Properties of Liquids and Solids

World of Chemistry
Chapter 14

14.1 Intermolecular Forces
Most substances made of small molecules are gases at normal temperature and pressure.

ex: oxygen gas, $\text{O}_2$; nitrogen gas, $\text{N}_2$; methane gas, $\text{CH}_4$; and carbon dioxide, $\text{CO}_2$

An exception to this is water.

Water is a small molecule but does not exist as a gas at normal temperature and pressure.

WHY? It has to do with something called intermolecular forces; forces that occur between molecules.

Don’t confuse intermolecular forces with intramolecular forces which are bonds that occur inside the molecules.
There are three types of intermolecular forces:

1. **dipole-dipole attraction**
   - polar molecules have a “center” of positive charge and a “center” of negative charge
   - we say that the polar molecule has a dipole moment
   - when these molecules are put together, they orient themselves so that positive and negative ends are close together (Figure a)

   ![Dipole-Dipole Interaction](image)

   (a)

   - this is called dipole-dipole attraction
   - this force is only about 1% as strong as covalent or ionic bonds; they become weaker as the distance between the dipoles increase
   - in liquids, the dipoles find the best compromise between attraction and repulsion (Figure b)
   - in gas phase, the molecules are far apart to these forces are unimportant

2. **hydrogen bonding**
   - a particularly strong dipole-dipole force can occur between molecules in which hydrogen is bound to a highly electronegative atom such as nitrogen, oxygen, or fluorine
   - strength of force is due to great polarity of the bond and the close approach of the dipoles
   - this force is called hydrogen bonding

   ![Hydrogen Bonding](image)

   • hydrogen bonding affects physical properties
   • increased boiling point because a large amount of energy is needed to overcome the strong intermolecular forces
3. London dispersion forces
   - these are forces that exist among noble gas atoms and nonpolar molecules like \( \text{H}_2 \), \( \text{N}_2 \), and \( \text{I}_2 \)
   - we usually assume that electrons of an atom are uniformly distributed about the nucleus
   - this is not true for every instance

— atoms must be greatly slowed down in order for London dispersion forces to lock the atoms into place to produce a solid
— this explains why noble gases have such low freezing points
— London forces become stronger as atom/molecule sizes increases since there are more electrons available to form the dipoles

Copy this on last page of notes:

**Strength of forces:**
Covalent > Ionic > Hydrogen Bond > Dipole-dipole > London dispersion

**Stronger forces have higher**… **but lower**…
Melting points
Boiling points
IMF WS #1
Focus Q’s#1-3 (pg. 449)
1. Intermolecular forces are found between molecules (ex. Hydrogen bond, London forces,..) and intramolecular forces are found between atoms (ex. ionic and covalent bonds).

Chapter Review Q#1-9 pg. 463
1. Dipole-dipole attraction is an attraction between polar molecules (atoms different in molecule). Attraction between SO₂, H₂S and CHCl₃ molecules would be large.
2. No, hydrogen bonding occurs between molecules so it is not considered a true bond.
3. Molecules without dipole moments form attractions via London dispersion forces which allows it to form liquid or solids.
4. Hydrogen bonding is an exceptionally strong dipole-dipole attraction when the molecule contains a H bonded directly to F, O, or N. Examples are H₂O, NH₃, HF.
5. Of the four compounds, only water is capable of hydrogen bonding (much stronger than dipole-dipole) so it takes more energy to change it from liquid to gas.
6. In gas phase, molecules are too far apart for dipole-dipole attraction to be important.
6. Noble gases can be liquified or solidified via attraction from London dispersion forces at very low temperatures.

7. a) Kr – London dispersion forces 
b) $S_8$ (all same element) - LDF 
c) $NF_3$ – LDF and dipole-dipole 
d) $H_2O$ – LDF and Hydrogen bonding

8. Noble gases are only capable of London dispersion. These forces become stronger as the size of the atom increases. (Larger noble gas has higher boiling points!)

9. The attraction between water and alcohol is stronger than when these substances are alone. Therefore the mixed molecules come close together and volume appears to get slightly smaller.

14.2 Water and Its Phases

What happens when we heat liquid water?

1. temperature of water rises 
2. motions of water molecule rises 
3. temperature of water reaches 100 °C

4. bubbles develop in the interior of liquid and floats to the surface and bursts; it has reached boiling point

5. water remains at 100 °C until all the liquid has changed into vapor (gas)

6. at 1 atm, normal boiling point of water is 100°C

7. see heating/cooling curve
14.2 Water and Its Phases

What happens when we cool liquid water?
1. temperature decreases until it reaches 0°C
2. temperature remains 0°C until all liquid has changed into ice (solid)
3. at 1 atm, normal freezing point of water is 0°C
4. see heating/cooling curve

14.3 Energy Requirements for the Changes of State

Remember, changes of state from solid to liquid and from liquid to solid are physical changes.

- No chemical bonds are broken in these processes.
- When water is boiled the water molecules separate but the molecule itself stays intact.

Density of Water

Solid water is less dense (0.917 g/mL) than liquid water (1.000 g/mL).

It takes more energy to vaporize a mole of water than to melt a mole of ice because in liquids, the particles are relatively close together so most of the intermolecular forces are still present. However, when particles go from liquid to gaseous state, they must be moved farther apart by removing all intermolecular forces.
14.4 Evaporation and Vapor Pressure

vaporization or evaporation: molecules of a liquid escape the liquid’s surface and form a gas; this process requires energy to overcome the intermolecular forces in the liquid.

Behavior of liquid in a closed container

– at first, it appears as if the amount of liquid decreases
– this is because of the transfer of molecules from liquid to gas phase

– then the amount becomes constant as the number of molecules returning from gas to liquid phase (rate of condensation) equals the number of molecules changing from liquid to gas phase (rate of evaporation)
– it has reached equilibrium; no net change in the two processes because they balance each other out
equilibrium vapor pressure or vapor pressure: the pressure of the vapor present at equilibrium with its liquid

- the substance is injected to the bottom of the mercury column
- because mercury is so dense, the substance rises to the top of the column
- when the substance produces vapor, some of the mercury is pushed out of the tube, and the change in the height of the mercury column is the vapor pressure.

We can use a barometer to measure vapor pressure.

Liquids with high vapor pressure are called volatile; they evaporate quickly. Substances like water have strong intermolecular forces that result in low vapor pressure.

Less volatile Lower vapor pressure

More volatile Higher vapor pressure

14.5 Boiling Point and Vapor Pressure

Boiling...what causes the bubbles to form at high temperatures?
- starts out as a tiny air bubble in the water
- bubble expands when high energy water molecule enters the bubble
• if the water molecule has enough kinetic energy, it can produce a pressure inside the bubble that is greater than the atmospheric pressure and push back the water surrounding the bubble
• so water boils at 100 °C because vapor pressure is 1 atm at 100 °C

What effect do you think elevation has on boiling point? Remember, atmospheric pressure changes with elevation.

**Boiling point decreases at higher elevations because the there is less atmospheric pressure to overcome.**

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Focus Q’s #1-3 pg.453

1. Perspiration is evaporation on your skin. This is an endothermic process; energy comes from your skin.

2. Methyl alcohol; high vapor pressure than water (weaker IMF). This means that more molecules leave surface and more evaporation means it feels cooler.

   Methyl alcohol
   \[
   \text{H} \quad \text{C} \quad \text{O} \quad \text{H}
   \]

   Water
   \[
   \text{H} \quad \text{O} \quad \text{H}
   \]

3. Evaporation happens below boiling point and occurs only at surface of liquid. Boiling happens throughout liquid and is much faster; requires energy.
53. Diethyl ether has higher vapor pressure because it doesn't exhibit hydrogen bonding like 1-butanol.

\[
\begin{align*}
\text{1-butanol} & : \quad \begin{array}{c}
\text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{O} & \text{C} & \text{H} \\
\end{array} \\
\text{Diethyl ether} & : \quad \begin{array}{c}
\text{H} & \text{H} & \text{H} \\
\text{H} & \text{C} & \text{C} & \text{C} & \text{C} & \text{O} & \text{H} \\
\text{H} & \text{H} & \text{H} \\
\end{array}
\end{align*}
\]

54. Higher boiling point means stronger IMF:
- a) \( \text{Ga} (\text{KBr}) \text{O}_2 \) (metallic ionic London)
- b) \( \text{Hg} (\text{NaCl}) \text{H}_2\text{O} \) (metallic ionic H-bond)
- c) \( \text{H}_2, \text{O}_2, \text{H}_2\text{O} \) (London London H-bond)

57. Dipole-dipole attractions are only 1% as strong as covalent bonds.

58. London forces arise from instantaneous dipoles; much weaker than dipole-dipole or covalent.

60. Volatile liquids evaporate easily. They have high vapor pressure due to weak IMF (usually London or dipole-dipole).

61. \( \text{NH}_3 \) can do hydrogen bond and \( \text{CH}_4 \) can only do London forces; therefore \( \text{NH}_3 \) has a much higher BP due to stronger IMF.

Relative Strength of Intermolecular Forces

\[
\begin{align*}
\text{London} & < \text{Dipole-Dipole} & < \text{Hydrogen Bond} \\
\text{Dispersion} & < \text{Attraction} & < \text{Bond}
\end{align*}
\]

How to determine predominant intermolecular forces between molecules (and noble gases):

- Is the first element a metal?
  - If yes, there will be ionic bonding
  - ex) \( \text{NaCl}, \text{KBr}, \text{Mg(NO}_3)_2 \)
  - If no, does the molecule have a dipole moment? (it will be made up of different elements)
    - If yes, is there a H bonded directly to F, O, or N in molecule?
      - If yes, Hydrogen bonding
      - If no, Dipole-dipole attraction
    - If no, there will be London dispersion force
      - ex) \( \text{Ar}, \text{H}_2, \text{N}_2, \text{Kr} \)
14.6 The Solid State: Types of Solids

crystalline solids
– regular arrangement of their components
– produces beautiful, regularly shaped crystal

three types:

1. ionic solids
   – components are ions
   – dissolves in water; conducts electricity
   – high melting point
   – example: NaCl

2. molecular solids
   – components are neutral molecules
   – dissolves in water; do NOT conduct electricity
   – low melting point due to weak intermolecular forces (dipole-dipole interaction of London dispersion forces)
   – example: glucose (sugar) and ice

3. atomic solids
   – components are atoms of one element covalently bonded to each other
   – properties vary because of the different ways in which the atoms can interact with each other
   – example: graphite and diamond (both are made of carbon)
Properties of solids are determined primarily by the nature of the forces that hold the solid together.

Bonding in Metals: electron sea model
- metal atoms are in a “sea” of valence electrons that are shared among the atoms in a nondirectional way and are quite mobile in the metal crystal

- mobile electrons can conduct electricity and heat
  - atoms can be moved rather easily (ex: metal can be hammered into sheets or pulled into wires)

Alloys
- a substance that contains a mixture of elements and has metallic properties

- Two types:
  1. substitutional alloy: some of the host metal atoms are replaced by other metal atoms of similar sizes; examples are brass in which 1/3 of copper host atoms are replaced by zinc atoms and sterling silver (93% silver and 7% copper)
2. **interstitial alloy**: some of the interstices (holes) among the closely packed metal atoms are occupied by atoms much **smaller** than the host atoms; example is steel (iron and carbon)

![Steel and Carbon Image]

Focus Questions #1-3 pg. 461

1. Salt is an ionic solid and has a high melting point due to its strong chemical bonding. Sugar is a molecular solid and has a lower melting point due to its much weaker intermolecular forces.

2. Molecular solids - intermolecular forces
   - Ionic solids - chemical bonds
   - Atomic solids - chemical bonds

   *** IMF are weaker than intramolecular forces so molecular solids would have the lowest boiling point.***

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3. Atomic solids have directed chemical bonds and metals have nondirectional bonding called “electron sea model”

Ch Rev Qs#31-42 pp.464-465

31. At higher altitudes, there is less atmospheric pressure so boiling point decreases. So, food must be cooked longer.

32. Boiling point is when vapor pressure is equal to atmospheric pressure.

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33. Crystalline solids have regular, repeating arrangement of particles; usually reflects the shape of the crystal.

34. Ionic Solid - NaCl (metal-nonmetal)
   Molecular Solid - H₂O (nonmetals only)
   Atomic Solid - C in diamonds

35. Ionic solids made up of ions (+|−) and molecular solids are made up of molecules.

36. Ionic solids have higher MP, BP, surface tension and lower vapor pressure as compared to molecular solids
37. Ionic bonds are much stronger than IMF that make up molecular solids so ionic solids are usually harder, have higher melting, etc. It takes much more energy to break a chemical bond than and IMF.

40. Kr- London dispersion forces (weakest IMF)
   C- strong chemical bond

41. An alloy is a mixture of metal atoms. Substitutional alloys contain host metal atoms that have been replaced with another metal. Interstitial alloys contain smaller atoms that fill the space between larger atom.

42. Alloying makes metals stronger because irregularities keep crystal from deforming.