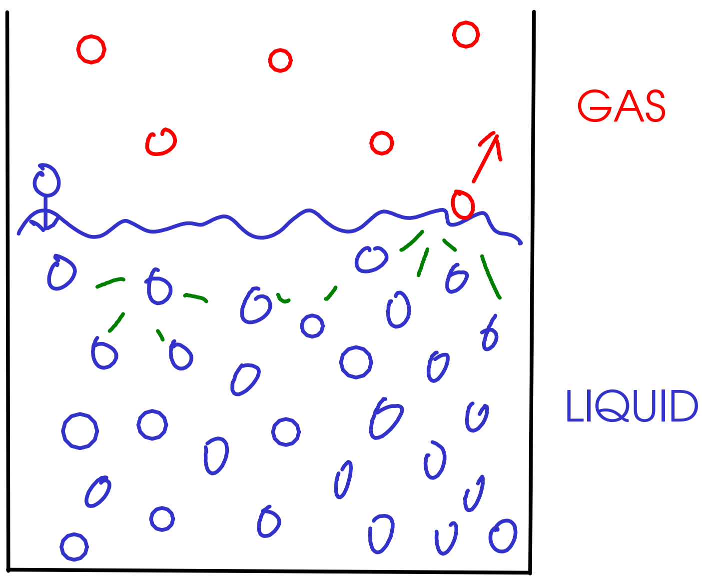
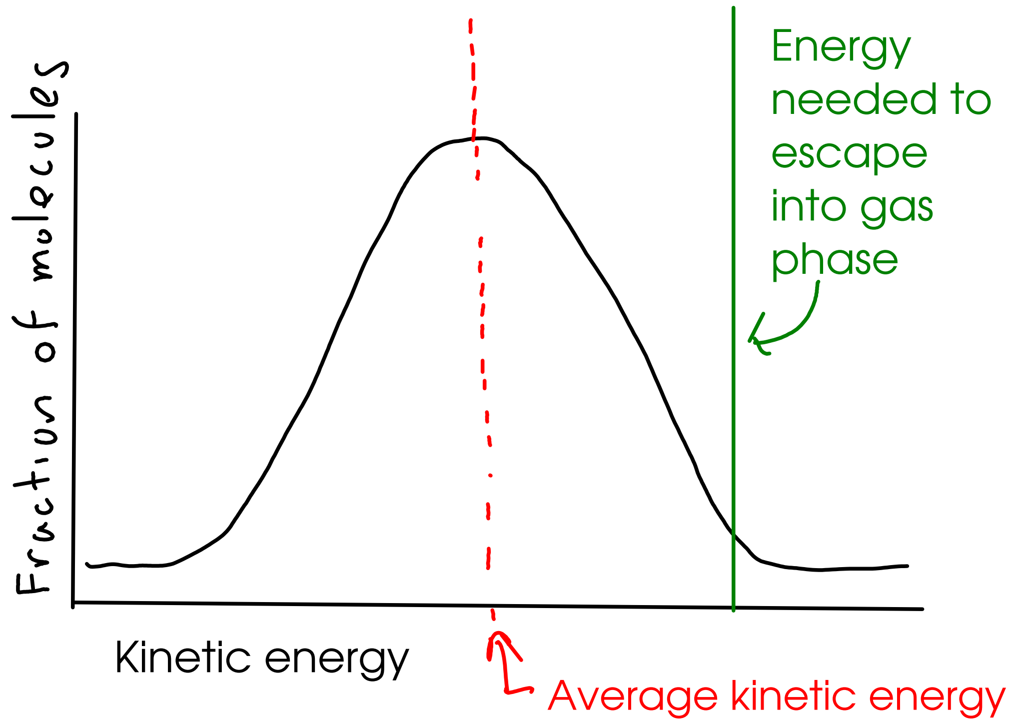
VAPORIZATION IN MORE DETAIL





- For a molecule to move from the liquid phase to the gas phase, it must acquire enough KINETIC ENERGY (which depends on molecular SPEED) to break away from the INTERMOLECULAR FORCES holding the molecule in the liquid.

- The AVERAGE KINETIC ENERGY of molecules is proportional to the TEMPERATURE. On average, molecules in both the liquid and the solid state move faster at higher temperatures.

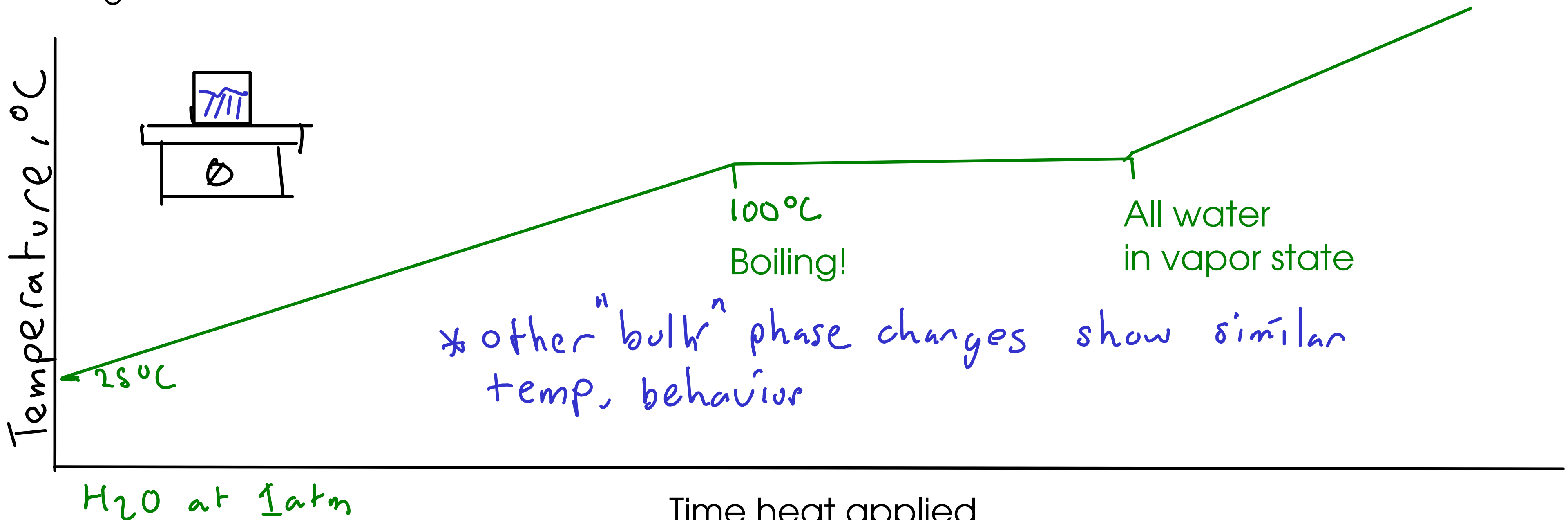
- Even at room temperature, some of the molecules have enough kinetic energy to escape into the gas phase. This accounts for the tendency of volatile liquids to evaporate from open containers even well away from the boiling point.

- As we increase temperature, the gaussian curve above shifts to the right, and a higher fraction of molecules have enough energy to vaporize.

TEMPERATURE PROFILE OF VAPORIZATION

- To increase the rate of vaporization, we can increase the temperature. At the BOILING POINT, the pressure of the vaporizing water is enough to push back against the liquid water and bubbles of vapor form in the liquid: BOILING

- If we look at the temperature changes up to and through the boiling point, we see something interesting...



DURING THE BOILING PROCESS, as long as you have some liquid water remaining, the temperature will remain constant - EVEN AS YOU CONTINUE TO APPLY HEAT!

How do we explain this behavior? ↙ ΔH_{vap}

- The VAPORIZATION itself requires an energy input. What's that energy doing? Breaking water molecules away from one another (breaking apart the water's intermolecular forces).

LIQUIDS

- FLUID, DENSE, INCOMPRESSIBLE

- Posess a few unique properties

(1) SURFACE TENSION

- a measure of the tendency of a liquid to minimize its surface area, or the resistance to the breaking of a liquid surface.

Liquid droplets tend to be spehrical...

... but friction causes falling droplets to take on the traditional "drop" shape.

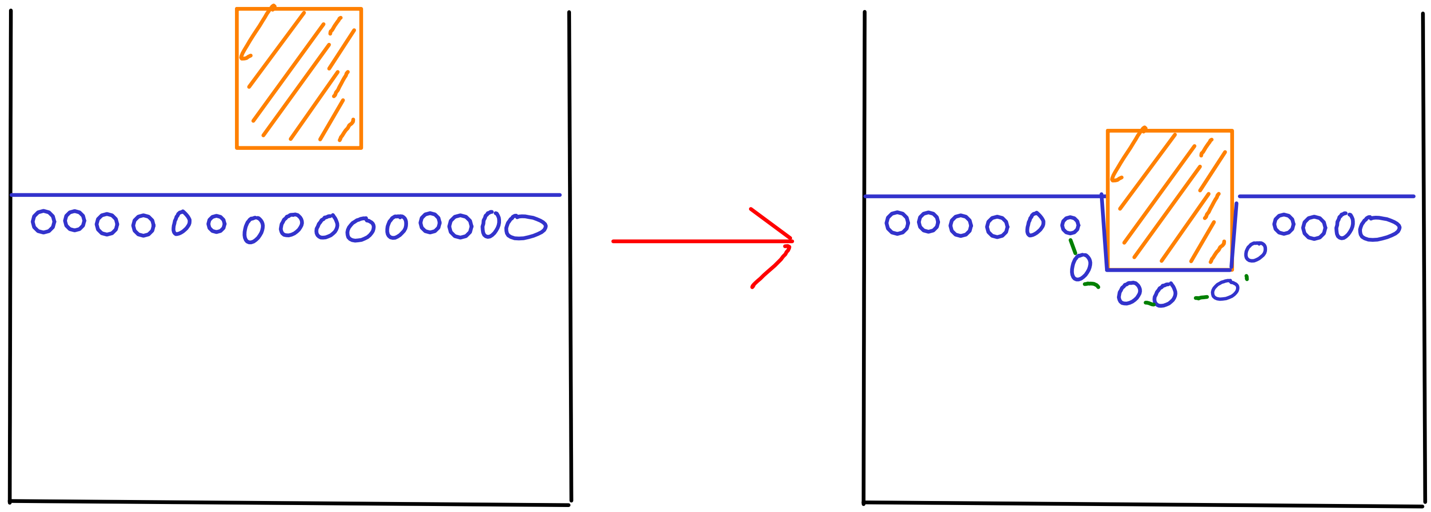
A sphere has the smallest surface area for a given volume.

(2) VISCOSITY (also for gases)

- a measure of a liquids resistance to flow, or "thickness"

SURFACE TENSION

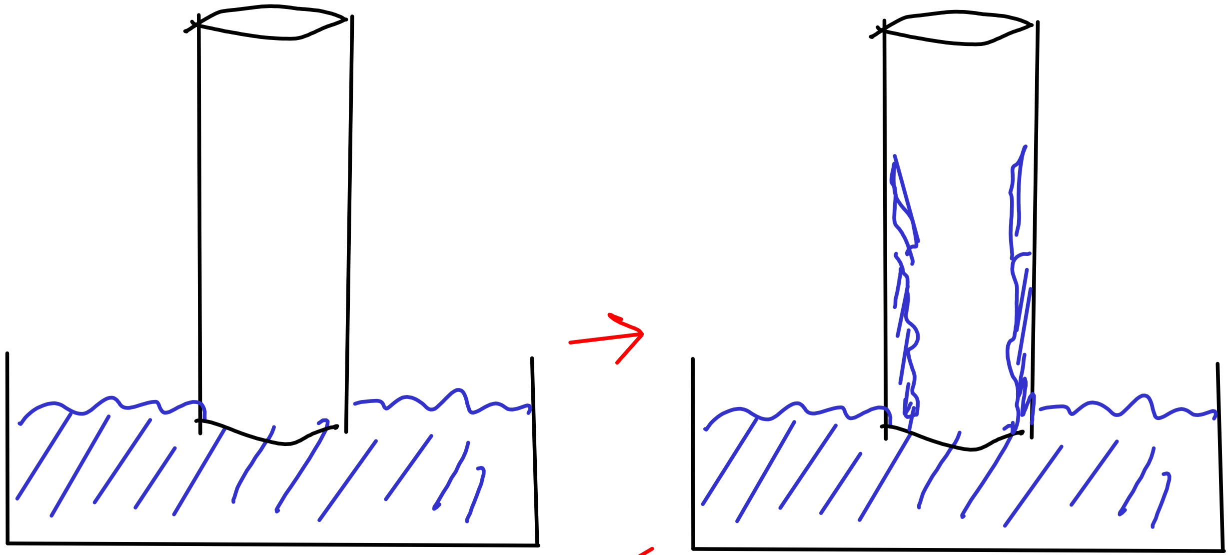
- Surface tension can be explained by looking at liquid molecules as being attracted to each other by INTERMOLECULAR FORCES.



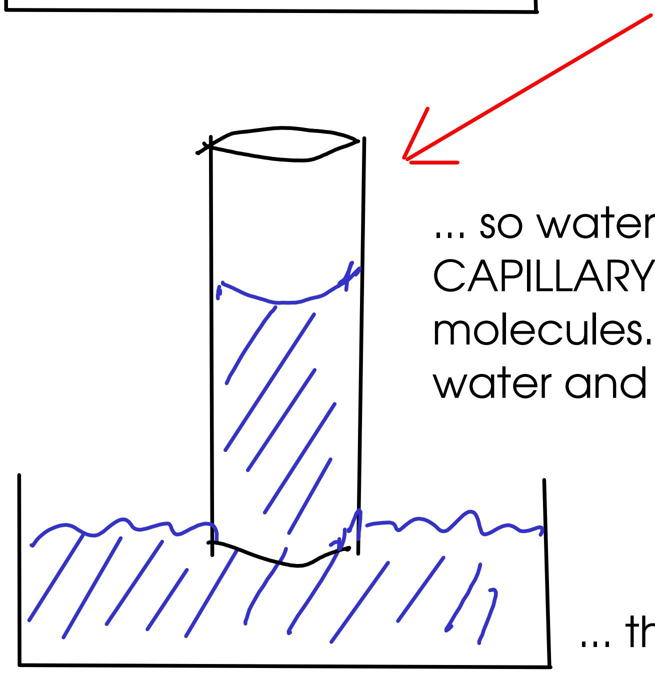
The object increases the surface area of the liquid!

For the object to penetrate the liquid surface, it must push water molecules at the surface apart. Since these water molecules ARE ATTRACTED TO ONE ANOTHER, the liquid will resist!

[39] Surface tension also explains CAPILLARY ACTION, the drawing up of WATER into a glass tube. Water is attracted to glass, and will climb up the surface of a glass tube.



... but this greatly increases the SURFACE AREA of the water, and pulls the water molecules farther from each other.



... so water is PULLED UP THE MIDDLE OF THE CAPILLARY by the attractive forcea between water molecules. This minimizes the surface area of the water and gets water molecules closer to one another.

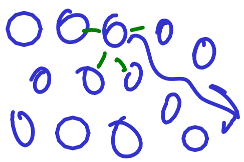
... the water is pulled up AGAINST GRAVITY!

VISCOSITY

- viscosity can also be explained (at least partially) by looking at INTERMOLECULAR FORCES!

- For a liquid to FLOW, its molecules must move past one another. This means that some of the molecules must move farther away from other molecules. Since the molecules in the liquid state are ATTRACTED TO ONE ANOTHER, thst means the flowing will be slowed.

- Viscosity is also determined by STRUCTURE. Liquids with large chains (like oils) which can rotate and tangle in one another will also be viscous.



Molecules have to move past one another to flow, and stronger attractions between molecules make that more difficult!

INTERMOLECULAR FORCES IN LIQUIDS

- "Intermolecular forces" is a generic term. It refers to any number of forces that exist between molecules!

- In liquids, there are three main types of intermolecular force

(1) DIPOLE-DIPOLE INTERACTIONS

- only for polar molecules

(2) LONDON DISPERSION FORCES

- exist in all liquids

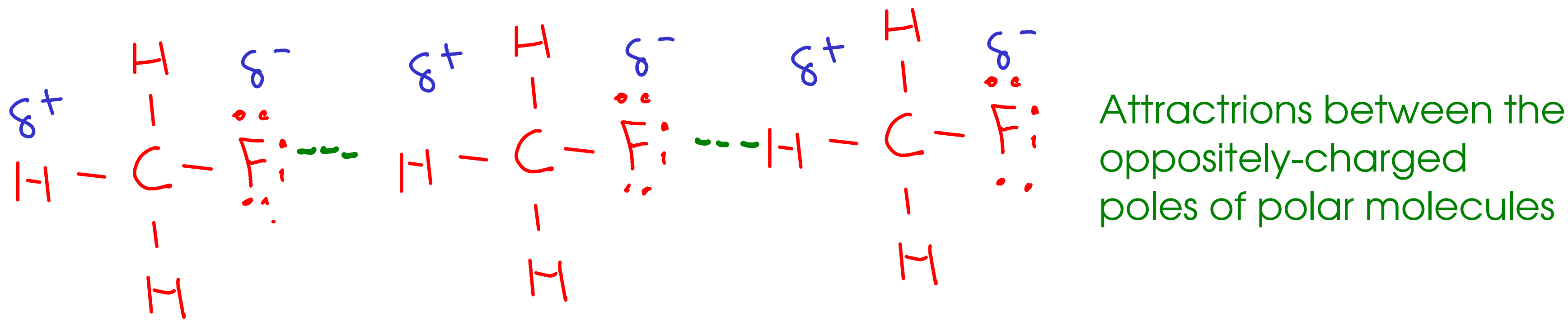
(1) and (2): van der Waals forces...

(3) HYDROGEN BONDS

- exist only when hydrogen is directly bonded to a highly electronegative atom.

DIPOLE-DIPOLE INTERACTIONS

- attractions between the positive pole of one polar molecule and the negative pole of another



- Dipole-dipole interactions occur only between POLAR molecules

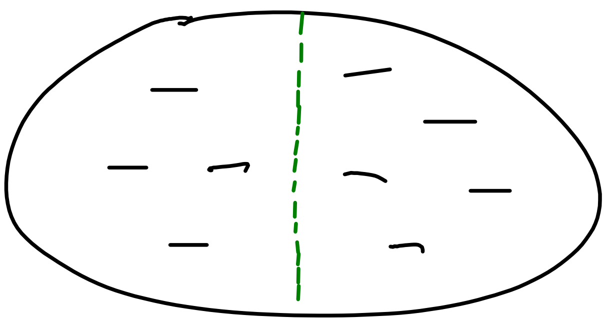
- Dipole-dipole interactions are weak relative to the other two kinds of intermolecular forces in liquids

- The more polar a molecule is (the larger the dipole moment), the stronger its dipole-dipole interactions.

LONDON DISPERSION FORCES

- often called "London forces" for short.

- occurs because electron density is - at any given point in time - likely to be uneven across a molecule due to the simple fact that electrons are MOVING!



Let's say that at one point in time, the electrons around a molecule are here. At this moment, there is a TEMPORARY (INSTANTANEOUS) DIPOLE.

This dipole can interact (and even induce) dipoles on other, nearby molecules.

What happens when the electrons move around some more? The dipole may flip, and the surrounding INDUCED DIPOLES flip along with it. This results in a net ATTRACTIVE FORCE between molecules.

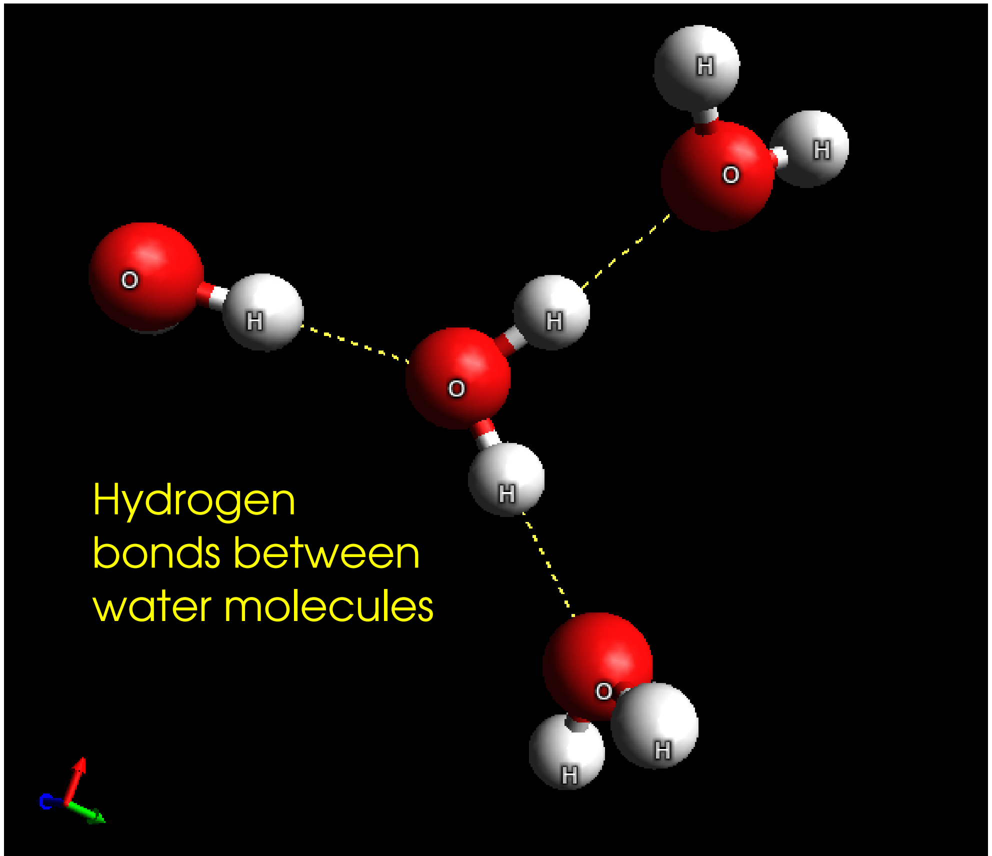
- London forces occur in all molecules, polar or nonpolar.

- London forces increase in strength as molecules get larger. The larger a molecule is, the more easily instantaneous dipoles form and the more easily dipoles can be induced.

- London forces are often the most important interaction between molecules in the liquid state, unless the molecules are capable of HYDROGEN BONDING.

HYDROGEN BONDING

- can happen when there is a HYDROGEN ATOM bonded DIRECTLY to either O, N, or F and that O, N, or F atom has at least one lone pair.



When hydrogen bonds to a very electronegative atom, electron density is pulled away from hydrogen.

Since hydrogen has no core electrons, this effectively exposes the hydrogen nucleus.

The exposed hydrogen nucleus is strongly attracted to any nearby electron density, such as the lone pairs on an adjacent molecule.

- Hydrogen bonds are the strongest type of intermolecular force in the liquid state. They are ALMOST as strong as a covalent bond.

- Hydrogen bonds form in several important molecules: water, ethanol, nucleic acids (they hold the DNA helix together)