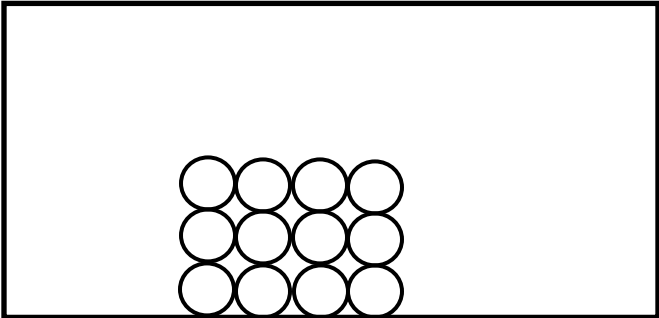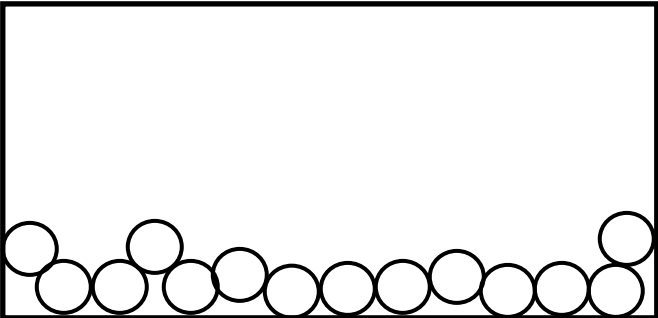# Phases BASICS: SOLIDS, LIQUIDS, GASES, and aqueous too
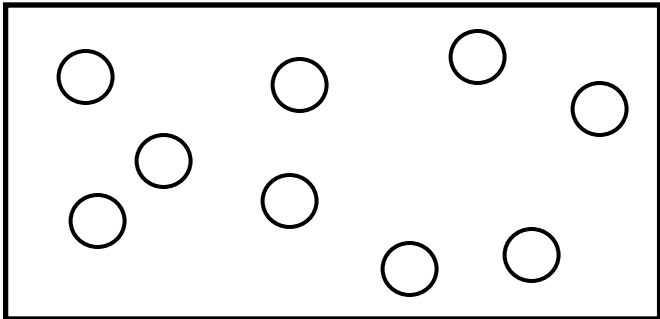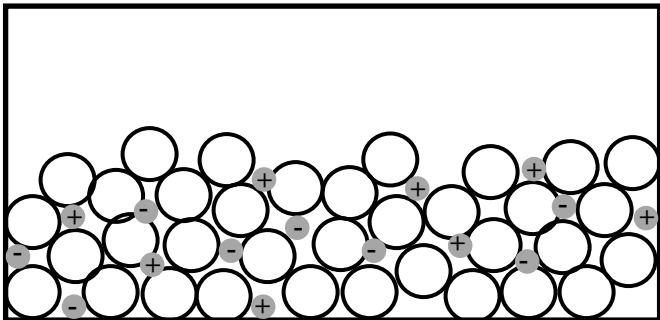
These diagrams must be in your head along with most of these facts. Work at it.
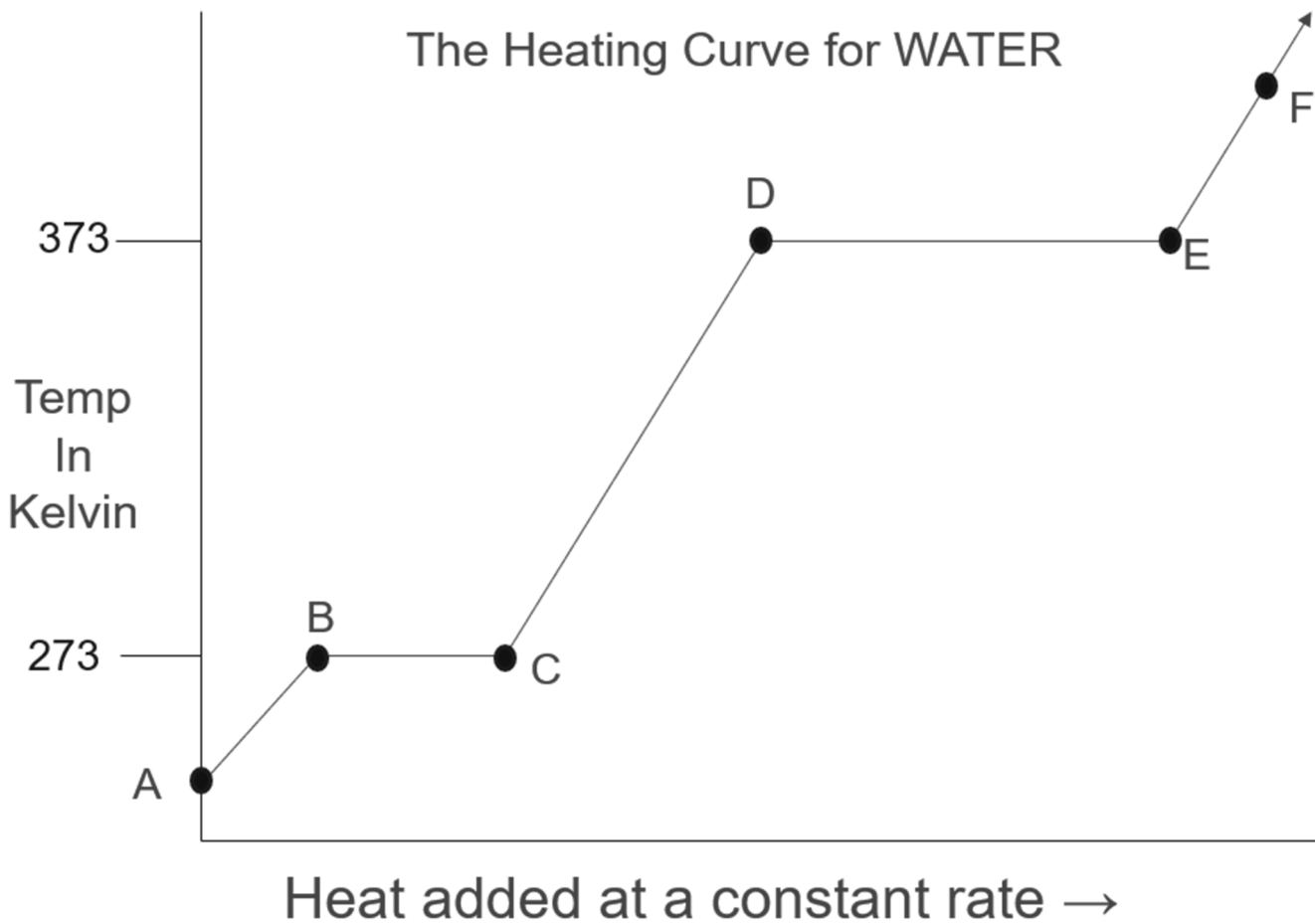
solid	Particles are strongly attracted to each other, other than some vibration there is nearly no movement of the particles. They have an organized, rigid or lattice arrangement. They hold their shape and hold their volume. Solids do not take the shape of the containers you put them into. Solids cannot be compressed very much because the particles are already packed very close together. Because of this, most solids have the highest density of all 3 phases. When energy or heat is added, the particles will vibrate more, which often makes solids only slightly expand but remain solid. Particles in the solid phase have the lowest kinetic energy. Add sufficient energy and they will vibrate so much that they break apart and turn into liquids. This shows a particle diagram of a solid in a rectangle container. NOTE: the solid particles are little circles, they are arranged in an orderly, repetitive way. They hold their shape in the container, they do not conform to the bottom of the container, as a liquid would. This particle diagram shows the particles as sticky AND STUCK TOGETHER.
liquid	Liquid particles have some attraction to each other but not enough to make them completely stick together. Liquid particles flow over themselves. The particles are in a constant random motion. Liquids do not have a definite shape which means they take the shape of the bottom of any container you put them in. If you spill liquids, the force of gravity spreads them out. The hotter liquids get when you add energy, the faster the particles move, Liquids expand only slightly when heated. Liquids are rather dense, but usually not as dense as solids. Heat a liquid enough, the particles move so fast that they explode apart turning into gases. This shows a particle diagram of a liquid in a rectangle container. NOTE: the particles are little circles, they are arranged in a disorderly pattern, but they do touch each other, and they do take the shape of the bottom of the container. (like water in the bottom of a beaker or cup). This shows these particles as "sticky, but not stuck" together. Their disordered arrangement indicates they are not a solid.

gas	There is virtually no attractive or repulsive force between the gas particles. The gas 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 spread out and take the shape of any container that you put them in. Any amount of a gas will fill any container that you put it in (the gas would change pressure when expanding or squishing). The collisions between gas particles are considered to be elastic, meaning there is no loss of kinetic energy due to the collisions—the energy is transferred from particle to particle. Heated gases make the particles move faster & have more collisions, causing the gas to expand. If the gas is in a hard container and expansion is not possible, then the pressure increases due to more and stronger collisions. Gas Particles have the highest kinetic energy. Gases have (by far) the lowest density of the three phases. This shows a particle diagram of a gas in a rectangle container. NOTE: the gas particles are little circles, they are appear to be "flying" around the container, which is what they do actually do. They do not as a rule stick together (that's liquid or worse, solid!). They might once in a while bounce off of the bottom of the container, but we do not draw them on the bottom, we show the gas particles "up" in the container.
aqueous	Although aqueous is not a "real phase" many substances (solid, liquid or gas) can dissolve into water. It no longer is the original phase if it dissolves, it is called aqueous. Aqueous means dissolved in water. Solid salt that dissolves into water is not solid anymore. Carbon dioxide gas that dissolves into water to form seltzer or soda is not a gas anymore. Anything that is dissolved into water is called aqueous. The trickiest phase change is solid → aqueous because it seems like something special is happening even though not much really occurs. Here: water molecules are the circles and the ions are grey dots — positive and negative, spread out in the water. They might be Na^{+1} and Cl^{-1} if it's table salt that is aqueous in this water. Any ionic compound that is AQ will ionize or dissociate this way: $NaCl_{(S)} \rightarrow Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$ Or, this way: $Co(NO_3)_{3(S)} \rightarrow Co^{+3}_{(AQ)} + 3NO_3^{-1}_{(AQ)}$

6 Phase Changes	Phases change when energy is added or removed from a substance. A phase change is a physical change, not a chemical change. Ice ⇆ water ⇆ steam, but it stays H_2O, with the same chemical properties.
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.
Freezing	The opposite of melting, when a liquid loses enough energy that the attractive forces between particles is greater than the kinetic energy of the particles, they are able to lock into a lattice or network, and a solid results.
Vaporization Can be Evaporation	Liquids have enough attraction to hold themselves together, but not enough to be stuck solid. These liquid particles are all moving and have what's called average kinetic energy. Not all the particles have exactly the same kinetic energy, and when the more energetic ones get to the surface of a liquid, some particles can make the leap to the gas phase. This is evaporation, and it happens only at the surface of a liquid. Warmer water evaporates faster than cold. The water can be liquid, but more molecules will have the kinetic energy to leap into the air as a gas.
or it can be 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 attractiveness of the molecules and the air pressure on the liquid. Boiling occurs throughout the liquid not just at the surface. The bubbles in boiling water are gas H_2O 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.
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	In this process a solid jumps directly to the gas phase. Not all substances normally do this. Iodine solid can be heated directly to an iodine gas. Solid CO_2 (dry ice) will go directly to the gas phase as well. Water can sublimate— 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 become dry because $H_2O_{(S)} \rightarrow H_2O_{(G)}$ and goes into the air.
Deposition	This is the opposite of sublimation. Gases become solids without becoming liquids first. Iodine will do this well.

Heating curves

The HEATING CURVE for water shows how water is heated from below
freezing through two phases changes until it is steam.
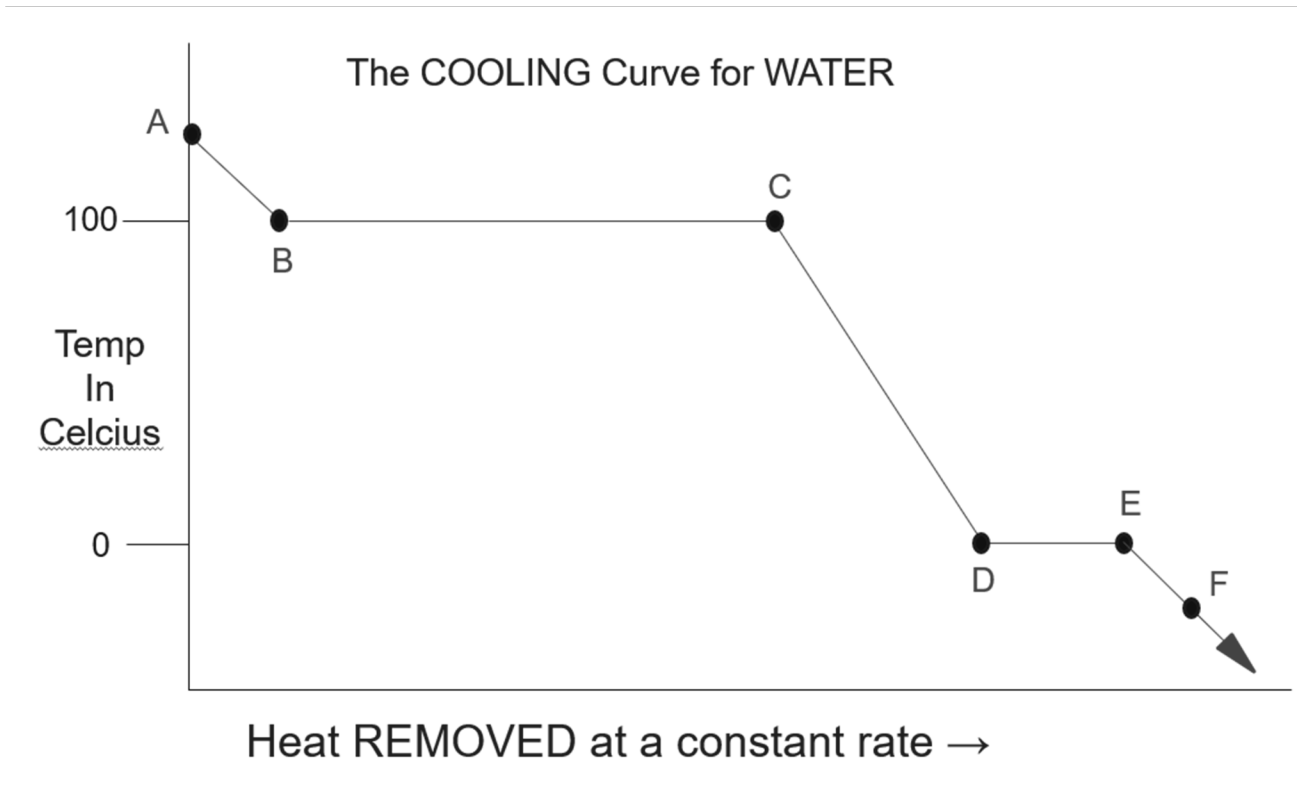All heating curves look like this one (more or less)

The Heating Curve for WATER

Temp In Kelvin (y-axis): 373, 273 marked
Heat added at a constant rate → (x-axis)
Points labeled A, B, C, D, E, F

Segment	Phase or phases present in segment	What is the temperature doing?	What happens to the KE? Kinetic Energy is the Energy of MOTION	What happens to PE? Potential Energy is the Energy of PHASE
AB	Solid Ice only	Temp Increases	Kinetic Energy Increases	Potential Energy Steady
BC	Solid Ice melts to Liquid Water	Temp Steady	Kinetic Energy Steady	Potential Energy Increases
CD	Liquid Water only	Temp Increases	Kinetic Energy Increases	Potential Energy Steady
DE	Liquid Water vaporizes to Steam gas	Temp Steady	Kinetic Energy Steady	Potential Energy Increases
EF	Steam Gas only	Temp Increases	Kinetic Energy Increases	Potential Energy Steady

Things to remember: 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.

The COOLING Curve for WATER

Temp
In
Celcius

100

0

Heat REMOVED at a constant rate →

A → B: steam gas cools as heat is removed. Only gas water exists as the temperature is still above 100°C.
  Kinetic energy decreases along with temperature. Potential energy is steady as phase does NOT change.

B → C: the HOT phase change occurs, steam gas condenses into liquid water. Both phases exist at 100°C. The
  removal of energy does not lower the temperature, because during the hot phase change all energy removed
  allows the gas to condense. Since the phase "moves lower" down from liquid to solid,
  the POTENTIAL ENERGY decreases. Steady temp = steady KE.

C → D: the liquid only phase, temperature drops and so does the kinetic energy. Since the water stays liquid,
  there is no change in phase, and not change in potential energy either.

D → E: the cold phase change occurs here. Note that the temperature stays constant as all the energy removed allows
  the liquid water to bond into solid ice. The liquid water becomes solid water (ice). No temperature change
  means no change in Kinetic Energy, but since the phase "moves lower" down from liquid to solid,
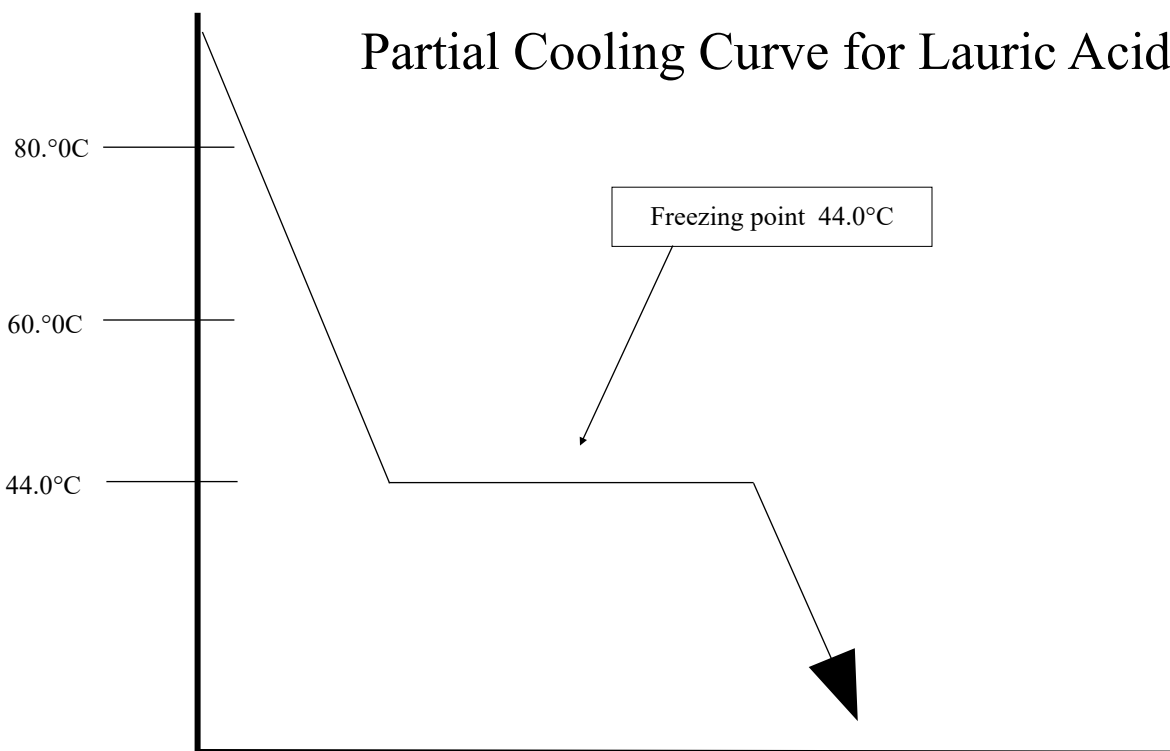 the POTENTIAL ENERGY decreases.

E → F: the solid only phase. Although liquid water can NEVER get colder than 0°C (at normal pressure) solids can
  get as cold as you can make them (but never so cold as to be at zero Kelvin). As the temperature decreases
  so does the Kinetic Energy. Since the phase is contast (solid only) the POTENTIAL ENERGY is constant.

With a cooling curve, a substance starts hot, usually hot enough to be gas. As it cools we can graph the temperature change over time. Cooling curves run opposite the heating curve. As temperatures drop, so does the kinetic energy. Kinetic energy and temperature always change together, or remain steady together (during phase changes).

On a graph, the phase changes are always parallel with the horizontal axis, because they only occur at a one temperature. Water freezes **at** the freezing point of 273 K or 0°C. That is one exact temperature. A parallel line represents ONE temperature.

Below is a PARTIAL cooling curve for lauric acid (solid at room temperature but melts in a hot water bath.

Lauric acid freezes at 44.0°C. Above that temperature it is liquid, below that it is a solid.

Partial Cooling Curve for Lauric Acid

Freezing point 44.0°C
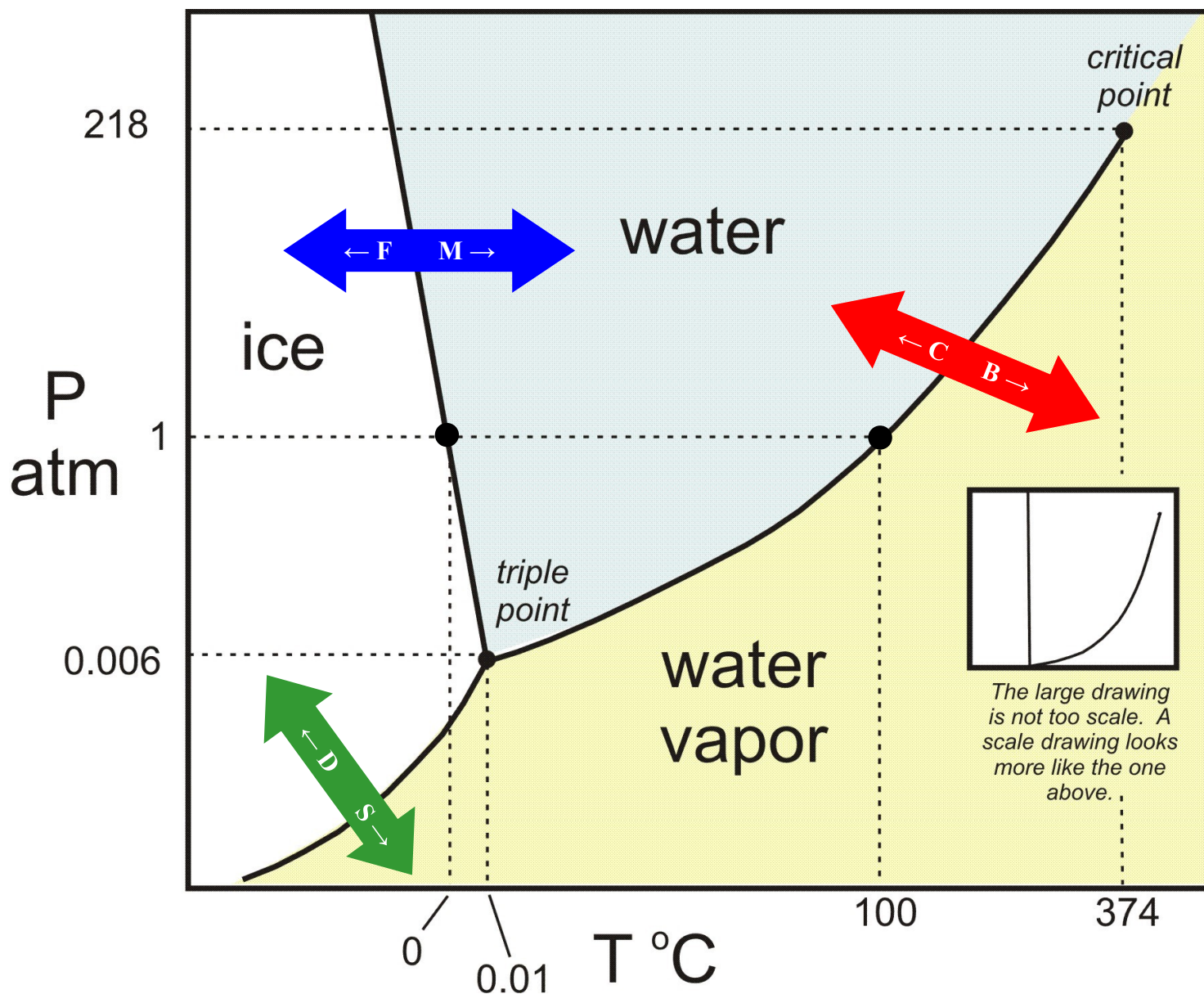
80.°0C

60.°0C

44.0°C

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 that data into graph form. You should memorize the phase diagram for water. You should be able to draw a phase diagram for water without notes. All substances have a phase diagram, and you should be able to interpret them if given to you.

ALL points crossing from Ice to Water are the melting/freezing point at various pressures. All points on the curve between water and gas are the boiling/condensing points at various pressures. In the bottom corner you get the solid ice to gas sublimation/deposition points at different pressures.

Plotting a point on the graph allows us to know the phase exists at a particular temperature and pressure.

The small white box on the graph is more to scale. The sublimation/deposition portion of the graph is very small. The pressure scale goes from approximately zero to 1 atm, then JUMPS to ~218 atmospheres!

Normal Freezing point (0°C) and Normal Boiling point (100°C) are at NORMAL PRESSURE. The Triple point is the pressure/temperature where all six phase changes occur simultaneously, and all 3 phases coexist. The critical point is where liquid and gas H_2O can not be distinguished, they become the same thing.

The Kinetic Molecular Theory and Phases

Is used to explain how solids and liquids and gases exist differently from one another. Solids, liquids and gases are made up of the exact same particles, the difference is how they act together. That is largely based upon temperature, colder particles have lower kinetic energy and move less. Hotter particles move a lot faster and break apart. If particles have sufficient kinetic energy (hot enough temperature) they break apart into gases. Solids have very slow moving particles, they don't move much, but they vibrate. As energy is added, the solid particles shake themselves loose from the intermolecular forces that bind them as solids. Liquid particles have much more kinetic energy, and they are in constant motion, but are still sticky together. Adding enough energy makes them move so fast they can no longer stay close together, they "explode" apart into gases.

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

You should also know by now, how to do pressure conversion problems (mm Hg ↔ kPa ↔ atm ↔ psi). Use Table A to help you through that.

Normal or STANDARD TEMPERATURE + PRESSURE, known as STP, are listed in table A as well.

Table H and Vapor Pressures (and intermolecular forces of attraction)

Vapor pressure is the extra pressure caused by the vaporization of a liquid into a gas, inside of a sealed system (like a can of pork and beans or a sealed bottle. Table H shows 4 liquids and their vapor pressures at different temperatures. To determine vapor pressure, find the temperature, and slide your finger UP to the proper curve, then slide to the left to read the added vapor pressure caused on this liquid by this temperature.

To determine the boiling point of any of these liquids at different pressures, find the pressure, slide right to the proper curve, then slide down to the boiling point temperature.

To decide if one of these four substances is a gas or liquid at any pressure and temperature, put your finger on the "point" of the graph in question and look at the one curve you are deciding about. If your finger is to the right of the curve, it's a gas. If your finger is "behind" the curve then it's a liquid. If it is ON the curve, then both phases exist at this point.

In your pre-chemistry minds water always boiled at 100°C, and that was that. Now you need to realize that water boils at 100°C only if the air pressure is exactly at STANDARD, or normal: Water boils at 100°C only at 101.3 kPa, 760. mm Hg, 14.7 psi, or 1.0 atm. If the pressure were to change, so would the boiling point.
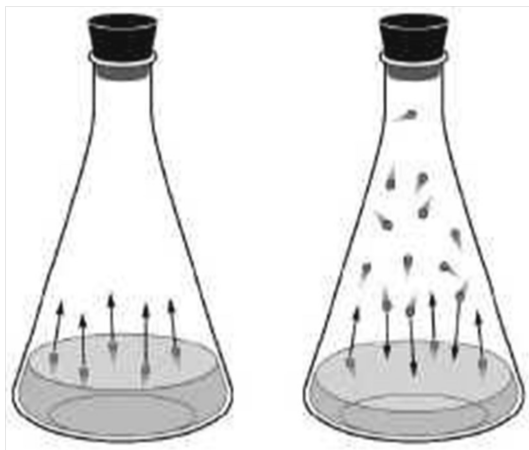
A higher pressure results in higher boiling points. Lower pressure would allow water to boil at a lower temperature.

Liquids can never get hotter than their boiling point. If more heat is applied, the water boils faster but it cannot get hotter. Water actually boils at 97°C in Boulder, Colorado. In fact, it can't get to 100°C in Boulder because the air pressure is lower there than in Vestal.

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 the molecules of propanone do not stick together as well as the other three substances. Ethanoic acid has the lowest vapor pressure, which means it evaporates the least, because it sticks together the strongest.   The term for "stickiness" is intermolecular forces of attraction. Ethanoic acid has the strongest IMF of the four substances. Water has a higher IMF than ethanol. Vapor pressure is determined by temperature and what liquid is present. Propanone, ethanol, water and ethanoic acid have different intermolecular forces of attraction.

At first,
air pressure
exists inside
the bottle

After a bit,
air pressure PLUS
vapor pressure exists
inside the bottle

In a closed system, like a corked bottle, there is air pressure above the liquid. As the liquid evaporates it adds more gas to that space above the liquid, which increases the pressure. This extra pressure, over the air pressure that is already there, is called the vapor pressure. Hotter makes more evaporation, hence more vapor pressure.
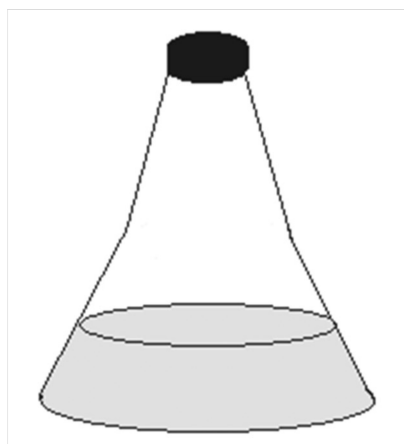
How easily a liquid evaporates, depending upon it's intermolecular attractions, also affects the vapor pressure.

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.
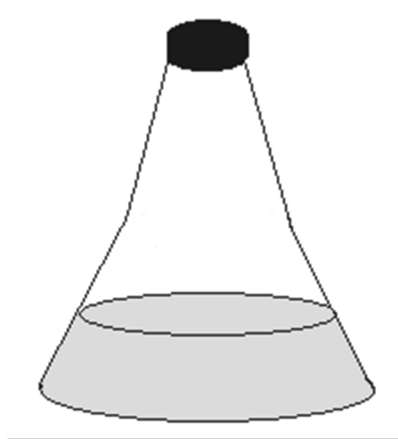
Water

Propanone

STP

Two identical bottles, corked in the same room at standard temperature and pressure.

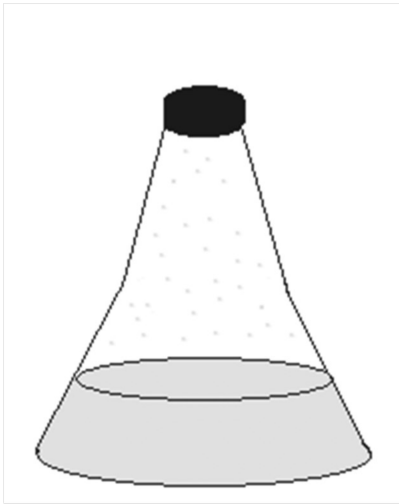Both will evaporate & produce vapor pressure, added to the 101.3 kPa provided by the air already in the bottles.

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

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

Water

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

Propanone

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

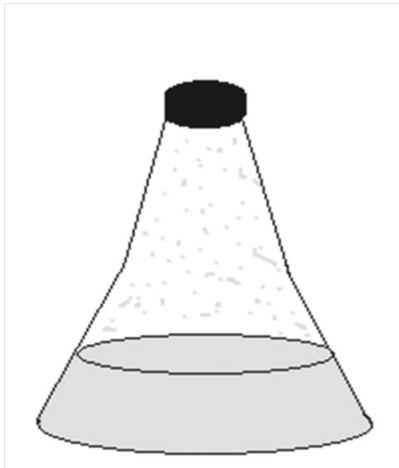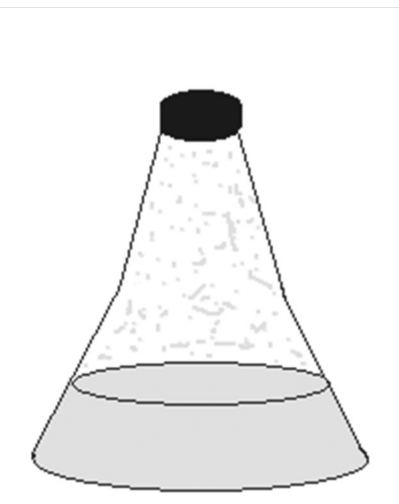Vapor pressure increases
too.

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
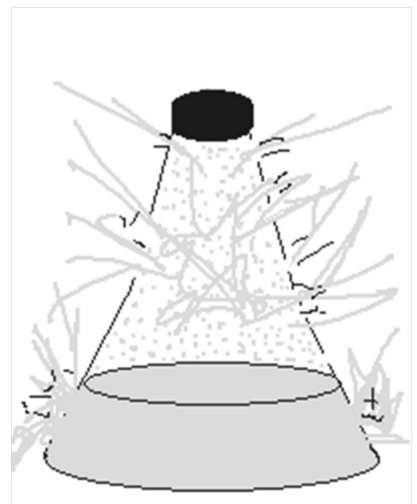= about 139 kPa total pressure

Temperature increases
to 75°C

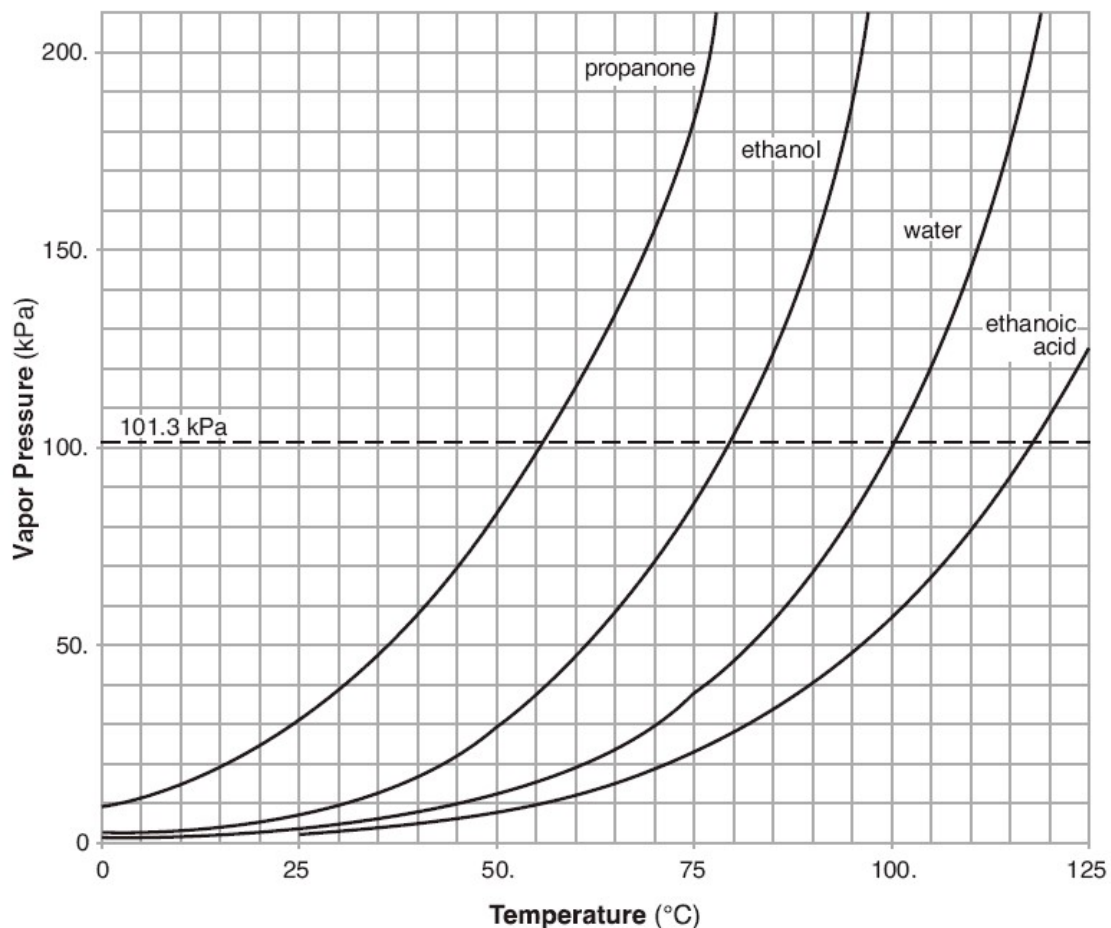Vapor pressure increases
too.

At this temp:

Water VP = 38 kPa

Propanone VP = 82 kPa

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

Table H

Find 50°C and slide your finger up. The first curve you touch is for Ethanoic Acid, then water, then ethanol, and lastly propanone. At any temperature it is the ethanoic acid with the lowest vapor pressure and propanone the highest. This indicates the most "internal stickiness", it evaporates the worst.

Find normal pressure (the dotted line). Slide your finger to the right, and the normal BP of propanone first. Keep sliding and you will get to the normal BP for ethanol, the normal BP for water, and finally, the normal BP for ethanoic acid.

	propanone	water	ethanol	ethanoic acid
VP at 50°C	8 kPa	12 kPa	30 kPa	83 kPa
Normal BP	55°C	79°C	100°C	117°C

Lower vapor pressure is due to less internal attraction.
Lower Boiling points also are due to this lesser internal attraction.

To phase change from liquid to gas the substance must overcome the air pressure pushing down on it's surface, and the internal attraction holding it together.

Both vapor pressure and boiling point "work" with the same principle: internal attraction of the particles, due to the polarity of the molecules. More polar means stronger attraction between them.

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. In the USA we use pounds per square inch, just because. That's equal to the other, more scientific units too.

Standard pressure is equal to 1.0 atm of pressure (1 atmosphere). 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. 101.3 kPa is standard and metric!

In total, all standard pressures are the same, so:

$$1.0 \text{ atm} \ = \ 760. \text{ mm Hg} \ = \ 101.3 \text{ kPa} = 14.7 \text{ psi}$$

Because these quantities are all equal to each other, you can
make many conversion factors from these, to convert one unit to another.

All of these conversion factors are equal to the others

All of these conversion factors are equal to 1.

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

$$\frac{1.0 \text{ atm}}{14.7 \text{ psi}} = \frac{14.7 \text{ psi}}{101.3 \text{ kPa}} = \frac{14.7 \text{ psi}}{760. \text{ mm Hg}} =$$

All of these conversion factors = ONE
All are equal to each other,
and all have unlimited SF in our math.

Practice problems for pressure conversion
(use correct SF each, all 3 SF except for the last one!)

Using the equalities you put onto Table A,
you can make any conversion factors you need.

Convert 744 mm-Hg to atm.	$\dfrac{744 \text{ mm Hg}}{1}$ X $\dfrac{1.0 \text{ atm}}{760. \text{ mm Hg}}$ = 0.979 atm
Convert 744 mm-Hg to kPa.	$\dfrac{744 \text{ mm Hg}}{1}$ X $\dfrac{101.3 \text{ kPa}}{760. \text{ mm Hg}}$ = 99.2 kPa
Convert 155 kPa to mm-Hg	$\dfrac{155 \text{ kPa}}{1}$ X $\dfrac{760. \text{ mm Hg}}{101.3 \text{ kPa}}$ = 1160 mm Hg
Convert 155 kPa to atm	$\dfrac{155 \text{ kPa}}{1}$ X $\dfrac{1.0 \text{ atm}}{101.3 \text{ kPa}}$ = 1.53 atm
Convert 0.724 atm to kPa	$\dfrac{0.724 \text{ atm}}{1}$ X $\dfrac{101.3 \text{ kPa}}{1.0 \text{ atm}}$ = 73.3 kPa
Convert 0.724 atm to mm-Hg	$\dfrac{0.724 \text{ atm}}{1}$ X $\dfrac{760. \text{ mm Hg}}{1.0 \text{ atm}}$ = 550. mm Hg
Convert 0.724 atm to psi	$\dfrac{0.724 \text{ atm}}{1}$ X $\dfrac{14.7 \text{ psi}}{1.0 \text{ atm}}$ = 10.6 psi
Convert 25.6 psi to mm-Hg	$\dfrac{25.61 \text{ psi}}{1}$ X $\dfrac{760. \text{ mm Hg}}{14.7 \text{ psi}}$ = 1324 mm Hg