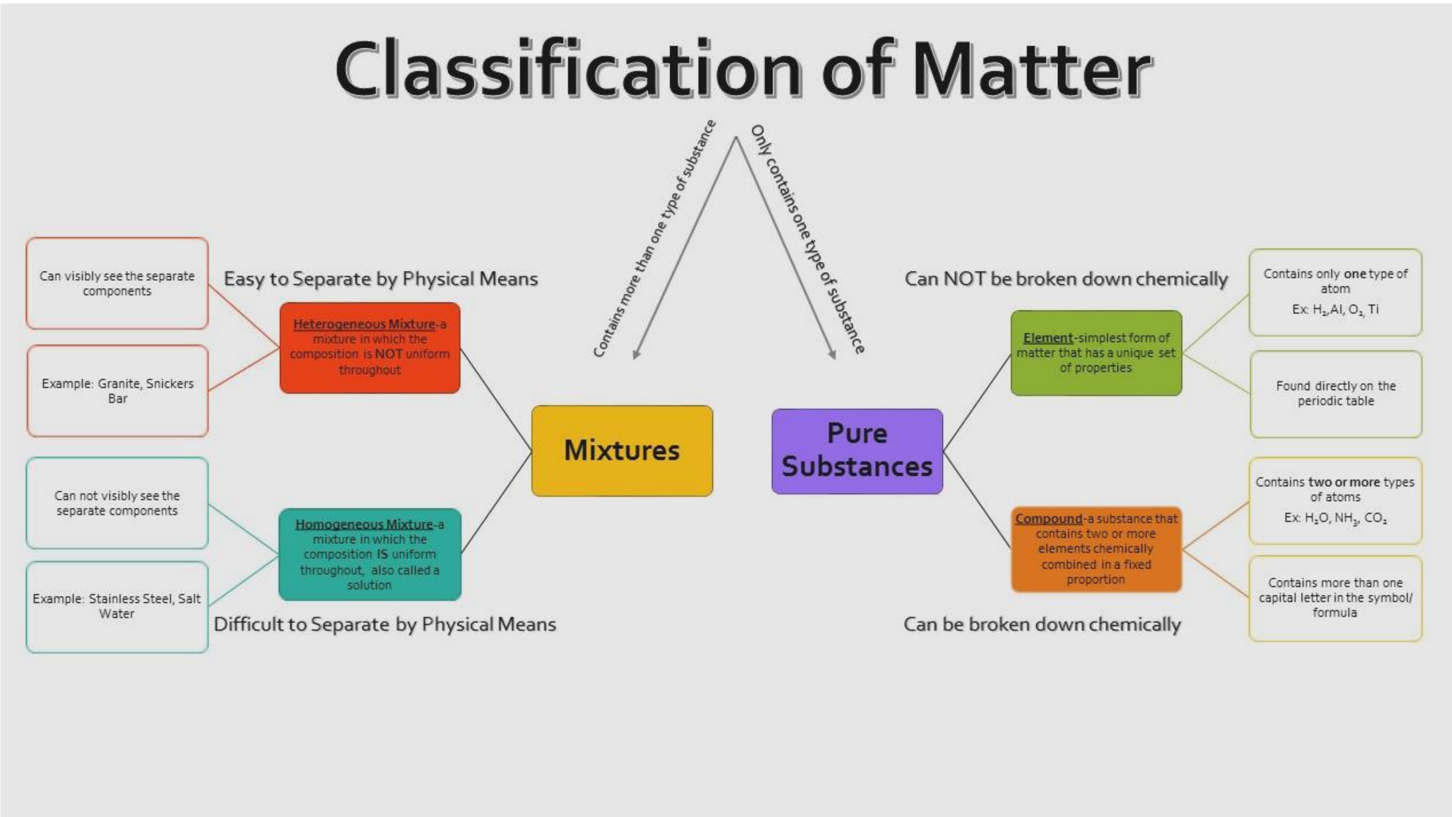
Unit 2 (Matter, Energy, & Change) Test Review Sheet

Matter

Chemistry is the study of matter. What it is, what is it made up of and how does matter react with other matter. Matter is the "stuff" that makes up the whole universe. It's the stuff that takes up space and has mass. Matter is measurable, and matter reacts in predictable ways that we will be learning about all year.

Classifying Matter



The simplest forms of matter are the **elements** which are listed in the periodic table of elements. Elements are the unique types of matter that cannot be broken down into simpler substances by any chemical or physical process. Examples include mercury, iron, carbon, and uranium. The smallest part of an element is called an atom.

Each element on the periodic table has unique properties that can be measured by us in class, or by scientists. Each element reacts in ways that are known, and which can be relied upon. We will learn later in the year how the atoms of certain elements can chemically combine, or bond, with other atoms to form new substances called **compounds**. These new substances have their own measurable properties which are different from the properties of the atoms that make them up. Compounds can only be broken back down into elements by chemical means, usually requiring energy as well.

Pure substances are types of matter made up of only one type of particle. Examples include elements which are each made up of only one kind of atom. Another example are the compounds. Compounds are made up of a specific ratio of at least two different kinds of atoms. Examples include pure water (H_2O), or glucose ($C_6H_{12}O_6$).

Pure substances are always the same throughout, which is called homogeneous. The properties of the elements are always the same in every sample, and within any one sample. All samples of pure water have the same density (for example), and all of the properties water has are the same for all samples of pure water.

Matter can be chemically combined into compounds, or just mixed together, into **mixtures**. A mixture is a physical blend of pure substances. Two or more elements can be mixed, two or more compounds can be mixed, or elements and compounds together can be mixed. Mixtures have no definite ratio of the component parts like compounds have. Because of this, mixtures are not always homogeneous, they can also be heterogeneous, or different throughout.

Mixtures are just stirred up, and the pure substances that make them up keep their properties. No new substances are formed, rather there is just a rearrangement of the atoms or particles. Compounds are new pure substances, with new properties. Since these mixtures are just physical blends, they can be separated easily, by physical means (no chemical reactions required). The processes used to separate these mixtures work because they take advantage of differences in certain physical properties of the parts of the mixture. Here are several examples…

Physical and Chemical Changes

Know the difference between physical and chemical changes.

A physical change does not result in a new substance being created which means the composition of the substance DOES NOT CHANGE. Examples of physical changes are phase changes, dissolving a solid substance in a liquid, changing the shape or size of a substance, etc.

A chemical change creates a new substance as the composition changes of the original substances. A match burning is a chemical change as the substances of the wood combine with oxygen to create new substances such as the gas CO_2.

Energy

Energy exists in two forms: 1) kinetic energy (motion) and 2) potential energy(stored).

Thermal energy (**heat**) is the energy due to the motion of particles. Heat is measured by the transfer of energy and flows from a high temperature to a lower temperature.

Chemical potential energy is a more difficult concept, but it relates to the stored energy that is in a substance based on its composition and the forces that that keep the particles of a substance together.

Temperature is a measurement of average kinetic energy,
Example: solid silver at 60° Celsius has greater average kinetic energy than a gas at 45° Celsius

Two scales of Temperature are used in chemistry: Celsius and Kelvin. The relationship between Celsius and Kelvin is found in Table T in the Reference Table: $K = C + 273$

Exothermic: A change that releases heat to the surroundings is referred to as "exothermic". For example, when a match is on fire, the change is releasing heat to the surroundings.

Endothermic: A change that absorbs energy from the surroundings is referred to as "endothermic". For example, when you heat water to boil it, the process of boiling is "endothermic" since the change requires that water absorbs heat from the surroundings.

Phases

All Matter is made up of tiny particles, and it is important to remember the following characteristics when describing a phase or phase change:

I. The particles of matter have space in between them
II. The particles of matter are ***continuously moving***
III. The particles of matter **attract each other**
IV. Diffusion (movement of particles from a high concentration to a low concentration) occurs in gases and liquids
V. The particles of a gas move according to the concept known as Brownian motion. (Constant, random, straight line movement)

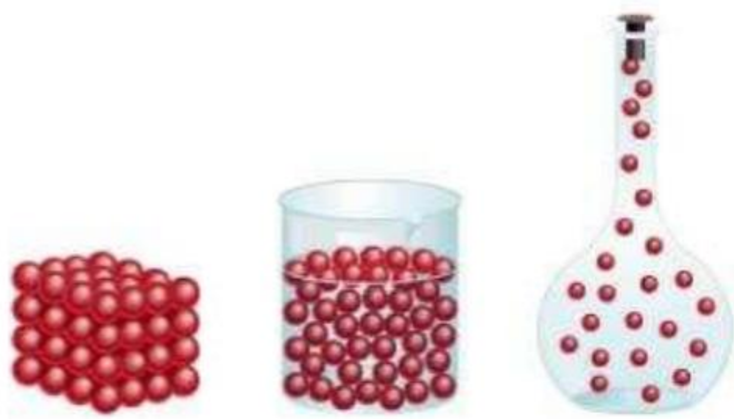
PURE Matter comes in 3 states, or phases: **Solids, liquids and gases.**

Solid matter has it's atoms (or molecules) packed in very close together. These particles are in a rigid arrangement that does not change. Because of this the solids are hard, and hold their shape and volume. Solids cannot be compressed practically at all because of the closeness of the particles.

Liquids also have their atoms (or molecules) packed close together, but they are not locked in place, rather they constantly move, or flow over each other. The closeness of the particles means that liquids have a definite volume. Liquids do not hold their own shape, they take the shape of any container they are put into. Liquids cannot be compressed practically at all because of the closeness of the particles.

Gases are very different than either solids or liquids. Most different is the proximity of the atoms (or molecules). Gas particles are very far apart from each other, always moving very rapidly, bouncing around off of each other and the walls of any container that holds them. Gases have no definite shape and will expand to fill any size container they're put into. Gases have no definite volume either, they can be greatly compressed into small containers, there is plenty of room between the particles.

Here is a diagram of a solid, liquid, and a gas. Each little red ball represents a particle (atom or molecule).

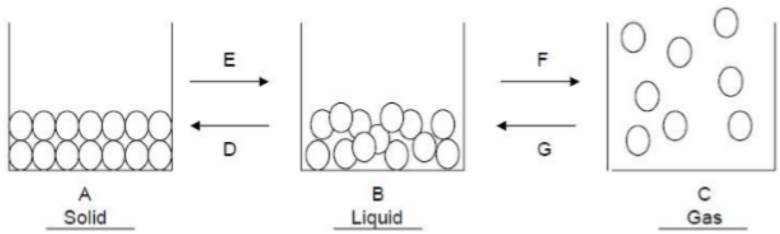


The atoms of a solid are very close together, and have a definite shape and volume. The liquid has close packed particles moving around, but have the shape of the container they are in. The gas has a lot of space between the particles, and fills any shape or sized container it's put into.

Phase Changes

a. Phase changes are examples of physical changes.

b. Almost all substances can be made to change between the 3 phases, simply by altering the temperature.

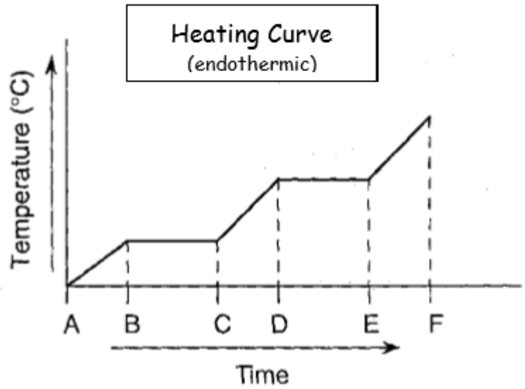


i. Solid to liquid (Arrow E) = melting
ii. Liquid to solid (Arrow D) = freezing
iii. Liquid to gas (Arrow F) = evaporation (open container); vaporization (closed container)
iv. Gas to liquid (Arrow G) = condensation
v. Solid to gas = sublimation
vi. Gas to solid = deposition

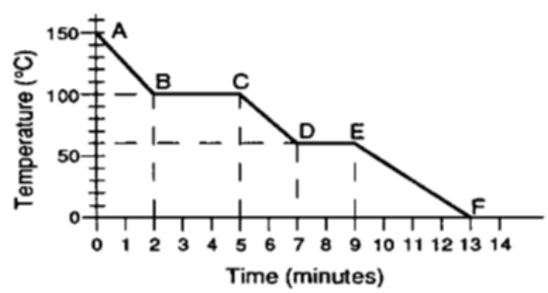
Phase Change Diagrams: Heating & Cooling Curves

a. **Heating Curve**: Temperature vs. Time is graphed while a substance is being HEATED at a constant rate.
b. **Cooling Curve**: Temperature vs. Time is graphed while a substance is being COOLED at a constant rate.
c. The freezing point and the melting point of a substance are the same.

Heating Curve
(endothermic)



Cooling Curve
(exothermic)



A→B = Solid
B→C = Melting (Fusion)
C→D = Liquid
D→E = Boiling (Vaporization)
E→F = Gas

A→B = Gas
B→C = Condensation
C→D = Liquid
D→E = Freezing
E→F = Solid

Note: There are kinetic energy changes when there **is** a change in temperature (sloped portions of the graph, i.e., A→B). There are potential energy changes when there is a change in phase with no change in temperature (flat portions of the graph, i.e., B→C).

Measurement of Heat Energy (Potential Energy)

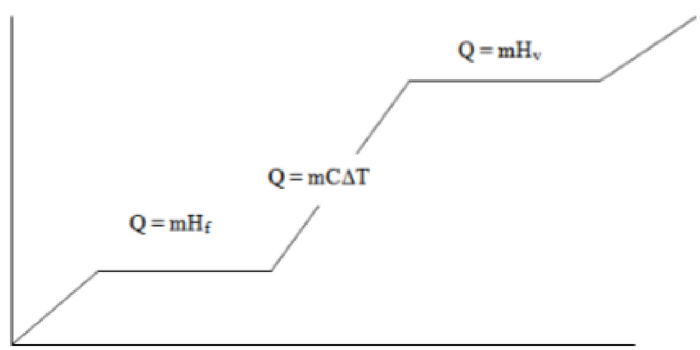
a. **Heat**: Energy transferred due to a difference in temperatures.
b. **Specific Heat (c)**: the amount of heat needed to raise the temperature of 1 gram of water by 1 C.
c. The amount of heat lost or gained can be calculated using the following equation:

$$\boxed{q = mc\Delta T}$$

q = heat (units = Joules or J)
m = mass of sample
c = heat capacity of sample (see Table B for H_2O)
ΔT = change in temperature (T_f-T_i)

This formula **cannot** be used to calculate heat change during phase changes because $\Delta T = 0$.

d. The formulas for calculating heat in joules during changes of phase are listed on Table T.
e. **Heat of Fusion**: the energy absorbed (endothermic) to melt a sample of a solid to a liquid or the energy released (exothermic) when a sample of liquid freezes to a solid. (Table B for water)
   i. $Q = mH_F$
f. **Heat of Vaporization**: the energy absorbed (endothermic) to vaporize a sample of a liquid to a gas or the energy released (exothermic) when a sample of a gas condenses to a liquid. (Table B for water)
   i. $Q = mH_V$



Behavior & Properties of Gases

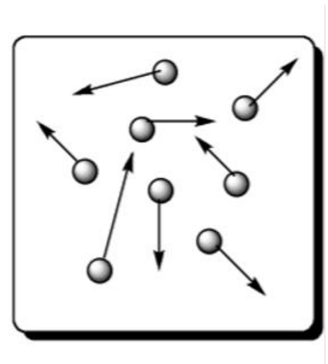
Pressure

All gases exert pressure.

At any point, a gas exerts an equal **pressure** in all directions at any point within a gas.

Pressure is defined as the *force per unit area*.

Please note the diagram showing a sample of gas molecules enclosed in a container. The arrows indicate the velocities of the molecules.



Kinetic Molecular Theory of Gases

It's the Kinetic Molecular Theory that allows us to think about, discuss, and understand gases.

How do gases behave? In order to study gases, chemists have devised a model. The model is called an **ideal gas** (a gas which explains the behavior of all gases). This Ideal Gas model is based on the following assumptions, and can be applied only under conditions of **LOW PRESSURE** AND **HIGH TEMPERATURE**.

POSTULATES OF THE KINETIC MOLECULAR THEORY

1. Gases consist of tiny particles (atoms or molecules).
2. These particles are so small, compared with the distances between them that the volume (size) of the individual particles can be assumed to be negligible (zero).
3. The particles are in constant random motion, colliding with the walls of the container. These collisions with the walls cause the pressure exerted by the gas.
4. The particles are assumed to not attract nor repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature of the gas.

Gas Laws

Boyle's Law

The relationship between **PRESSURE and VOLUME**

If we were to take the pressure and the volume of ANY GAS at ANY CONDITION, and multiply them together, we would get a particular gas constant for that particular gas sample. This relationship is:

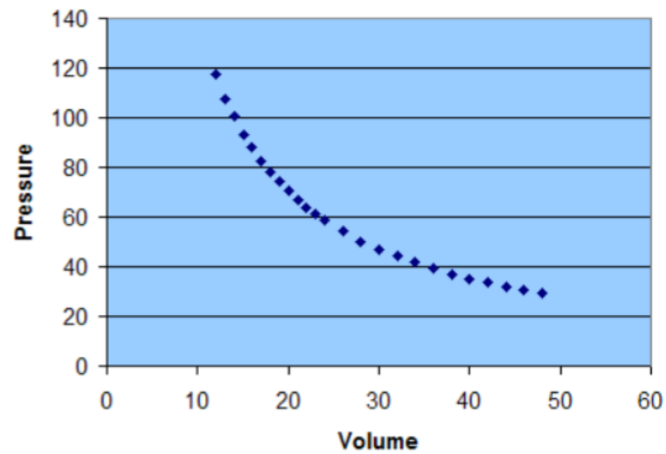
PV = Gas Constant

Any gas constant is the CONSTANT JUST FOR THAT SAMPLE.

But if you change the pressure, you can calculate the new volume since PV = that constant.

If you were to change to volume, you could calculate the new pressure as well.

At any point, the PRESSURE + the VOLUME multiplied together will give the SAME ANSWER.



Since this is true for every point where pressure and volume exist, not only is PV = a constant, any PV point will equal the same constant, so we can also say:

$$P_1V_1 = P_2V_2$$

Pressure and volume are inversely proportional. That means as one goes up, the other must go down.

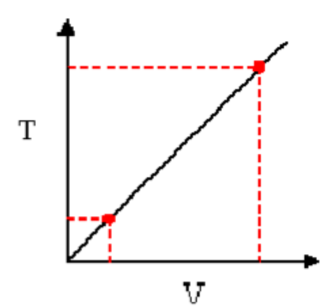
Charles' Law

The relationship between **VOLUME and TEMPERATURE**

When it comes to volume and temperature, this relationship is directly proportional. As one goes up, so does the other, and the reverse too, as volume decreases, so does temperature.

Directly Proportional ALWAYS is a STRAIGHT LINE GRAPH
The formula showing this relationship is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$



Gay-Lussac's Law

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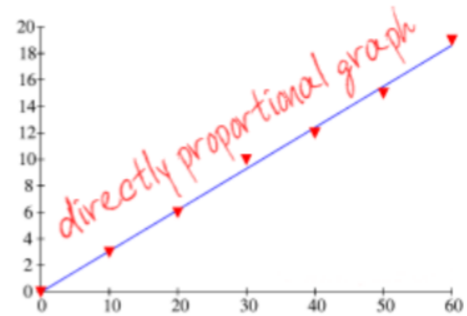


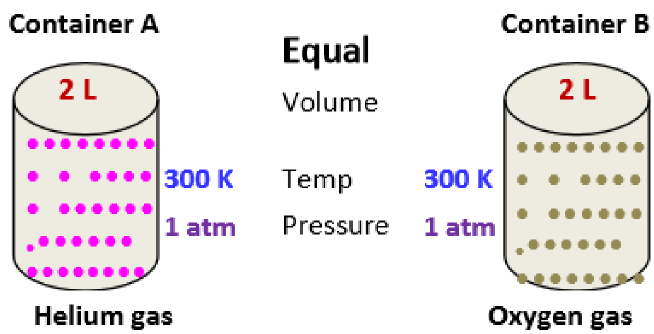
As temperature rises, so does pressure, because the increase in temperature is an increase in Kinetic Energy of the gas particles. Since they are now moving faster, they have more collisions PER SECOND, and the collisions they have are stronger because they are moving faster. This results in a greater pressure of gas if the volume is held constant.

Avogadro's Law

Avogadro's Law states: Under the same conditions of temperature and pressure, gases of equal volume contain equal number of molecules (particles).

In the example below, container A contains helium gas and container B contains oxygen gas.

NOTE that both containers have the same volume , and are at the same temperature and pressure.



If helium gas molecules are counted in Container A and oxygen gas molecules are counted in Container B, ***the number of molecules of He in A will be the same as the number of molecules of O_2 in B.***

The Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Because of the three different relationships between pressure, volume and temperature of gases, we can combine the three into one formula. The left side, the "1" side, is the initial conditions of a gas. The "2" side is the changed conditions of a gas. In a gas problem you will always be given the initial conditions, and 2 of 3 conditions of the final changed ones. Deciphering the questions is less tricky if you always start with the formula, then fill in the parts you do know, and then solving for the missing part.

STP means standard temperature and pressure. Those are in Table A on the reference table.

Pressure can be done in kPa or atm. As long as you use the SAME UNITS on both sides of the equal sign for pressure, it does NOT MATTER which unit you do use.

For Temperature, standard is 273 Kelvin. You MUST USE KELVIN, for this reason: If the temperature ever is zero, then it can only mean absolute zero. SO, always use Kelvin, even if the problem uses Celsius (to try to throw you off!). The conversion from Celsius to Kelvin is discussed above.

Volumes can also be in ANY UNIT, as long as the same units are on both sides of the equal sign! Liters, mL, deciliters, cubic centimeters, etc. are all fine units of volume. They don't matter, as long as you use the same one in the whole problem.