MATTER

Chapter Three (Sec. 3.1-3.10)

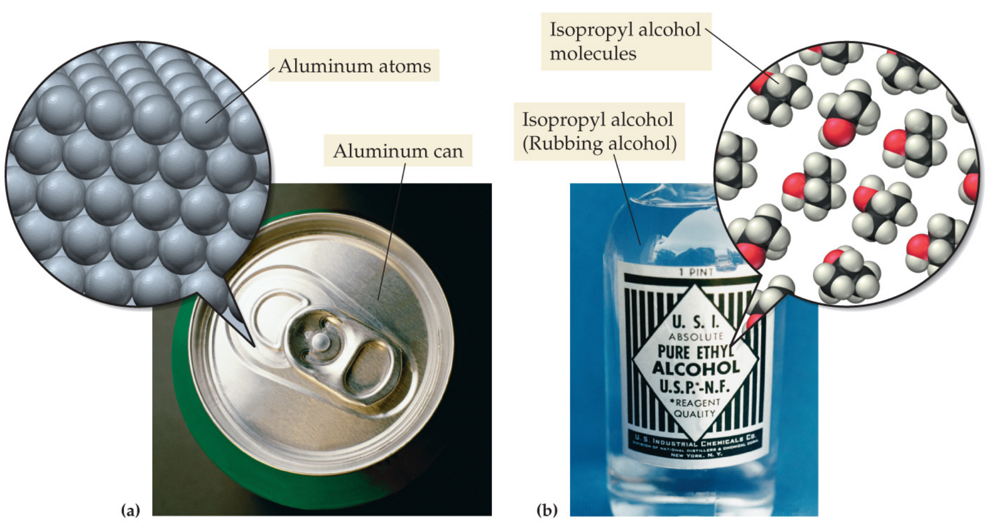
Matter

- Recall that matter
  - Has mass
  - Takes up space (i.e. has volume)
- Matter also has one of four states
  - Gas
  - Liquid
  - Solid
  - Plasma
    - We will <u>not</u> be concerned with the plasma state in this course!

Atoms and Molecules

- The fundamental unit of matter is the atom, particles which are incredibly small and come in numerous "varieties"
  - Atoms themselves are composed of smaller particles (which are themselves composed of even smaller particles), however, as we shall see, there is a certain uniqueness possessed by atoms which does not exist for smaller particles
- Molecules are groups of two or more atoms held together by covalent bonds
  - We will discuss these bonds later in the semester
- Molecules generally contain a well defined number of atoms which can be expressed in a whole-number ratio
- We shall see that other organizations of atoms do exist which we do not call molecules which, while they have a specific atom-to-atom ratio, contain more atoms than can be practically counted
  - While not an *infinite* number of atoms in reality, the number of atoms in these structures would certainly *seem* infinite if we tried to count them.

Atoms and Molecules



Atoms are represented as spheres in this (and most other) textbooks. Different colors and sizes represent different "types" of atoms.

States of Matter

- The states of matter are classified by two parameters
  - Does it take the shape of its container?
  - Does it completely fill its container?
- Gases
  - Take the shape of their container
  - Completely fill their container

States of Matter

- Liquids
  - Take the shape of their container
  - Do not fill their container completely
- Solids
  - Do not take the shape of their container
  - Do not fill their container completely

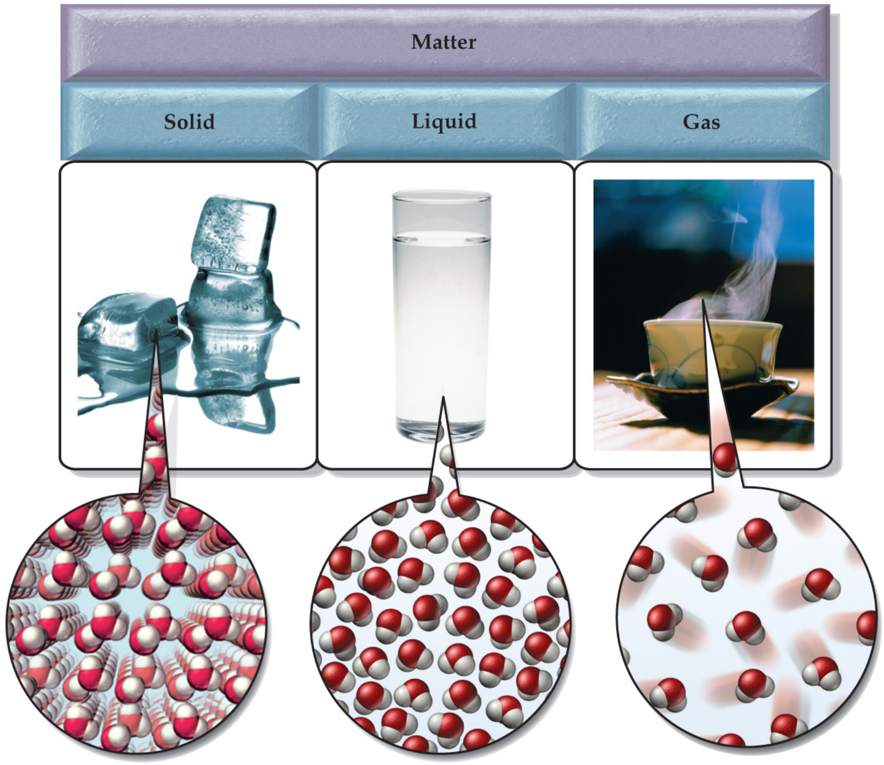


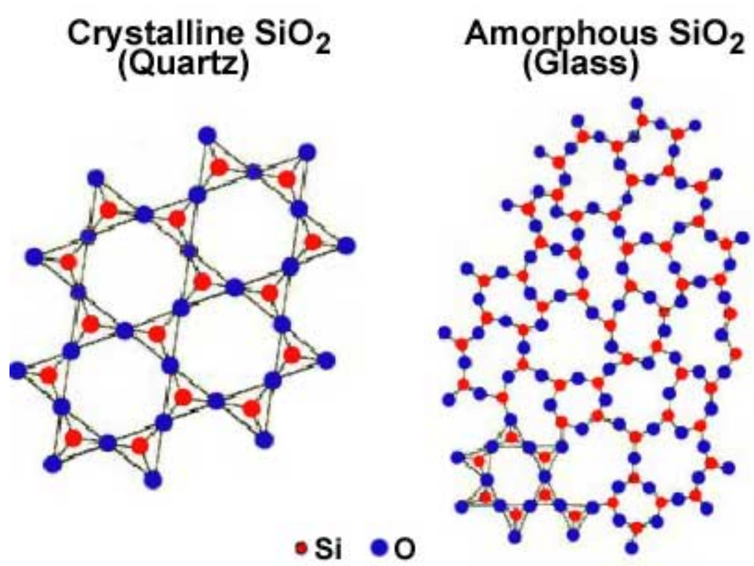
TABLE 3.1 **Properties of Liquids, Solids, and Gases**

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

Crystalline vs. Amorphous Solids

- If we were able to view solids at a submicroscopic level, we would see that two distinct types of solids occur
- Crystalline solids are composed of atoms in highly-ordered arrangements
  - Common examples include table salt and diamonds
- In amorphous solids, there is generally much less organization among atoms. In fact, the positions of atoms may in some cases appear to be quite random
  - Common examples include rubber, glass, plastics, and Styrofoam

Crystalline vs. Amorphous Solids

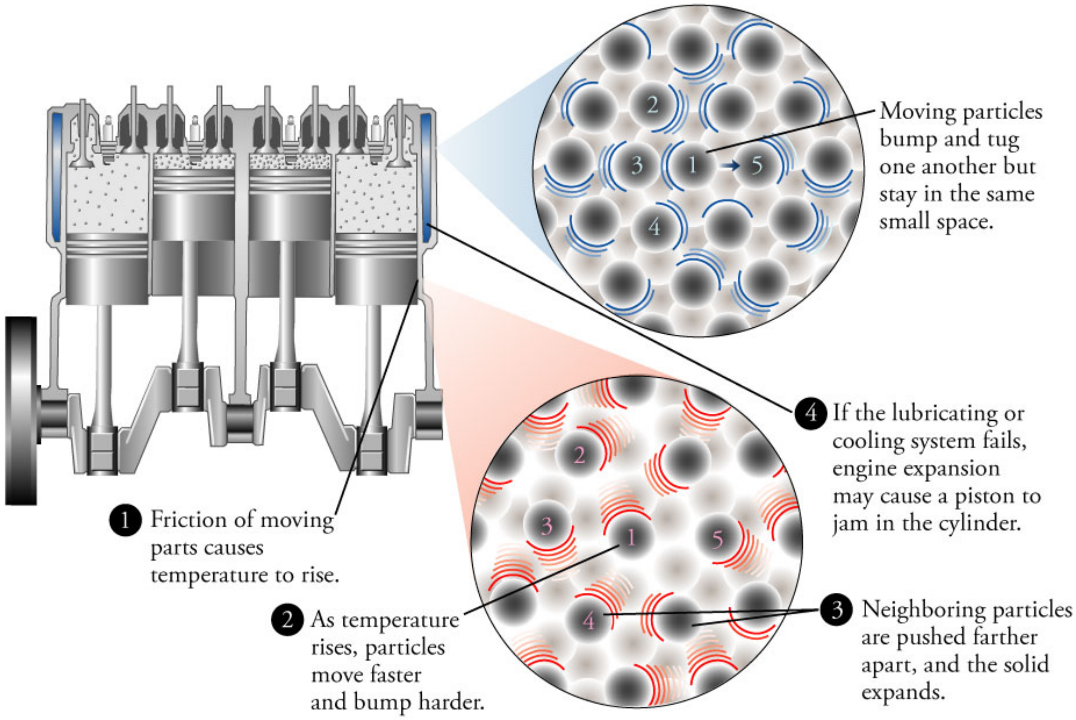


States of Matter

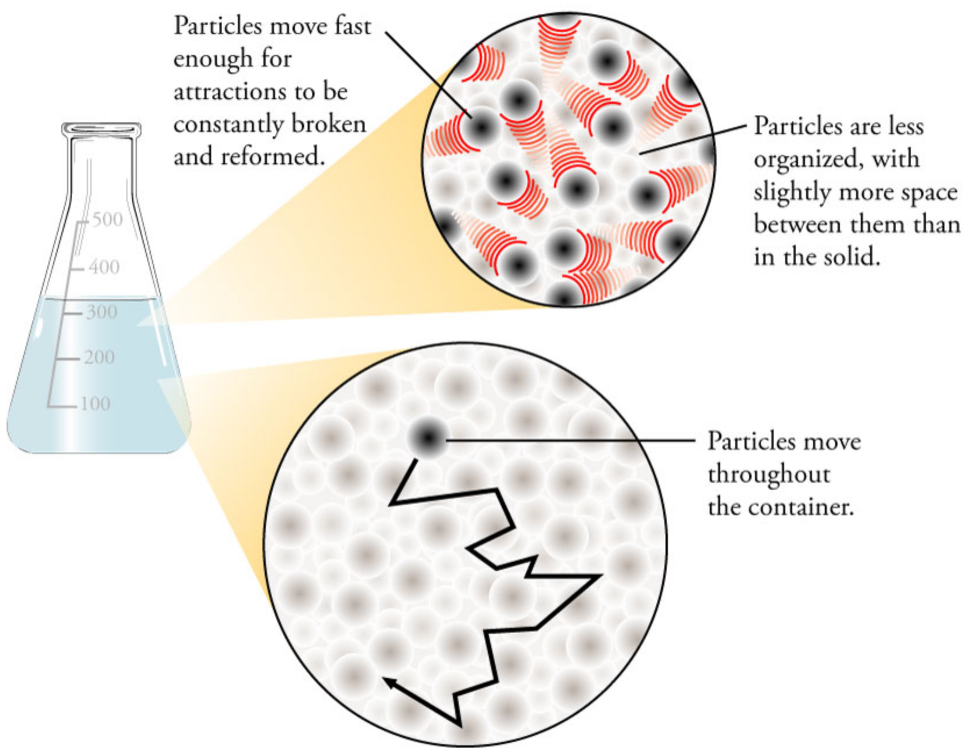
- Which state a substance is in depends on two factors
  - Temperature
  - Pressure (we will discuss this later in the course)
- Gases are comprised of molecules which generally are far apart from one another, and travel in random paths.

States of Matter

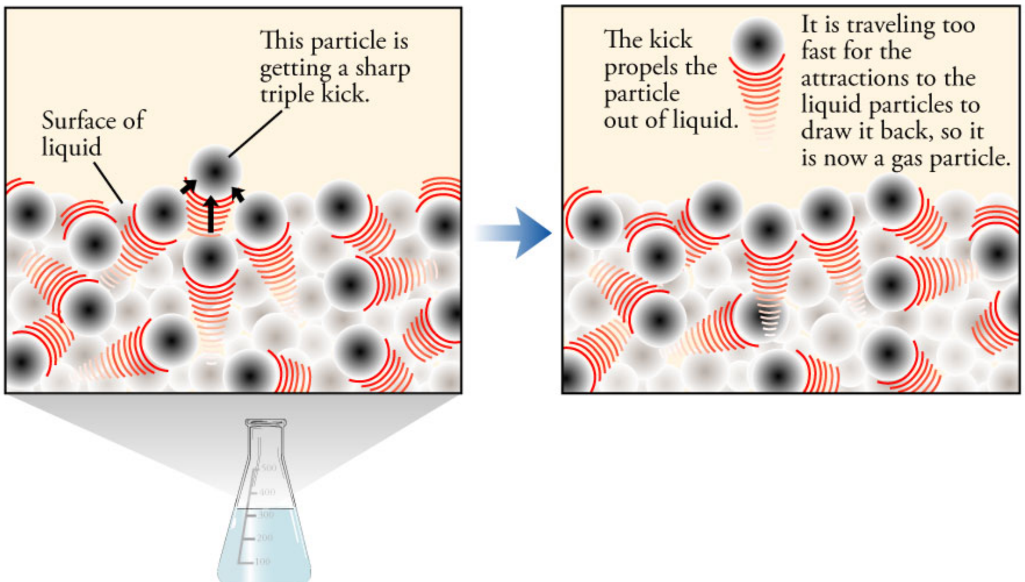
- The particles in liquids and solids are much closer together than those of gases
- The particles of solids are generally "stuck in place", but are able to vibrate
- Liquid particles freely move across one another

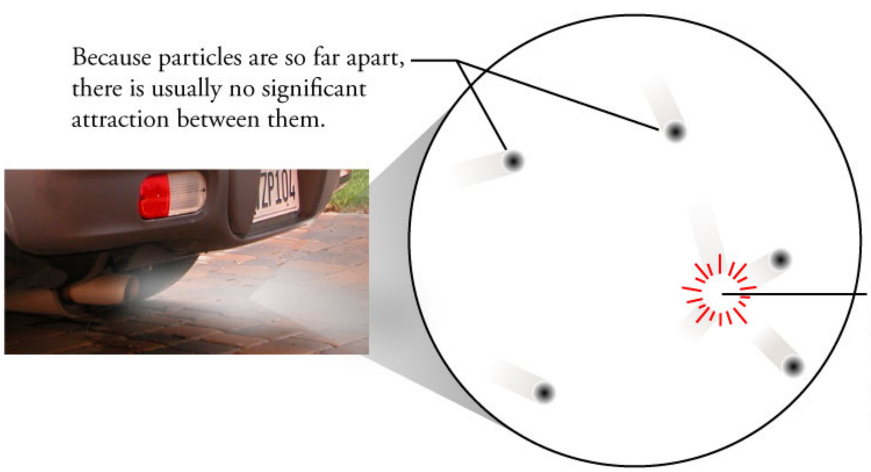


Moving particles bump and tug one another but stay in the same small space.
2
3
1
5
4
4 If the lubricating or cooling system fails, engine expansion may cause a piston to jam in the cylinder.
1 Friction of moving parts causes temperature to rise.
2 As temperature rises, particles move faster and bump harder.
3 Neighboring particles are pushed farther apart, and the solid expands.



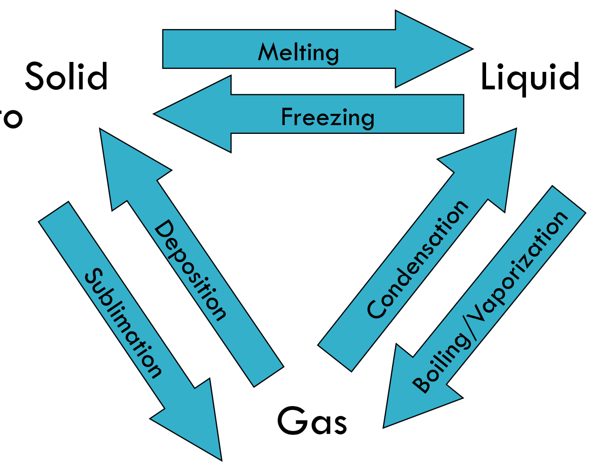
Particles move fast enough for attractions to be constantly broken and reformed.
Particles are less organized, with slightly more space between them than in the solid.
500
400
300
200
100
Particles move throughout the container.


This particle is getting a sharp triple kick.
Surface of liquid
The kick propels the particle out of liquid.
It is traveling too fast for the attractions to the liquid particles to draw it back, so it is now a gas particle.


Because particles are so far apart, there is usually no significant attraction between them.
ZP104
Particles move in straight paths, changing direction and speed when they collide.

Changes of State

- Special terms are associated when matter changes from one state to another.
  - You may be familiar with the term "freezing," which describes a change from liquid to solid.



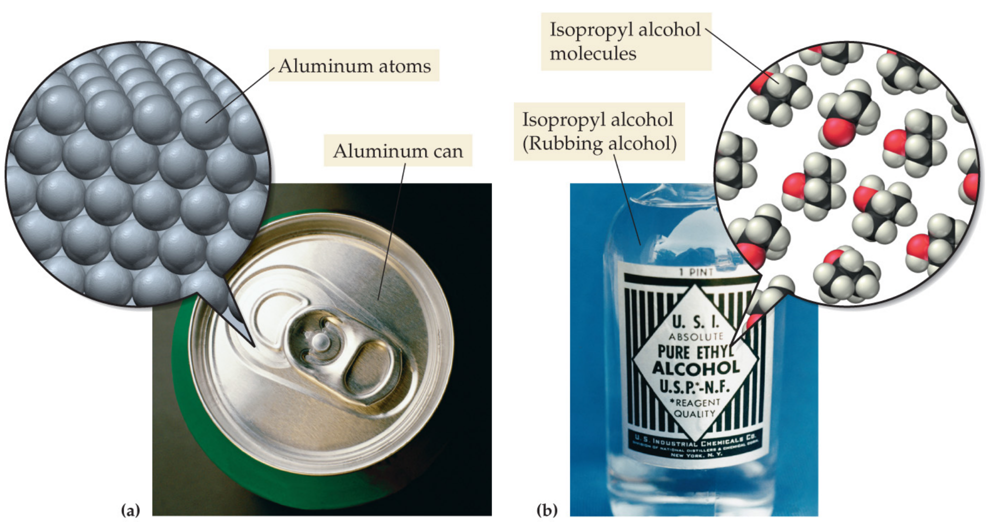
Classification of Matter

- Matter can generally be classified as either a <u>pure substance</u> or a <u>mixture</u>
- Pure substances include
  - Elements
  - Compounds
- Mixtures may be
  - Homogeneous, or
  - Heterogeneous

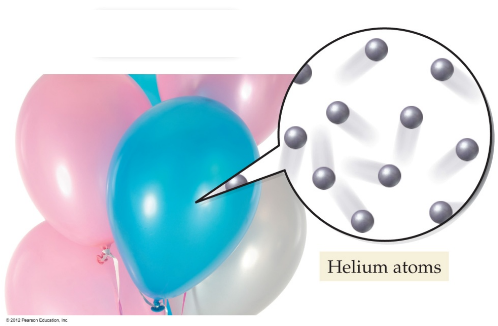
Pure Substances

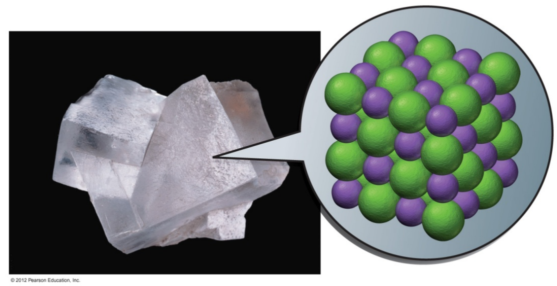
- Pure substances have a *consistent, uniform* composition throughout. They include
  - Elements
  - Compounds
- Elements are made up solely of the same type of atom (ex. nitrogen, neon, copper)
- Compounds are made up of atoms in a fixed, whole number ratio.

Pure Substances



Pure Substances





Mixtures

- The composition of a *mixture* is not fixed.
  - Consider salt water, a mixture of H_2O and NaCl.
    - Is salt water always found in the same proportion? The same atom-to-atom ratio?
- Mixtures can be classified into two types:
  - Homogeneous mixtures have all parts in the same state (gas, liquid or solid) <u>and</u> all parts must be mixed together.
    - If the parts of the mixture are *visually* inseparable, we will call the mixture homogeneous.
  - Heterogeneous mixtures are simply those which are not homogeneous.

(a)

(b)

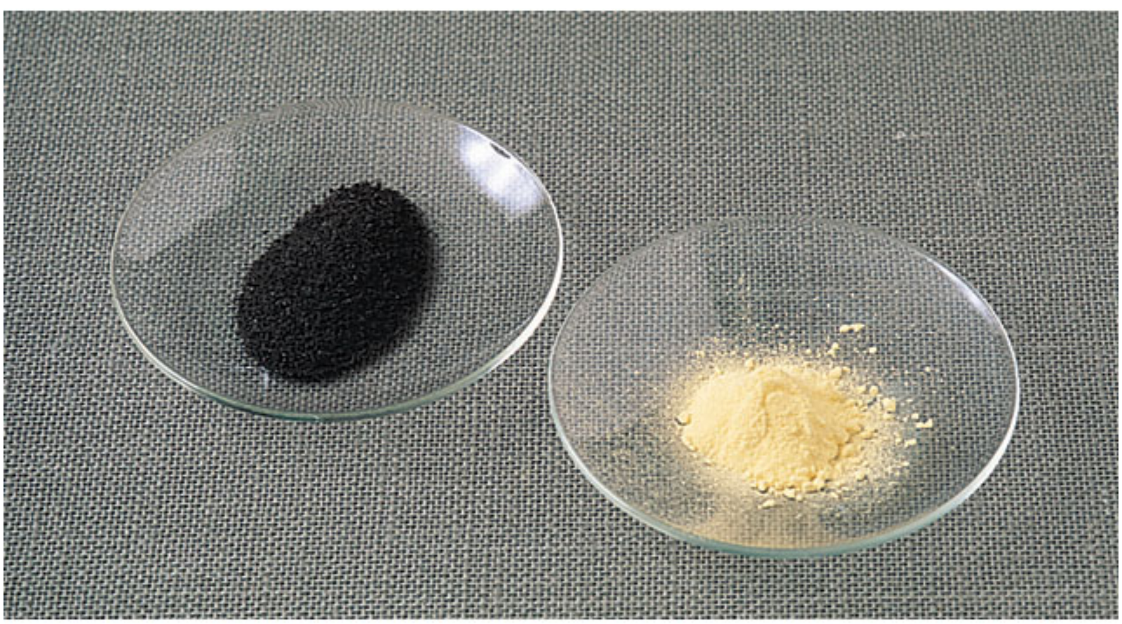
Mixtures

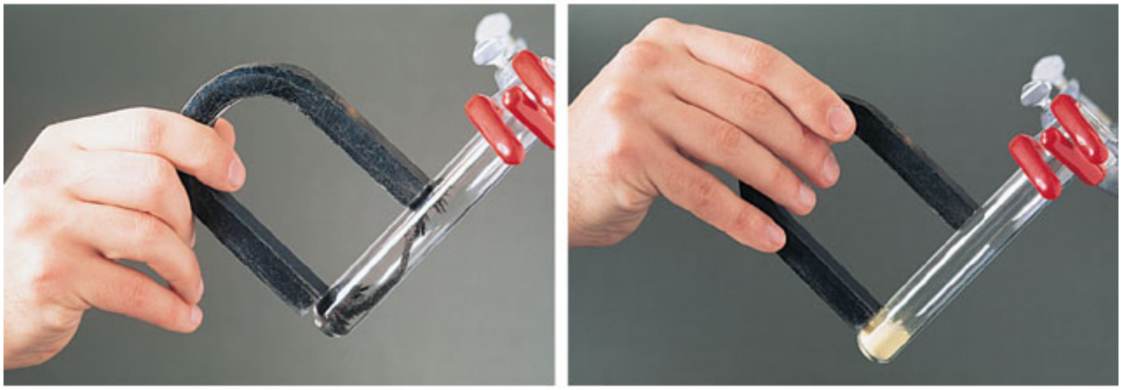
Are each of these mixtures homogeneous or heterogeneous? Why?

- Vodka (a mixture of water and ethyl alcohol)
- Cheerios in milk
- A mixture of oil and water
- A salt water solution
  - Note that the term <u>solution</u> is often used to refer to homogeneous mixtures, especially for compounds dissolved in water.
- Air
- Dirty, dust-filled air

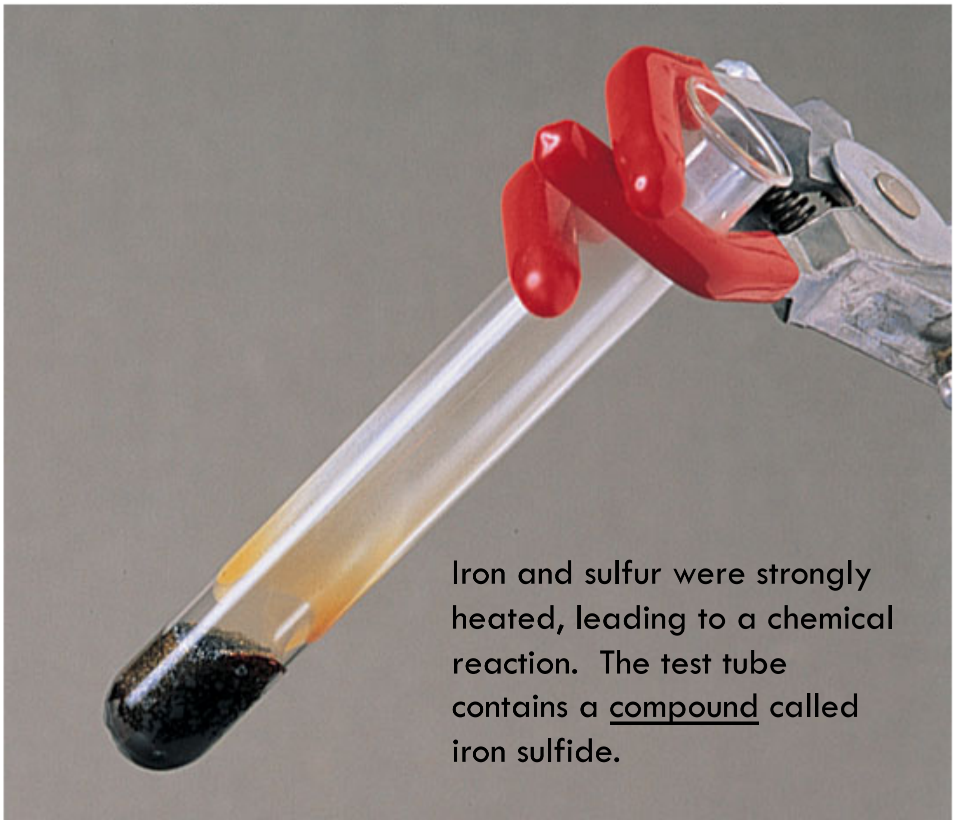
Separations

- Mixtures can be separated by *physical* methods
  - A heterogeneous mixture of coffee grounds and water can be separated by filter paper
  - The water can be removed from a salt water solution (a homogeneous mixture) by boiling the water off. The salt will remain in the container.
- Compounds can only be separated into their individual elements by *chemical* means (i.e. through the result of a chemical reaction)



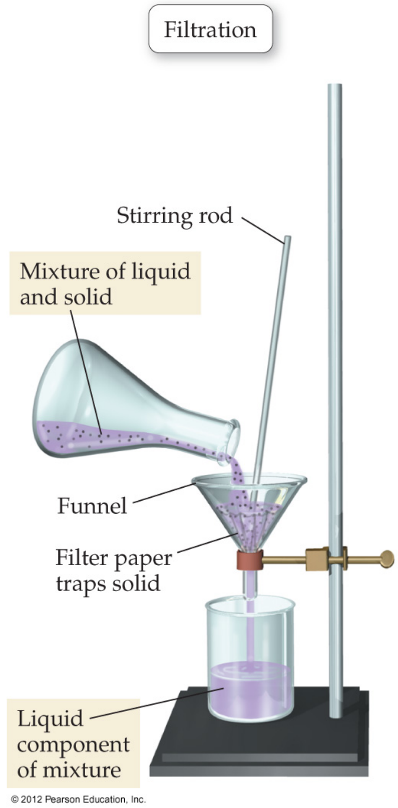


a mixture of iron and sulfur



Iron and sulfur were strongly heated, leading to a chemical reaction. The test tube contains a compound called iron sulfide.

the compound iron sulfide


Filtration
Stirring rod
Mixture of liquid and solid
Funnel
Filter paper traps solid
Liquid component of mixture
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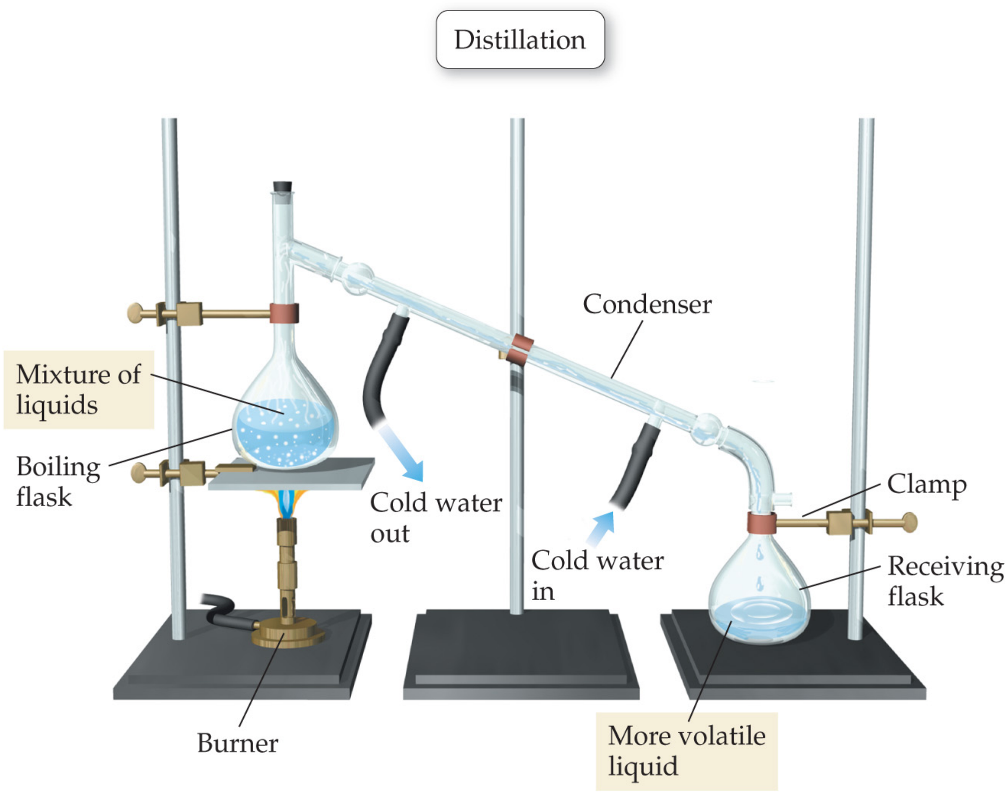
Distillation
Condenser
Mixture of liquids
Boiling flask
Cold water out
Cold water in
Clamp
Receiving flask
Burner
More volatile liquid
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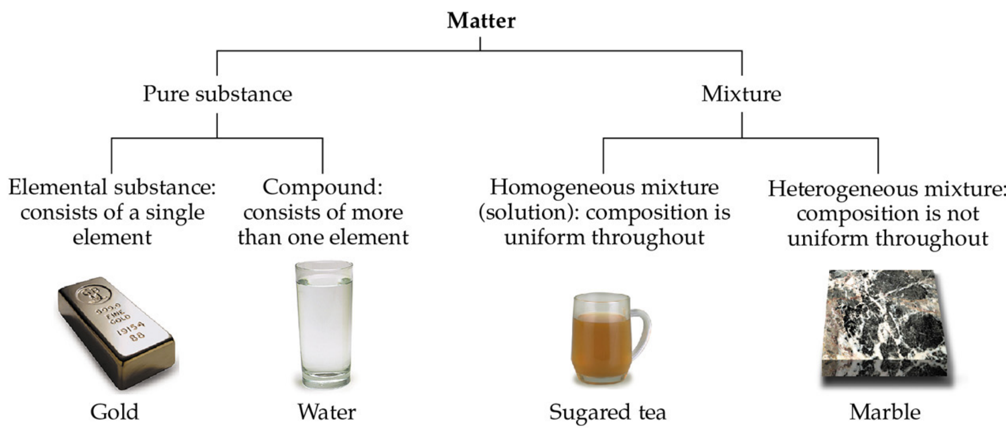


Classification Problems

Classify each of these as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

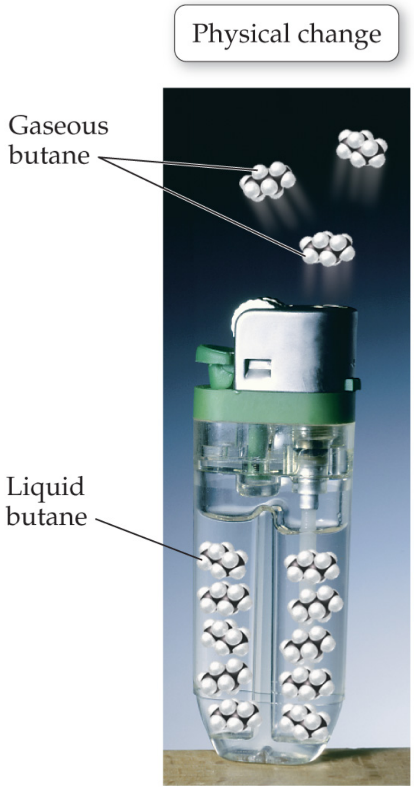
- Tap water
- Steel (an alloy of several metals)
  - Note that alloys can be mixed in different proportions.
- Helium
- Mud
- Carbon dioxide

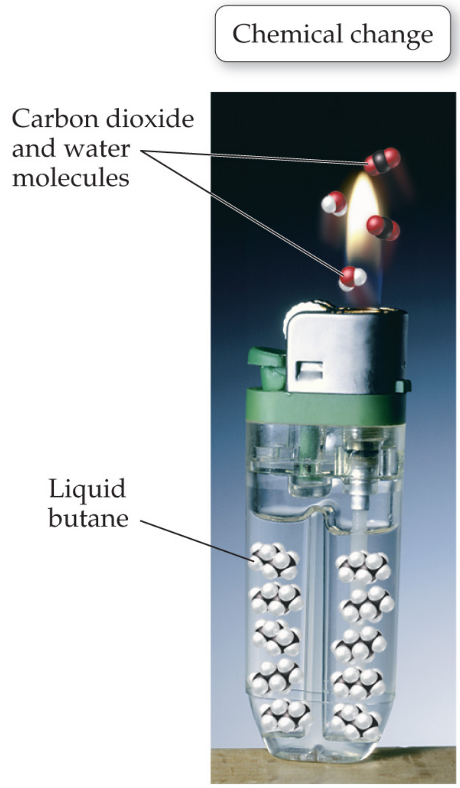
Properties of Matter

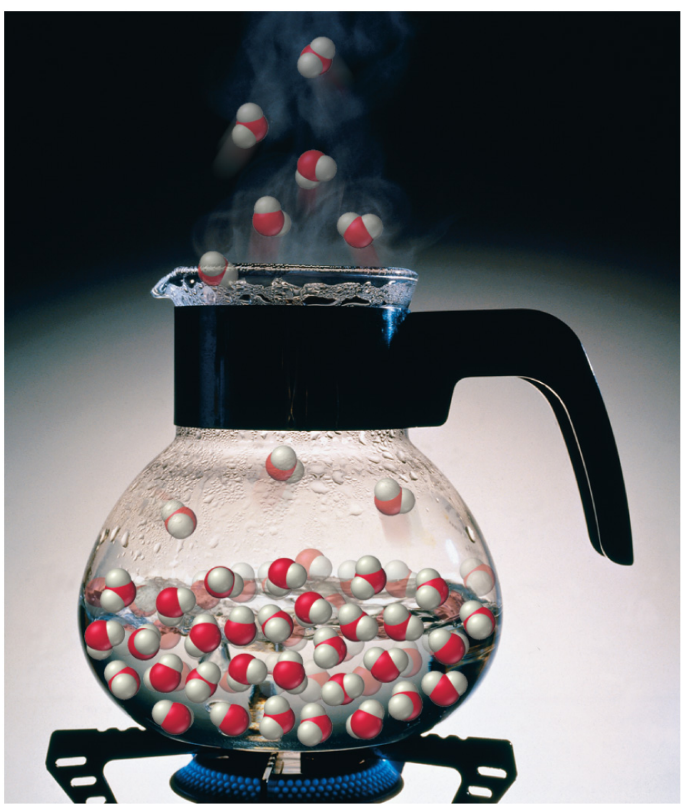
- We can describe matter in two ways:
  - By its chemical properties, which describe how a type of matter interacts (or "reacts") with another type of matter.
    - For example, hydrogen is able to react with oxygen to form water. Helium reacts with virtually nothing.
  - By its physical properties, which include all non-chemical properties.
    - For example, water is a liquid at room temperature, freezes at 0 °C, has a density of 1.0 g/mL, and is both clear (we can see through it) and colorless.

Changes of Matter

- Similarly, changes can be classified in the same way as properties:
  - Chemical changes involve a rearrangement of atoms, producing chemical compounds that were not there before.
    - When iron (Fe) is exposed to oxygen on a wet day, rust (Fe_2O_3) is formed.
  - Physical changes are those which do not involve a chemical change.
    - Boiling water (liquid H_2O changes to gaseous H_2O), glass shattering, a chemical evaporating.


Physical change
Gaseous
butane
Liquid
butane


Chemical change
Carbon dioxide
and water
molecules
Liquid
butane



The Law of Conservation of Mass

- One of the most fundamental statements about matter is described by this law, which states:

  "Matter can neither be created nor destroyed in any chemical process"
- Another consequence of this law is that one type of atom cannot be changed into another through a chemical reaction.
  - If you start a chemical process with 10 million hydrogen atoms and 10 million oxygen atoms, then you will end the process with 10 million hydrogen atoms and 10 million oxygen atoms. ALWAYS!

Energy

- Energy is an important subject in chemistry, as chemical reactions either give off or take in energy.
- The Law of Conservation of Energy states that "Energy can neither be created nor destroyed in a process."
- Energy can be transferred between two systems in two ways
  - As work (w), which we will consider as energy put to some specific use, like making a motor run.
  - As heat (q), which will be random energy, like that given off by your cars engine (it serves no useful purpose).

Energy

- So any change in energy is just the sum of the work and heat changes.
- The SI unit of energy is the Joule (J). Other units include the kJ, and the calorie (cal).
- For example, suppose that burning a certain amount of gasoline produces 50 kJ of energy. If only 20 kJ of it goes into doing work in a car engine, the remaining 30 kJ must be lost as heat.
- Remember, energy cannot be destroyed!

Energy

- An <u>exothermic</u> process is one which gives off more heat than it takes in.
  - For an exothermic reaction, $q < 0$.
  - For example, consider burning gasoline.
- An <u>endothermic</u> process is one which brings in more heat than it gives off.
  - For an endothermic reaction, $q > 0$.

Specific Heat

- Some substances require more energy to raise their temperatures than others.
  - For example, it requires much less energy to raise the temperature of 50. g of aluminum by 10 °C than it would to raise 50. g of water by the same amount.
- This difference is represented by a constant called the specific heat, c.
- Its units are J/(g ·°C) or cal/(g ·°C) .

Transferring Heat

- The amount of heat (*q*) absorbed or given off by a substance when it changes temperature can be found using the following equation:

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

m is the mass of the substance

ΔT is the change in the temperature; it equals

(final *T* – starting *T*)

c is the specific heat of the substance

- Note that heat always flows from an area of high temperature to one of lower temperature!

Example

How much heat is required to raised the temperature of 50.0 g of aluminum from 32 °C to 47 °C? The specific heat of aluminum is 0.903 J/(g ·°C).

A More Difficult Example

A 50.0 g aluminum block is heated to 75 °C, then dropped into a sealed container containing 350. g of water at 15 °C. What will be the temperature of the block when it is finished cooling?

More On Conservation

- Note that in the last problem, heat was transferred from the aluminum to the water, but never destroyed.
- Einstein is credited with the famous equation:

  $$E = mc^2$$

  Where E is energy, m is mass, and c is a constant (the speed of light, *not* the specific heat!)
- This shows that matter *can* be converted to energy.

More on Conservation

- Note however, this will not apply to chemical reactions. We see this with nuclear reactions.
- In *this class*, mass is always conserved.
- We can now state the Law of Conservation of Mass-Energy:

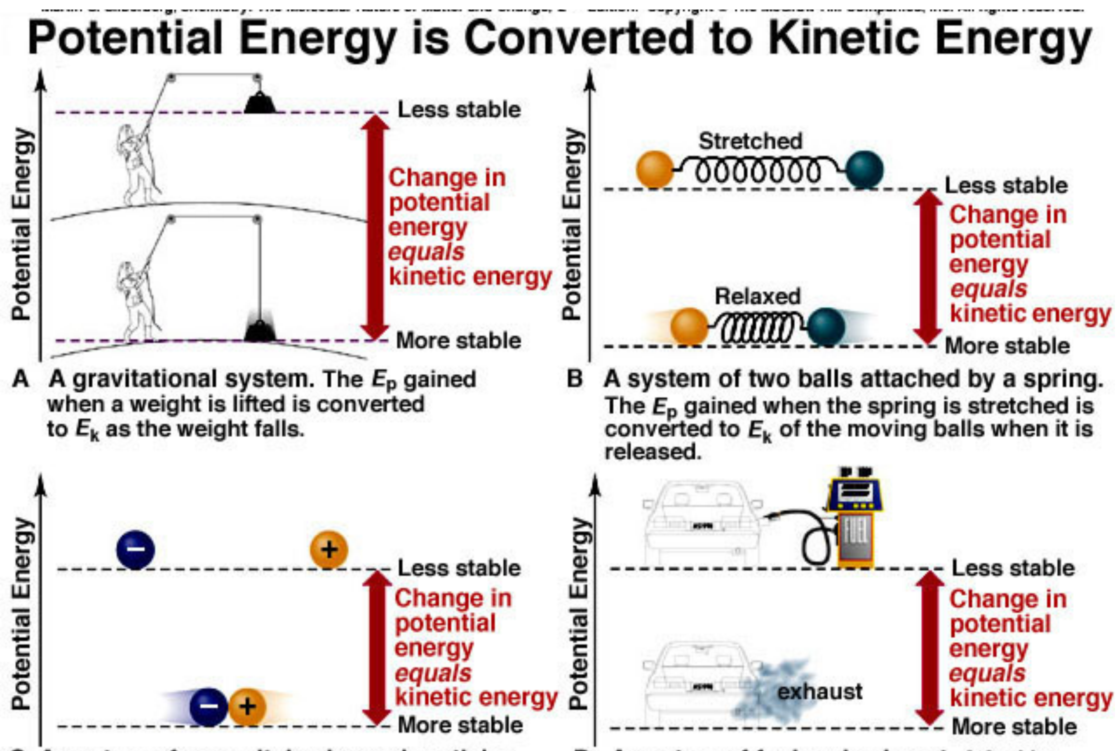
  "Mass and energy can neither be created nor destroyed in a process, though they may be interchanged."

Kinetic and Potential Energy

- *Kinetic Energy (K.E.)* is the energy of motion.
- The amount of *K.E.* an object has is described by the formula

$$K.E. = \frac{1}{2}mv^2$$

    where *m* is the object's mass, and *v* is its velocity(speed).
- *Potential Energy* is the energy of position.
  - Objects may be in a "high energy" position, and can convert this potential energy to kinetic energy while moving to a "lower energy" position.
- Examples:



A A gravitational system. The E_p gained when a weight is lifted is converted to E_k as the weight falls.

B A system of two balls attached by a spring. The E_p gained when the spring is stretched is converted to E_k of the moving balls when it is released.

C A system of oppositely charged particles. The E_p gained when the charges are separated is converted to E_k as the attraction pulls them together.

D A system of fuel and exhaust. A fuel is higher in chemical E_p than the exhaust. As the fuel burns, some of its E_p is converted to E_k of the moving car.