Experiment 22
Intermolecular Forces and Solubility

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Name:

Date:

Key Objectives

1. What is responsible for each of the four Intermolecular Forces (IMF's)
2. Assign Intermolecular Forces to organic molecules.
3. Understand the relationship between Intermolecular Forces and physical properties (specifically solubility in this lab)
4. The relationship between Functional Groups and Intermolecular Forces
5. Why size matters

Discussion

Review: Intramolecular Forces (Chemical Bonds) vs Intermolecular Forces

Chemical bonds are the attractive forces holding atoms together to form molecules. This include forming ionic compounds (ionic bonds - gain and lose electrons to complete octets) or molecular compounds (covalent bonds - share electrons to complete octets). Chemical bonds are responsible for chemical properties, mainly the chemical reactions we study. Chemical bonds are much stronger than intermolecular forces.When drawing molecules we show bonds as solid lines.

Intermolecular forces (IMF's) are the attractive forces between molecules. They are responsible for many of the physical properties of substances. Physical properties include boiling points, melting points, and solubility. When drawing molecules we show IMF's as dashed lines. There are four important IMF we will consider in this experiment.

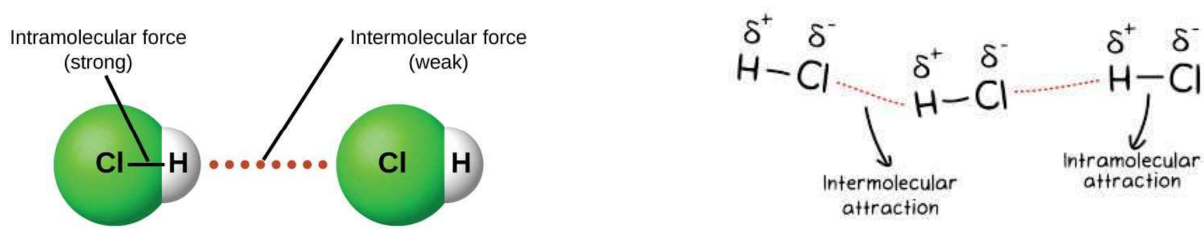


Figure 22.1: Intramolecular (Chemical Bonds) vs Intermolecular (Between molecules) credit: top - OpenStaxCollege. License: CC BY 4.0 bottom - https://cdn.kastatic.org/ka-perseus-images/2596bfb5d1f501a78d6693d0a9bd79c17631ffb0.png

Types of Intermolecular Forces (IMF)

For a detailed discussion of Intermolecular forces review your notes from last semester, look at the linked video on (www.chemhaven.org/che102) or borrow a textbook (Tro Section 11.2).

London Dispersion Forces (LDF)

London Dispersion Forces (LDF) occur between non-polar molecules due to the formation of instantaneous dipoles. The dipoles are temporary (very) lasting a fraction of second, therefore the attraction between molecules is very weak. They are the weakest IMF but are proportional to the size (or molecular weight) of the molecule so in large molecules can outweigh the effects of stronger IMF's in smaller molecules. Non-polar molecules are often symmetrical.

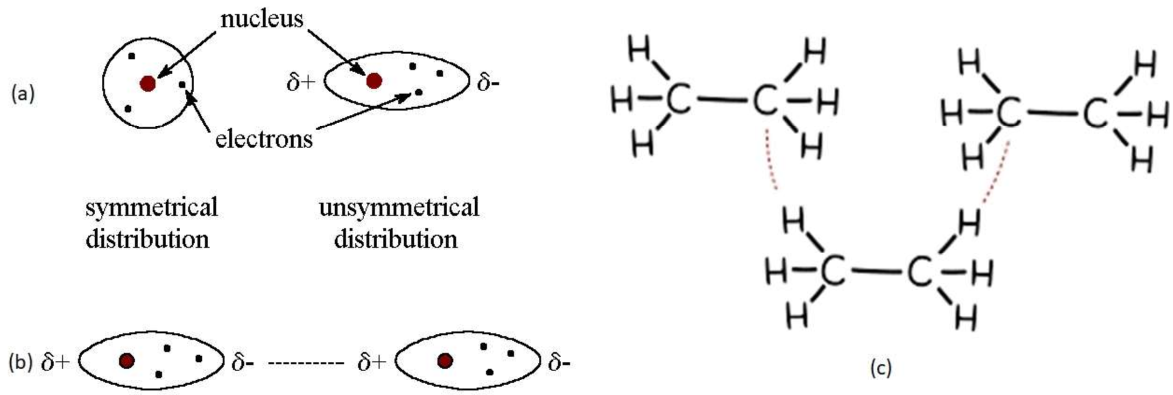


Figure 22.2: (a) Top - Symmetric and asymmetric distribution of electrons (b) Bottom - Temporary dipoles can be attracted to each other and are called London Dispersion Forces (LDF). (c) Right - For larger molecules that are symmetrical the attraction is generally shown as dotted lines between the molecules. credit: (a) and (b) - `https://www.chem.purdue.edu/gchelp/liquids/disperse.html` (c) `https://www.khanacademy.org/science/class-11-chemistry-india/xfbb6cb8fc2bd00c8:in-in-states-of-matter/xfbb6cb8fc2bd00c8:in-in-intermolecular-forces/a/intramolecular-and-intermolecular-forces`

Dipole-Dipole (DD)

Dipole-Dipole forces occur between polar molecules due to the asymmetric distribution of charge producing a partial charge (δ^{+}/δ^{-}), and electrostatic attraction between the molecules. Molecules that are not symmetrical often form DD interactions. The attractive force is stronger than LDF, but weaker than Hydrogen Bonds.

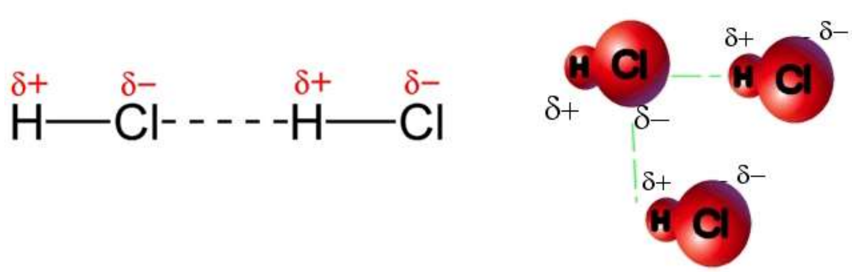


Figure 22.3: Dipole-Dipole (DD) interaction between two hydrochloric acid molecules. credit: left - `https://courses.lumenlearning.com/introchem/chapter/dipole-dipole-force/` right - `https://wyomingapchemistry.weebly.com/101-intermolecular-forces.html`

Hydrogen Bonds (HB)

A special case of Dipole-Dipole attraction in which a Hydrogen bonded to Oxygen, Nitrogen or Fluorine is attracted to the lone pair electrons on or a partial negative charge on a molecule. The asymmetric distribution of charge when H is bonded to O,N, or F is substantially larger than a normal dipole due to Hydrogen only having one electron, resulting in an attractive force much larger than the typical DD interaction.

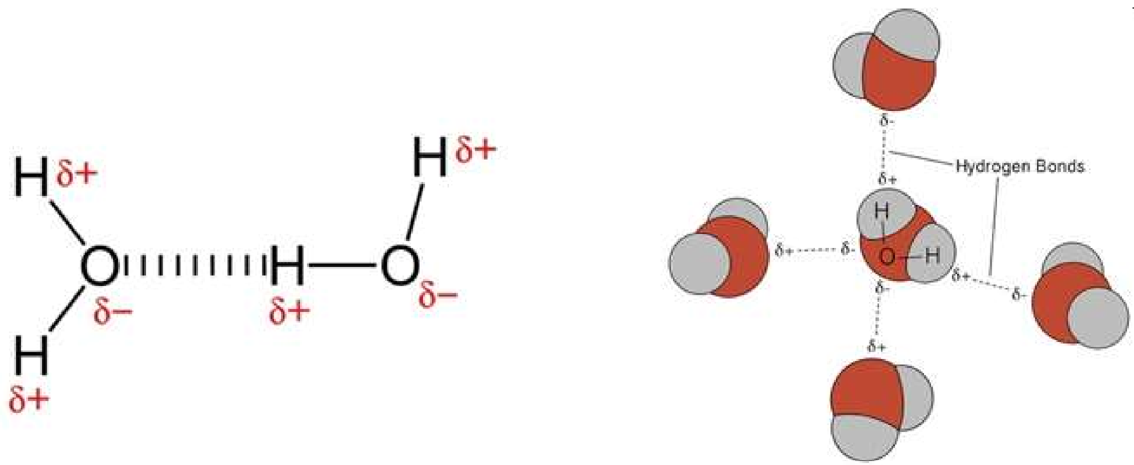


Figure 22.4: Hydrogen Bond (HB) between two water molecules. credit: `https://courses.lumenlearning.com/cheminter/chapter/hydrogen-bonding/`

Ion-Dipole(ID)

The electrostatic attraction between a molecule with a partial charge and an ion with a full charge. Not relevant in this experiment as we are dealing with pure solutions.

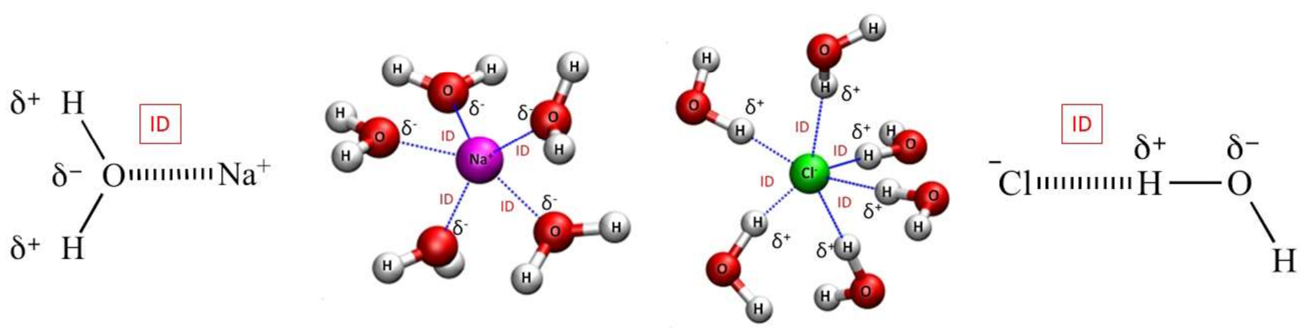


Figure 22.5: Ion Dipole interaction (ID) a cation (Na^+) and an anion (Cl^-) and water molecules. credit: author

IMF's and Solubility

IMF's are responsible for the attractive forces between molecules. They are important because they determine many of the physical properties we will explore for each functional group in organic chemistry. In this lab we will focus on the Solubility of various functional groups and molecules.

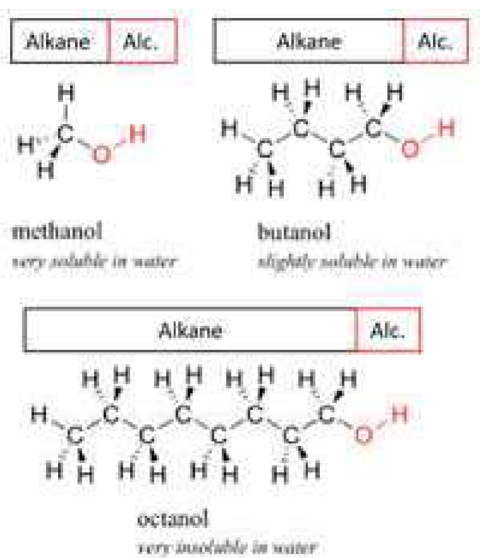
For review read about solubility in Hein Chapter 14.2-14.3. Recall that a **Solution** is formed when a **Solute** is dissolved in a **Solvent.** The solute is generally the compound that is smaller in quantity while the solvent is the compound that is larger in quantity. If the solutions tested are liquids we generally use the terms **Miscible** to mean that two solutions mix to form a **homogeneous** solution, and **Immiscible** to describe two solutions that do not mix (are insoluble, visible phase boundary).

The general rule we learned first semester is that "like dissolves like" or its opposite "oil and water don't mix". We learned that what had to be "like" was the Intermolecular Forces (IMF's) between the molecules.

We can expand on that by realizing that Polar Substances (DD/HB) can generally dissolve in other Polar substances while Non-Polar Substances (LDF) can generally only dissolve in other Non-Polar Substances. A second rule is that generally a substance will only dissolve in a group higher or lower than it on the chart to the right.

Ranking of IMF's
Ionic or ID
HB
DD
LDF

For molecules that have more than one functional group solubility can depend on the size of the molecule. For example methanol is infinitely miscible with water, but as the size of the Alkane part of an alcohol increases the solubility decreases until 1-butanol has a solubility of 9 g alcohol/100 g H_2O and 1- octanol is only 0.054 g alcohol/100 g H_2O and after 1-decanol alcohols are no longer soluable in water. Figure 22.6 shows the structure of each of the molecules



Formula	Name	Solubility in Water (g/100 g)
CH_3OH	methanol	infinitely soluble
CH_3CH_2OH	ethanol	infinitely soluble
$CH_3(CH_2)_2OH$	propanol	infinitely soluble
$CH_3(CH_2)_3OH$	butanol	9
$CH_3(CH_2)_4OH$	pentanol	2.7
$CH_3(CH_2)_5OH$	hexanol	0.6
$CH_3(CH_2)_6OH$	heptanol	0.18
$CH_3(CH_2)_7OH$	octanol	0.054
$CH_3(CH_2)_8OH$	decanol	insoluble in water

Figure 22.6: As the size of the Alkane portion of an alcohol increases, the solubility decreases. credit: author

Procedure

In this experiment you will be measuring the Solubility of several organic substances in both water and hexane to see the effects of different Functional Groups (IMF's) and the effect of size on solubility.

Part I - Solubility Tests in Water

1. Test the solubility of each of the listed substance with water by adding 1 mL (or 0.5g for solids)of the first organic compound to 5 mL water in a test tube.

2. Mix each of the test tubes vigorously for 15 seconds. Wait 30 seconds. Sketch a picture of your results. In the sketch, label each liquid in the test tube. Note which pairs are miscible and which are not.

3. For any solutions that are insoluble note the relative density of the hydrocarbons with respect to water.

4. Dispose of the solutions in the waste bottle labeled "Volatile Waste" .

Part 2 - Solubility Tests in Hexane

1. Test the solubility of each of the listed substance with hexane by adding 1 mL (or 0.5g for solids)of the first organic compound to 3 mL hexane (or second hydrocarbon) in a test tube.

2. Mix each of the test tubes vigorously for 15 seconds. Wait 30 seconds. Sketch a picture of your results. In the sketch, label each liquid in the test tube. Note which pairs are miscible and which are not.

3. For any solutions that are insoluble note the relative density of the hydrocarbons with respect to water.

4. Dispose of the solutions in the waste bottle labeled "Volatile Waste" .

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Name: ____________________ Class: ________________ Date: ________________

Results

Mixture	**Solubility in Water**		**Solubility in Hexane**	
	Observation/Sketch	**M or I**	**Observation/Sketch**	**M or I**
Hexane				
1-Hexanol				
Ethoxy Ethane				
Butanal				
3-Pentanone (Acetone)				
Benzoic Acid (Propanoic Acid)				
Ethyl Butanoate				
Diethyl amine				

1. List any of the results in Table 1 that break the rules established in the discussion section of the lab? Explain

2. For each of compounds miscible in water from Table 1 sketch a picture of the interaction between the organic compound tested and water. Label the IMF's present.

Name: ______________________ Class: ________________ Date: ________________

Prelab Questions

1. Complete the table below illustrating the differences between Chemical Bonds and Intermolecular Forces (IMF's).

Property	Chemical Bonds	Intermolecular Forces
Attraction between:		
Relative Strength:		
Represented by (in LS):		
Determine Properties like:		

2. Define (using sentences) each of the following Intermolecular Forces (IMF's) including their relative strengths. In addition draw an example illustrating the attractive force between **TWO** molecules. Properly label all charges (+/-) and partial charges (δ^{+}/δ^{-}).

(a) London Dispersion Forces (LDF)

(b) Dipole-Dipole Forces (DD)

(c) Hydrogen Bonding (HB)

(d) Ion-Dipole (ID)

3. Draw the generic structure (Lewis or Line) of each of the compounds used in lab **AND** indicate which IMF are present. Hint: Table 19.1 (or Lewis Structure Lab)

Hexane	Alcohol
Ether	Aldehyde
Ketone	Carboxylic Acid
Ether	Amide
Amine	