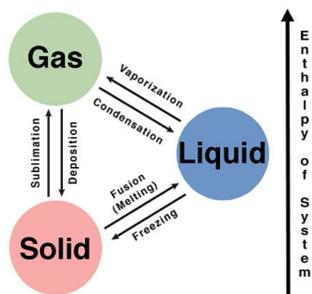


Chapter 10: Phases of Matter Diary



Phase describes a physical state of matter. The key word to notice is physical. Things only move from one phase to another by physical means. If energy is added (like increasing the temperature or increasing pressure) or if energy is taken away (like freezing something or decreasing pressure) you have created a physical change.

Solids

Particles that are strongly attracted to each other are called solids. Other than some small vibration, there is nearly no movement of the atoms or molecules. They have a rigid or lattice arrangement of the particles and they keep their shapes and volumes. Solids do not take the shape of their containers. Solids cannot be compressed very much because the particles are very close together. Because of this most solid matter has a high density compared to their liquids or gases. When energy or heat is added, the particles will vibrate more, which often makes solids expand when heated. Particles in solids have the lowest kinetic energy. Give solids enough energy (at the proper pressure) and they will vibrate so much that they break apart and turn into

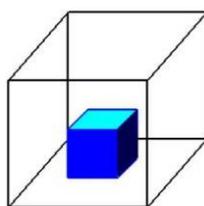
Liquids

Particles that have some attraction to each other but not enough to make them stuck to each other are called liquids. Liquids flow over themselves. The particles are in constant random motion. Liquids do not have a definite shape which means they take the shape of the container you put them in. If you spill liquids, the force of gravity spreads them out quite well. The hotter liquids get when you add energy, the faster the particles move, and liquids too expand slightly when heated. Liquids are dense as well, but usually not as dense as solids are. Heat a liquid enough, the particles move so much that they turn into...



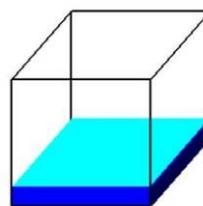
Phases of Matter

Glenn
Research
Center



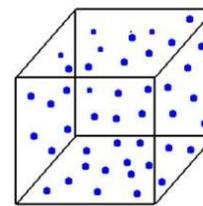
Solid

Holds Shape
Fixed Volume



Liquid

Shape of Container
Free Surface
Fixed Volume



Gas

Shape of Container
Volume of Container

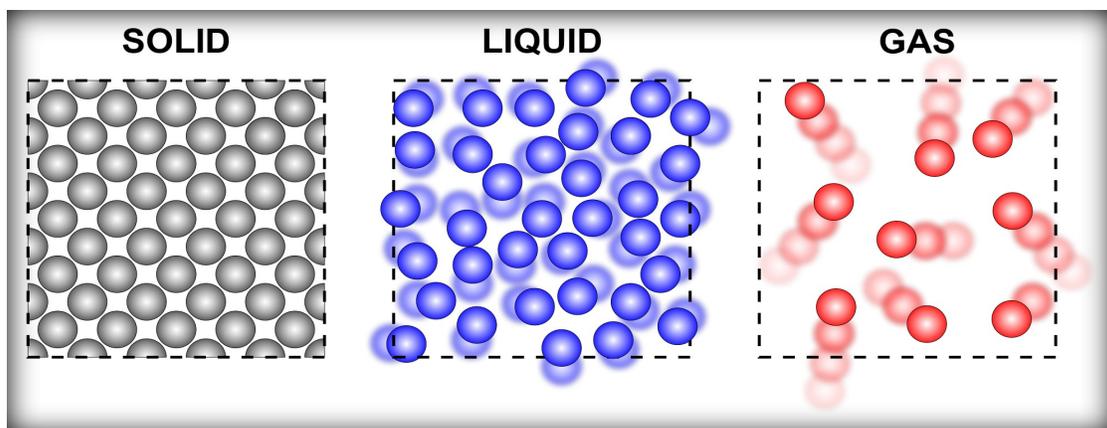
Gases

There is virtually no attractive or repulsive force between particles of a gas. The particles move in straight lines and very fast. They collide with other particles all of the time. These collisions will cause gas (or air) pressure. Gases take the shape of the container that you put them in. Any amount of a gas will fill any container that you put it in. The collisions are considered to be elastic, meaning there is no loss of kinetic energy due to the collisions. Heating gases make the particles move faster and have more collisions, causing expansion if possible, or greater pressures if contained in a definite volume. Gas Particles have the highest kinetic energy. Gases have very low density.

Changing Phases

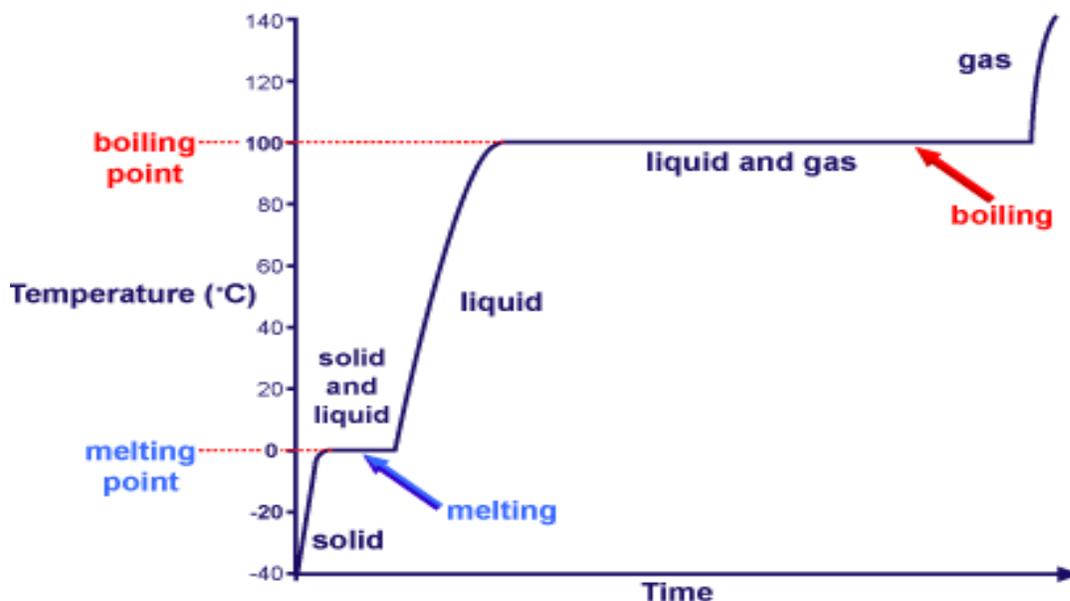
Phases change when energy is added or removed from a substance. Phase change is a physical change, not a chemical change.

Melting	Solids have many intermolecular bonds. The added energy makes the rigid network of particles shake so much that the vibrations reach the point that they exceed the attractive forces keeping them solid. The solid then breaks up and turns into a liquid.
Evaporation	Liquids have an average kinetic energy level that keeps them from flying apart into a gas. Although the average kinetic energy of all the water molecules is such that it is a liquid, some molecules will get enough energy and they will be able to jump to the gas phase. This is evaporation, which happens only at the surface of a liquid. Warmer water will still be liquid, but more molecules will have the energy to leap into the air as a gas. Once heated enough, all of the molecules can escape the remaining forces of attraction and change into a gas.
Boiling	When liquid water gains enough kinetic energy to have ALL the particles have enough energy to jump to the gas phase, this is boiling. The boiling point is when the amount of kinetic energy is sufficient to overcome the attractive forces between the molecules and the air pressure being exerted against the liquid. Boiling occurs throughout the liquid not just at the surface. The bubbles in boiling water are gaseous H ₂ O expanding away from each other. In this gas phase they are of course LESS DENSE than the liquid water, so the bubbles flow upwards and break into the air, releasing water molecules as gas.



Freezing	The opposite of melting, when a liquid loses enough energy that the attractive forces between molecules is greater than the kinetic energy in the molecules, they are able to lock into a lattice or network, and a solid results.
Condensing	The opposite of boiling, when a gas collapses into a liquid. The gas phase has by far the highest kinetic energy, and when a gas condenses a lot of energy is released. Steam condensing on your hand (for example) is a much more energetic process than spilling hot water on your hand. Steam burns are much worse than water burns.
Sublimation	A process where a solid jumps directly to the gas phase. Not all substances do this. Iodine solid can be heated directly to an iodine gas. Solid carbon dioxide (dry ice) will go directly to the gas phase as well. Water can sublime too, imagine putting wet clothes on a clothesline in the cold winter. The wet water freezes almost immediately. Later on, even though the temperatures never rise above the melting point of water, the clothes are dry.
Deposition	Is the opposite of sublimation. Gases become solids directly without becoming liquids. Iodine will do this well.

Heating and cooling curves



Remember: liquid water exists only at temperatures between 0°C - 100°C*. Below 0°C water is a solid called ice. Above 100°C water is a gas. Water in the liquid form cannot exist at colder or warmer temperatures (except under unusual pressure conditions).

Note that this curve is made up of five line segments. What happens at each line segment is special, and you should be able to look at a heating curve for any substance and grasp what phases are present. We will assume that NORMAL PRESSURE conditions exist unless we are told otherwise.

Kinetic energy (KE) and Temperature are directly related. As the temperature rises, so does kinetic energy. As the temperature decreases, kinetic energy decreases. If the temperature is CONSTANT, kinetic energy remains constant.

So, look at the curve and find the line segments listed below, and think about what is written.

A to B: solid ice warms as heat is added. Only solid water exists as the temperature is still below 0°C. Kinetic energy rises as does the temperature.

B to C: the cold phase change happens, ice melts to water. Both phases exist at 0°C. The increase of energy does not raise the temperature, because during this part of the graph, all the energy gain is used up breaking up the bonds between water molecules, making the solid a liquid. No temperature increase = no increase in KE.

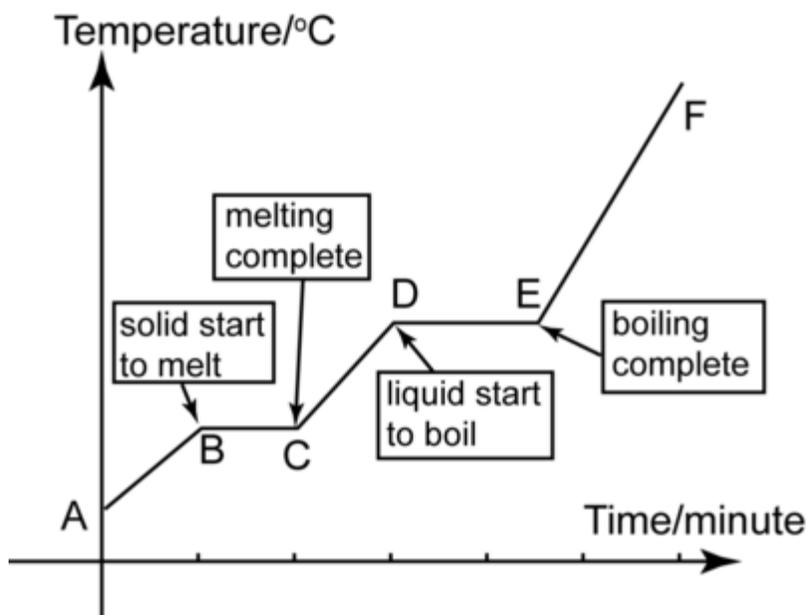
C to D: the liquid only phase* until 100°C. As the temperature increases so does the kinetic energy.

D to E: The hot phase change occurs here. Note that the temperature becomes constant again as all the energy gain is used up to break all of the bonds that keep the molecules stuck together. This makes the water become gas. Kinetic energy is constant, so is temperature. Both liquid phase & gas phase exist at exactly 100°C + 1 atm.

As any substance cools from a liquid to a solid (or even from a gas to a liquid to a solid), a graph can be put together that shows the different temperatures and phases at any point.

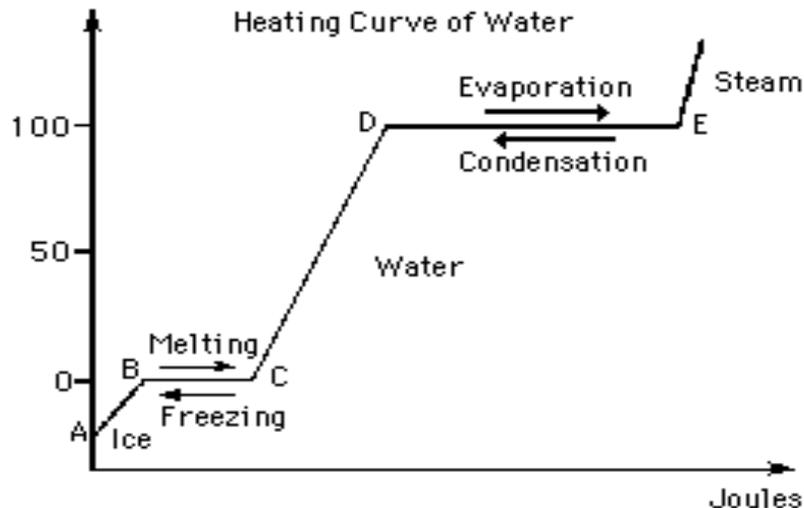
For water, the heating curve shows solid water at a temperature below zero centigrade rising up to zero (the melting point) which is where the COLD phase change happens from solid to liquid. Then the water heats up until 100 centigrade, the temperature of the HOT phase change. Here the liquid becomes gas. Over 100 centigrade liquid water cannot exist at normal pressure. But, steam can be heated much hotter than 100 degrees C.

When the temperature is rising so is the kinetic energy. When the temperature is steady, so is the kinetic energy. During the phase change the kinetic energy stops rising and something else is going on. The POTENTIAL ENERGY is rising. Where the temperature is rising with kinetic energy, in those places the potential energy is steady. Energy is added at a constant rate, so when the water's temperature is not rising, all the energy is being used to break up the bonds (solid to liquid, or, liquid to gas). This is where potential energy is increasing. Remember, Breaking of bonds is an endothermic process. Heat energy must be added to break the intermolecular forces of attraction.



With a cooling curve, the substance starts out hot enough to be a gas even. As it cools we can graph the temperatures over time. The exact opposite that happens during the heating curve happens here. As temperatures drop, so of course does the kinetic energy.

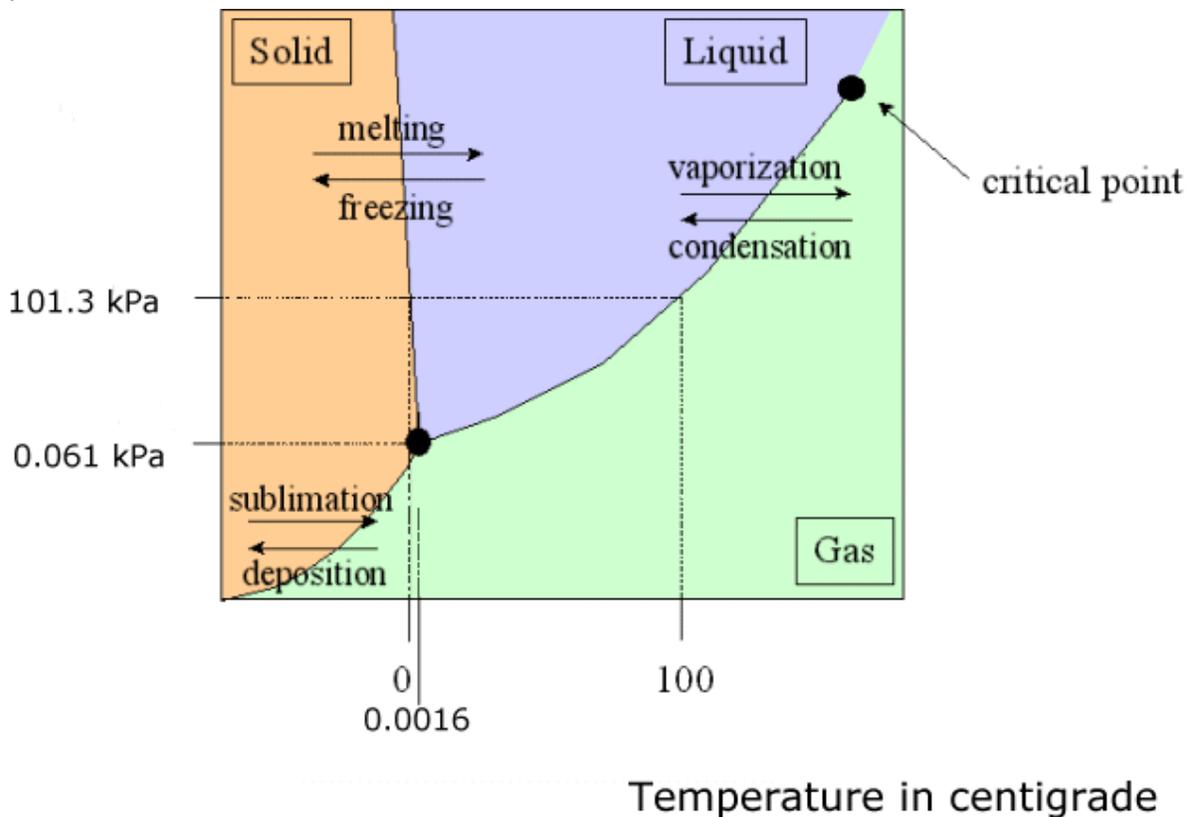
During phase changes, the temperature gets steady, and so does the kinetic energy. When you undergo a cooling phase change, energy is being released. This is an exothermic process. This should result in lower temperatures, but only if kinetic energy is being released. During a phase change



Phase Diagrams

If we look at a single substance (water for example), at any temperature and pressure it will exist as a solid or liquid or gas. A phase diagram puts into picture or graph form this idea. Phase diagrams for water need to be part of your chemistry understanding. You should be able to draw a phase diagram for water. All substances can have a phase diagram, and you should be able to interpret them if given to you as well.

The pressure is always on the vertical axis, and temperature is always horizontal. The lines represent the phase changes from solid to liquid or gas, from liquid to solid or gas, or from gas to liquid or solid.



Points of interest to us on these graphs are the Normal boiling point, Normal freezing point, Triple Point, and Critical point.

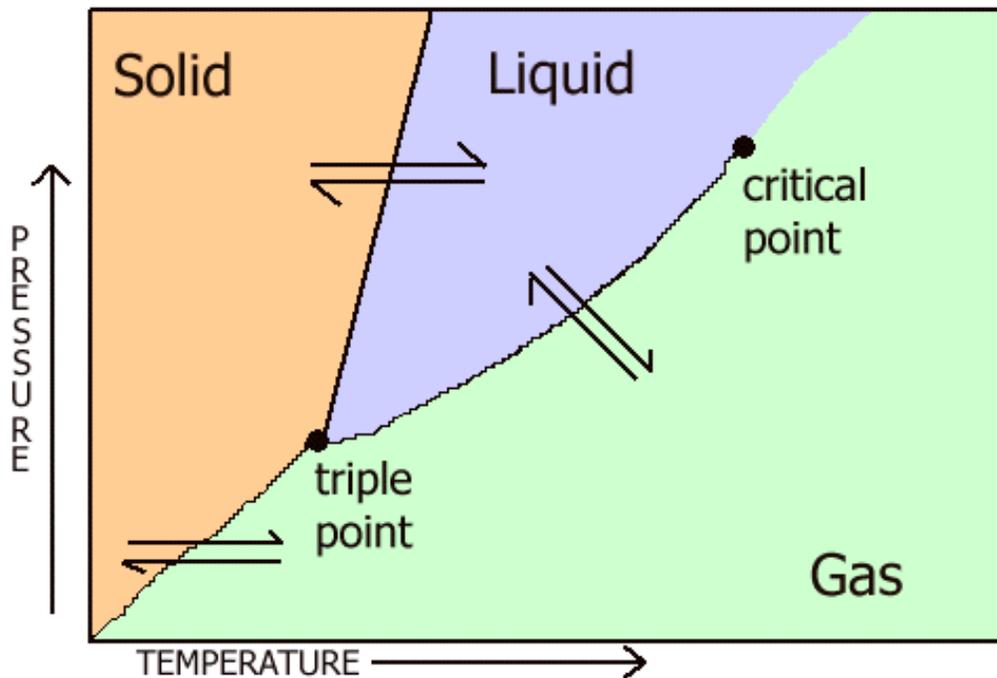
Normal boiling point and normal freezing point are the temperatures of freezing or boiling at NORMAL PRESSURE.

The Triple point is the one pressure/temperature combination that allows all three phases to exist at once.

The Critical point is the "end" of the line between liquid and gas.

It is the temperature and pressure at which the liquid and gaseous phases of a pure stable

PHASE DIAGRAM



The Double Arrows

- Between solid and gas is sublimation & deposition.
- Between gas and liquid is condensation & boiling.
- Between liquid and solid is freezing & melting.

The Kinetic Theory and Phases

The Kinetic Molecular Theory is used to explain how solids, liquids and gases exist differently from one another. Solids are made up of the exact same particles as liquids and gases. As stated earlier, phase changes are ONLY PHYSICAL, not chemical. So, solids have very slow moving particles, with minimal vibration (Brownian movement). As energy is added, the particles shake themselves loose from the intermolecular forces that bind them as solids. As liquids the particles have much more kinetic energy, they move lots more. The energy results in many more collisions between the particles. As more energy is added, the particles gain enough kinetic energy to jump apart from each other and become gases.

This same theory will help explain the RATES of chemical reactions later in the year.

Gas Pressure Conversion Problems

Table A shows us standard temperature and standard pressure. Air or gas pressure is measured with several different units for various reasons. We need to be able to go from one unit to another. We'll use several basic conversion factors (similar to mole island conversions, but these are way easier).

Standard pressure means normal pressure. The pressure of the air at sea level on Earth, measured all around the world, over and over, and finally it was standardized. Standard pressure is equal to 1.0 atm, or 1.0 atmospheres of pressure. It's an old fashioned unit, but it works. Standard pressure was also measured with the mercury barometers of years gone by. At standard pressure, the column of mercury was pushed up the tube to 760. millimeters in height. So, standard pressure also equals 760. mm Hg. Metric standard pressure is measured in kilo-pascals, a finely named unit if there ever was one! 101.3 kPa is standard (so much for metric!).

Table A
Standard Temperature and Pressure

Name	Value	Unit
Standard Pressure	101.3 kPa 1 atm	kilopascal atmosphere
Standard Temperature	273 K 0°C	kelvin degree Celsius

$$\text{Kelvin} = \text{degree Celsius} + 273$$

$$\text{degree Celsius} = \text{Kelvin} - 273$$

$$1 \text{ atm} = 101,300 \text{ Pa} = 101.3 \text{ kPa} = 760 \text{ mmHg} = 760 \text{ torr}$$

$\frac{1.0 \text{ atm}}{760. \text{ mm Hg}}$	$\frac{1.0 \text{ atm}}{101.3 \text{ kPa}}$	$\frac{760. \text{ mm Hg}}{101.3 \text{ kPa}}$	$\frac{760. \text{ mm Hg}}{1 \text{ atm}}$	$\frac{101.3 \text{ kPa}}{1.0 \text{ atm}}$	$\frac{101.3 \text{ kPa}}{760. \text{ mm Hg}}$
----------------------------------------------	---------------------------------------------	------------------------------------------------	--------------------------------------------	---------------------------------------------	------------------------------------------------

Practice problems for pressure conversion (3 SF each)

Convert 744 mm-Hg to atm and to kPa.

$$\frac{744 \text{ mm Hg}}{1} \times \frac{1.0 \text{ atm}}{760.0 \text{ mm Hg}} = 0.979 \text{ atm}$$

$$\frac{744 \text{ mm Hg}}{1} \times \frac{101.3 \text{ kPa}}{760.0 \text{ mm Hg}} = 99.2 \text{ kPa}$$

Convert 155 kPa to mm-Hg and atm

$$\frac{155 \text{ kPa}}{1} \times \frac{760.0 \text{ mm Hg}}{101.3 \text{ kPa}} = 1160 \text{ mm Hg}$$

$$\frac{155 \text{ kPa}}{1} \times \frac{1.0 \text{ atm}}{101.3 \text{ kPa}} = 1.53 \text{ atm}$$

Convert 0.724 atm to kPa and mm-Hg

$$\frac{0.724 \text{ atm}}{1} \times \frac{101.3 \text{ kPa}}{1.0 \text{ atm}} = 73.3 \text{ kPa}$$

$$\frac{0.724 \text{ atm}}{1} \times \frac{760.0 \text{ mm Hg}}{1.0 \text{ atm}} = 550. \text{ mm Hg}$$

Table H and Vapor Pressures

Table H in our reference tables shows four liquids and their vapor pressures at different pressures and temperatures. Look at only one line at a time to decide if a particular point (a pressure and a temperature combo) for THAT compound is a liquid or a gas. If the point is to the LEFT of the line, then at that point the compound is a liquid.

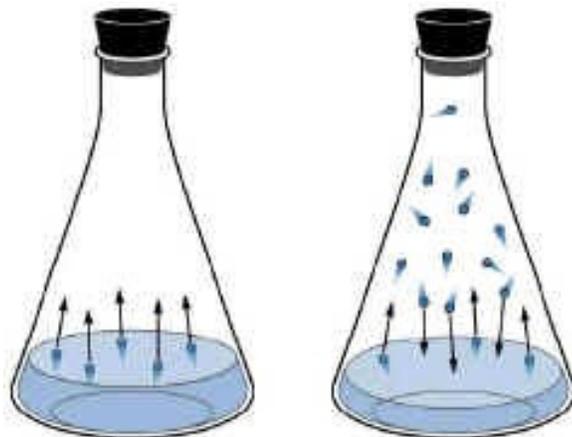
If the particular point of temperature and pressure is to the right of the curved line, then at that combination the compound is a gas.

The curved lines themselves represent ALL of the boiling points for the compounds AT ANY PARTICULAR pressure and temperature combination.

In our previous, pre-chemistry minds, water boiled at 100 C, and that was that. Now we realize that water boils at 100°C only if the pressure is exactly NORMAL or STANDARD, or 101.3 kPa, or 760. mm Hg, or 1.0 atm. pressure. If the pressure were to change, so would the boiling point.

Water can never get hotter than its boiling point. If more heat is applied, the water boils faster but it cannot get hotter. Water really boils at 97 C in Boulder, Colorado. In fact, it can't get to 100°C in Boulder. (except in a lab with a high pressure chamber).

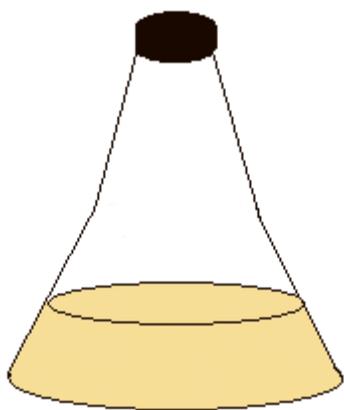
If we look at one particular temperature, say 20°C, we see that on Table H the compound with the HIGHEST VAPOR PRESSURE is propanone. That means that propanone will evaporate faster than the others because of its own properties. Vapor pressure is determined by temperature and what liquid is present. Propanone, ethanol, water and ethanoic acid are different from each other. They have different vapor pressures because of specific individual properties.



Vapor Pressure Explained (what the heck is Table H really for?)

Vapor pressure is the pressure caused by a liquid evaporating. The molecules of the gas from this evaporation add to the pressure of the air already in the bottle. Evaporation of liquids occurs only at the surface, and happens at all temperatures. The more particles of gas in container, the more collisions. The more collisions the more pressure. This pressure could become high enough to explode a sealed bottle.

Vapor pressure is controlled by two factors, the temperature which indicates how much kinetic energy the molecules of liquid have (more KE, more collisions, more evaporation), and the type of liquid you have in the bottle. Some liquids "stick together" very well because the molecules have more attraction to each other than other types of liquids. The more the attraction, the lower the vapor pressure (the harder it is to evaporate). Together these factors determine how much vapor pressure is inside a sealed system.

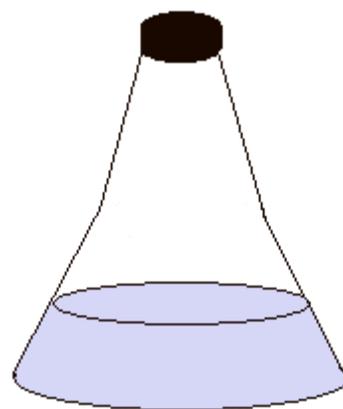


Air pressure 101.3 kPa + 3 kPa
from the evaporation of water
= about 104 kPa total pressure

STP

Two identical bottles,
corked in the same room
at standard temperature and
pressure. At left is water,
at right is propanone.

Both will evaporate & produce
vapor pressure,
added to the 101.3 kPa
provided by the air
already in the bottles.



Air pressure 101.3 kPa + 10 kPa
from the evaporation of propanone
= about 111 kPa total pressure



Air pressure 101.3 kPa + 8 kPa
from the evaporation of water
= about 109 kPa total pressure

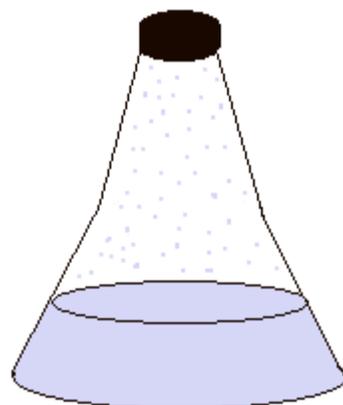
Temperature increases
to 25°C

Vapor pressure increases too.

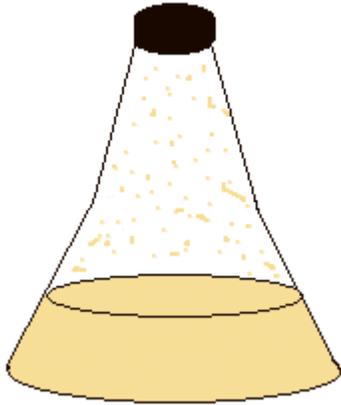
At this temp:

Water VP = 8 kPa

Propanone VP = 30 kPa



Air pressure 101.3 kPa + 30 kPa
from the evaporation of propanone
= about 131 kPa total pressure



Air pressure 101.3 kPa + 12 kPa
from the evaporation of water
= about 113 kPa total pressure

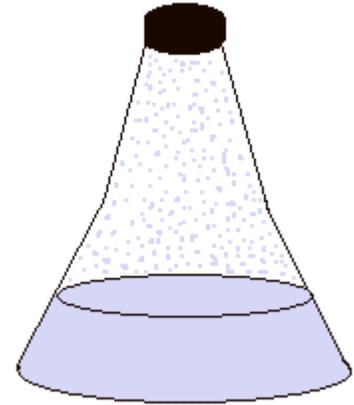
Temperature increases
to 50°C

Vapor pressure increases too.

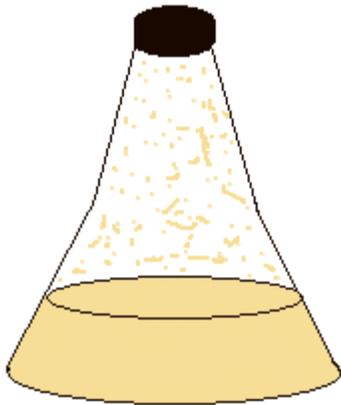
At this temp:

Water VP = 12 kPa

Propanone VP = 82 kPa



Air pressure 101.3 kPa + 82 kPa
from the evaporation of propanone
= about 183 kPa total pressure



Air pressure 101.3 kPa + 38 kPa
from the evaporation of water
= about 139 kPa total pressure

Temperature increases
to 75°C

Vapor pressure increases too.

At this temp:

Water VP = 38 kPa

Propanone VP = 82 kPa



Air pressure 101.3 kPa + 180 kPa
bottle explodes
due to excess pressure

Although we have not learned about gas pressure, the heating will also increase the air pressures inside the bottles. The air pressures will increase the same in each bottle, but the vapor pressures will increase at very different rates due to the fact that one contains water which is "very sticky to itself" and propanone which is much less attracted to itself.

The less attracted to itself propanone evaporates much easier, making much higher vapor pressures inside the bottles at similar temperatures than the water does.

Having a lower vapor pressure means that a liquid would evaporate faster than another at the same temperature. It would take LESS KINETIC energy to have particles jump from liquid to gas. Since that is true, it would also take LESS KINETIC energy to make ALL of them go from liquid to gas, hence higher vapor pressure liquids have lower boiling points. It takes less energy (less temperature) for the liquids to all become gas.

All of this is inside table H.

Normal Boiling point means the boiling point at exactly 101.3 kPa (normal pressure). Normal freezing point (which is NOT on table H) is the freezing point at NORMAL PRESSURE as well. Normal means normal (or standard).

Boiling points, anywhere on any of the 4 curved lines are a combination of kinetic energy measured as temperature and pressure measured in kPa.

Table H
Vapor Pressure of Four Liquids

