Kinetic Molecular Theory, Weak Interactions, States of Matter
(Why do liquids & solids exist?)

Reading: Chap 6 p 176-189, Chap 4 p.111-112
Problems: Ch 6.1, 2, 5, 9, 10, 15a, 23, 25, 29, 35, 37, 39a, 45a, c, e, 53, 55, 47, 51, 53, 55, 56

Themes for the chapter:
1. What is the Kinetic Molecular Theory?
2. How do #1 & weak interactions help us with (s), (l), & (g)?

Remember that KMT being a theory means it is considered to be at the highest level of development of an idea.
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 change through time. (Little in science is completely static, very much like life itself.)
   However, it is not common for theories to be modified substantially.

0. Matter can exist in different states
   The state depends on the temperature (and pressure).

   Melt          →
   Bose-Einstein
   Condensate    ←
   Solid         ←
   Liquid        ←
   Gas           ←
   Plasma        ←

   Of interest to living things.
   Freeze        ←
   Condense      ←

   What direction are we moving along diagram, if heat is being added? (l or r)

   What is sublimation?

   The state (of matter) also depends on the interactions that hold particles together.
   For molecules, the interactions that hold them together are called weak interactions.
   First we will introduce the Kinetic Molecular Theory (KMT), and then weak interactions.
   Then we will see how KMT & weak interactions relate to states of matter (s, l, & g), boiling point and melting point.

I. The Kinetic Molecular Theory

   A. This theory applies very directly to gases

   Gas: lots of movement,
   lots of space between
   molecules/atoms
B. Tenets (or assertions?) of the Kinetic Molecular Theory:

1. In the gas phase molecules move randomly in straight lines, at a range of speeds.

2. 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 volume of a gas molecule is negligible, and the space between the molecules is very large.

5. The gas molecules don’t stick to each other. (There are no attractive forces.)

One characteristic of a gas is that it exerts pressure on the walls of the container. With the “kmt”, we can understand pressure at the molecular level as being a result of gas molecules colliding with the container walls.

So amount of pressure is related to:

- the number of collisions per unit time, *and* how fast the molecules are moving.

Take a quick look at Boltzmann.

First note:

1. movement of atoms

2. collisions

3. speed of atoms

Variables to examine:

1. how many atoms?

2. temperature of system?

3. volume of system?

4. mass of molecules?

C. Temperature, *T*, is a measure of the kinetic energy of molecular motion.

1. Temperature scales:

<table>
<thead>
<tr>
<th></th>
<th>Fahrenheit, °F</th>
<th>Celsius, °C</th>
<th>Kelvin, K</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$_2$O(s) melts</td>
<td>32°F</td>
<td>0°C</td>
<td>273.15 K</td>
</tr>
<tr>
<td>H$_2$O(l) boils</td>
<td>212°F</td>
<td>100°C</td>
<td>373.15 K</td>
</tr>
</tbody>
</table>

i) At 0 K, molecules have a velocity of

“absolute zero”

ii) Must use Kelvin scale for all equations/formulas that have their origins in Kinetic Molecular Theory. There are many!
2. Conversion formulas:

\[ ^\circ C = \frac{5}{9} (^\circ F - 32) \]  
(Let’s solve this for \(^\circ F\))

So, \(^\circ F = \)

\[ K = ^\circ C + 273.15 \]

Problem to do on your own: Convert 98.6 \(^\circ F\) into \(^\circ C\) and \(K\).

D. Pressure, \(P\)

1. \(P\) is a function of the collisions between the gas and the \textit{container walls}.

   a) How many collisions per unit time?  
   \begin{align*}
   \text{More collisions} &= \text{(more or less?)} \quad P \\
   \end{align*}

   b) How fast is the molecule moving?  
   \begin{align*}
   \text{Higher speed} &= \text{(more or less?)} \quad P \\
   \end{align*}

   c) How much does the molecule weigh?  
   \begin{align*}
   \text{Higher weight} &= \text{(more or less?)} \quad P \\
   \end{align*}

   d) The pressures in your life are not a bad comparison.

2. Units of pressure

   a) atmospheres (atm)  
   Charleston, SC \(\approx 1\) atm.

   b) (in medicine & science)  
   \begin{align*}
   1 \text{ atm} &= 760.00 \text{ mm Hg}, \\
   1 \text{ mm Hg} &= 1 \text{ torr} \\
   \end{align*}

E. The Ideal Gas Law (re. Avogadro’s Hypothesis)

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

The \textbf{Ideal Gas Law} is: \(PV = nRT\), where

\begin{align*}
P &= \text{pressure} \\
V &= \text{volume} \\
n &= \# \text{ of moles} \quad (1 \text{ mole} = 6.022 \times 10^{23} \text{ molecules } [\text{atoms}]) \\
R &= \text{the Universal Gas Constant} = 0.0821 \text{ (L·atm)/(mol·K)} \quad \textit{This helps with units!!!} \\
T &= \text{temperature} \\
\end{align*}

For \(P\) units must be \(\text{atm}\),  
For \(V\) units must be \(\text{L}\),  
For \(n\) units must be \(\text{mol}\),  
For \(T\) units must be \(\text{K}\).
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 well for gases not near the condensation point. Given any 3 variables, you can always calculate the other.

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

F. Other Gas Laws that are special cases of Ideal Gas Law (where one or more variables are kept constant.)

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: $P \times V$ is constant
   a) Assumes $T$ and $n$ are kept constant.
   b) A useful way to express this is: $P_1V_1 = P_2V_2$
   c) Derive (b) from the Ideal Gas Law

   $P_1V_1 = n_1RT_1$
   $P_2V_2 = n_1RT_1$

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

   d) If $V$ decreases going from state 1 to state 2, then $P$ _____ ($\uparrow$ or $\downarrow$)

Problem: If an O$_2$ tank used in a health care situation has a volume of 4.5 L and a pressure of 87.5 atm, what is the volume of the same amount of gas if the pressure is 0.98 atm? Assume the temperature is constant.

3. Charles’ Law: $V/T$ is a constant
   a) Assumes $P$ and $n$ are kept constant.
   b) A useful way to express this is: $V_1/T_1 = V_2/T_2$
   c) As $T$ increases, does $V$ increase or decrease? ______

   a) Assumes $n$ and $V$ must be constant! $P_1/T_1 = P_2/T_2$
   b) As $T$ increases, $P$ increases. (Industrial applications!)
5. The **Combined Gas Law** (a special case of the ideal Gas Law): \( P \times \frac{V}{T} \) is constant

G. Other Gas Laws (Dalton’s Law)

1. The total pressure is just the sum of the pressures exerted by each of the components in a mixture of gases.
2. Algebraically: \( P_T = P_1 + P_2 + P_3 + ... \)
3. This means that the Ideal Gas Law applies individually to each of the components in a mixture of gases.

\[
P_T = P_{\text{oxygen}} + P_{\text{nitrogen}} + P...
\]

*What other \( P \) would be important for a patient receiving respiratory gases?*

II. Weak Interactions (Important!!)

A. Introduction

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

2. All materials (even He) can exist in \((l)\) [or \((s)\)] state if temperature is low enough.

3. Discuss three weak interactions or inter-molecular forces, IMF
   a. dipole-dipole interactions
   b. London forces
   c. hydrogen bonding

   For “smallish” molecules, their relative strength is:
   H-bonding > dipole-dipole > London forces

B. Dipole-dipole interactions

1. Remember our previous interest in polarity?

   ![butane](image1.png)  ![acetone](image2.png)

   Polar? (Y or N)

   Can you locate \( \delta^+ \) and \( \delta^- \) on either structure?

2. A result of standard electrostatic attraction.

Draw a picture below showing how 2 acetone molecules would stick together:
C. Temporary dipole-induced dipole (a.k.a. London Forces, van der Waals interactions)

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

Imagine 2 “cold” helium atoms that happen to be next to each other: (see next page)

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. Note:
All atoms exhibit London Dispersion Forces.
D. Hydrogen Bonds *(This is important!!!)*

1. 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.

![Hydrogen bond components](image)

Notes:
  1. Use a dashed or dotted, **not solid**, line to show this bond type
  2. Why shouldn’t you use a solid line?
  3. Bond strength is greater with larger electronegativity differences
     (Which has higher electronegativity, N or O?)
  4. Attraction between molecules is stronger if more Hydrogen Bonds are involved. Two arm vs. one arm pullup?
  5. In a biologically interesting Hydrogen Bond, both X and Y have electronegativity values greater than or equal to 3.0.

  **Which elements have electronegativity $\geq 3.0$? which of these are covalently important in your body?**

3. Identify the donor (d) and acceptor (a) sites:

   ![Molecules with donor and acceptor sites](image)
4. Show one of the structures above forming a Hydrogen Bond:

5. Which would be stronger, and why:
   a) 
   b) 
   c) 

E. Comparison

<table>
<thead>
<tr>
<th></th>
<th>Non polar atoms/molecules</th>
<th>Polar molecules</th>
<th>Polar molecules w/ hydrogen bonding capabilities</th>
</tr>
</thead>
<tbody>
<tr>
<td>Example</td>
<td>Example: CH₄</td>
<td>Example: CH₂O</td>
<td>Example: H₂O</td>
</tr>
</tbody>
</table>
F. Relative strength (how much E to break bond): 

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Energy (kJ/mol)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Ionic bond(s) (lattice energy)</td>
<td>600-15,000</td>
<td>Na(^+) Cl(^-) (in a crystal)</td>
</tr>
<tr>
<td>2) Covalent bonds</td>
<td>300-450</td>
<td>O — H or C — C</td>
</tr>
<tr>
<td>3) Ionic bonds(aq):</td>
<td>80-90</td>
<td>—COO(^-)-----(^+)NH(_3) — (in H(_2)O)</td>
</tr>
<tr>
<td>4) Hydrogen Bond</td>
<td>20</td>
<td>O — H-----NH(_2) —</td>
</tr>
<tr>
<td>5) Dipole-dipole</td>
<td>9</td>
<td>C=O (\delta^-) ----- (\delta^+)C=O</td>
</tr>
<tr>
<td>6) London Dispersion</td>
<td>0.3</td>
<td>C — H ----- H — C</td>
</tr>
</tbody>
</table>

III. States of Matter

A. Gas (g) (See Boltzmann)

1. Microscopically: lots of movement, lots of space between molecules/atoms
2. Macroscopically: fills the container, whatever its size (7 molecules)
3. Is density high or low in a gas? (Compressibility!)

B. Liquid (l)

1. Microscopically: lots of movement, (Molecules sliding around each other.) very little space between molecules/atoms
2. Macroscopically: adapts to shape of container, and has a 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 occurs (at molecular level) when H₂O boils?
1. Defining vapor pressure

   a) Start w/ 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
      
         ii) Which molecules are most likely to escape liquid (Boltzmann distribution?)

   c) As molecules build up in the vapor phase, they start colliding with the surface of the liquid
      i) This process is called condensation
      ii) When the evaporation_rate = 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?
IV. Weak interactions and boiling points

All compounds have $58.1 < \text{MW} < 60.1$

<table>
<thead>
<tr>
<th>Compound</th>
<th>Boiling Point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2-methylpropane</td>
<td>$-42^\circ$</td>
</tr>
<tr>
<td>Butane</td>
<td>$-0.4^\circ$</td>
</tr>
<tr>
<td>Acetone</td>
<td>$56^\circ$</td>
</tr>
<tr>
<td>Propanal</td>
<td>$49^\circ$</td>
</tr>
<tr>
<td>1-Propanol</td>
<td>$92.7^\circ$</td>
</tr>
<tr>
<td>Acetic Acid</td>
<td>$117.8^\circ$</td>
</tr>
</tbody>
</table>

Based on the b.p., do the 2-methylpropane molecules require high or low kinetic energy to escape from each other? Will they be moving away from each other at high or low speeds? Is 2-methylpropane polar or non-polar? What type of IMF does this molecule have?

Compared to 2-methyl propane, does a butane molecule require more or less kinetic energy to escape from another butane molecule? What does this say about the relative strength of the attractive forces between molecules of 2-methylpropane compared to butane? What type of IMF does butane have?

Is acetone polar or non-polar? Can two acetone molecules form a H-bond? What type of IMF exist with acetone?

Is propanal polar or non-polar? Can two propanal molecules form a H-bond? What type of IMF exist with propanal?

Is 1-propanol polar or non-polar? Can two 1-propanol molecules form a H-bond? What type of IMF exist with 1-propanol?

Based on bp, does an acetic acid molecule require more or less kinetic energy to escape from another acetic acid molecule as compared to molecules of propanal? What does this say about the relative strength of the attractive forces between molecules of propanal compared to acetic acid? Is acetic acid polar or non-polar? Can two acetic acid molecules form a H-bond? What type of IMF exist with acetic acid?