Chapter 3

Matter and Energy

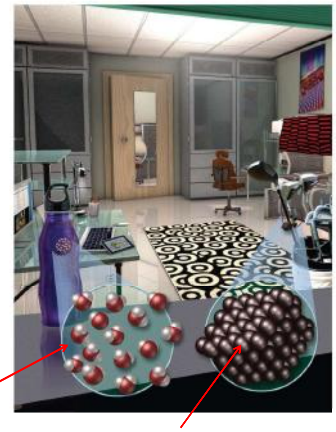
Based on slides provided with Introductory Chemistry, Fifth Edition
Nivaldo J. Tro

By Dr. Sylvia Esjornson
Southwestern Oklahoma State University
Weatherford, OK

With modifications and additions by
Dr. Deniz Cizmeciyan-Papazyan and Dr. Arno Papazyan

Atoms and Molecules 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.



Defining Matter

- **Matter** is defined as anything that has mass.
  -- Typically it also occupies space
- 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.

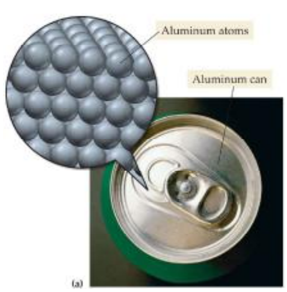
Matter Is Composed of Atoms and Molecules

- Matter may appear smooth and continuous, but actually it is not.
- Matter is ultimately composed of **atoms,** submicroscopic (impossible to see with an optical microscope) 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.

Independent Atomic Particles in Aluminum

Atoms and molecules

All matter is ultimately composed of atoms.
In some substances, such as aluminum, the atoms exist as independent particles. Then the particle characterizing the substance is an <u>atom</u>.

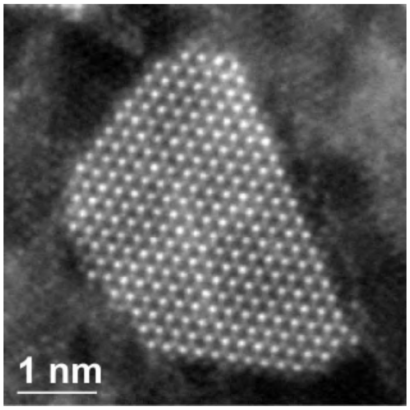


Well-Defined Molecular Particles in Isopropanol

Atoms and molecules

All (ordinary) matter is ultimately composed of atoms. In some substances, such as rubbing alcohol, several atoms bond together in well-defined structures called <u>molecules</u>. Then the particle characterizing the substance is a <u>molecule</u>.

Atoms in a molecule always “travel together”, even when the substance is dissolved in another substance, or is melted.



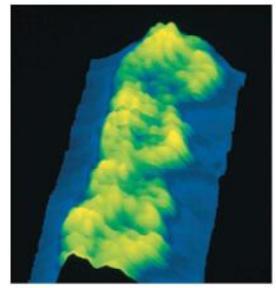
Individual atoms can now be imaged using various techniques.

This one is taken by using a very high resolution electron microscope.

A single atom thick molybdenum disulphide nanoparticle on a thin graphite supportImage: Q M Ramasse (SuperSTEM Laboratory), L Hansen and S Helveg (Haldor Topsoe A/S)

Image of DNA by Scanning Tunneling Microscope

- **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 is discernible.

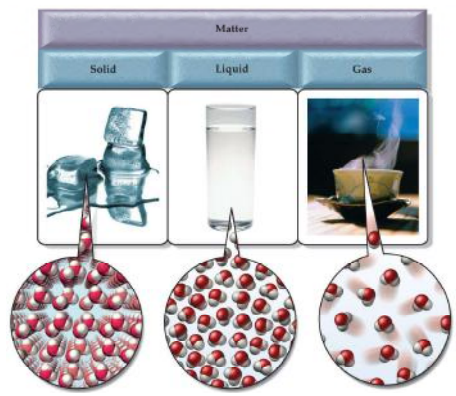


The Common States of Matter

- The common **states of matter** are **solid, liquid,** and **gas.**

- **Matter can be classified according to its states: 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.

Three States of Matter



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 are arranged in geometric patterns with long-range, repeating order.
- **Amorphous solid:** Atoms or molecules do not have long-range order.

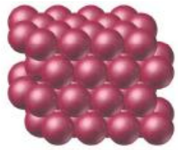
- Examples of *crystalline* solids include salt and diamond.
- The well-ordered, geometric shapes of salt and diamond crystals reflect the well-ordered geometric arrangement of their atoms.

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

States of Matter: Types of Solids

(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.
-- Like a snapshot of a liquid

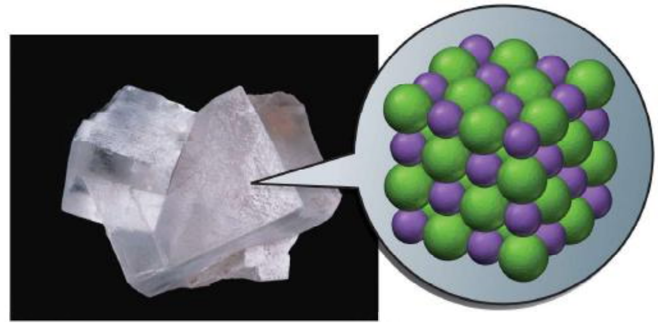


(a) Crystalline solid

(b) Amorphous solid

States of Matter: Crystalline Solid

- 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

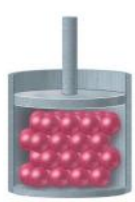
- In liquid matter, atoms or molecules are close to each other but are free to move around and by each other.
- Liquids have a fixed volume because their atoms or molecules are in close contact.
- Liquids assume the shape of their containers because the atoms or molecules are free to move relative to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

States of Matter: Gas

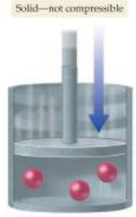
- In gaseous matter, atoms or molecules are separated by large distances and are free to move relative to one another.
- Since the atoms or molecules that compose gases are not in contact with one another, gases are **compressible.**
- Gases always assume the shape and volume of their containers.
- Oxygen,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

Gases Are Compressible

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



Solid—not compressible



Gas—compressible

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

Classifying Matter According to Its Composition

- 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.

Classifying Matter: Composition of Elements

- **Element:** A pure substance that cannot be broken down into simpler substances
- No chemical transformation can decompose an element into simpler substances.

-- That was Dalton's definition; OK but not the best

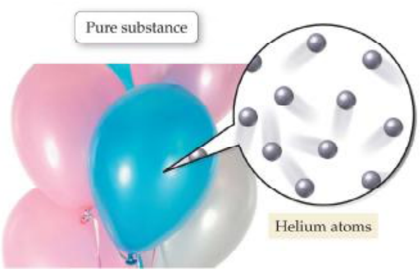
A more modern definition:

- **Element:** A substance that contains only one kind of atom

- All known elements are listed in the periodic table.
- A periodic table can be found on the inside front cover of your book, and the elements are listed in alphabetical order on the inside back cover.

Classifying Matter: Composition of the Element Helium

Helium: A pure substance composed of only one type of atom—helium atoms

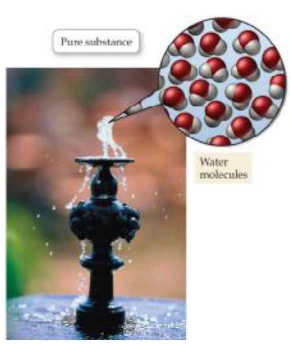


Classifying Matter: Composition of Compounds

- **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 (i.e. elements).

Classifying Matter: Composition of Compound Water

- Water: A pure substance composed only of water molecules
- Two elements in a fixed, definite proportion
  -- and we have no control over it; it is dictated by what it is



Classifying Matter: Compounds and Mixtures

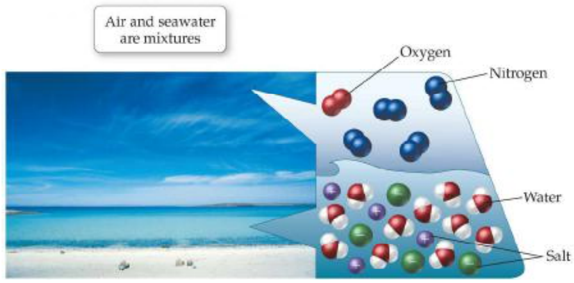
When matter contains two (or more) types of atoms, it may be a ***pure substance*** or a ***mixture.***

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.
-- We can vary the composition of a mixture
-- More than one substance occupying the same volume.

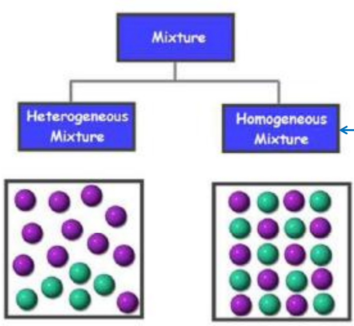
Classifying Matter: Mixtures

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



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 water) or a ***heterogeneous mixture*** (such as oil and water).



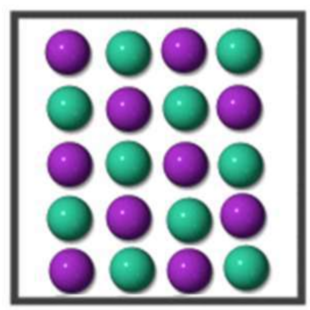
Synonymous with "solution"

Components can be solid, liquid, or gas

Homogeneous Mixture

So intimately mixed that no matter how closely we look into it, the mixture looks uniform, and we find the same proportion of the constituents, until we come to the individual, characteristic particles of the substances.

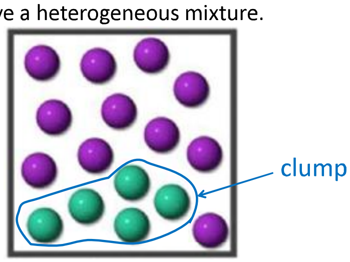
Also called a **"solution"**



Heterogeneous Mixture

Individual components clump together. If we look closely, we can see the individual components forming distinct pieces or regions.

A mixture may look homogeneous, but at close enough inspection (say, using microscopes), if we can discern regions with different compositions, we have a heterogeneous mixture.

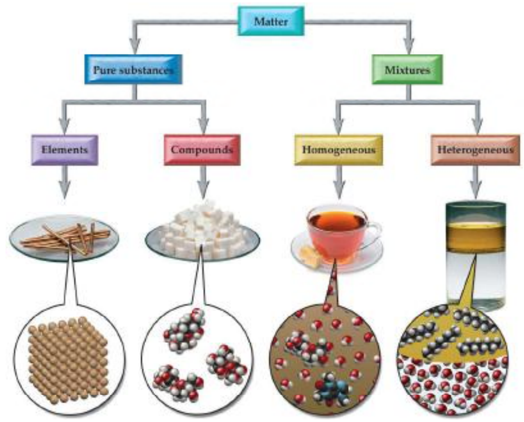


As long as the clumps (regions) are made up of many molecules (or individual atoms as the case may be), the mixture is heterogeneous.

Visible light scatters off the "clumps" in a heterogeneous mixture when the clumps are larger than the wavelength (basically the size of the light waves)

If a mixture is not "clear" (if it is "milky", or "cloudy") it is heterogeneous.
-- unless the individual molecules, even when not clumped together, are so huge that they act like "clumps"

If a mixture is "clear" (not "milky" or "cloudy") there is a good chance it is homogeneous (therefore a "solution")
-- but it may be a heterogeneous mixture with clumps too small for visible light to scatter off



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.

Practice
Check your solution on next page

Decide between pure substances and mixture

- steam
- orange juice
- cappuccino
- oxygen
- vegetable soup
- air
- sand

Practice

Decide between pure substances and mixture

- steam pure
- orange juice mixture of water and various organic materials
- Cappuccino mixture
- oxygen pure
- vegetable soup mixture
- Air mixture
- sand mixture of quartz and crushed stones

Practice:
Check your solution on next page

Decide between homogeneous and heterogeneous mixture

- mud
- salad dressing
- salsa
- air

Practice

Decide between homogeneous and heterogeneous mixture

- Mud Heterogeneous
- salad dressing Heterogenous
- Salsa Heterogeneous
- Air Homogenous

Different Kinds of Matter Have Different Properties

- A **physical property** is one that a substance displays without changing what it is
- A **chemical property** is one that a substance displays only through changing what it is through a chemical process.

Difference in Physical and Chemical Properties

- The ***density*** of gasoline is a ***physical*** property—we don't need to conduct chemical reactions when we measure it.

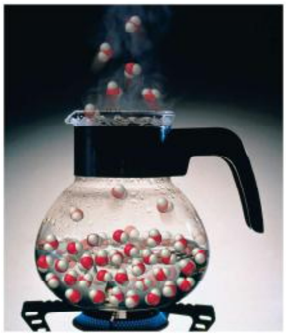
- The ***flammability*** of gasoline is a ***chemical*** property—it must change (by a chemical reaction) to other substances when it burns.

Physical Change of Water

- The atomic or molecular ***composition*** (i.e. its characteristic units that make it what it is) of a substance ***does not change*** when the substance displays its physical properties.
- The boiling point of water—a physical property—is 100 °C.
- When water boils, it changes from a liquid to a gas, but the gas is still water.

Physical Change of Water from Liquid to Gas

- **A physical property**
- The boiling point of water
- Boiling is a physical change.
- When water boils, it turns into a gas, but the water molecules are the same in both the liquid water and the gaseous steam.



Changes of State Are Physical Changes

State changes—transformations from one state of matter to another, such as from solid to liquid—are always physical changes.

State change is commonly known as "phase change" e.g. "going from liquid phase to gas phase" is a "phase change".

No chemical bonds are made or broken in physical changes.

- We don't officially know what a "chemical bond" is yet, right?
- You want to think about what they are?

More on Physical Changes

- 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.
  - Nothing changed at the atomic scale

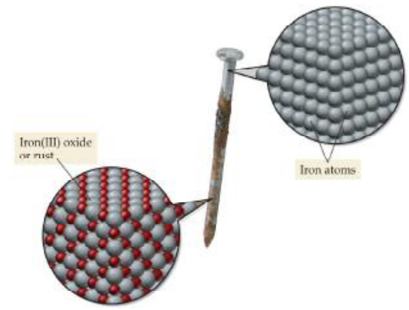
- But not all physical changes are phase changes like melting or boiling.

More on Physical Changes

- Crushing a crystal into powder is a physical change. Nothing changes at the atomic scale.
- Even when atoms are rearranged without breaking/making chemical bonds, it would be a physical change.
- Changing temperature or pressure would be a physical change

Kinds of Chemical Properties from Chemical Changes

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



How Matter Changes in Chemical Reactions

- Matter undergoes a chemical change when it undergoes a **chemical reaction.**
  - Chemical bonds are made and/or broken
- In a chemical reaction, the substances present before the chemical change are called **reactants.**
- The substances present after the change are called **products.**

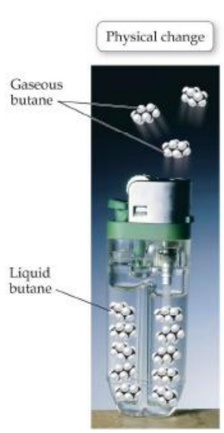
$$\text{Reactants} \xrightarrow[\text{Chemical Change}]{} \text{Products}$$

The Difference Between 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.
  -- *Chemical bonds* <u>are not</u> made and/or broken
- In chemical changes, atoms do change their fundamental associations, resulting in matter with a new identity.
  -- *Chemical bonds* <u>are</u> made and/or broken
- *A chemical change results in a new substance*.

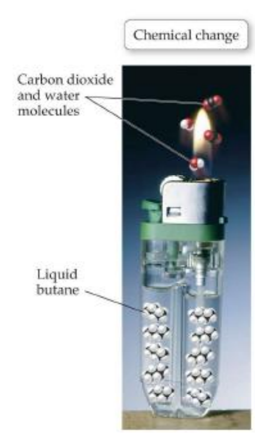
Vaporization: A Physical Change

- 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.



Burning: A Chemical Change

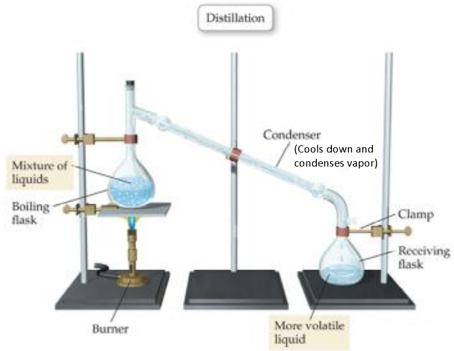
- Push the button *and* turn the flint to create a spark.
- Produce a flame.
- The butane molecules react with oxygen molecules in air to form new molecules, carbon dioxide and water.
- This is a chemical change.



Distillation Separates Mixtures Through Physical Changes

Distillation

Separating substances using their difference in volatilities



Filtration Separates Mixtures Through Physical Changes

Filtration

Separation of components based on their particle size

- Liquid "particles" are tiny molecules that easily pass through the filter's pores
- Solid particles are trapped if we use a filter with small enough pores

Practice:
Check your solution on next page

Decide between Physical and Chemical Property

- corrosiveness of sulfuric acid
- toxicity of cyanide
- flammability of gasoline
- ability of antacid to neutralize stomach acid
- lead becomes a liquid when heated to 601°C
- water freezes at 0^oC

Practice

Decide between Physical and Chemical Property

- corrosiveness of sulfuric acid Chemical Property
- toxicity of cyanide Chemical Property
- flammability of gasoline Chemical Property
- ability of antacid to neutralize stomach acid Chemical Property
- lead becomes a liquid when heated to 601°C Physical Property
- water freezes at 0^oC Physical Property

Practice:
Check your solution on next page

Decide between Physical and Chemical Change

- water, when heated, forms steam Physical
- bleach turns hair yellow Chemical
- sugar, when heated, becomes brown Chemical
- milk turns sour Chemical
- Chemical
- apples, when exposed to air, turn brown

Practice:

Decide between Physical and Chemical Change

- water, when heated, forms steam Physical
- bleach turns hair yellow Chemical
- sugar, when heated, becomes brown Chemical
- milk turns sour Chemical
- apples, when exposed to air, turn brown Chemical

Law of 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. $E = mc^2$; big change in energy E (the heat generated by the nuclear reaction) leads to measurable change in m.
- 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.

Law of Conservation of Mass

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.

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

- Keep this law in mind.
- Sometimes impossible-looking problems are easily solved by applying this simple law.

The Behavior of Matter Is Driven by Energy

- **Energy** is a major component of our universe.
- *Energy is the capacity to do work.*
- **Work** is defined as the result of moving against a force
- So **energy is the ability to move against a force**
- The behavior of matter is driven by energy.
- Understanding energy is critical to understanding chemistry.

Law of Conservation of 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

Each of those energies involve directly or indirectly a “capacity to move against a force”

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.

But the first person to deduce that energy can be converted from one form to another was a German physician Julius Robert von Mayer on a Dutch ship in the tropics, by observing the change in the color of blood (oxygenation level) with the surrounding temperature. !!!!!!!!

Units of Energy

- 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 °C. A calorie is a larger unit than a joule.
- A related energy unit is the nutritional or *capital C* **Calorie (Cal),** (food calories) equivalent to 1000 *little c* calories (regular, scientific calories).

Units of Energy

- The **kilowatt-hour (kWh)** is 1 kJ per second times 3600 seconds (which is an hour).
  - -- Watt(W) = J/s (energy per time)
  - -- kWh = (1000 J/s)(3600 s) = 3.6×10^6 J exactly
- The average cost of residential electricity in the United States is about $0.12 per kilowatt-hour.

Various Energy Units and Their Conversion Factors

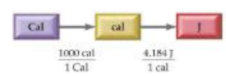
Conversion factors below are exact (their values are defined)

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

"Food calorie"

Example: Conversion of Energy Units

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



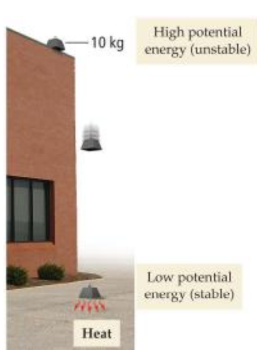
1000 calories = 1 Cal (Table 3.2)
4.184 J = 1 cal (Table 3.2)

- GIVEN: 225 Cal
- FIND: J

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

Potential Energy of Raised Weight

A weight lifted off the ground has a high potential and will fall toward the ground to lower its potential energy.



Chemical Changes Change Potential Energy

- *Systems with high potential energy have a tendency to change in a way that lowers their potential energy.*
- When the potential energy is different at different points in space, there is a force that pushes the object in the direction of lower potential energy
  - That's pretty much the definition of force
- Objects or systems with high potential energy tend to be *unstable.*
  - Forces that develop push the particles/objects in the direction of lower potential energy

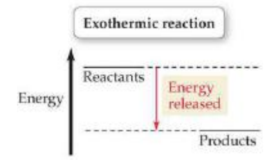
Chemical Changes Change Potential Energy

- Some chemical substances, such as the molecules that compose TNT (trinitrotoluene), have a relatively high potential energy.
  -- Their chemical bonds are not very stable
- TNT molecules tend to undergo rapid chemical changes that lower their potential energy (producing compounds with more stable bonds), which is why TNT is explosive.

Exothermic and Endothermic Reactions

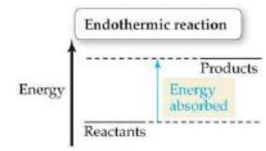
In an exothermic reaction, energy is released.

— Bonds in products are more stable (lower in energy) than bonds in reactants



In an endothermic reaction, energy is absorbed.

— Bonds in products are less stable (higher in energy) than bonds in reactants



Thermal Energy: Constant Random Motion of Matter

- 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 thermal energy.
- The hotter an object, the greater the random motion of the atoms and molecules that compose it, and the higher its temperature.

Do 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 thermal energy of matter (not the exchange of thermal energy).

Background of the Fahrenheit Temperature Scale

The Fahrenheit scale assigns

- 0 °F to the freezing point of a concentrated saltwater solution; and
- 98 °F to normal body temperature.

On the **Fahrenheit (°F) scale,**

- water freezes at 32 °F; and
- water boils at 212 °F.

Different Temperature Scales: Celsius, Kelvin

On the **Celsius (°C) scale,**

- water freezes at 0 °C; and
- water boils at 100 °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.

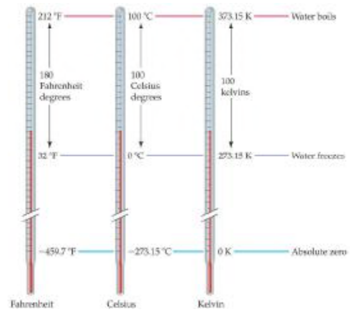
On the **Kelvin (K) scale,**

- water freezes at 273 K; and
- water boils at 373 K.

“Room temperature” definition varies.
In scientific contexts, it is often taken to be 25 °C, or 298 K.

Comparison of the 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.
  - -- we don't say "degree Kelvin",
  - -- we just say "Kelvin"



Converting Between Temperature Scales

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

$$T_K = T_C + 273.15$$

$$T_C = \frac{T_F - 32}{1.8} \qquad T_F = 1.8\,T_C + 32$$

Converting Between Temperature Scales

But note that change (Δ) in temperature only cares about the relative sizes of the different "degrees".

$\Delta T_K = \Delta T_C$ Kelvin and Celsius "degrees" are the same size

$\Delta T_F = 1.8\ \Delta T_C$ Fahrenheit degree size is 1.8 times
$\Delta T_F = 1.8\ \Delta T_K$ larger than Celsius and Kelvin

Practice:
Check your solution on next page

- Lead melts at 601.0°C. What temperature is this in °F?
- The outside air temperature is 30°F, what is the temperature in Kelvin?
- Helium boils at 4K what is the temperature in °C
- If you raise the thermostat by 5°F, by how much will the temperature change in °C?
- Temperature change between day and night on the moon is around 540 °F. What is it in Kelvins?

Practice

- Lead melts at 601.0°C. What temperature is this in °F?

$T_F = 1.8\,T_C + 32$ plug in numbers to get 1114 °F

- The outside air temperature is 30°F, what is the temperature in Kelvin? First lets change 30°F, to °C using $T_C = \frac{T_F - 32}{1.8}$ leads -1.1 °C now $T_K = T_C + 273.15$ leads 272 °F,

- Helium boils at 4K what is the temperature in °C

$T_K = T_C + 273.15$ rearrange the formula to get -269.15°C

- If you raise the thermostat by 5°F, by how much will the temperature change in °C? $\Delta T_F = 1.8\ \Delta T_C$ rearrange the formula and plug in numbers to get 2.8 °C.

- Temperature change between day and night on the moon is around 540 °F. What is it in Kelvins? $\Delta T_F = 1.8\ \Delta T_K$ rearrange the formula to get 300°C

Temperature Changes: Heat Capacity

- **Specific heat capacity** (or the **specific heat**): The quantity of heat required to change the temperature of **1 gram** of a substance by 1 °C (or 1 Kelvin).

- It has has units of joules per gram per degree Celsius
  J/(g °C) or J/(g K).

Temperature Changes: Heat Capacity

The temperature in the units of heat capacities always corresponds to the change in temperature, not temperature itself. Therefore

J/(g °C) = J/(g K).

What would be the relationship between J/(g °C) and J/(g °F)?

Specific Heat Capacity for Several Substances

- Notice that water has the highest specific heat capacity on the list.
- This data table is needed 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

Energy Temperature Change and Heat Capacity

$$\text{Heat} = \text{Mass} \times \text{Specific Heat Capacity} \times \text{Temperature Change}$$
$$q = m \times C \times \Delta T$$

- 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 ΔT means *the change in temperature.*
- $\Delta T = T_{final} - T_{initial}$

When the substance warms up:

- ΔT is positive
- Therefore q is positive $\quad \overset{(+)}{q} = m\,C\,\overset{(+)}{\Delta T}$
- Heat is absorbed by the substance
- "Heat absorbed" = q

When the substance cools down:

- ΔT is negative
- Therefore q is negative $\quad \overset{(-)}{q} = m\,C\,\overset{(-)}{\Delta T}$
- Heat is released from the substance

- When q is negative, and we are asked "heat released", we report the absolute value of q, as a positive number.
  -- just as we say "we lost $1000" when our profit is –$1000

Practice:
Check your solution on next page

Calculate the amount of heat released when 7.40 g of water cools from 49.0° to 29/.0 °C

(c = 4.18 J/g·°C)

How much heat is absorbed by a copper penny with mass 3.10 g whose temperature rises from −8.0 °C to 37.0 °C? (c = 0.385 J/g·°C)

Practice

Calculate the amount of heat released when 7.40 g of water cools from 49.0° to 29.0 °C

(c = 4.18 J/g·°C)

q=mcΔT

= 7.40g x 4.18 J/g·°C x (29.0-49.0)^{0}C= -619J

or 619J of heat released

How much heat is absorbed by a copper penny with mass 3.10 g whose temperature rises from −8.0 °C to 37.0 °C? (c = 0.385 J/g·°C)

q=mcΔT

= 3.10g x 0.385 J/g·°C x [37.0 -(-8.0)] °C = 53.7 J

or 53.7 J of heat absorbed

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 is changed by a chemical reaction.

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 release 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 Learning Objectives

1. Classify matter as element, compound, or mixture.
2. Distinguish between physical and chemical properties.
3. Distinguish between physical and chemical changes.
4. Apply the law of conservation of mass.
5. Identify and convert among energy units.
6. Convert between Fahrenheit, Celsius, and Kelvin temperature scales.
7. Relate energy, temperature change, and heat capacity.