

## Worksheet 15 - Intermolecular Forces

Chemical bonds are **intramolecular** forces which hold *atoms* together as molecules. The forces that hold *molecules* together in the liquid and solid states are called **intermolecular** forces.

Intermolecular forces (IMF) can be qualitatively ranked using Coulomb's Law:

$$\text{force} \propto \frac{Q_1 Q_2}{r^2}$$

where  $Q_1$  and  $Q_2$  are charges and  $r$  is the distance between them.

The obvious consequences of this relationship are that:

- the bigger the charges, the stronger the interaction
- the closer the charges, the stronger the interaction

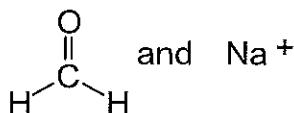
The charges in chemical species can be classified as:

- ionic** - full positive and negative charges
- dipole** - partial positive or negative charges, resulting in polarized bonds in which there is unequal sharing of valence electrons. This arises when electronegative elements (O, N, halogens) are bonded to less electronegative species. **Molecular geometry** also plays a role in determining the net **dipole moment** of a molecule.
- London Dispersion Forces (LDF)** - instantaneous dipole moments resulting from the motion of electrons in an atom or molecule. All atoms and molecules possess LDF, which increase with the number of electrons in the molecule, and its shape. These are very small, very short-lived charges.

1. List **all** types of IMF which would be exhibited between the pairs of species shown below?

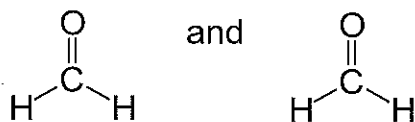
a)  $\text{O}=\text{C}=\text{O}$  and  $\text{O}=\text{C}=\text{O}$                       LDF only (nonpolar)

b)    LDF and ion-dipole



c)

LDF and dipole-dipole



d)

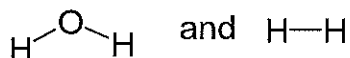
LDF and ion-dipole



There is one type of **dipole-dipole** interaction which is especially strong. It is called a **hydrogen bond** (H-bond). The partial negative charge ( $\delta^-$ ) comes from one of the electronegative elements **F**, **O** or **N**. These are called H-bond **acceptors**. The partial positive charge ( $\delta^+$ ) comes from an **H** bonded **directly** to one of these elements (**O-H**, **N-H**, **F-H**). These are called H-bond **donors**. Both a donor and an acceptor are needed to form an H-bond. These are stronger IMF than dipole-dipole interactions, but not as strong as ion-dipole interactions.

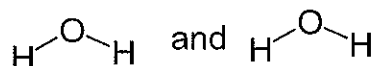
2. Which of the following pairs of compounds can form H-bonds? For those that can, mark the position of the partial positive ( $\delta^+$ ) and negative ( $\delta^-$ ) charges in the molecules and indicate where the H-bonds will form. For those that can't form H-bonds, describe the strongest IMF available to that pair of compounds.

a)



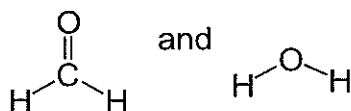
$\text{H}_2$  has no H-bond acceptor;  
LDF is highest force present

b)



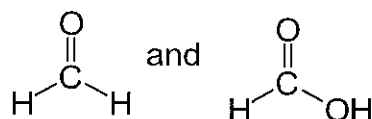
oxygen =  $\delta^-$ ; hydrogens =  $\delta^+$ ;  
H-bond forms between O and H on  
neighboring molecules

c)



oxygen =  $\delta^-$ ; hydrogens =  $\delta^+$ ;  
H from  $\text{H}_2\text{O}$  (directly bonded to O)  
donates to O in the other molecule

d)



oxygen =  $\delta^-$ ; hydrogens =  $\delta^+$ ;  
H from  $\text{CH}_2\text{O}_2$  (directly bonded to O)  
donates to O in the other molecule

3. What is the nature of the major attractive IMF in the following molecules?  
Please draw a Lewis structure and determine the shape of each molecule.

- |   |  |
|---|--|
| a) $I_2$<br>linear; nonpolar;<br>LDF  | d) HF<br>linear; polar; contains H-F bond;<br>hydrogen-bonding |
| b) MgO<br>$Mg^{2+}$ and $O^{2-}$ ions<br>ion-ion  | e) HCl<br>linear; polar;<br>dipole-dipole                      |
| c) $CH_3CN$<br>C bonded to all H and the other C; 2 <sup>nd</sup> C is triple-bonded to the nitrogen; polar molecule; dipole-dipole |  |

It requires energy to pull molecules apart in going from the solid to liquid to gaseous states. The stronger the IMF, the more energy must be added, the higher the phase transition temperature will be.

4. Which of the following substances is most likely to exist as a crystalline solid at room temperature? Of the gases, which would be the hardest to condense to a liquid under pressure?

Rank these substances from lowest to highest intermolecular forces

- |                                    |                                   |
|------------------------------------|-----------------------------------|
| a) HF<br>hydrogen bonding          | d) $SO_2$<br>polar; dipole-dipole |
| b) $PCl_3$<br>polar; dipole-dipole | e) $F_2$<br>nonpolar; LDF         |
| c) $FeCl_2$<br>ionic compound      |                                   |

$F_2 < PCl_3/SO_2 < HF < FeCl_2$ ;  $FeCl_2$  has the highest melting point and is most likely to exist as a solid at room temperature.  $F_2$  has the lowest boiling point and would be the last to condense to a liquid

5. Which of the following is expected to have the highest boiling point?  
Which is expected to have the lowest?

Rank these substances from lowest to highest intermolecular forces

- |                            |                             |                            |
|----------------------------|-----------------------------|----------------------------|
| a) $CO_2$<br>nonpolar; LDF | b) Ar<br>nonpolar; LDF      | c) $CF_4$<br>nonpolar; LDF |
| d) LiCl<br>ionic           | e) $SiF_4$<br>nonpolar; LDF |                            |

LiCl is ionic and is expected to have the highest boiling point. The other substances are all limited to LDF, so they're ranked based on number of electrons ( $CO_2 = 22$ ,  $Ar = 18$ ,  $CF_4 = 42$ ,  $SiF_4 = 50$ ). Ar has the fewest electrons, so it has the weakest LDF and therefore the lowest boiling point.

In general, be careful when comparing molecules with very different structures and/or molecular weights. For example,  $\text{H}_2\text{O}$  is a liquid at room temperature and  $\text{I}_2$  is a solid. This doesn't mean that H-bonding, in general, is weaker than LDF. It's just that the structures are too different to compare.

6. Give an explanation in terms of IMF for the following differences in boiling point.

a)  $\text{HF}$  ( $20^\circ\text{C}$ ) and  $\text{HCl}$  ( $-85^\circ\text{C}$ )

Both are polar.  $\text{HF}$  has hydrogen bonding, which is a stronger force, so it will have the higher boiling point.

b)  $\text{CHCl}_3$  ( $61^\circ\text{C}$ ) and  $\text{CHBr}_3$  ( $150^\circ\text{C}$ )

Both are polar and would have dipole-dipole interactions. In this case, we can compare LDF instead.  $\text{CHBr}_3$  has more electrons, so it has the higher boiling point.

c)  $\text{Br}_2$  ( $59^\circ\text{C}$ ) and  $\text{ICl}$  ( $97^\circ\text{C}$ )

$\text{Br}_2$  is nonpolar and has only LDF.  $\text{ICl}$  is polar and has dipole-dipole interactions, so it will have the higher boiling point.

7. Indicate which of the following properties will increase, decrease or remain unaffected by an increase in the strength of the intermolecular forces?

a) vapor pressure (pressure of gas above a liquid sample in a closed container)

**decreases with increased intermolecular forces**

b) normal boiling point (boiling point at 1 atmosphere pressure)

**increases with increased intermolecular forces**

c) heat of vaporization (heat requires to take a liquid sample to the gaseous phase)

**increases with increased intermolecular forces**

d) surface tension (adhesion of molecules)

**increases with increased intermolecular forces**