# ATTRACTIVE FORCES

## INTRAMOLECULAR FORCES

### 1. IONIC Bonds

A ionic solid is held together by electrostatic attractions between ions that are next to one another in a lattice structure. They are strong, and substances held together by ionic bonds have high mp and bp.

Coulomb's Law- the more highly charged the ions in an ionic bond, the stronger the bond. The smaller the ions in an ionic bond, the stronger the bond. (Why?- Because a small ionic radius allows the charges to get closer together and increases the force between them)

In an ionic solid, each e- is LOCALIZED around an atom, so e- do not move around the lattice, this makes ionic solids poor conductors of electricity. Ionic liquids to conduct electricity because ions move about freely in the liquid phase.

*2. COVALENT (NETWORK) Bonds*

In a network solid, atoms are held together in a lattice of covalent bonds. Network solids are very hard and have very high mp and bp. The e- in a network solid are localized in covalent bonds between particular atoms, so they are not FREE to move about the lattice. This makes network solids poor conductors of electricity.

*3. METALLIC BONDS*

Metallic substances can be compared with a group of nuclei surrounded by a sea of mobile e-. They too (along with ionic and network substances) can be visualized as 1 large molecule. Most metals have high mp and bp.

Metals composed of atoms with smaller nuclei tend to form stronger bonds than metals made up of atoms with larger nuclei. This is because smaller sized nuclei allow the positively charged nuclei to be closer to the negatively charged e-, increasing the attractive force from Coulomb's law.

The e- in a metallic a substance are delocalized and can move freely throughout the substance. The freedom of the e- makes it a Great conductor of heat and electricity.

## *INTERMOLECULAR FORCES (IMFS)*

1. London Dispersion Forces: These occur between neutral, nonpolar molecules. These very weak attractions occur because of random motions of e- on atoms within molecules. At a given moment, a nonpolar molecule might have more e- on one side than on the other, giving it an instantaneous polarity. For that fleeting instant, the molecule will act as a very weak dipole.

Since London dispersion forces depend on the random motions of e-, molecules with more e- will experience greater forces, the one with more e- will generally have a higher melting and boiling pts. London dispersion forces are even weaker than dipole-dipole forces, so substances that experience only London dispersion forces melt and boil at extremely low temps and tend to be gases at room temp.

2. Dipole-dipole forces occur between neutral, polar molecules: The + end of 1 polar molecule is attracted to the – end of another polar molecule.

Molecules with greater polarity will have greater dipole-dipole attraction, so molecules with larger dipole moments tend to have higher melting and boiling pts. Dipole-dipole attractions are weak, however, and these substances melt and boil at very low temperatures. Most substances held together by dipole-dipole attraction are gases or liquids at room temperature.

3. HYDROGEN "BONDS"

They are much stronger than dipole-dipole forces because when a hydrogen atom gives up its lone electron to a bond, its positively charged nucleus is left virtually unshielded. In a hydrogen bond, the positively charged hydrogen end of the molecule is attracted to the negatively charged end of another molecule containing an extremely electronegative element.

Substances that have H bonds (like water and ammonia), have higher mp and bp than substances that are held together by dipole-dipole forces.

**Water is less dense as a solid than as a liquid because its hydrogen bonds force the molecules in ice to form a crystal structure, which keeps them farther apart than they are in liquid form.

4. Ion-Dipole - A force **between** an **ion** and a polar molecule.

An ion-dipole force is an attractive force that results from the electrostatic attraction between an ion and a neutral molecule that has a dipole.

***Most commonly found in solutions. Especially important for solutions of ionic compounds in polar liquids