COMPETENCY 1.0 KNOWLEDGE OF CONCEPTS OF MATTER

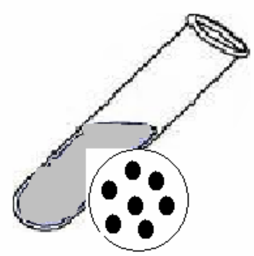
“Nothing in life is to be feared, it is only to be understood. Now is the time to understand more, so that we may fear less.” - Marie Curie

Skill 1.1 Differentiate between pure substances, homogeneous mixtures and heterogeneous mixtures.

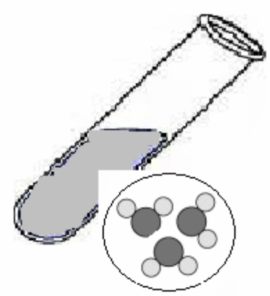
The word "matter" describes everything that has physical existence, i.e. has mass and takes up space. However, the make-up of matter allows it to be separated into categories. The two main classes of matter are **pure substance and mixture.** Each of these classes can also be divided into smaller categories such as element, compound, homogeneous mixture or heterogeneous mixture based on composition.

PURE SUBSTANCES: A pure substance is a form of matter with a definite composition and distinct properties. This type of matter can not be separated by ordinary processes like filtering, centrifuging, boiling or melting.

Pure substances are divided into elements and compounds.

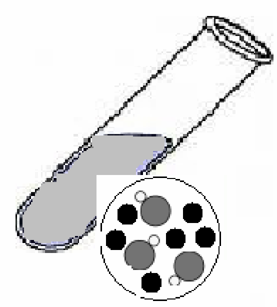


Elements: A single type of matter, called an atom, is present. Elements can not be broken down any farther by ordinary chemical processes. They are the smallest whole part of a substance that still represents that substance. Examples are hydrogen and oxygen.



Compounds: Two or more elements chemically combined are present. A compound may be broken down into its elements by processes that cause chemical change such as electrolysis. Compounds have a uniform composition regardless of the sample size or source of the sample.

Examples are salt and pure water.



MIXTURES: Two or more substances that are not chemically combined. Mixtures may be of any proportions of the components and can be physically, or mechanically, separated by processes like filtering, centrifuging, boiling, or melting. The components do not form new chemical bonds to each other.

Mixtures are either homogeneous or heterogeneous: They can be classified according to particle size.

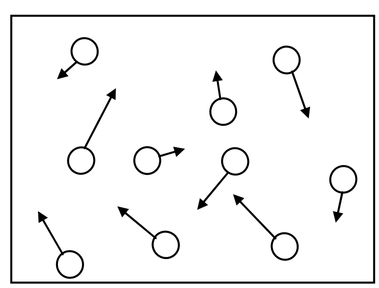
> **Homogeneous Mixtures:** Homogeneous mixtures have the same composition and properties throughout the mixture and are also known as solutions. They have a uniform color and distribution of solute and solvent particles throughout the mixture. A homogeneous mixture exists in a single phase. **Solutions** are homogeneous mixtures. See Skill 1.4. Particle size is less than 1 nm, about the size of an atom or molecule. Examples are salt dissolved in water and air which is a mixture of gases.
>
> **Heterogeneous Mixtures:** Heterogeneous mixtures do not have a uniform distribution of particles throughout the mixture. The different components of the mixture can easily be identified and separated.**Colloidal dispersions** are heterogeneous mixtures. Colloids have particles in the 1 to 1000 nm size, larger than in a solution but smaller than in a suspension. Jelly, milk and blood are examples. **Suspensions** are heterogeneous mixtures in which the particles of at least one component are larger than 1000 nm. The components will settle out if left long enough. These particles show the **Tyndall effect** of scattering light. An example is sand in water.

Skill 1.2 Determine the effects of changes in temperature, volume, pressure or quantity on an ideal gas.

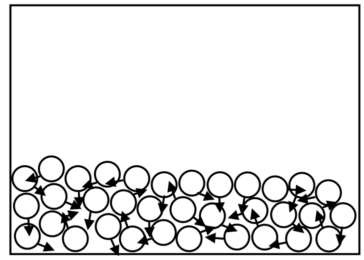
These relationships were found by experimental observation and may be explained by the kinetic molecular theory. See Skill 2.2.

Molecules have **kinetic energy (**they move around), and they also have **intermolecular attractive forces** (they stick to each other). The relationship between these two determines whether a collection of molecules will be a gas, liquid, or solid.

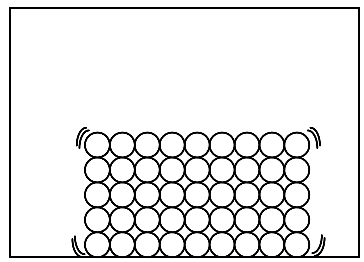
A **gas** has an indefinite shape and an indefinite volume. The kinetic model for a gas is a collection of widely separated molecules, each moving in a random and free fashion, with negligible attractive or repulsive forces between them. Gases will expand to occupy a larger container so there is more space between the molecules. Gases can also be compressed to fit into a small container so the molecules are less separated. **Diffusion** occurs when one material spreads into or through another. Gases diffuse rapidly and move from one place to another.



A **liquid** assumes the shape of the portion of any container that it occupies and has a specific volume. The kinetic model for a liquid is a collection of molecules attracted to each other with sufficient strength to keep them close to each other but with insufficient strength to prevent them from moving around randomly. Liquids have a higher density and are much less compressible than gases because the molecules in a liquid are closer together. Diffusion occurs more slowly in liquids than in gases because the molecules in a liquid stick to each other and are not completely free to move.



A **solid** has a definite volume and definite shape. The kinetic model for a solid is a collection of molecules attracted to each other with sufficient strength to essentially lock them in place. Each molecule may vibrate, but it has an average position relative to its neighbors. If these positions form an ordered pattern, the solid is called **crystalline**. Otherwise, it is called **amorphous**. Solids have a high density and are almost incompressible because the molecules are close together. Diffusion occurs extremely slowly because the molecules almost never alter their position.



In a solid, the energy of intermolecular attractive forces (such as ionic or covalent bonds) is much stronger than the kinetic energy of the molecules. As temperature increases in a solid, the vibrations of individual molecules grow more intense and the molecules spread slightly further apart, decreasing the density of the solid.

In a liquid, the energy of intermolecular attractive forces (such as dipole=dipole and London dispersion forces) is about as strong as the kinetic energy of the molecules. Therefore, both play a role in the properties of liquids. Liquids will be discussed in detail in Skill 1.8.

In a gas, the energy of intermolecular forces is much weaker than the kinetic energy of the molecules. Kinetic molecular theory is most commonly used to understand gases and is best applied by imagining ourselves shrinking down to become a molecule and picturing what happens when we bump into other molecules and into container walls.

Gas **pressure** results from molecular collisions with container walls. The **number of molecules** striking an **area** on the walls and the **average kinetic energy** per molecule are the only factors that contribute to pressure. A higher **temperature** increases speed and kinetic energy. There are more collisions at higher temperatures, but the average distance between molecules does not change, and thus density does not change in a sealed container.

Kinetic molecular theory explains how pressure and temperature influences behavior of gases by making a few assumptions, namely:

1) The energies of intermolecular attractive and repulsive forces may be neglected.
2) The average kinetic energy of the molecules is proportional to absolute temperature.
3) Energy can be transferred between molecules during collisions and the collisions are elastic, so the average kinetic energy of the molecules doesn't change due to collisions.
4) The volume of all molecules in a gas is negligible compared to the total volume of the container.

Strictly speaking, molecules also manifest some kinetic energy by rotating or experiencing other motions. The motion of a molecule from one place to another is called **translation**. Translational kinetic energy is the form that is transferred by collisions, and kinetic molecular theory ignores other forms of kinetic energy because they are not proportional to temperature.

The following table summarizes the application of kinetic molecular theory to an increase in container volume, number of molecules, and temperature:

Effect of an **increase** in one variable with other two constant	Impact on gas: – = decrease, **0** = no change, + = increase						
	Average distance between molecules	Density in a sealed container	Average speed of molecules	Average translational kinetic energy of molecules	Collisions with container walls per second	Collisions per unit area of wall per second	Pressure (*P*)
Volume of container (*V*)	+	–	**0**	**0**	–	–	–
Number of molecules	–	+	**0**	**0**	+	+	+
Temperature (*T*)	**0**	**0**	+	+	+	+	+

Additional details on the kinetic molecular theory may be found at http://hyperphysics.phy-astr.gsu.edu/hbase/kinetic/ktcon.html. An animation of gas particles colliding is located at http://comp.uark.edu/~jgeabana/mol_dyn/.

The kinetic molecular theory describes how an ideal gas behaves when conditions such as temperature, pressure, volume or quantity of gas are varied within a system. An **ideal gas** is an imaginary gas that obeys all of the assumptions of the kinetic molecular theory. While an ideal gas does not exist, most gases will behave like an ideal gas except when at very low temperatures or very high pressures.

Charles's law states that the volume of a fixed amount of gas at constant pressure is directly proportional to absolute temperature, or increasing the temperature causes a gas to expand, in a mathematically proportional manner:

$$V \propto T.$$

Or V=*k*T where k is a constant. This gives a mathematical equation $\frac{V_1}{T_1} = \frac{V_2}{T_2}$.

Changes in temperature or volume can be found using Charles's law.

Example: What is the new volume of gas if 0.50 L of gas at 25ºC is allowed to heat up to 35ºC at constant pressure?

Solution: This is a volume-temperature change so use Charles's law. Temperature must be on the Kelvin scale. K= ºC + 273.

T_1= 298K (273 + 25)
V_1= 0.50 L
T_2= 308K (273 + 35)
V_2 = ?
Use the equation: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ and rearrange for $V_2 = \frac{T_2 V_1}{T_1}$.
Substitute and solve V_2=0.52 L.

Boyle's law states that the volume of a fixed amount of gas at constant temperature is inversely proportional to the gas pressure, or increasing the pressure causes a gas to contract in a mathematically proportional manner:

$$V \propto \frac{1}{P}.$$

Or V=*k*/P where k is a constant. This gives a mathematical equation $P_1V_1=P_2V_2$. Pressure or volume changes (at a constant temperature) can be determined using Boyle's law.

Example: A 1.5 L gas has a pressure of 0.56 atm. What will be the volume of the gas if the pressure doubles to 1.12 atm at constant temperature?

Solution: This is a pressure-volume relationship at constant temperature so use Boyle's law.

P_1= 0.56 atm
V_1= 1.5 L
P_2= 1.12 atm
V_2=?

Use the equation $P_1V_1=P_2V_2$, rearrange to solve for $V_2 = \frac{P_1V_1}{P_2}$.

Substitute and solve. V_2=0.75 L

Gay-Lussac's law states that the pressure of a fixed amount of gas in a fixed volume is proportional to absolute temperature, or increasing the temperature causes the pressure to increase in a mathematically proportional manner:

$$P \propto T.$$

Or P=*k*T where k is a constant. This gives a mathematical equation $\frac{P_1}{T_1} = \frac{P_2}{T_2}$.

Changes in temperature or pressure (with a constant volume) can be found using Gay-Lussac's law.

Example: A 2.25 L container of gas at 25˚C and 1.0 atm pressure is cooled to 15˚C. How does the pressure change if the volume of gas remains constant?

Solution: This is a pressure-volume change so use Gay-Lussac's law.

P_1= 1.0 atm
T_1= 298 K (273 +25)
T_2= 288 K (273 + 15)

Change the temperatures to the Kelvin scale. K=˚C+ 273.

Use the equation $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ to solve. Rearrange the equation to solve for P_2, substitute and solve.

$P_2 = \frac{P_1T_2}{T_1}$ = 0.97 atm

The **combined gas law** uses the above laws to determine a proportionality expression that is used for a constant quantity of gas:

$$V \propto \frac{T}{P}.$$

The combined gas law is often expressed as an equality between identical amounts of an ideal gas at two different states (n_1=n_2):

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}.$$

Example: 1.5 L of a gas at STP is allowed to expand to 2.0 L at a pressure of 2.5 atm. What is the temperature of the expanded gas?

Solution: Since pressure, temperature and volume are changing use the combined gas law to determine the new temperature of the gas.

P_1= 1.0 atm
T_1= 273K
V_1= 1.5 L
V_2= 2.0L
P_2= 2.5 atm
T_2=?

Using this equation, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$**,** rearrange to solve for T_2.

T_2= $\frac{P_2V_2T_1}{P_1V_1}$

Substitute and solve T_2=910 K or, after subtracting 273: 637 ºC

Avogadro's hypothesis states that equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules. **Avogadro's law** states that the volume of a gas at constant temperature and pressure is directly proportional to the quantity of gas, or:

$V \propto n$ where *n* is the number of moles of gas.

Avogadro's law and the combined gas law yield $V \propto \frac{nT}{P}$.

The proportionality constant *R*--the **ideal gas constant**--is used to express this proportionality as the **ideal gas law**:

$$PV = nRT.$$

The ideal gas law ($PV = nRT$) is useful because it contains all the information of Charles's, Avogadro's, Boyle's, and the combined gas laws in a single equality.

If pressure is given in atmospheres and volume is given in liters, a value for *R* of **0.08206 L-atm/(mol-K)** is used. If pressure is given in Pascal (newtons/m^2) and volume in cubic meters, then the SI value for *R* of **8.314 J/(mol-K)** may be used. This is because a joule is defined as a Newton-meter. A value for *R* of **8.314 m^3-Pa/(mol-K)** is identical to the ideal gas constant using joules.

Many problems are given at "**standard temperature and pressure**" or "**STP**." Standard conditions are *exactly* **1 atm** (101.325 kPa) and **0 ºC (273.15 K)**. At STP, one mole of an ideal gas has a volume of:

$$V = \frac{nRT}{P}$$

$$= \frac{(1\text{ mole})\left(0.08206\ \frac{\text{L-atm}}{\text{mol-K}}\right)(273\text{ K})}{1\text{ atm}} = 22.4\text{ L.}$$

The value of 22.4 L is known as the **standard molar volume of any gas at STP**.

Solving gas law problems using these formulas is a straightforward process of algebraic manipulation. **Errors commonly arise from using improper units**, particularly for the ideal gas constant *R*. An absolute temperature scale must be used (never ºC) and is usually reported using the Kelvin scale, but volume and pressure units often vary from problem to problem. Temperature in Kelvin is found from:

$$\text{T (in K)} = \text{T(in }^\circ\text{C)} + 273.15$$

Tutorials for gas laws may be found online at: http://www.chemistrycoach.com/tutorials-6.htm. A flash animation tutorial for problems involving a piston may be found at: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/gasesv6.swf.

Example: What volume will 0.50 mole an ideal gas occupy at 20.0 °C and 1.5 atm?

Solution: Since the problem deals with moles of gas with temperature and pressure, use the ideal gas law to find volume.

$$PV = nRT$$

$$V = nRT/ P$$

$$V= 0.50 \text{ mol } (0.0821 \text{ atm L/mol K}) \ 293 \text{ K}/1.5 \text{ atm}$$

$$V= 8.0 \text{ L}$$

Example: At STP, 0.250 L of an unknown gas has a mass of 0.491 g. Is the gas SO_2, NO_2, C_3H_8, or Ar? Support your answer.

Solution: Identify what is given and what is asked to determine.

Given: T= 273K
P=1.0 atm
V= 0.250 L
Mass= 0.419 g

Determine: Identity of the gas. In order to do this, must find molar mass of the gas. $n = \frac{\text{mass}}{\text{MM}}$. Find the number of moles of gas present using PV=nRT and then determine the MM to compare to choices given in the problem.

Solve for $n=\frac{PV}{RT}$ = 0.011 moles

MM= $\frac{\text{mass}}{n}$ = 38.1 g/mol

Compare to MM of SO_2 (96 g/mol), NO_2 (46 g/mol), C_3H_8 (44 g/mol) and Ar (39.9 g/mol). It is closest to Ar, so the gas is probably Argon.