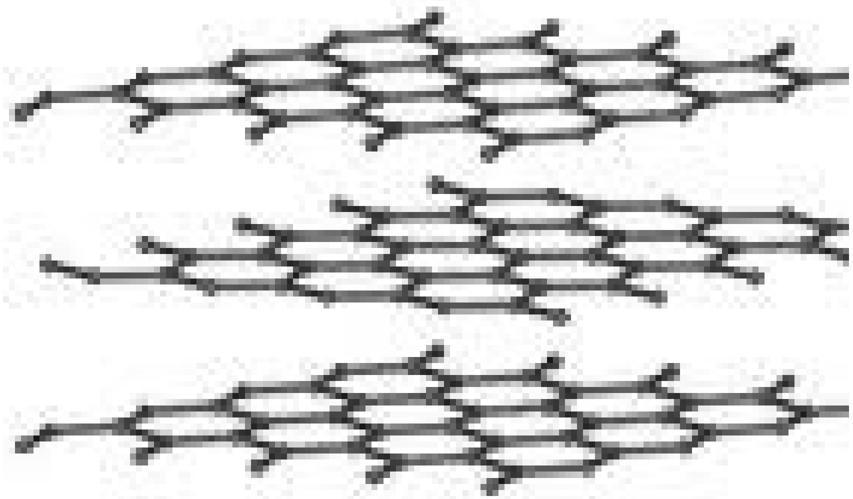
3.2 What Is Matter?

- Matter may appear smooth and continuous, but actually it is not.
- Matter is ultimately composed of **atoms**, submicroscopic particles that are the fundamental building blocks of matter.
- In many cases, these atoms are bonded together to form **molecules**, two or more atoms joined to one another in **specific geometric arrangements**.
- Recent advances in microscopy have allowed us to image the atoms and molecules that compose matter.

Compare two materials made of carbon atom:

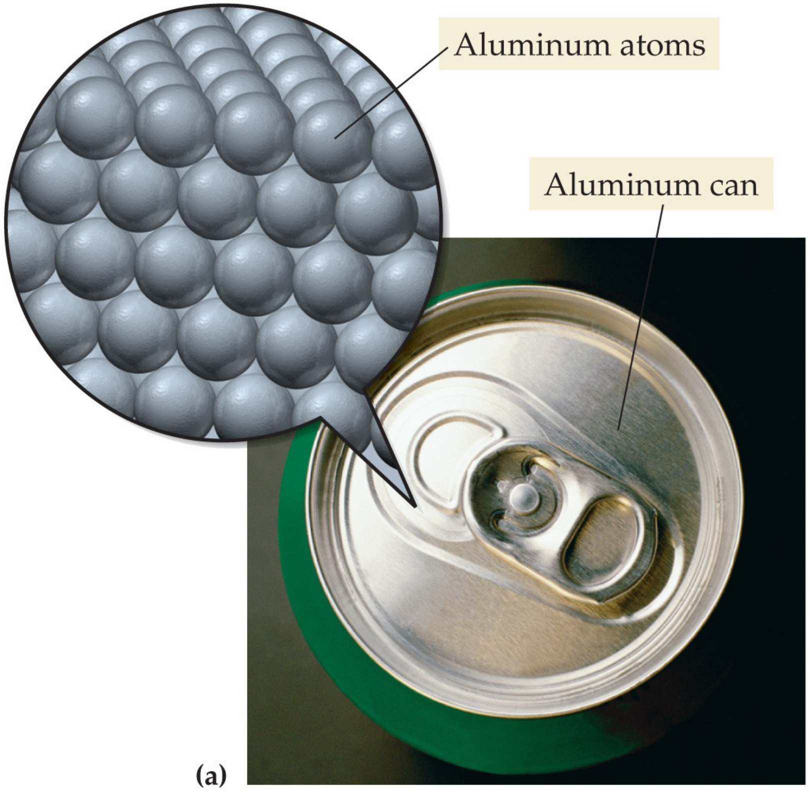
Diamond and Graphite





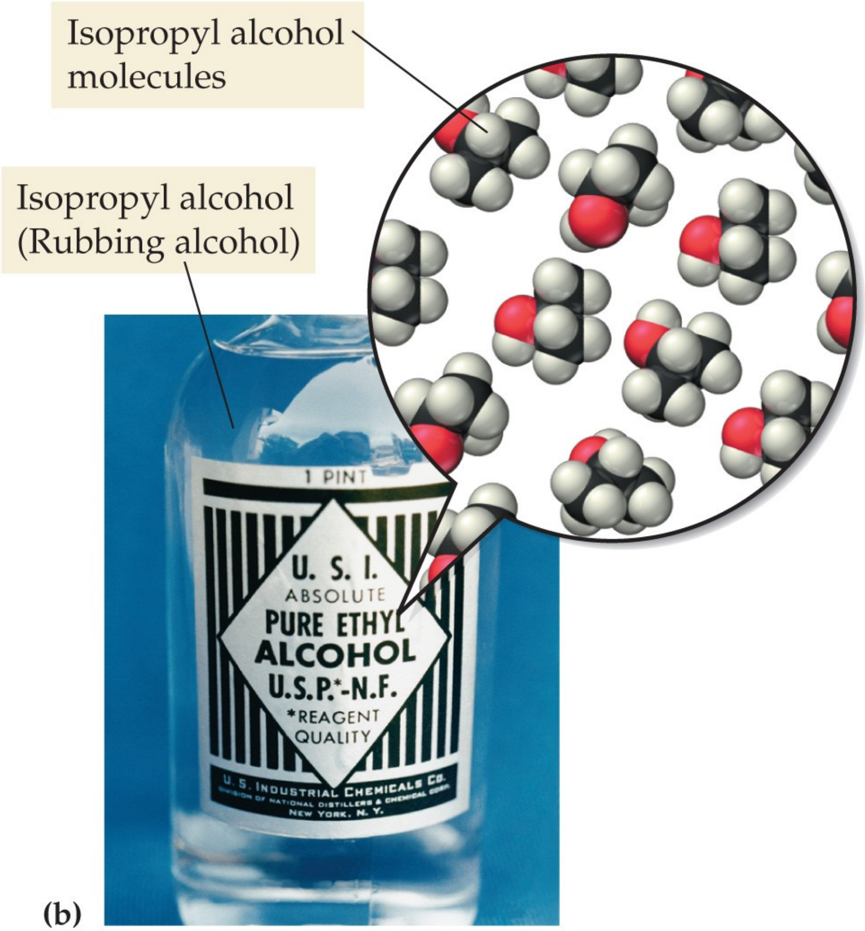
3.2 What Is Matter?

- **FIGURE 3.1 Atoms and molecules**
- All matter is ultimately composed of atoms. **(a)** In some substances, such as aluminum, the atoms exist as independent particles.



3.2 What Is Matter?

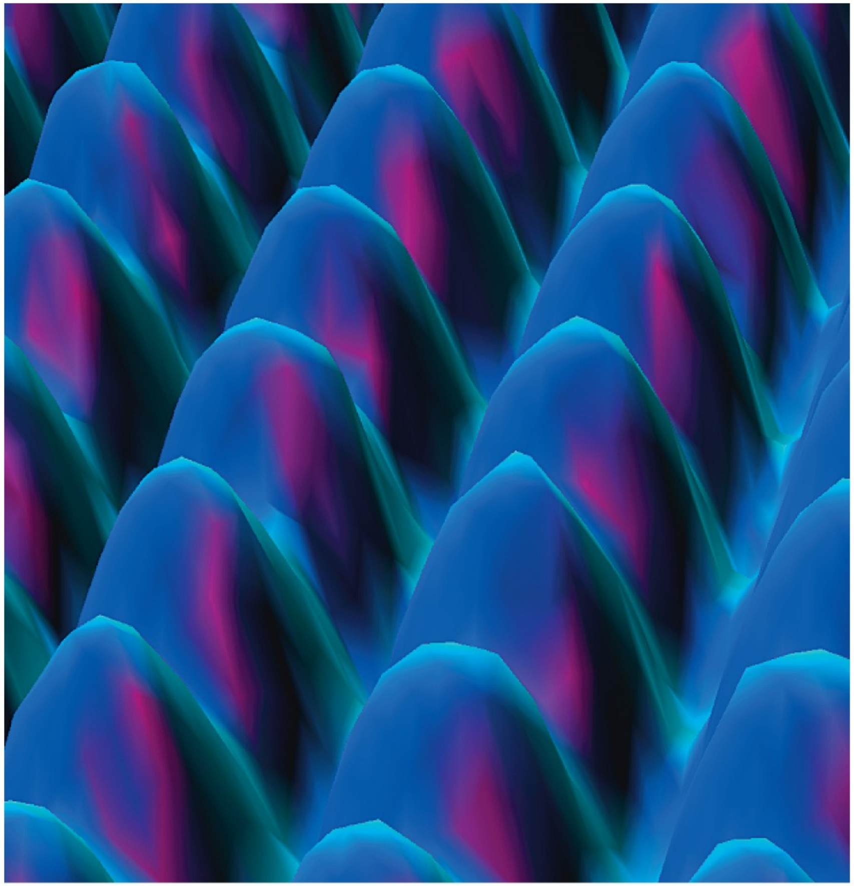
- **FIGURE 3.1 Atoms and molecules**
- **Most of matter is**
- **ultimately composed of atoms.**
- **(b)** In other substances, such as rubbing alcohol, several atoms bond together in **well-defined structures called molecules.**



(b)

3.2 What Is Matter?

- **FIGURE 3.2 Scanning tunneling microscope image of nickel atoms**
- A scanning tunneling microscope (STM) creates an image by scanning a surface with a tip of atomic dimensions.
- It can distinguish individual atoms, seen as blue bumps, in this image.
- (*Source:* Reprint Courtesy of International Business Machines Corporation, copyright © International Business Machines Corporation.)

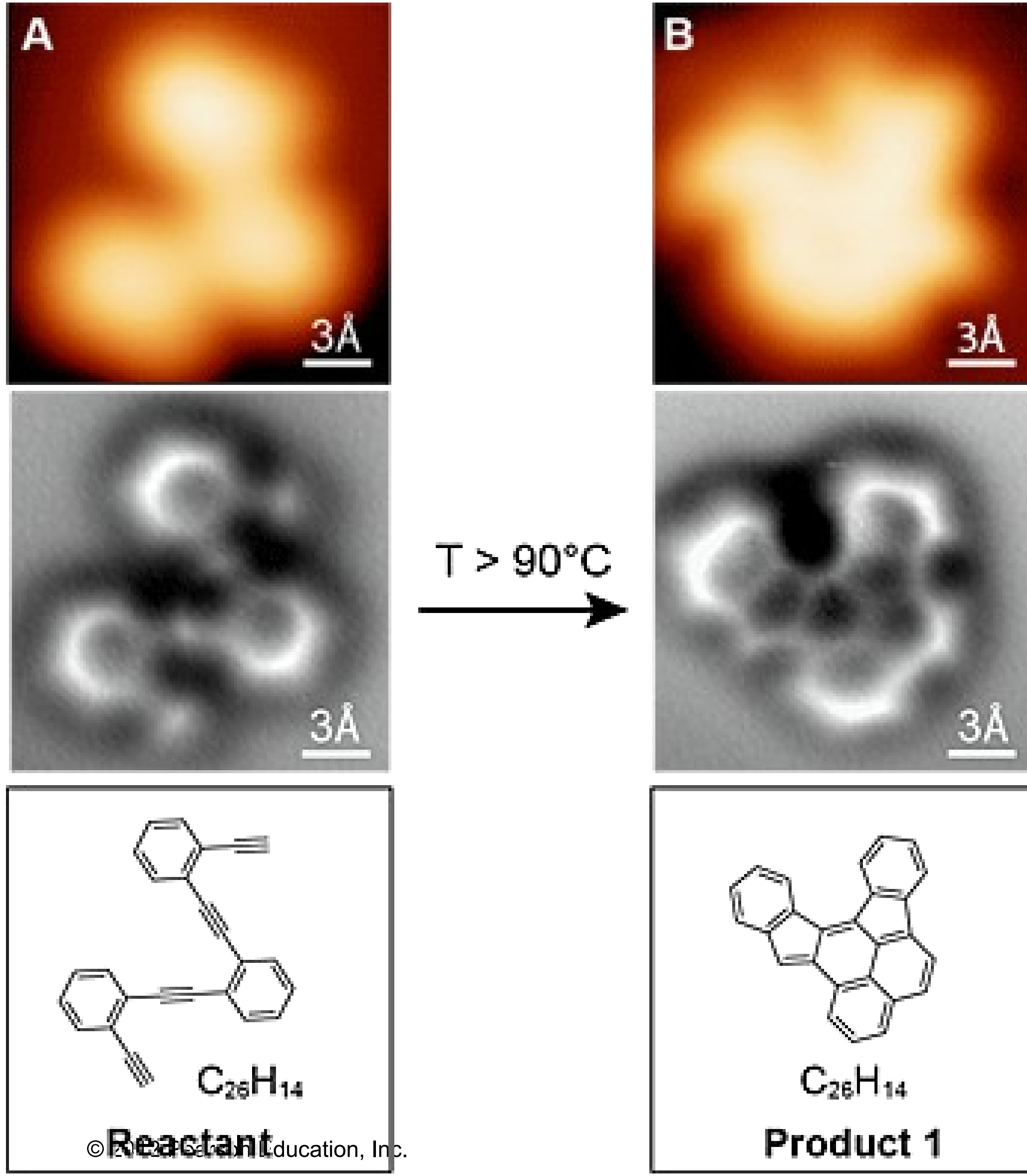


Can We See Molecules?

Non-contact atomic force microscope (nc-AFM) images (center) of a molecule before and after a reaction improve immensely over images (top) from a scanning tunneling microscope and look <u>just like the classic molecular structure diagrams (bottom).</u>

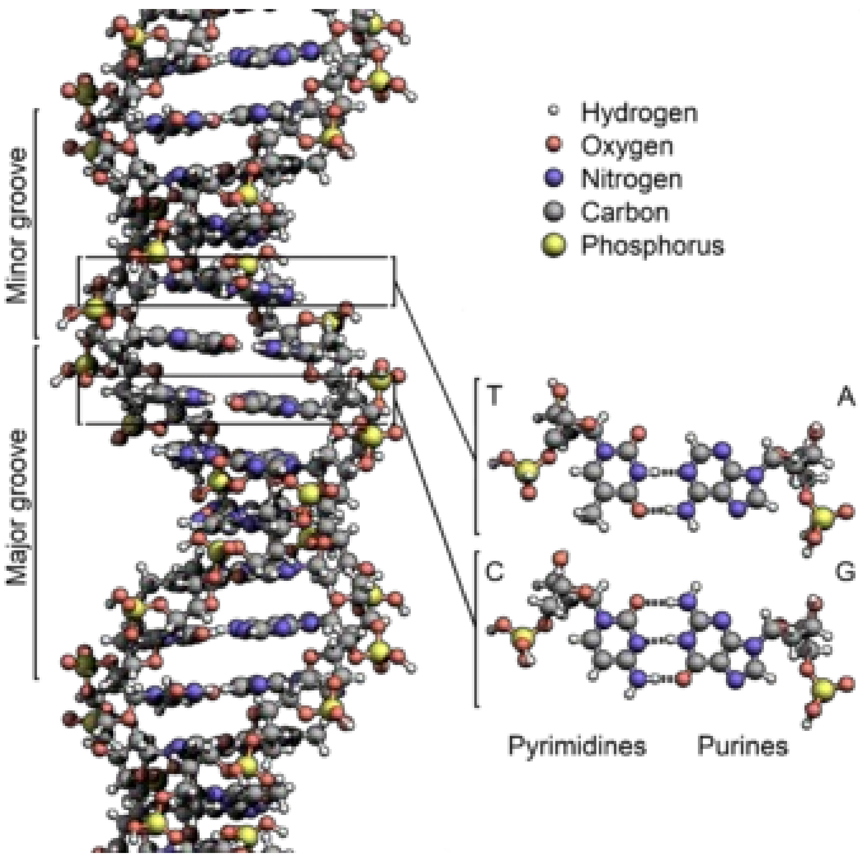
Berkeley news:
Oteyza et al, Science 340, 1434 (2013).

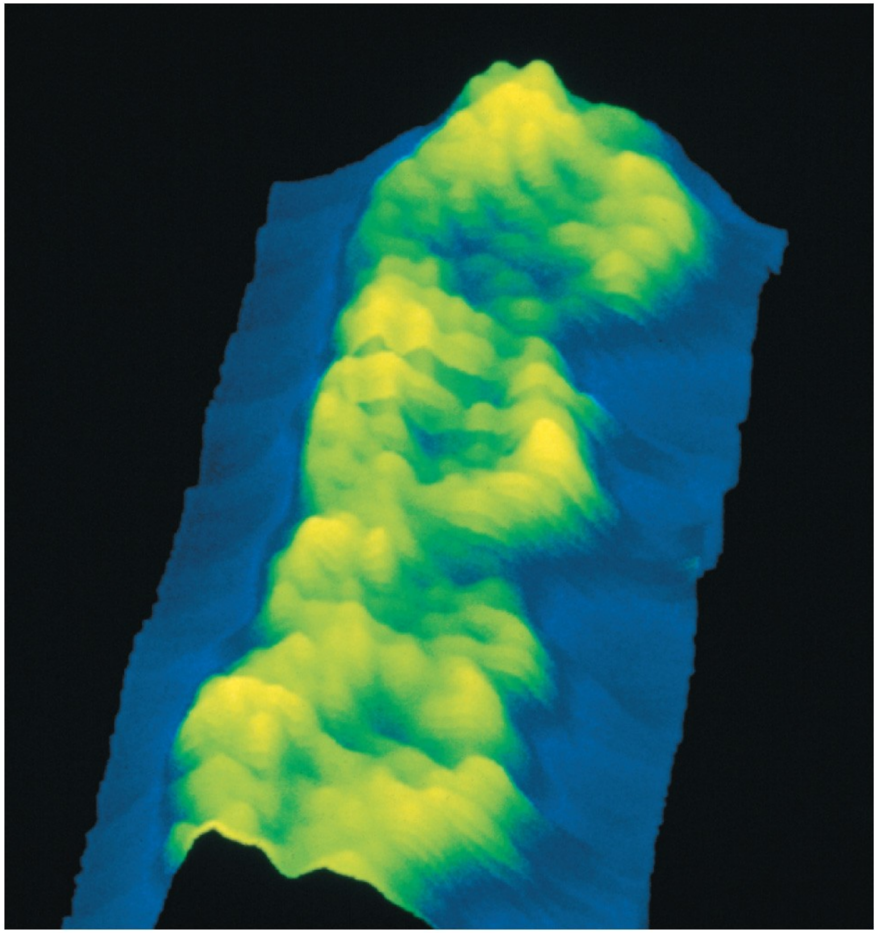
1 angstrom = 1.0×10^{-8} centimeters



3.2 What Is Matter?

- **FIGURE 3.3 Scanning tunneling microscope image of a DNA molecule**
- DNA is the hereditary material that encodes the operating instructions for most cells in living organisms.
- In this image, the **DNA molecule is yellow.**
- The double-stranded structure of DNA can be observed.

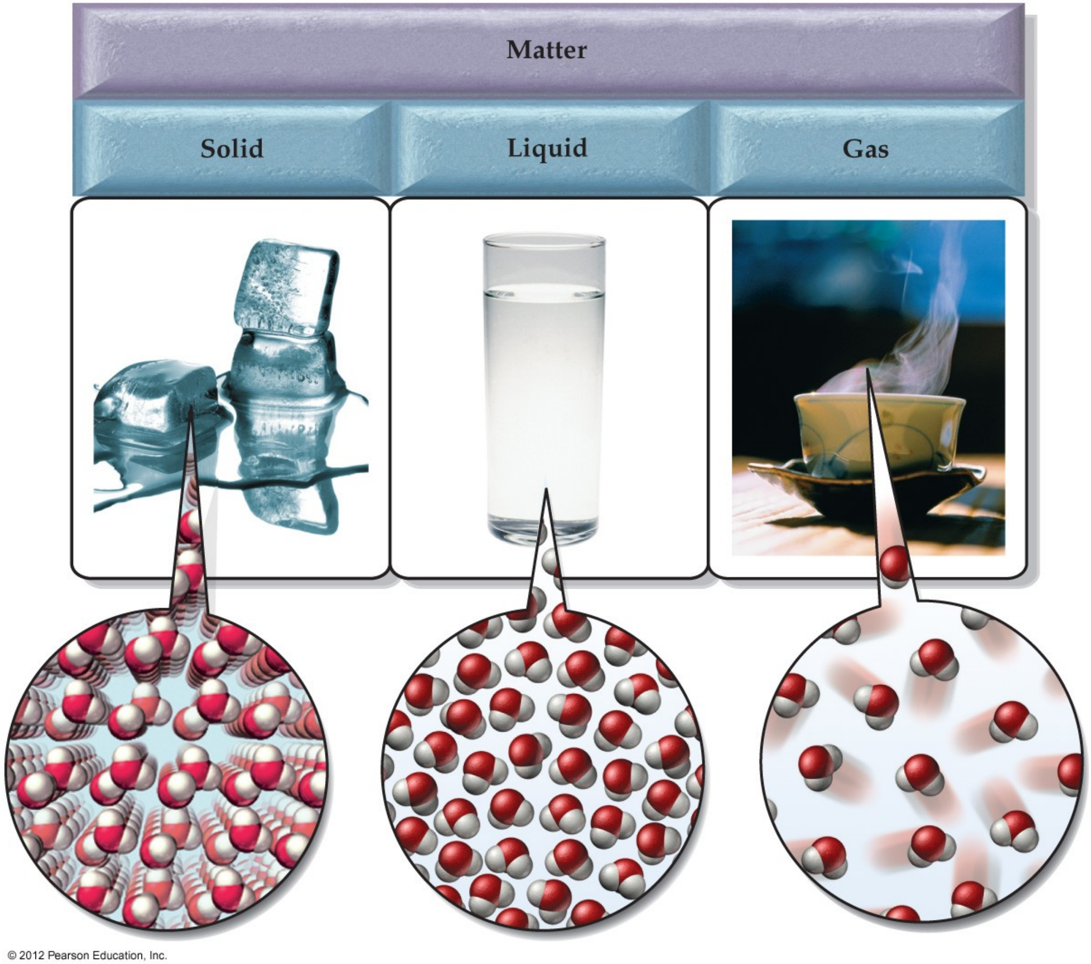




3.3 Classifying Matter According to Its State: Solid, Liquid, and Gas

- **The common states of matter** are **solid**, **liquid**, and **gas**.
- **Water** exists as **ice (solid), water (liquid), and steam (gas).**
- **In ice,** the water molecules are closely spaced and, although they vibrate about a fixed point, they do not generally move relative to one another.
- **In liquid water,** the water molecules are closely spaced but are free to move around and past each other.
- **In steam,** water molecules are separated by large distances and do not interact significantly with one another.

FIGURE 3.4 Three states of matter



A nice You Tube video about **different phases, phase transitions** is:

"Molecular Motions CHEM Study" published by Lawrence Hall of Science (13 min):

https://www.youtube.com/watch?v=qIfsx6gsG_Q

States of Matter: Solid

- In **solid matter**, atoms or molecules pack close to each other in fixed locations.
- Neighboring atoms or molecules in a solid may vibrate or oscillate, but they do not move around each other.
- Solids have fixed volume and rigid shape.
- Ice, diamond, quartz, and iron are examples of solid matter.

States of Matter: Types of Solids

- **Crystalline solid:** Atoms or molecules arranged in geometric patterns with long-range, repeating order.

- **Amorphous solid:** Atoms or molecules do not have long-range order.

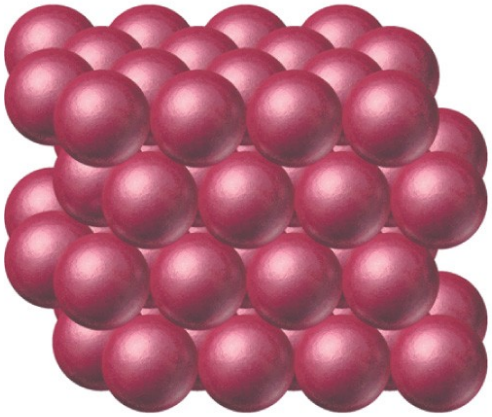
- **Examples of *crystalline*** solids include **salt and diamond**.

- The well-ordered, geometric shapes of salt and diamond crystals reflect the well-ordered geometric arrangement of their atoms.

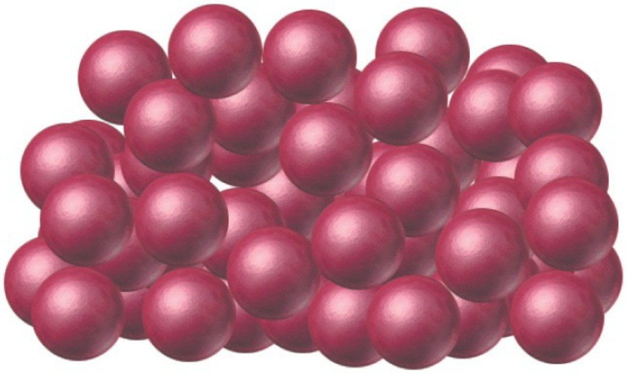
- **Examples of *amorphous* solids** include **glass, rubber, and plastic.**

States of Matter: Types of Solids

- **FIGURE 3.5 (a) In a crystalline solid,** atoms or molecules occupy specific positions to create a well-ordered, three-dimensional structure.
- **(b)** In an amorphous solid, atoms do not have any long-range order.



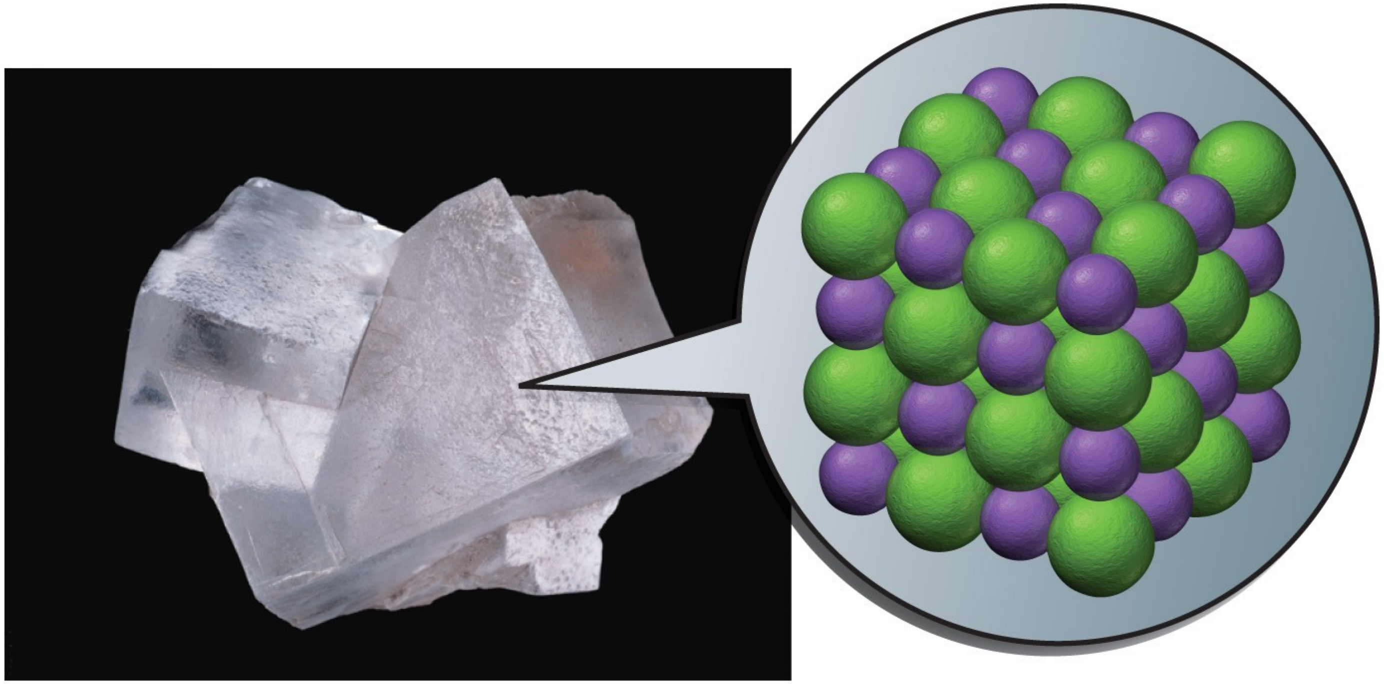
(a) Crystalline solid



(b) Amorphous solid

States of Matter: Crystalline Solid

- **FIGURE 3.6**
- Sodium chloride is an example of a crystalline solid.
- The well-ordered, cubic shape of salt crystals is due to the well-ordered, cubic arrangement of its atoms.



States of Matter: Liquid

- **In liquid matter,** atoms or molecules are close to each other but are free to move around and by each other.
- Liquids have a fixed volume because their atoms or molecules are in close contact.
- Liquids assume the shape of their container because the atoms or molecules are free to move relative to one another.
- **Water, gasoline, alcohol, and mercury are all examples of liquid matter.**

States of Matter: Gas

- **In gaseous matter,** atoms or molecules are separated by large distances and are free to move relative to one another.
- Since the atoms or molecules that compose gases are not in contact with one another, gases are **compressible**.
- Gases always assume the shape and volume of their containers.
- **Oxygen, helium, and carbon dioxide are examples of gases.**

FIGURE 3.7 Gases are compressible.

Since the atoms or molecules that compose gases are not in contact with one another, gases can be compressed.

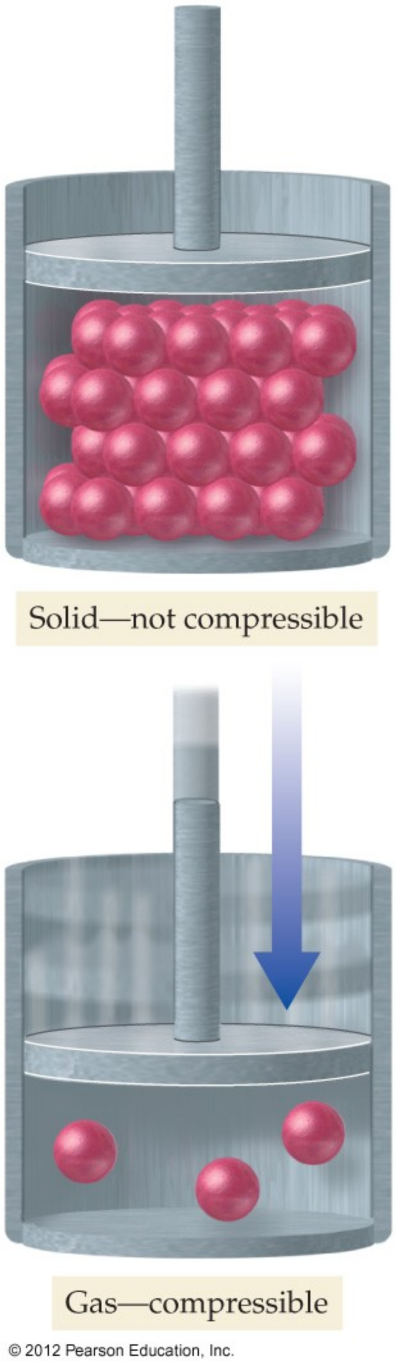


Table 3.1 summarizes the properties of solids, liquids, and gases.

TABLE 3.1 Properties of Liquids, Solids, and Gases

State	Atomic/Molecular Motion	Atomic/Molecular Spacing	Shape	Volume	Compressibility
Solid	Oscillation/vibration about fixed point	Close together	Definite	Definite	Incompressible
Liquid	Free to move relative to one another	Close together	Indefinite	Definite	Incompressible
Gas	Free to move relative to one another	Far apart	Indefinite	Indefinite	Compressible

3.4 Classifying Matter According to Its Composition:

Elements, Compounds, and Mixtures

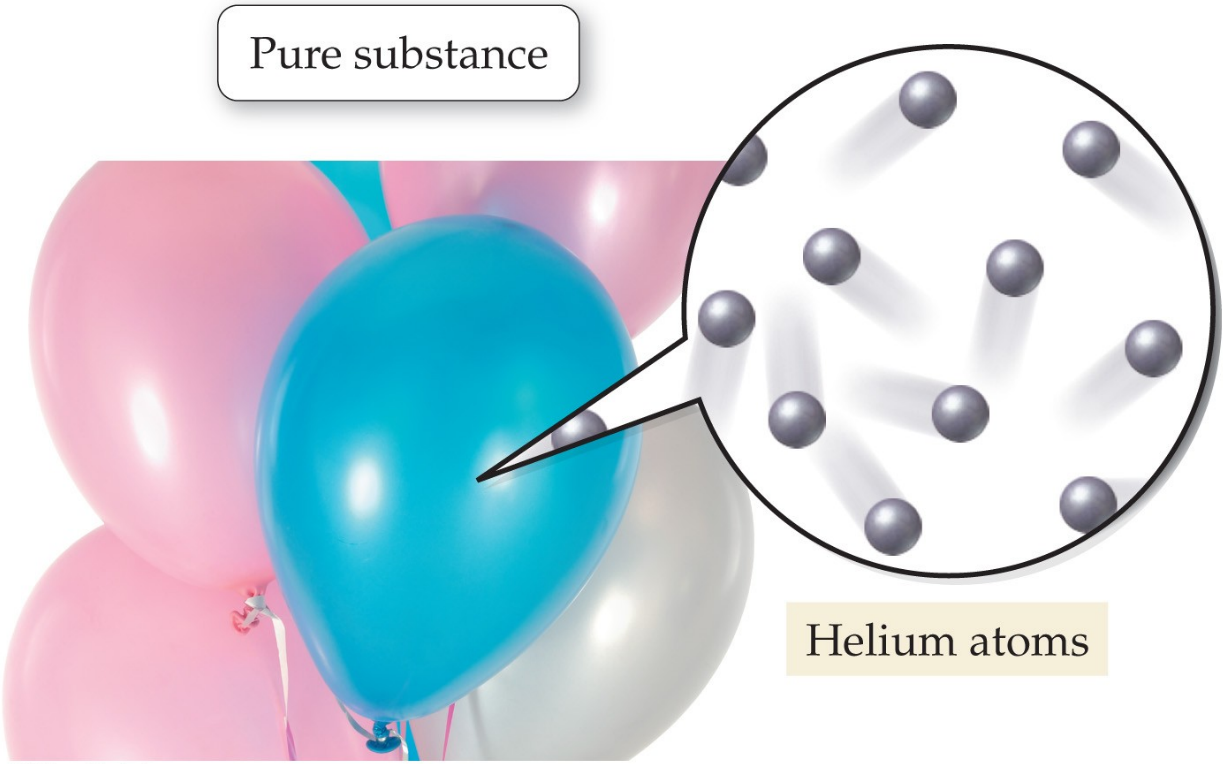
- A **pure substance** is composed of only one type of atom or molecule.
- A **mixture** is composed of two or more different types of atoms or molecules combined in variable proportions.

3.4 Classifying Matter According to Its Composition: Elements

- **Element**: A pure substance that cannot be broken down into simpler substances.
- No chemical transformation can decompose an element into simpler substances.
- **All known elements are listed in the periodic table** in the inside front cover of this book and in alphabetical order on the inside back cover of this book.

3.4 Classifying Matter According to Its Composition: Elements

- **Helium:** A pure substance composed only of one type of atom: helium atoms.

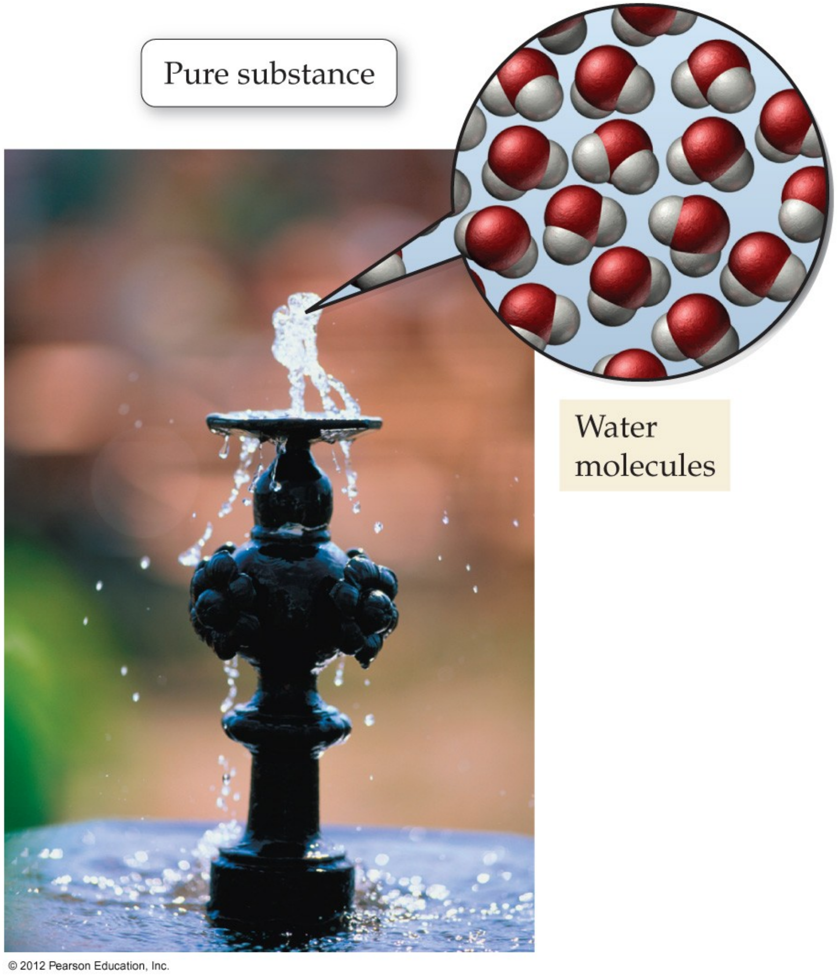


3.4 Classifying Matter According to Its Composition: Compounds

- **Compound:** **A pure substance** composed of **two or more elements** in fixed definite proportions.
- Compounds are more common than pure elements.
- Most elements are chemically reactive and combine with other elements to form compounds.
- Water, table salt, and sugar are examples of compounds.
- **Compounds can be decomposed into simpler substances.**

3.4 Classifying Matter According to Its Composition: Compounds

- **Water:** A pure substance composed only of water molecules.
- **Two elements** in fixed, definite proportions.



Classification of Matter: Compounds and Mixtures

- When matter contains **two types of atoms** it may be a ***pure substance*** or a ***mixture***.

A *compound* is a **pure substance** composed of different atoms that are chemically united (bonded) in fixed definite proportions.

A *mixture* is composed of **different substances** that are not chemically united, but simply mixed together.

3.4 Classifying Matter According to Its Composition: Mixtures

Air and seawater are examples of mixtures. Air contains primarily nitrogen and oxygen. Seawater contains primarily salt and water.

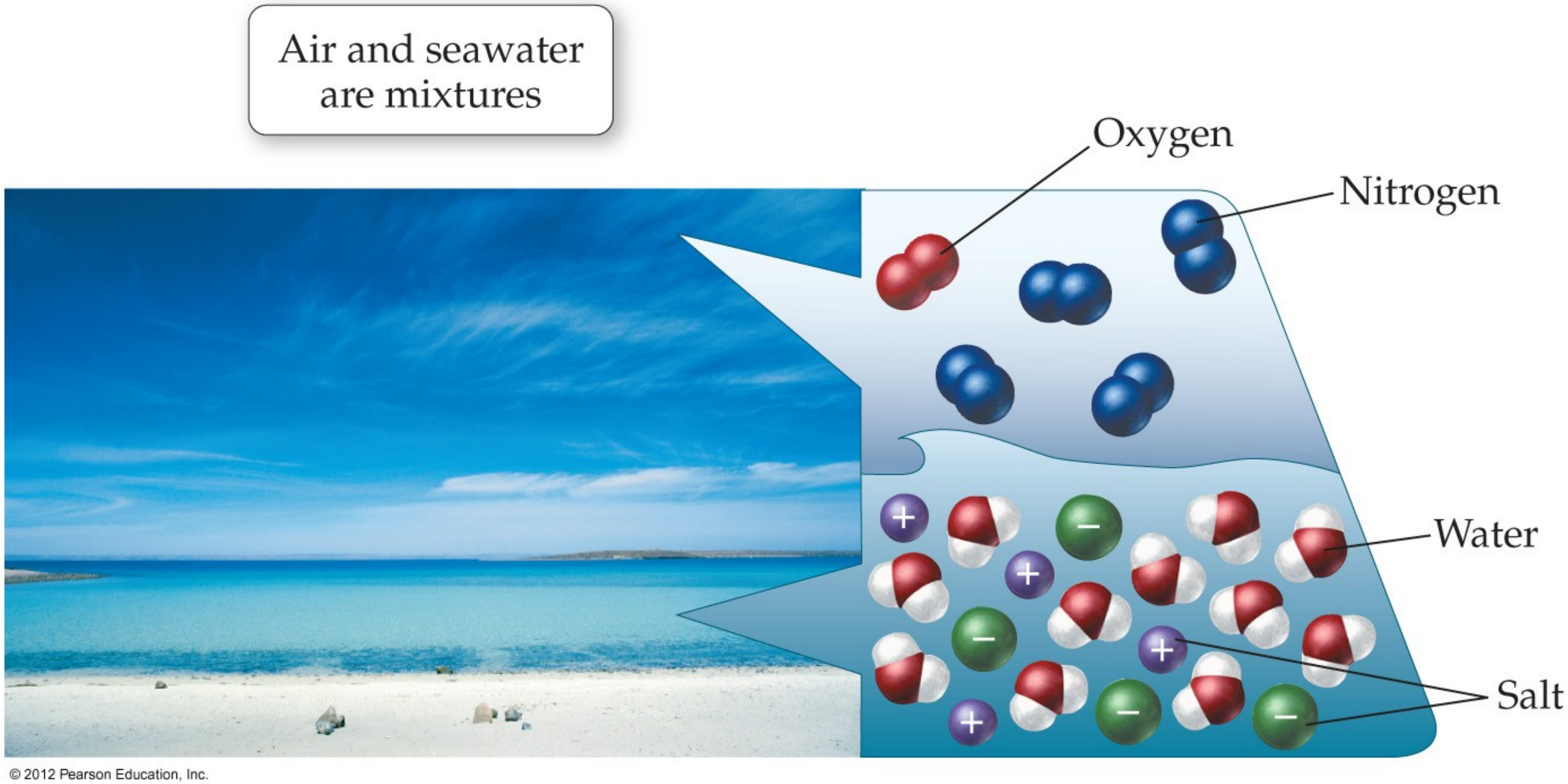
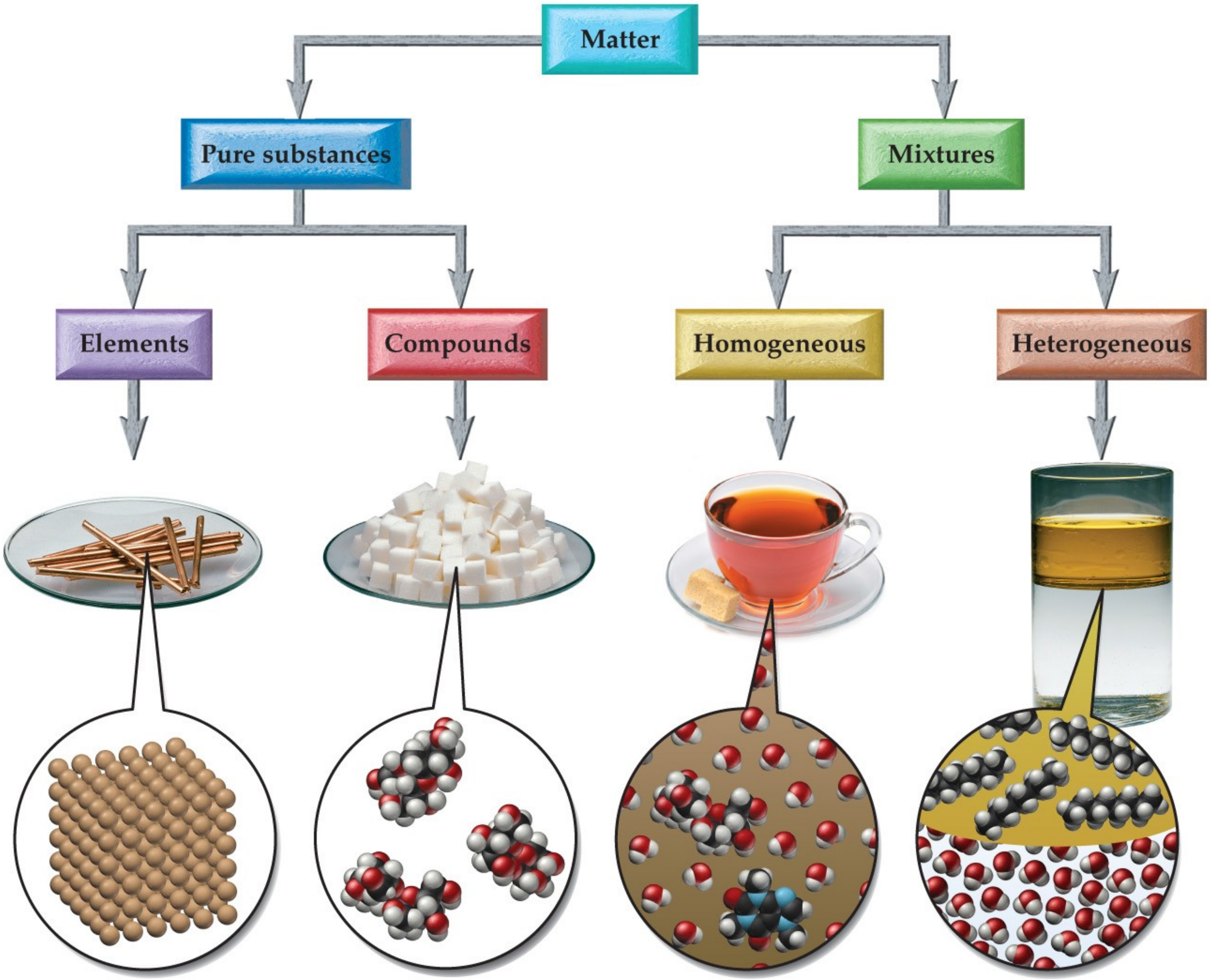


FIGURE 3.8 Classification of matter

- **A pure substance** may be either an ***element*** (such as copper) or a ***compound*** (such as sugar).

- **A mixture** may be either a ***homogeneous mixture*** (such as sweetened tea) or ***heterogeneous mixture*** (such as hydrocarbon and water).


Matter
Pure substances
Mixtures
Elements
Compounds
Homogeneous
Heterogeneous

- **To summarize, as shown in Figure 3.8:**

- Matter may be a pure substance, or it may be a mixture.
- A pure substance may be either an element or a compound.
- A mixture may be either homogenous or heterogeneous.
- Mixtures may be composed of two or more elements, two or more compounds, or a combination of both.

3.5 How We Tell Different Kinds of Matter Apart: Physical and Chemical Properties

- A **physical property** is one that a substance displays without changing its composition.
- A **chemical property** is one that a substance displays only through changing its composition.

Recommended video:

- Physical and Chemical Changes
- (by Bozeman Science, 11 minutes)
- https://www.youtube.com/watch?v=X328AWaJXvI

3.5 How We Tell Different Kinds of Matter Apart: Physical and Chemical Properties

- The characteristic ***odor*** of gasoline is a ***physical*** property —gasoline does not change its composition when it exhibits its odor.
- The ***flammability*** of gasoline is a ***chemical*** property—gasoline does change its composition when it burns.

3.5 How We Tell Different Kinds of Matter Apart: Physical Properties

- The atomic or molecular ***composition*** of a substance ***does not change*** when the substance displays its **physical properties**.
- The **boiling point of water — a** physical property—is 100 degrees Celsius.
- When water boils, it changes from a liquid to a gas, but the gas is still water.

3.6 How Matter Changes: Physical Change

- **State changes—transformations from one state of matter to another** (such as from solid to liquid)—

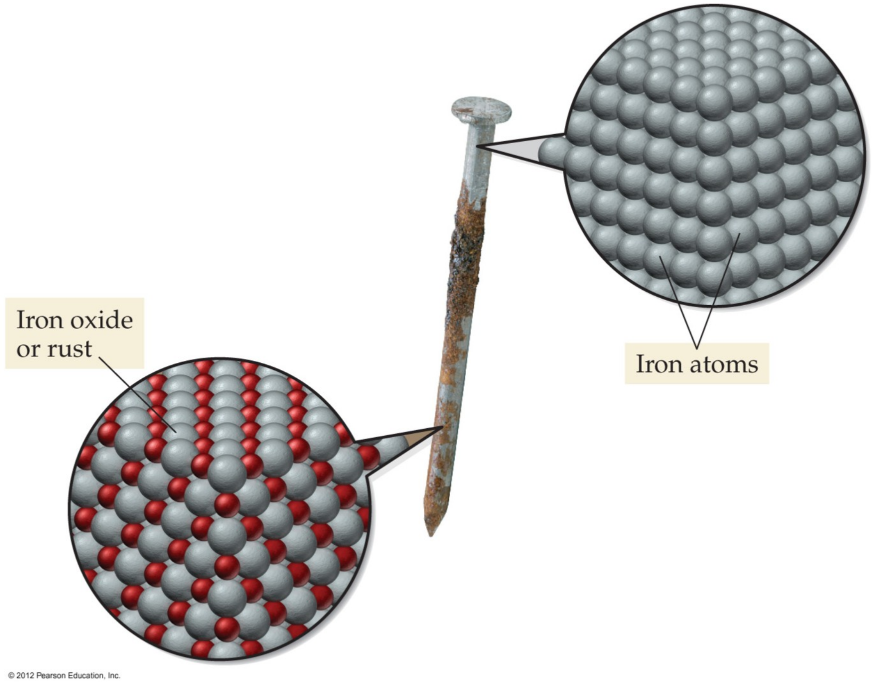
  are **always** **physical changes.**

3.6 How Matter Changes: Physical and Chemical Changes

- In a **physical change**, matter changes its appearance but not its composition.
- When ice melts, it looks different but its composition is the same. Solid ice and liquid water are both composed of water molecules, so melting is a physical change.
- In a **chemical change**, matter *does* change its composition.
- For example, copper turns green upon continued exposure to air because it reacts with gases in air to form new compounds. This is a chemical change.

3.5 How We Tell Different Kinds of Matter Apart: Chemical Properties

- **A chemical property:** The susceptibility of iron to rusting.
- Rusting is a chemical change.
- When iron rusts, it turns from iron to iron oxide.



3.6 How Matter Changes: Chemical Change

- Matter undergoes a **chemical change** when it undergoes a **chemical reaction.**
- In a chemical reaction, the substances present **before the chemical change** are called **reactants.**
- The substances present **after the change** are called **products.**

$$\text{Reactants} \xrightarrow[\text{Chemical Change}]{} \text{Products}$$

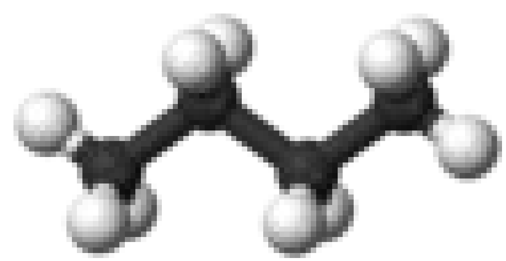
3.6 How Matter Changes: Physical and Chemical Changes

- The difference between chemical and physical changes is seen at the molecular and atomic level.
- **In physical changes**, the atoms that compose the matter ***do not* change** their fundamental associations, even though the matter may change its appearance.
- **In chemical changes**, atoms **do change** their fundamental associations, resulting in matter with a new identity.

- *A **physical change** results in a different form of the **same substance.***
- *A **chemical change** results in a completely **new substance.***

FIGURE 3.11 Vaporization: a physical change

- Push the button on a lighter without turning the flint.
- The **liquid butane** vaporizes to gaseous butane.
- The liquid butane and the gaseous butane are both composed of butane molecules; this is a **physical change**.



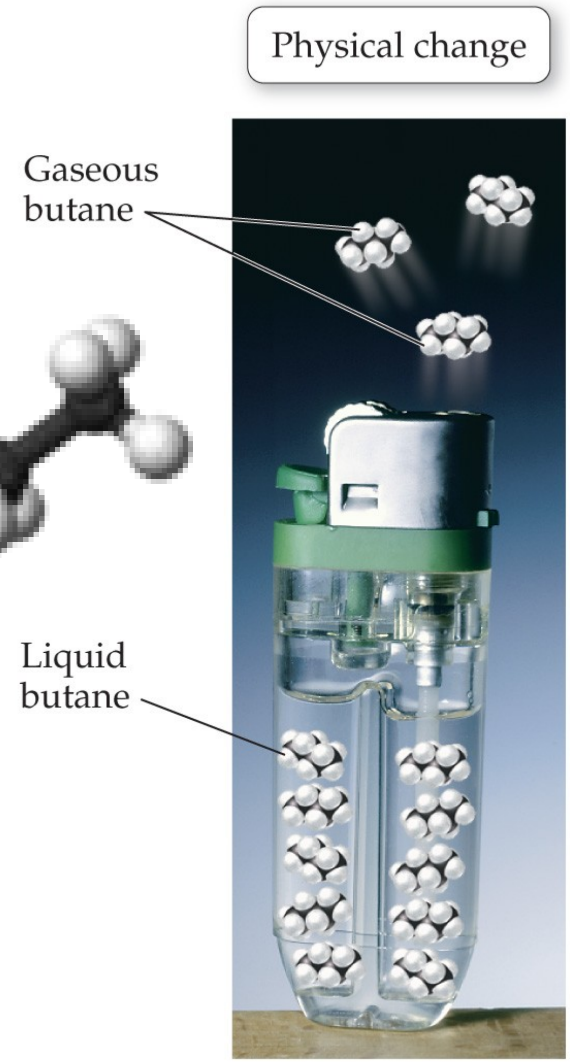
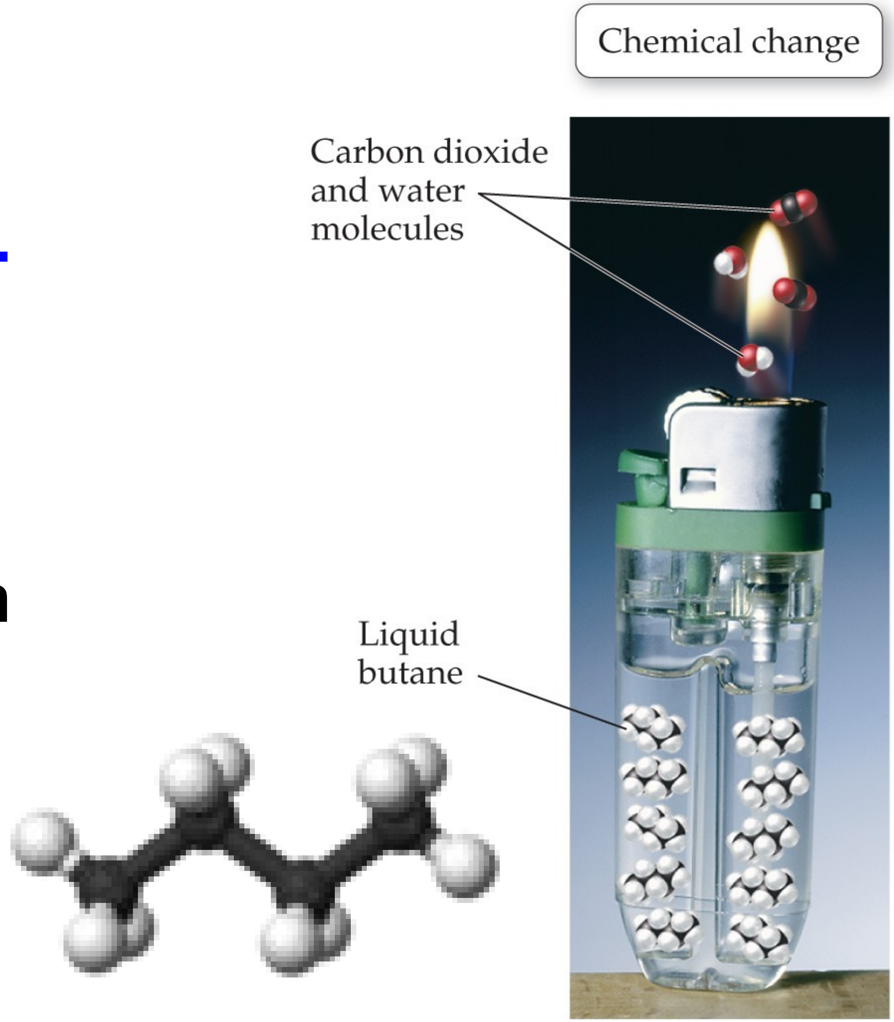


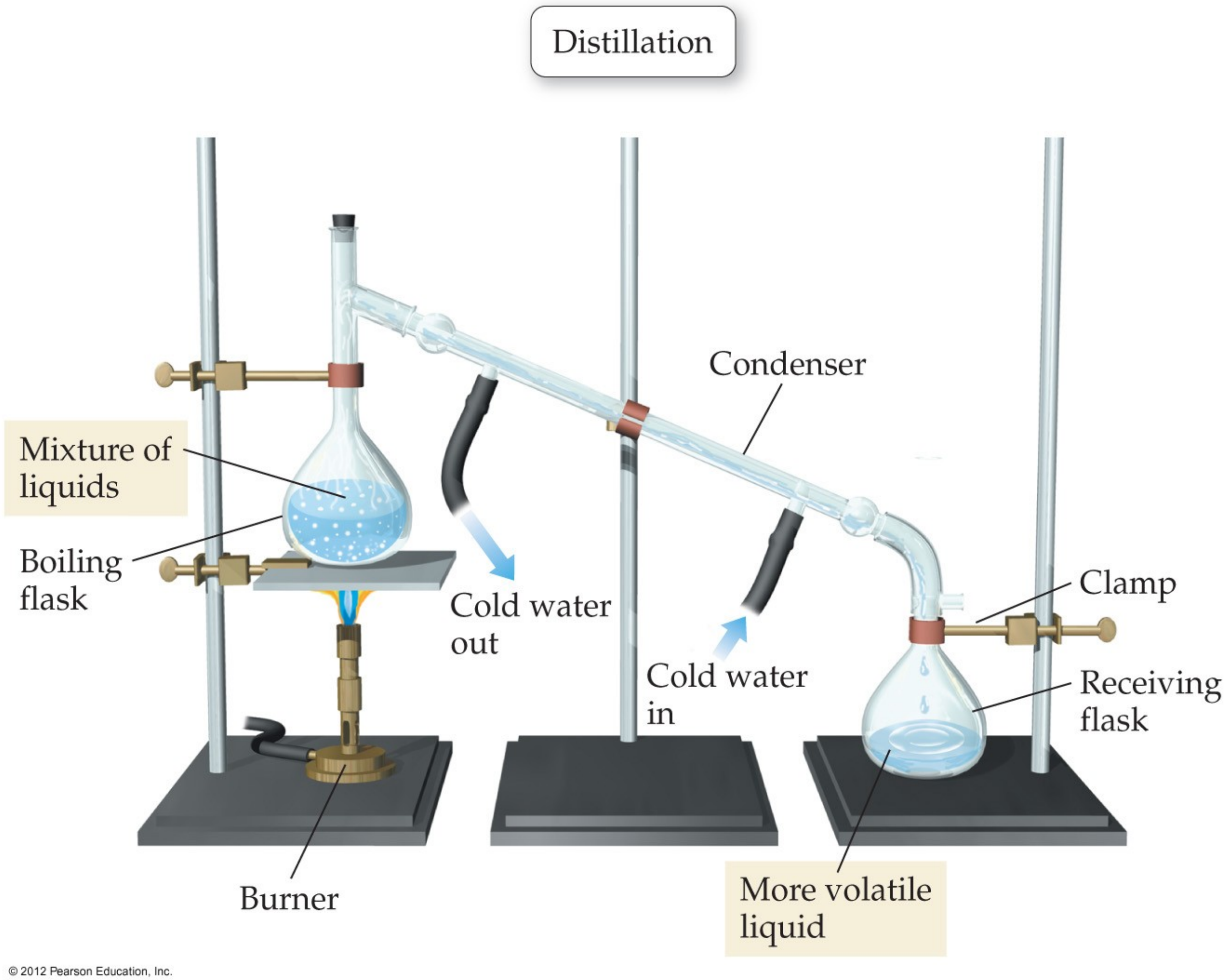
FIGURE 3.12 Burning: a chemical change

- Push the button *and* turn the flint to create a spark: **Produce a flame.**

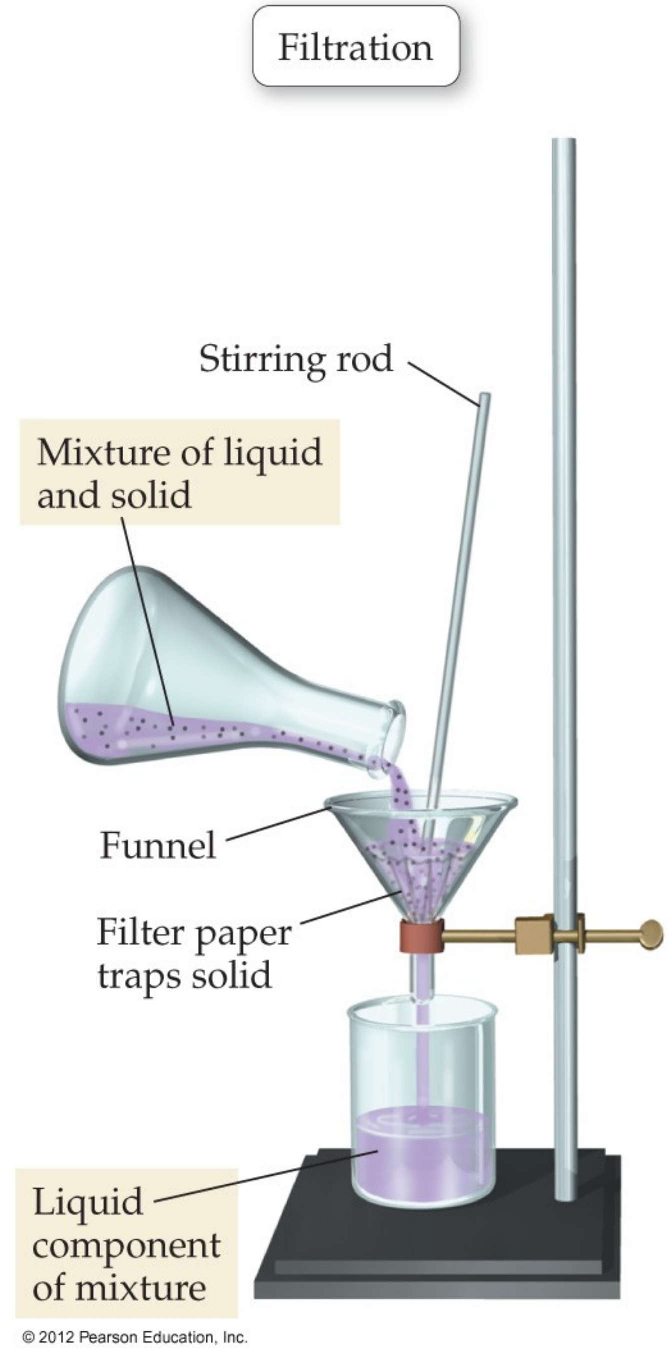
- <u>The butane molecules</u> react with oxygen molecules in air to form <u>new molecules</u>, carbon dioxide and water.
- This is a **chemical change.**



Separating Mixtures through Physical Changes by Distillation



Separating Mixtures through Physical Changes by Filtration



3.7 Conservation of Mass: There Is No New Matter

- Matter is neither created nor destroyed in a **chemical reaction**.
- In a **nuclear reaction**, significant changes in mass can occur.
- In **chemical reactions**, however, the changes in mass are so minute that they can be ignored.
- During physical and chemical changes, the total amount of matter remains constant.

Suppose that we burn 58 g of butane in a lighter. It will react with 208 g of oxygen to form 176 g of carbon dioxide and 90 g of water.

$$\begin{array}{ccc} \text{Butane} + \text{Oxygen} & & \text{Carbon Dioxide} + \text{Water} \\ \underbrace{58\text{ g} + 208\text{ g}}_{\boxed{266\text{ g}}} & \longrightarrow & \underbrace{176\text{ g} + 90\text{ g}}_{\boxed{266\text{ g}}} \end{array}$$

3.8 Energy

- **Energy** is a major component of our universe.
- ***Energy*** *is the capacity to do work.*
- **Work** is defined as the result of a force acting on a distance.
- **The behavior of matter is driven by energy**.
- Understanding energy is critical to understanding chemistry.

Recommended videos:

- **What is Energy?** **(Neil deGrasse Tyson explains)**
- **you-tube video (16 min)**
- https://www.youtube.com/watch?v=b69LGTX-c7M

- **What is Energy ?** **(by Fermilab, 10 min)**
- https://www.youtube.com/watch?v=u36H4Uo3rPM

3.8 Energy

- **Like matter, energy is conserved.**
- The **law of conservation of energy** states that *energy is neither created nor destroyed.*
- **The total amount of energy is constant.**
- Energy can be changed from one form to another.
- Energy can be transferred from one object to another.
- **Energy cannot be created out of nothing, and it does not vanish into nothing.**

There are different forms of energy.

- **The total energy** of a sample of matter is **the sum** of its **kinetic energy**, the energy associated with its motion, and its **potential energy**, the energy associated with its position or composition.

- **Electrical energy:** The energy associated with the flow of electrical charge.
- **Thermal energy:** The energy associated with the random motions of atoms and molecules in matter.
- **Chemical energy:** A form of potential energy associated with the positions of the particles that compose a chemical system.

Units of Energy

- The SI unit of energy is the **joule (J),** named after the English scientist James Joule (1818–1889), who demonstrated that energy could be converted from one type to another as long as the total energy was conserved.
- A second unit of energy is the **calorie (cal),** the amount of energy required to raise the temperature of 1 g of water by 1 degree Celsius.
- A calorie is a larger unit than a joule: A related energy unit is the nutritional or *capital C* **Calorie (Cal),** equivalent to 1000 *little c* calories.
- **kilowatt-hour (kWh).**
- The average cost of residential electricity in the U.S. is about $0.12 per kilowatt-hour.

Table 3.2 lists various energy units and their conversion factors.

TABLE 3.2 Energy Conversion Factors

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

EXAMPLE Conversion of Energy Units

- A candy bar contains 225 Cal of nutritional energy. How many joules does it contain?

Solution Map

$$\text{Cal} \xrightarrow{\dfrac{1000\ \text{cal}}{1\ \text{Cal}}} \text{cal} \xrightarrow{\dfrac{4.184\ \text{J}}{1\ \text{cal}}} \text{J}$$

Relationships Used

1000 calories = 1 Cal (Table 3.2)

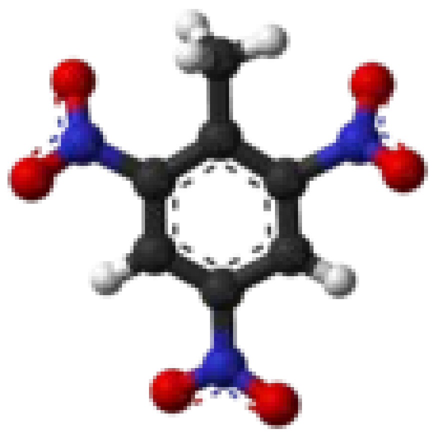
4.184 J = 1 cal (Table 3.2)

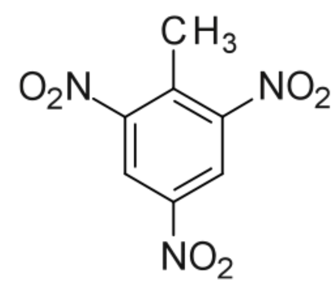
Solution

$$225\ \cancel{\text{Cal}} \times \frac{1000\ \cancel{\text{cal}}}{1\ \cancel{\text{Cal}}} \times \frac{4.184\ \text{J}}{1\ \cancel{\text{cal}}} = 9.41 \times 10^5\ \text{J}$$

3.9 Energy and Chemical and Physical Change

- ***Systems with high potential energy*** *have a* ***tendency to change*** *in a way that* ***lowers their potential energy.***
- **Objects or systems with high potential energy tend to be *unstable*.**
- Some chemical substances, such as the molecules that compose **TNT (trinitrotoluene), have a relatively high potential energy.**
- Upon detonation, TNT decomposes as follows:
- **$2\ C_7H_5N_3O_6 \rightarrow 3\ N_2 + 5\ H_2O + 7\ CO + 7\ C$**
- **$2\ C_7H_5N_3O_6 \rightarrow 3\ N_2 + 5\ H_2 + 12\ CO + 2\ C$**





- **TNT molecules** tend to undergo rapid chemical changes that lower their potential energy, which is why **TNT is explosive**.

System & Environment

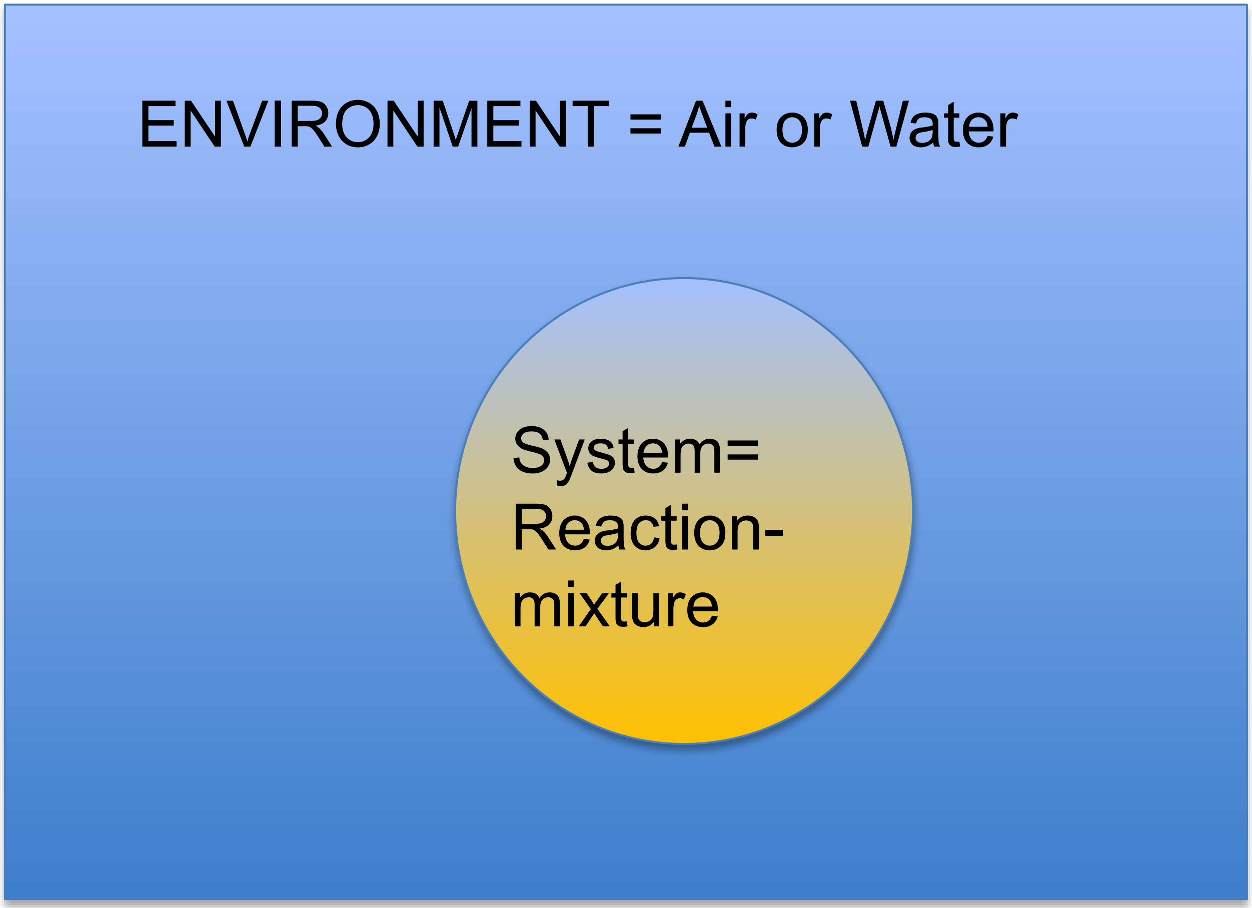
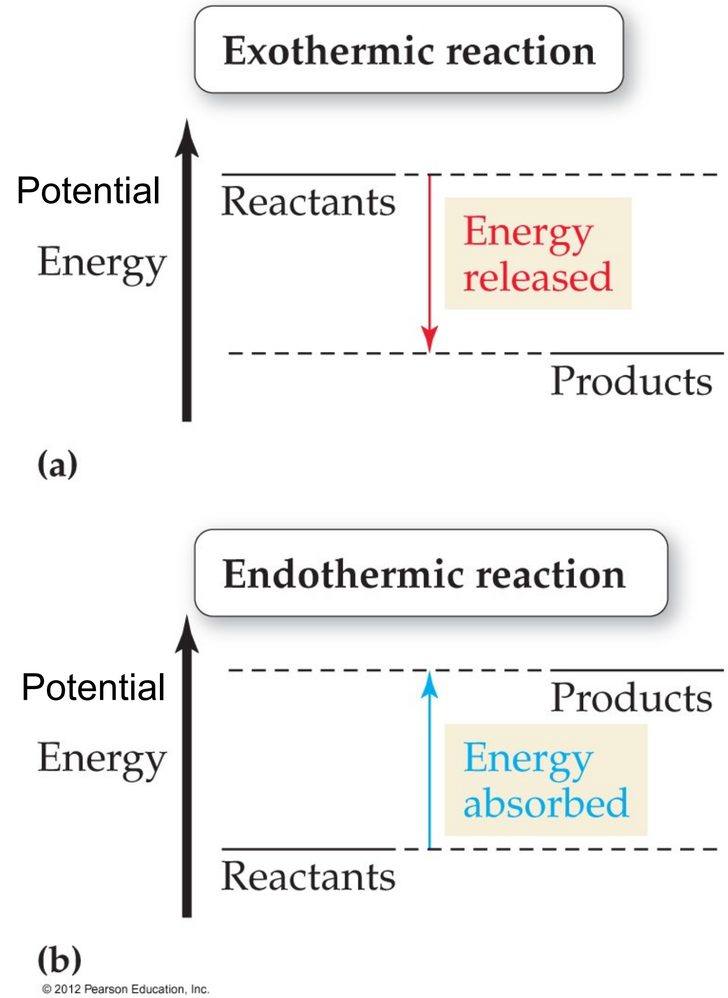
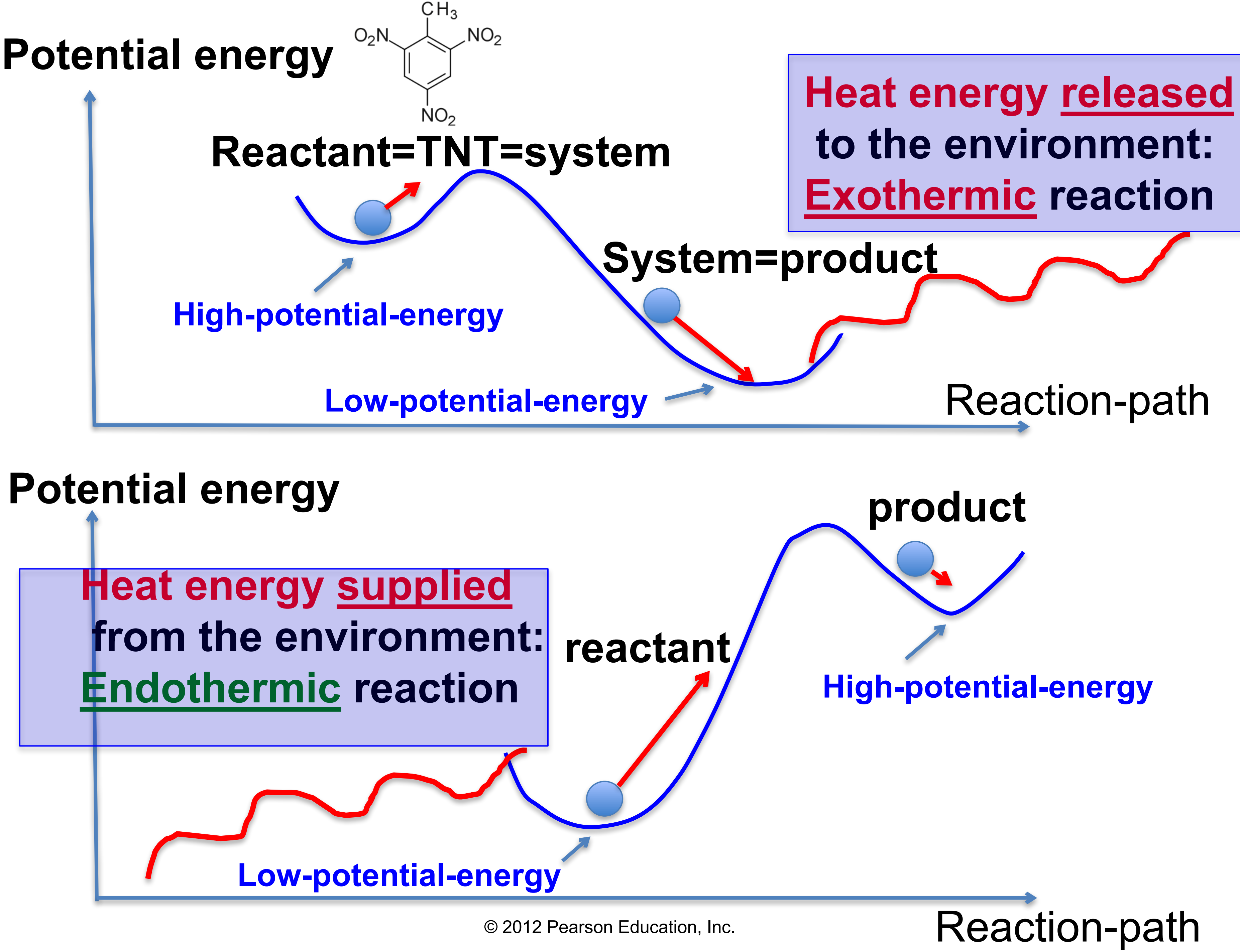


FIGURE 3.16 Exothermic and endothermic reactions

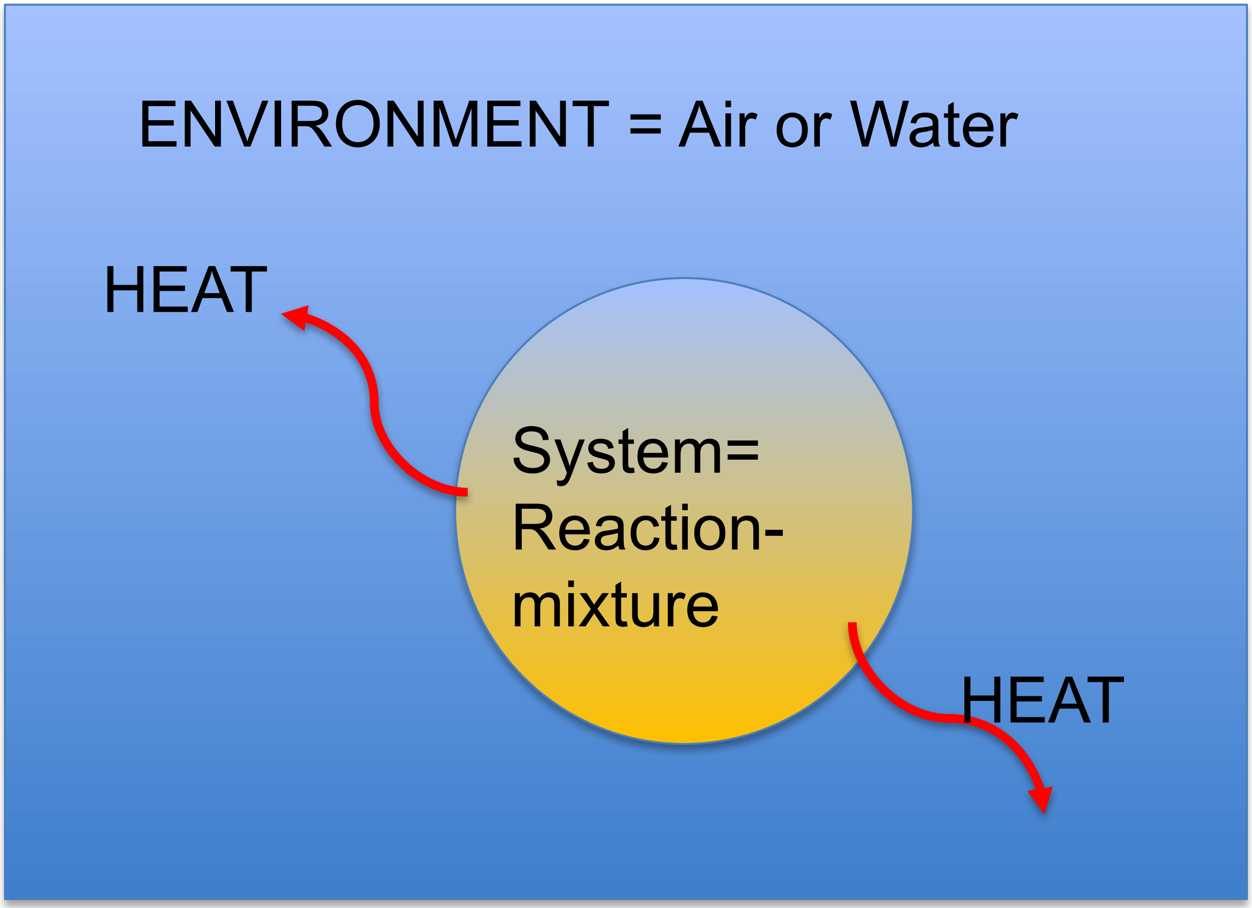
- (a) In an **exothermic** reaction, **thermal energy is transferred from the reaction mixture (the system) to the environment (air or bath).**

- (b) In an **endothermic** reaction, **thermal energy is transferred from the environment (air or bath) to the reaction mixture (the system).**

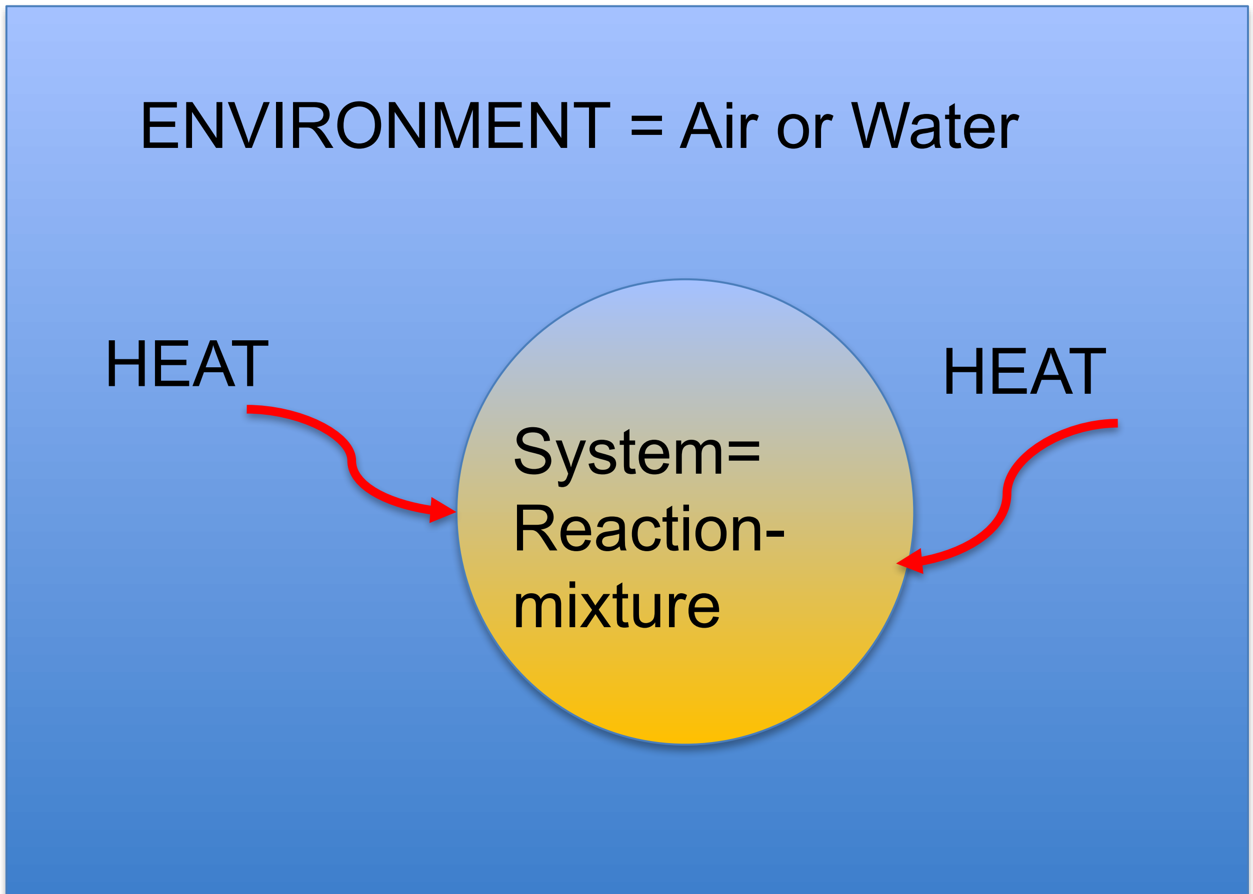



Potential energy
CH_3
O_2N
NO_2
NO_2
Reactant=TNT=system
Heat energy released
to the environment:
Exothermic reaction
System=product
High-potential-energy
Low-potential-energy
Reaction-path
Potential energy
product
Heat energy supplied
from the environment:
Endothermic reaction
reactant
High-potential-energy
Low-potential-energy
Reaction-path

Exothermic Reaction =
System transfers heat energy to Environment



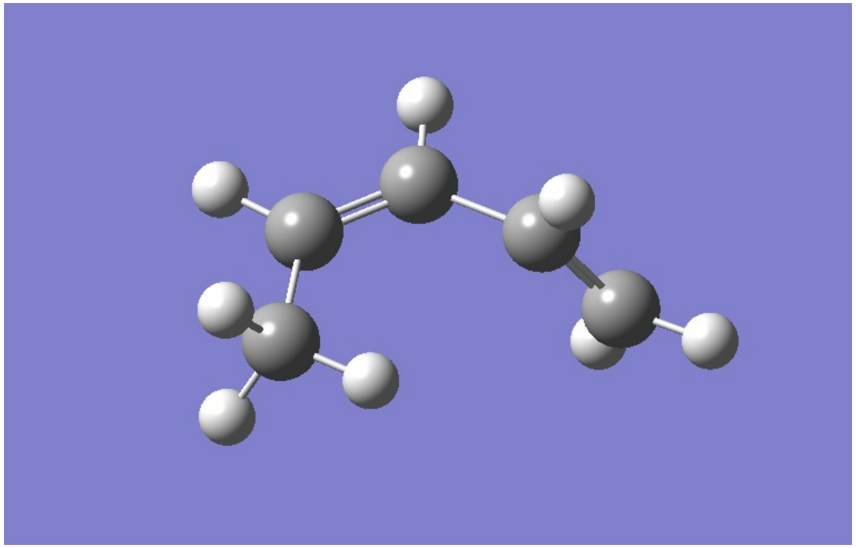
Endothermic Reaction =
System absorbs heat energy from Environment

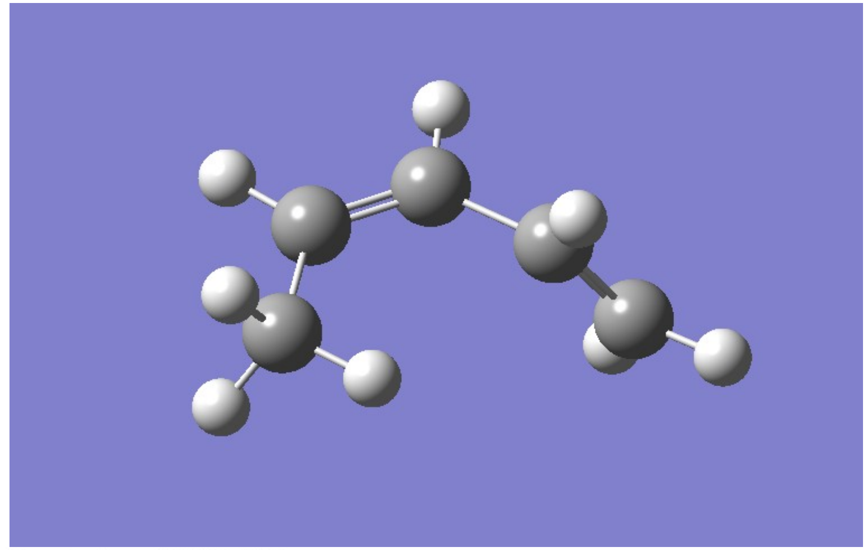


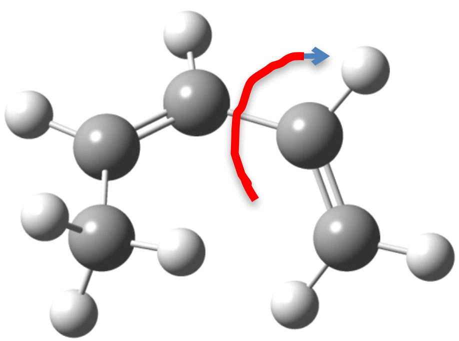
An Example

- **A Chemical system (a molecule) transitioning from:**
- High Potential Energy (Cis-structure)
- to
- Low Potential Energy (trans-structure)

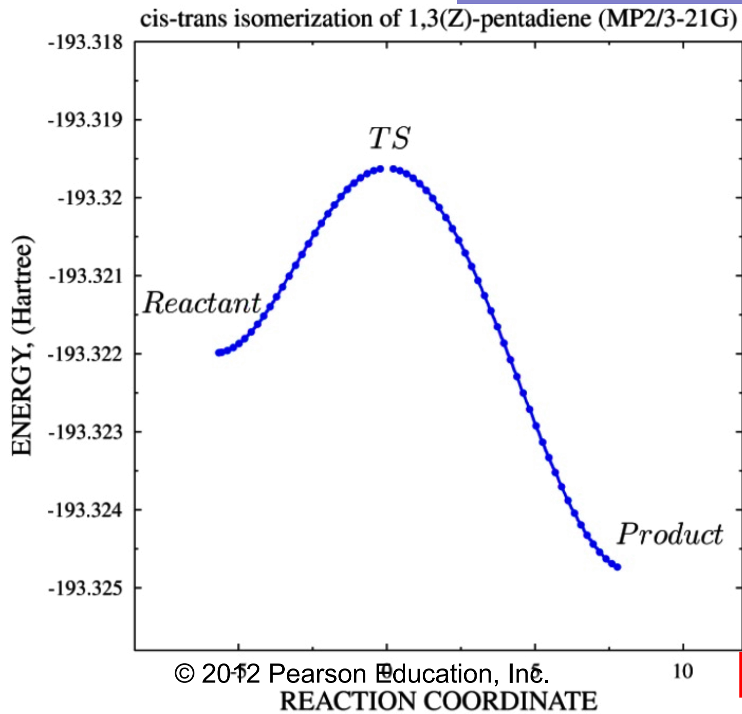
Structural Changes in Chemical Reaction (from Dr. Bytautas research)

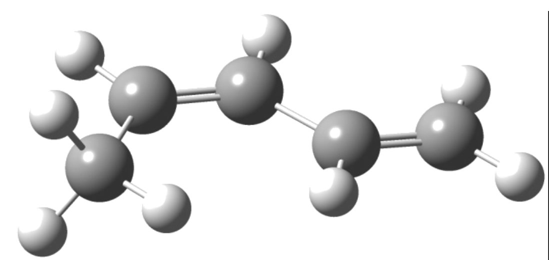






Reactant = **cis**





Product = **trans**

3.10 Temperature: Random Motion of Molecules and Atoms

- The atoms and molecules that compose matter are in constant random **motion** — they contain ***thermal energy.***
- The **temperature** of a substance is a **measure of its average thermal energy.**
- The **hotter** an object, the greater the random motion of the atoms and molecules that compose it, and **the higher its temperature.**

- **Recommended TUTORIAL You-Tube VIDEOS:**

- **TUTORIAL about TEMPERATURE**

You Tube video (Bozeman Science, 5min):

https://www.youtube.com/watch?v=B6hAwZH2mmA

- **TUTORIAL about Thermal Energy, Temperature and Heat. You Tube video (by Khan Academy, 12 min):**
- https://www.youtube.com/watch?v=-7GI-yKF6Y4

We must be careful to not confuse *temperature* with *heat*.

- **Heat,** which has **units of energy**, is the ***transfer or exchange* of thermal energy** caused **by a temperature difference.**

- For example, when a piece of cold ice is dropped into a cup of warm water, heat (thermal energy) is transferred from the water to the ice.

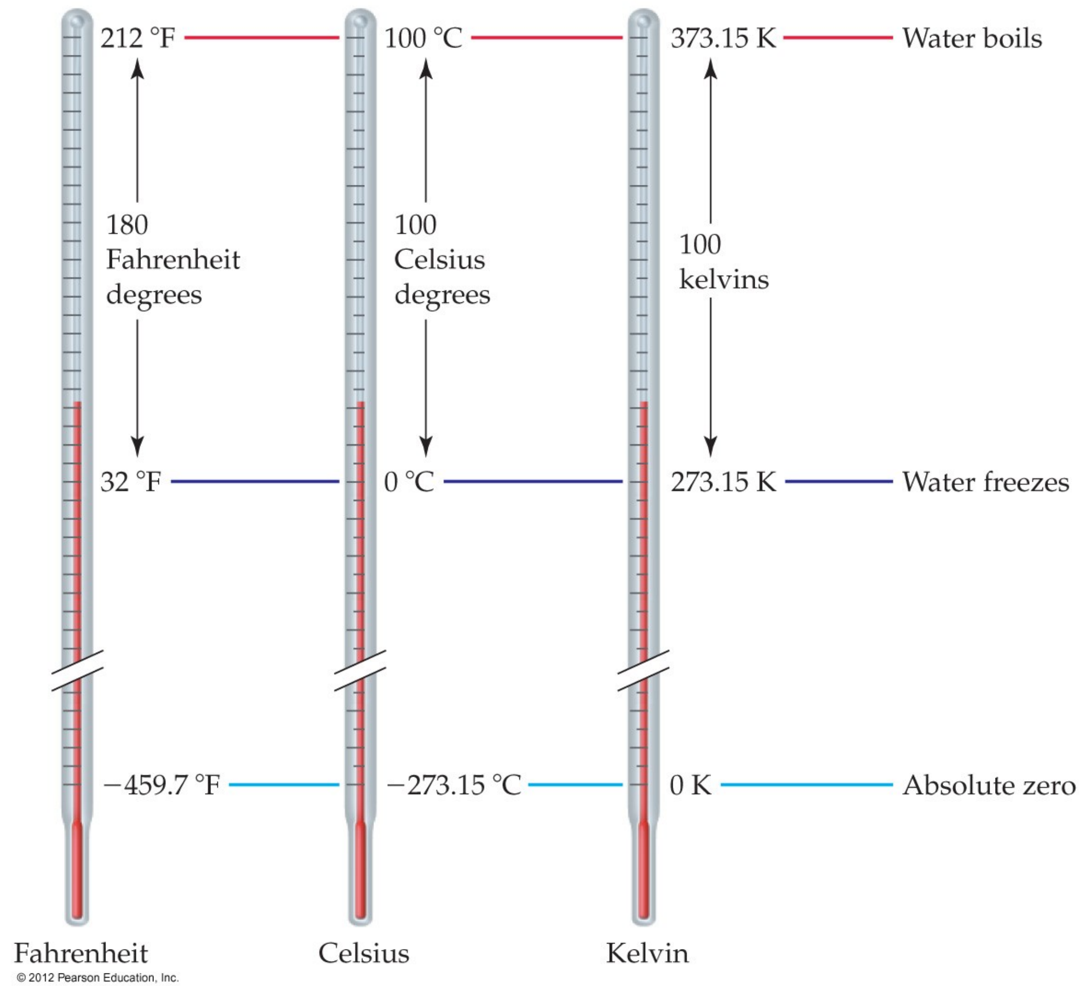
- **Temperature**, by contrast, is **a *measure* of the average thermal energy of matter** (not the exchange of thermal energy).

Three different temperature scales are in common use: Fahrenheit, Celsius, and Kelvin.

- The Fahrenheit scale assigns 0 °F to the freezing point of a concentrated saltwater solution and 96 °F to normal body temperature.
- On the **Fahrenheit (°F) scale**, water freezes at 32 °F and boils at 212 °F. Room temperature is approximately 72 °F.
- On the **Celsius (°C) scale**, water freezes at 0 °C and boils at 100 °C. Room temperature is approximately 22 °C.
- The **Kelvin (K) scale** avoids negative temperatures by assigning 0 K to the coldest temperature possible, absolute zero.
- **Absolute zero** is the temperature at which molecular **motion virtually stops.**

FIGURE 3.17 Comparison of the **Fahrenheit, Celsius, and Kelvin temperature scales**

- The **Fahrenheit** degree is **five-ninths** the size of a Celsius degree.
- The **Celsius** degree and the **Kelvin** degree are **the same size.**



Converting between Temperature Scales

- We can convert between Fahrenheit, Celsius, and Kelvin temperature scales using the following formulas.

$$K = {}^{\circ}C + 273.15$$

$${}^{\circ}C = \frac{({}^{\circ}F - 32)}{1.8}$$

3.11 Temperature Changes: Heat Capacity

- **Heat capacity:** The **quantity of heat (usually in joules)** required to change the temperature of a **given amount** of the substance by **1 °C (also in Kelvin).**

- **When the amount of the substance is expressed in grams**, the **heat capacity** is called the **specific heat capacity** (or the **specific heat**) and has units of joules per gram per degree Celsius, **J/g °C** (also in Kelvin).

- **Conceptual Question ?**

- **What is the difference between Heat capacity**

And

Specific Heat Capacity ?

Table 3.4 lists the values of the specific heat capacity for several substances.

- Notice that water has the **highest specific heat capacity** on the list.
- This table is useful when solving homework problems.

TABLE 3.4 Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity (J/g °C)
Lead	0.128
Gold	0.128
Silver	0.235
Copper	0.385
Iron	0.449
Aluminum	0.903
Ethanol	2.42
Water	4.184

Heat capacity demonstration

- **(Perimeter Institute for Theoretical Physics)**
- **Heat capacity demonstration:**
- https://www.youtube.com/watch?v=KtGV5WUXrGc

- **Heat capacity experiment:**
- https://www.youtube.com/watch?v=vmJXud8vHtg

Conceptual Question ?

<u>What Factors</u> in different materials that have <u>same mass</u> determine the magnitude of the Specific Heat Capacity ?

- For Example,

1g of Liquid Water will absorb 4.184 Joules after its temperature increases by 1 C degree.

- 1g of Aluminum will absorb 0.903 Joules after its temperature increases by 1 C degree.
- 1g of Copper will absorb 0.385 Joules after its temperature increases by 1 C degree.
- 1g of Gold will absorb 0.128 Joules after its temperature increases by 1 C degree.

Specific Heat Capacity <u>is directly proportional</u> to which factor?

- Hint: Possible factors to consider:
- mass of each atom,
- the number of moles of atoms in each sample,
- density of material.

Conceptual Question ?

What Factors in different materials that have same mass determine the magnitude of the Specific Heat Capacity ?

- For Example,

1g of Liquid Water will absorb 4.184 Joules after its temperature increases by 1 C degree.

- 1g of Aluminum will absorb 0.903 Joules after its temperature increases by 1 C degree.
- 1g of Copper will absorb 0.385 Joules after its temperature increases by 1 C degree.
- 1g of Gold will absorb 0.128 Joules after its temperature increases by 1 C degree.

Specific Heat Capacity is directly proportional to which factor?

- Hint: Possible factors to consider (Will return in CHAPTER 6):
- mass of each atom,
- the number of moles of atoms in each sample,
- density of material.

- **HINT (After Chapter 6 is covered):**

- MAKE A PLOT ON EXCEL TABLE TO SEE WHICH OF THESE THREE FACTORS
- Correlate the best with the specific heat capacity!
- First, you need to find out the quantities needed from periodic table of elements
- And other sources.

3.12 Energy and Heat Capacity Calculations

The equation that relates these quantities is

$$\text{Heat} = \text{Mass} \times \text{Specific Heat Capacity} \times \text{Temperature Change}$$

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

- ***q*** **is the amount of heat in joules.**
- ***m*** **is the mass of the substance in grams.**
- ***C*** **is the specific heat capacity in joules per gram per degree Celsius.**
- ***ΔT*** is the **temperature change in Celsius.**
- The symbol Δ means *the change in*, so Δ*T* means *the change in temperature*.

Relating Heat Energy to Temperature Changes

Gallium is a solid at 25.0 °C but melts at 29.9 °C.

- If you hold gallium in your hand, it melts from your body heat.
- How much heat must 2.5 g of gallium absorb from your hand to raise its temperature from 25.0 °C to 29.9 °C?
- The specific heat capacity of gallium is 0.372 J/g °C.

Given:

2.5 g gallium (m)

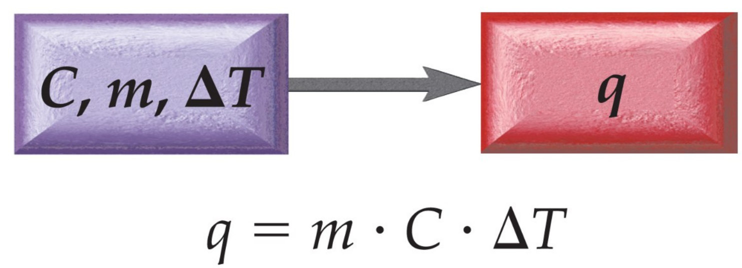
T_i = 25.0 °C

T_f = 29.9 °C

C = 0.372 J/g°C

Find: q

Solution Map



Relationships Used

$q = m \cdot C \cdot \Delta T$

Relating Heat Energy to Temperature Changes

Gallium is a solid at 25.0 °C but melts at 29.9 °C.

$$C = 0.372\,\mathrm{J/g\,°C}$$

$$m = 2.5\,\mathrm{g}$$

$$\Delta T = T_f - T_i$$

$$= 29.9\,°\mathrm{C} - 25.0\,°\mathrm{C}$$

$$= 4.9\,°\mathrm{C}$$

$$q = m \cdot C \cdot \Delta T$$

$$= 2.5\,\cancel{\mathrm{g}} \times 0.372\,\frac{\mathrm{J}}{\cancel{\mathrm{g}}\,°\cancel{\mathrm{C}}} \times 4.9\,°\cancel{\mathrm{C}} = 4.557\,\mathrm{J} = 4.6\,\mathrm{J}^*$$

An additional 200 J would be needed to melt 2.5 g of gallium once it reaches the melting point. ΔH_{fus}=5.59 kJ/mol.

Recommended You Tube video **problem solving session:**

- **CHEMISTRY TUTORIALS (Self-check with solutions):**
- **Heat and Specific heat capacity (VIDEO tutorials)**

- https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:thermodynamics/x2eef969c74e0d802:heat-capacity-and-calorimetry/v/heat-capacity

Self Check!

- What is the mass of the gold coin that
- Undergoes 40.0 C temperature increase after receiving q=55.0 Joules of heat?

- ***Given: C(Au)* = 0.128 J/g°C**
- **m = ?**

Self-check Problem:

(25) An unknown metal is suspected to be gold. When 2.8 J of heat was added to 5.6 grams of this sample its temperature increased by 3.9 degrees Celsius.

Are these data consistent with the sample being gold?

TABLE.

==

Metal	Specific Heat Capacity(J/g·°C)
Aluminum	0.903
Copper	0.385
Gold	0.128
Iron	0.449
Silver	0.235
Lead	0.128
copper	0.385
ethanol	2.42
water	4.18

==

Self-check Problem (Solution):

Chapter 4

(25) Using the formula relating heat:

$$q = m \times c_x \, (T_f - T_i).$$

Given:

$q = 2.8\,J$

$\Delta T = T_f - T_i = 3.9\,^{\circ}C$

$m = 5.6\,g$

Solution: gold has a specific heat capacity ($c_{Au} = 0.128 \frac{J}{g\,^{\circ}C}$).

If the data yields to c_x (unknown) being $= 0.128 \frac{J}{g\,^{\circ}C}$ then it would be consistent with gold.

Let's check!

$q = m \cdot c_x \cdot \Delta T \Rightarrow$

$$\Rightarrow c_x = \frac{q}{m \cdot \Delta T} = \frac{2.8\,J}{5.6\,g \cdot 3.9\,^{\circ}C} = \boxed{0.128 \frac{J}{g\,^{\circ}C}}$$

Yes, it matches: $c_x = c_{Au} = \boxed{0.128 \frac{J}{g\,^{\circ}C}}$. ✓

Chapter 3 in Review

- **Matter:** Matter is anything that occupies space and has mass. Matter can exist as a **solid**, **liquid**, or **gas**. Solid matter can be either **amorphous** or **crystalline**.
- **Classification of matter: Pure** matter is either an **element** (a substance that cannot be decomposed into simpler substances) or a **compound** (a substance composed of two or more elements in fixed definite proportions).
- **Mixtures:** Two or more different substances, the proportions of which may vary from one sample to the next.
- **Homogeneous mixture**: Having the same composition throughout.
- **Heterogeneous mixture**: Having a composition that varies from region to region.

Chapter 3 in Review

- **Properties and Changes of Matter:**
- The **physical properties** of matter do not involve a change in composition.
- The **chemical properties** of matter involve a change in composition.
- In a **physical change**, the appearance of matter may change, but its composition does not.
- In a **chemical change**, the composition of matter changes.

Chapter 3 in Review

- **Conservation of mass:** Matter is always conserved. In a chemical change, the sum of the masses of the reactants must equal the sum of the masses of the products.
- **Energy:** Energy is conserved—it can be neither created nor destroyed. Units of energy are the *joule (J)*, the *calorie (cal)*, the nutritional *Calorie (Cal)*, and the *kilowatt-hour (kWh)*.
- Chemical reactions that emit energy are **exothermic**; those that absorb energy are **endothermic**.
- **Temperature:** The temperature of matter is related to the random motions of the molecules and atoms. Temperature is measured in three scales: *Fahrenheit (°F)*, *Celsius (°C)*, and *Kelvin (K)*.
- **Heat capacity:** The temperature change that matter undergoes upon absorption of heat is related to the heat capacity of the substance composing the matter.

Chemical Skills

- Classifying Matter
- Physical and Chemical Properties
- Physical and Chemical Changes
- Conservation of Mass
- Conversion of Energy Units
- Converting between Temperature Scales
- Energy, Temperature Change, and Heat Capacity Calculations