Kinetic Molecular Theory, Weak Interactions, States of Matter *(a.k.a.: Why do liquids and solids exist?)*

Themes for the day:

1. What is the Kinetic Molecular Theory?
2. How do #1 & weak interactions help us understand liquids, gases, and solids?

Comment on _______ "being a theory, not a fact" or ________ being "just a theory."

1. A theory is supported by many, many experiments (facts?)
2. A theory provides an intellectual framework that often leads us to new intellectual territory.
3. A theory may be changed somewhat through time. (Little in science is completely unchanging, very much like life itself.) However, it is not common for theories to be modified substantially.

I. The Kinetic Molecular Theory

A. While many aspects of quantum mechanics were a bit beyond our reach, the Kinetic Molecular Theory certainly is not. This theory applies:

1. very directly to gases
2. somewhat less well to liquid solutions
3. not well at all to solid state systems

B. Assumptions (assertions?) of the Kinetic Molecular Theory:

1. In the gas phase molecules move randomly in straight lines, at a range of speeds.
2. The average energy of the molecules is related to the ***temperature*** (**Kelvin units!!!**) of the gas.
3. The molecules exchange energy when they hit each other, but the total energy is conserved. (Elastic collisions, momentum transfer.)
4. The space between the molecules is very large.
5. The gas molecules don't stick to each other.
6. The molecules exert ***pressure*** on the surface of the container they are in. The pressure is related to the number of collisions per unit time, and how fast the molecules are moving. (The container has some finite ***volume***.)

Take a quick look at Boltzmann.

Variables to examine: 1) how many atoms?
2) temperature of system?
3. volume of system?

C. Temperature: proportional (in Kelvin units) to energy content of a sample

1. Symbol: *T* Most people "have a feel" for temperature.

2. Units:
   a) Fahrenheit- $H_2O(s)$ melts at 32° F, $H_2O(l)$ boils at 212° F (Just US?)
   b) Celsius- $H_2O(s)$ melts at 0° C, $H_2O(l)$ boils at 100° F (centigrade?)
   c) Kelvin- $H_2O(s)$ melts at 273.15 K, $H_2O(l)$ boils at 373.15 K
      a) At 0 K, molecules have a velocity of ____________________
         Comment on "absolute zero."
      b) Must use Kelvin scale for all equations/formulas that have their origins in Kinetic Molecular Theory. There are many!

3. Conversion formulas:

   °C = 5/9 (°F−32) (Let's solve this for °F)

   So, °F =

   K = °C + 273.15

Problem to do on your own: Convert 98.6° F into °C and K.

D. Pressure

1. Symbol: *P* This is often harder to get a feel for.

2. *P* is a function of the collisions between the gas and the container walls.
   a) How many collisions per unit time? More collisions = (more or less?) *P*
   b) How fast is the molecule moving? Higher speed = (more or less?) *P*
   c) How much does the molecule weigh? Higher weight = (more or less?) *P*
   d) The pressures in your life are not a bad comparison.

3. Units of pressure

a) atmospheres (atm) Charleston, SC is ≈ 1atm.

b) often in medicine and science, 1.00 atm = 760. mm Hg, 3 s.f. &

1 mm Hg = 1 torr No need to memorize!

E. The Ideal Gas Law (from Avogadro's Hypothesis)

1. The Ideal Gas Law comes directly from kinetic molecular theory.

The Ideal Gas Law is: $PV = nRT$, where

P = pressure

V= volume

n= number of moles (1 mole = 6.022×10^{23} molecules [atoms?])

R= the Universal Gas Constant = 0.0821 (L•atm)/(mol•K)

T= temperature

For P units must be _____ ? For V units must be _____ ?

For n units must be _____ ? For T units must be _____ ?

2. Run through logic associated with $PV=nRT$ Take a closer look at Boltzmann.

a) Why does P increase as n increases? ___________________________________

b) Why does P increase as T increases? ___________________________________

c) Why does P decrease as V increases? ___________________________________

3. The Ideal Gas Law works pretty well for gases if you are not to close to the condensation point. If you know any 3 of the variables, you can always calculate the value of the other variable.

Problem: What volume will 17.00 mol of He occupy at 20.50° C and a pressure of 0.9836 atm?

F. Other Gas Laws that are special cases of Ideal Gas Law (constant n in all)

1. The following three laws are "2 condition problems." All assume n is constant. Begin at condition_1. How does changing one variable alter another?

2. Boyle's Law: $P_1V_1 = P_2V_2$

a) Derive this from the Ideal Gas Law when T and n are constant.

$P_1V_1 = n_1RT_1$ Change P & V, but keep n & T constant: $P_2V_2 = n_1RT_1$

Because obviously $n_1RT_1 = n_1RT_1$ then $P_1V_1 = P_2V_2$

b) If V decreases going from $state_1$ to $state_2$, P ________ (increases or decreases)
c) Is #2 a little like inhaling or exhaling? _____________ (Which?)
d) To which part of the breathing apparatus does this apply? _________________

3. Charles' Law: $V_1/T_1 = V_2/T_2$
   a) P must be constant.
   b) As T increases, V increases.
   c) Would this have many direct applications to the human body? _________
      Why or why not?

4. Gay-Lussac's Law: $P_1/T_1 = P_2/T_2$
   a) V must be constant!
   b) As T increases, P increases. (Many industrial applications.)

5. The Combined Gas Law is just a special case of the ideal Gas Law *(where n is constant). Don't worry too much about this.* $P_1V_1/T_1 = P_2V_2/T_2$

G. Other Gas Laws

1. Dalton's Law
   a) The total pressure is just the sum of the pressures exerted by each of the components in a mixture of gases.
   b) Algebraically: $P_T = P_1 + P_2 + P_3 + ...$
   c) This means that the Ideal Gas Law applies individually to each of the components in a mixture of gases. For air, roughly:

$$P_T = P_{oxygen} + P_{nitrogen}$$

II. Weak Interactions *(Important!!!)*

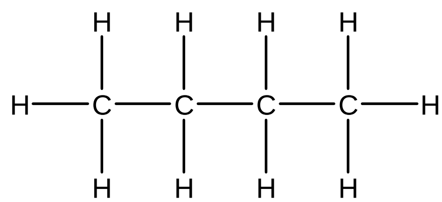
Liquids and solids form because item #5 of the kinetic molecular theory does not completely apply to real (as opposed to Ideal) gases.

All materials (even He) can exist in *(s)* or *(l)* state if temperature is low enough.

A. Dipole-dipole interactions

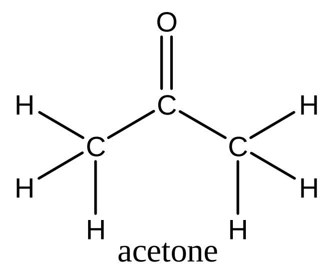
1. Remember our previous interest in polarity. Draw picture, see p. 10 & below

Polar? (Y or N)



butane

Polar? (Y or N)



acetone

2. This is your standard electrostatic attraction. Electronegativities? Spacial dependence?

Draw picture below showing how 2 acetone molecules would stick together:

B. Temporary dipole-induced dipole (a.k.a. London Forces, van der Waals interactions)

1. He can be liquified (b.p. ~ 4 K) How *do* He atoms stick together?

Imagine 2 helium atoms that happen to be next to each other:

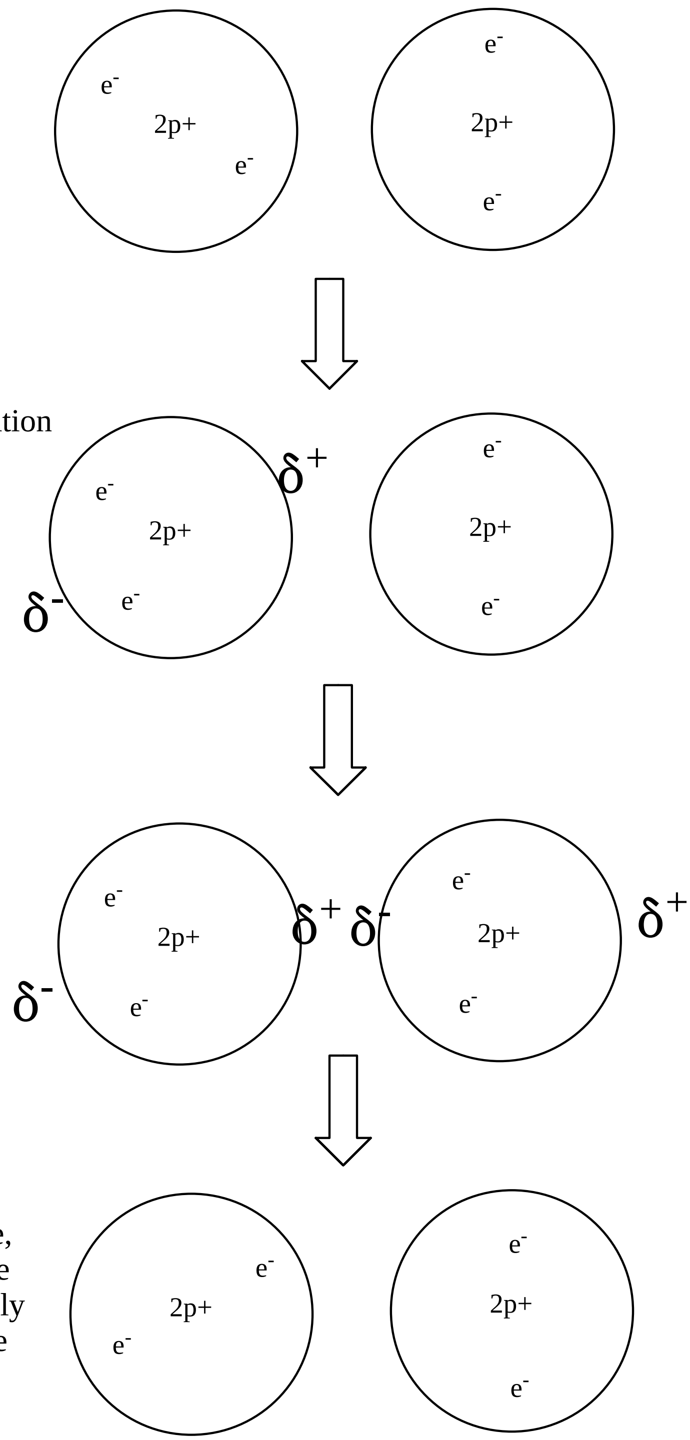
Even charge distribution

Every so often e^- distribution gets uneven. Here, the left atom forms a ***temporary dipole***.

The left hand dipole causes a dipole to form in the right atom. This is the ***induced dipole***.

Electrons are very mobile, so the temporary dipole goes away in a relatively short time. So does the attraction.

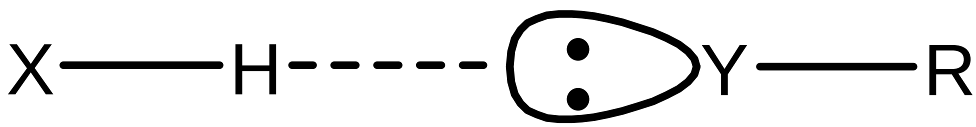
Soap opera link?



2. All atoms exhibit London Dispersion Forces.

C. Hydrogen Bonds ***(This is important!!!)***

1. This is sometimes called a “minor” valence, but remember, it is **not** a standard covalent bond.
2. To form a Hydrogen Bond, you must have:
   - a) a *highly* e^- deficient H atom (bonded to ______________?)
   - b) a non-bonding e^- pair on a N, O, or F atom.
   - c) appropriate orientation & distance of a) & b). Contrast w/ dipole-dipole interactions. *Bracket the Hydrogen bond!!!*



Notes:

- i) Use a dashed or dotted, not solid, line to show this bond type
- ii) Bond strength is greater with larger electronegativity differences
- iii) Attraction between molecules is stronger if more Hydrogen Bonds are involved.
- iv) In a biologically interesting Hydrogen Bond, both X and Y have electronegativity values greater than or equal to 3.0.

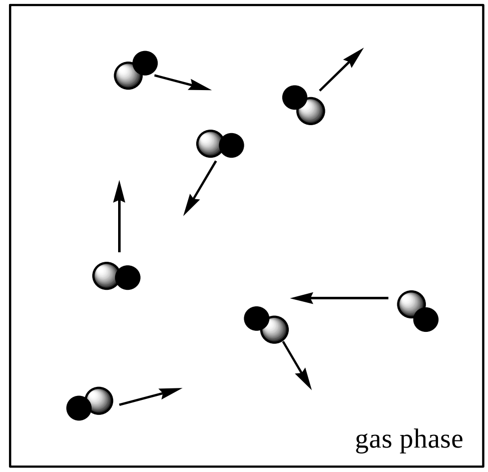
D. Relative interaction strength (how much energy to break bond):

		Example
1) Covalent bonds:	300-450 kJ/mol	O – H C – C
2) Ionic bonds in water:	80-90	–COO^------$^+NH_3$ –
3) Hydrogen Bond:	20	O – H------NH_2 –
4) Dipole-dipole:	9	C=O δ^- ------ δ^+C=O
5) London Dispersion	0.3	C – H ------ H – C

III. States of Matter

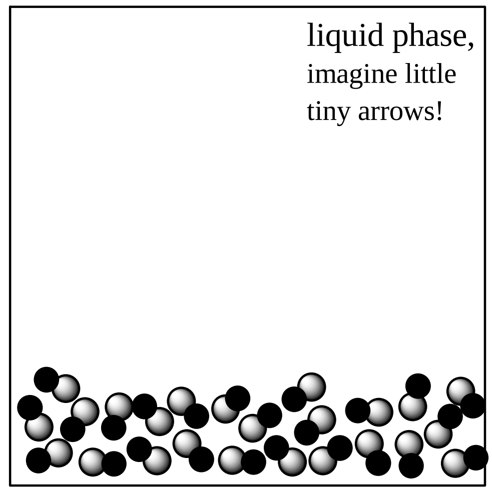
A. Gas *(g)*

1. Microscopically: lots of movement, lots of space between molecules/atoms
2. Macroscopically: fills the container, whatever its size (7 molecules)
3. Is density constant in gas phase?   Yes or No?



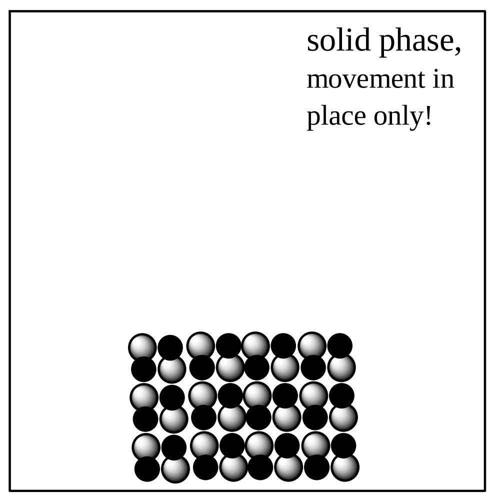
B. Liquid *(l)*

1. Microscopically: lots of movement, (Think of the molecules sliding around each other.) very little space between molecules/atoms
2. Macroscopically: adapts to shape of container, but fixed density (at a given *T*)   (24 molecules)
3. Compare density with gas phase above.



C. Solid *(s)* The one shown is crystalline.

1. Microscopically: only movement in place (oscillate, vibrate), very little space between molecules
2. Macroscopically: has fixed shape, fixed density is usually greater than liquid phase & relatively *T* independent. (24 molecules)
3. Apparent density relative to *(g)* and *(l)*?

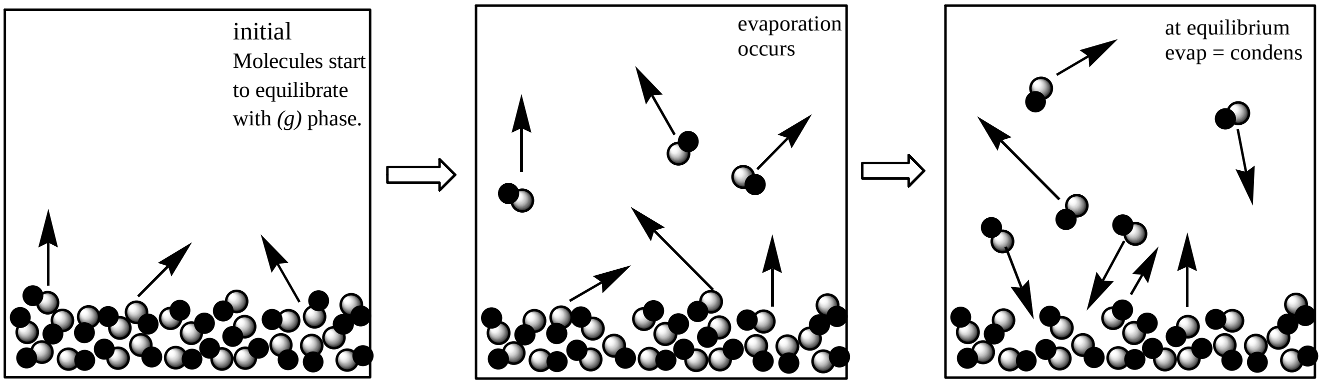


D. If you start (mentally) from the gas phase, why do liquids and solids form?

1. Intermolecular attractive forces (a.k.a., weak interactions)
2. What does it mean when a type of molecule in the gas phase condenses at a relatively high temperature? ______________________________________

E. What is occurring (at the molecular level) when you boil H_2O?

1. Define vapor pressure



a) Start with a liquid in a closed container. (left panel, above)

b) The liquid phase molecules are moving, and some are moving fast enough to escape from their neighbors and go into the gas phase (center panel)

i) This process is called *evaporation* (If container is open........?)

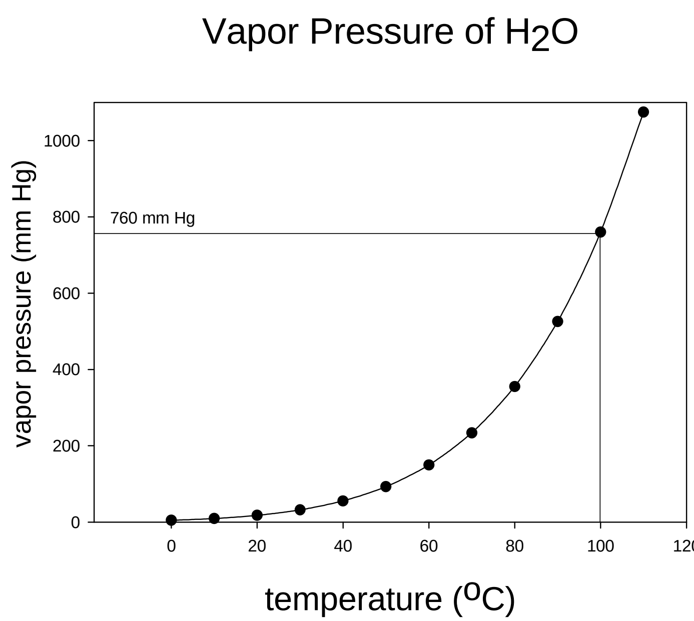
ii) Which molecules are most likely to escape (Boltzmann distribution?)

c) As molecules build up in the vapor phase, they start colliding with the surface of the liquid phase. Some of them stick to the liquid phase. (right panel)

i) This process is called *condensation*

ii) When the $\text{evaporation}_{rate} = \text{condensation}_{rate}$, we have reached equilibrium.

d) The pressure exerted by the gas phase at equilibrium is the ***vapor pressure*** of that liquid at that temperature.



2. How does vapor pressure vary as a function of *T*?

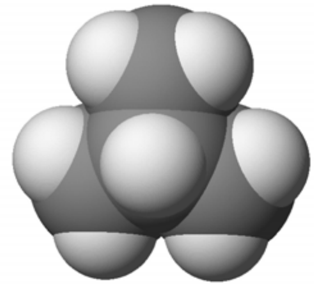
3. What happens when the vapor pressure = the external atmospheric pressure? (What is boiling?)

Can you use this page to explain evaporative cooling and why it takes longer to cook an egg by boiling in Denver, CO than in Greer, SC?

Weak interactions and boiling points

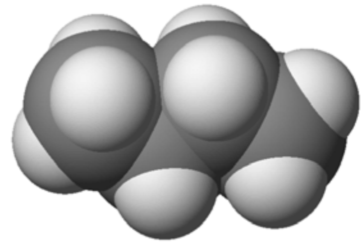
All compounds have 58.1< MW <60.1

2-methyl propane −42° C



Does this mean the 2-methyl propane molecules escape from each other at high speed or low? ____________

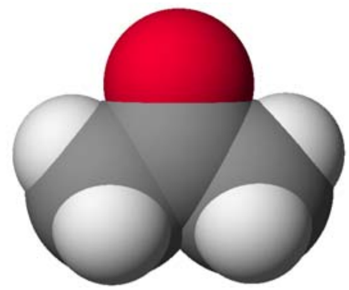
butane −0.4° C



Does this mean the butane molecules escape from each other at high speed or low? ______________________

What does that say about attractive forces?

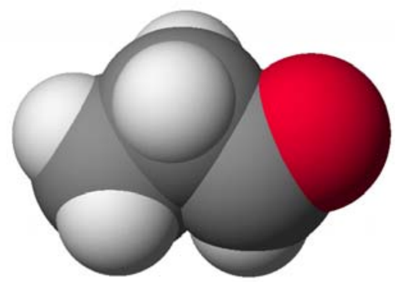
acetone 56° C



Are acetic acid molecules moving faster than butane molecules at −42° C?

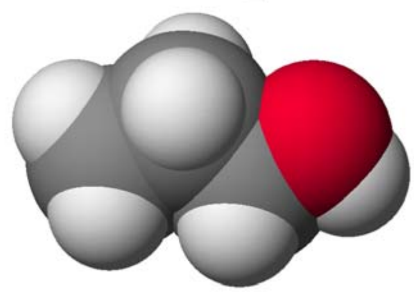
____

propanal 49° C



Why can butane, but not acetic acid molecules escape each other at −42° C? ___________________________

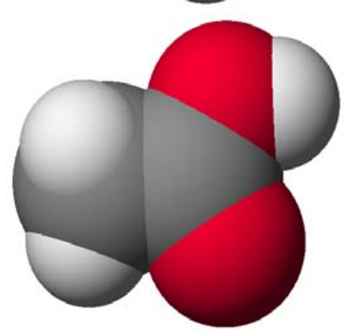
1-propanol 92.7° C



What can you say about the weak interaction attractive forces of acetic acid compared to butane?

____

acetic acid 117.8° C



Looking at the data/structures here, you should be able to determine what kinds of interactions maintain the liquid state for each compound.