Jan Wyszynski

Stream 11-1

Notes on Covalent Bonding

4.2.1 & 4.2.2: Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei & describe how a covalent bond is formed by sharing electrons

 When the atoms of two non-metals react, they form covalent bonds by sharing electrons in such a way that both atoms receive a stable octet. The pairs of electrons shared between two atoms are attracted to both nuclei. Atoms that are held together via a covalent bond are called molecules. Some pairs in a molecule may not be bonded.

4.2.3: Deduces the Lewis dot structure of molecules and ions up to four electron pairs on each atom

 When a molecule has one bond, if it attempts a stable octet, the atom will have 3 pairs of lone pairs. When a molecule has two bonds, if it attempts a stable octet, the atom will have 2 pairs of lone electrons. When a molecule has three bonds, if it attempts a stable octet, the atom will have 1 pair of lone electrons.

4.2.4: State and explain the relationship between number of bonds, and the number of bonds and bond length

 More bonds = Stronger bond = Shorter Bond

 Less Bonds = Weaker Bond = Longer Bond (Note that multiple bonds are treated a single negatively charged centers)

 Bond Strength:

 Triple Bond > Double Bond > Single Bond

 Bond Length:

 Single Bond > Double Bond > Triple Bond

4.2.5: Predict whether a compound of two elements would be covalent from the position of the elements in the Periodic Table or from their electronegativity values

 If the electronegativity difference is between 0 and 0.5 (inclusive), the bond is considered pure covalent.

 If the electronegativity difference is between 0.5 and 1.8 (exclusive), the bond is considered polar covalent. This means that one of the two atoms participating in a bond strongly attracts the shared electrons and claims a partially negative charge, while the other atom takes a partial positive charge. The higher the difference (within the interval) the more polar the bond, and the more ionic it is in character.

 With regards to the periodic table, bonds are covalent if the two bonding atoms are nonmetals. With regards to polarity, the more distant the bonded atom on the periodic table, the more polar they are.

4.2.6: Predict the relative polarity of bonds from electronegativity values:

 When a polar covalent bond is formed, the higher the difference in electronegativity, the more polar it is.

 When a polar covalent bond with high difference in electronegativity is formed, it is called a **dipole**.

4.2.7: Predict the shape and bond angles for species with four, three and two negative charge centers on the central atom using the VSEPR theory.

 Species with four charged centers form a tetrahedral shaped molecule where the bond angle is 109.5 degrees.

 Species with three charge centers form a trigonal planar (SAT name) or a planar triangular three dimensional form and a bond angle of 120 degrees.

 Species with two charge center form a linear arrangement.

 It is also important to note that with molecules like SO2 where Sulfur has a lone pair of electrons, the lone pairs of electrons distort the angle of the molecule, but do not count as part of the final structure. One popular molecule is water, where two lone pairs of electrons shrink the bond angle from its theoretical 180 degree arrangement to its bent arrangement of 105 bond angle.

 (Also note that multiple bonds are treated as one bond)

4.2.8: Predict whether or not a molecule is polar from its molecular shape and bond polarities

 The polarity of a molecule depends on

 -The polar bonds that it contains

 -The way in which the polar bonds are orientated with respect to each other.

If bonds are of equal polarity and are arranged symmetrically, their dipoles will oppose each other and thus cancel each other out (molecule is not polar).

The direction of the pull is also important, recall that electrons are drawn towards the most electronegative atom thus while BF3 is non-polar because all Fluorine atoms attract atoms away symmetrically from Boron, while NH3 is polar because it has a net dipole pointing towards nitrogen because nitrogen attracts electrons form all three hydrogen’s.

4.2.9: Describe the structures of all three allotropes of carbon (diamond, graphite, and Fullerene):

 The following are three allotropes of carbon each of which is a giant covalent compound (macro-covalent structure)

In **diamond**, each carbon is **covalently bonded to four other carbons** in a **tetrahedral** structure where the bond angle is **109.5 degrees**. The density of carbon is 3.51 g cm-3 and it is the hardest natural substance. **All electrons are bonded** and it **doesn’t conduct electricity**.

 - Lustrous Crystal

In graphite, each carbon is covalently **bonded to three other carbons** and forms parallel arrangements with **120 degree bond angles**. The Carbon’s form hexagons and are held together by very weak van der Waal’s forces. The density is 2.26 g cm-3. Graphite contains **one non-bonded delocalized electron per atom**. It is a conductor of electricity.

 - Non-lustrous grey solid

Each carbon atom is covalently bonded to three other carbons in a sphere of 60 carbon’s that includes 12 pentagons and 20 hexagons. Fullerene has the lowest density at 1.72 g cm-3 and **easily accepts electrons from negative ions.** Fullerene **reacts with K to make superconducting crystalline material**

 - Yellow crystalline solid, soluble in benzene

4.2.10: Describe the structure of silicon and silicon dioxide:

In an elemental state, each silicon atom is covalently bonded to **four other silicon atoms** in a **tetrahedral** arrangement.

Silicon Dioxide forms a tetrahedral giant covalent structure but here **each silicon is bonded to four oxygen atoms and** each **oxygen atom** is bonded to **two other silicon atoms**. Note that the molecular formula of Silicon Dioxide merely shows the ratio of atoms within the giant molecule while the actual number is a large multiple of it.

 SiO2 is insoluble in water and has a high melting point, and does not conduct electricity of heat. Glass and sand are different forms of silica.

4.3.1: Describe the types of intermolecular forces (attractions between molecules that have temporary dipoles or hydrogen bonding) and explain how they arise from the structural features of molecules.

The diatomic molecule of chlorine has no permanent separation of charges and both chlorine atoms attract electrons equally therefore the molecule is non-polar.

However it is important to note that electron positions are not constant, rather, they function as mobile clouds of negative charge, thereby having the capacity to create temporary polarity and form a **weak and temporary dipole**. When this temporary electron distribution effects the positioning of neighboring molecules, their temporary dipoles are called **induced dipoles**. (The attraction involved in creating this induced dipole is an intermolecular force whose strength increases with increasing numbers of electrons) 🡪 Therefore with increase molecular mass, the probability of this attraction increases.

Substances that are held together by intermolecular forces (whose strength increases with increasing molecular mass) have **generally low melting and boiling points** because there isn’t much energy required to separate the molecules.

Dipole-Dipole Attraction

Molecules with permanent dipoles such as HCl form dipole-dipole attractions. This is the intermolecular attraction between the partial positive charge of one molecule and the partial negative charge of a neighboring one (or vice versa).

 The strength of the intermolecular force depends on the degree of polarity of the bond. Therefore the strength of intermolecular bonding in the following:

 HCl > HBr > HI

 The melting and boiling points of polar molecules with dipole-dipole attractions are higher than those of non-polar substances with similar molecular mass. (Non-polar substances have van der Waal’s forces; the ones that create induced dipoles)

 Hydrogen Bonding

With strongly polar molecules that include hydrogen, the intermolecular forces of attraction are called hydrogen bonds. It is an exceptionally strong dipole-dipole attraction because the highly electronegative anion draw hydrogen electron giving it properties similar to that of a proton. The hydrogen attracts the neighboring anions very strongly.

 Hydrogen bonding only occurs in molecules containing hydrogen bonded to **Oxygen, Flourine, and Nitrogen**.

 Hydrogen bonding is particularly strong and thereby increase the boiling and melting points far above what would be expected from their molecular mass.

 When observing the trends in hydride of elements in groups 4-7, the boiling points increase down the group as the molecular mass increases. The only anomalies are **NH3, HF, and H2O** which have higher boiling points that would be expected.

 When comparing organic molecules, we see higher boiling points from molecules that include hydrogen bonding.

 Each water molecule can form up to four hydrogen bonds because it has 2 hydrogen’s and two lone pairs (liquid water contains the fewer, but in solid state each water molecule is maximally with the four hydrogen bonds).

 The result of the strong bonding in the solid is a tetrahedral arrangement that holds the molecules in a fixed distance from each other. The remarkable thing here is that because of this fixed distance, solid water is actually **less dense** that liquid water. Evidence for this is that solid water (ice) floats in liquid water. This means that the water expands upon freezing.

 For answering questions in which you must determine the molecules in increasing boiling point first determine the molecular mass of all molecules. If they are similar, determine the type of bonding that will occur:

 Non-polar molecules 🡪 van de Waal’s (weakest intermolecular forces)

 Terminal Polar Bond Exists 🡪 dipole-dipole force

 If there exists an HO, HF, or HN 🡪 hydrogen bond

Physical Properties

 4.5.1: Compare and explain the properties of substances resulting from different types of bonding

Ionic Substances:

 - High Melting Points and High Boiling Points (forces of electrostatic attraction are strong and ions are in a lattice)

 - Thus ionic compounds are solids at room temperature.

 - Macromolecular or Macro-covalent structures have high melting points and high boiling points as the covalent bonds must be broken.

 - Covalent substances have lower melting points and lower boiling points that ionic compounds as the forces that are needed to desperate the molecules are relatively weak due to weak intermolecular forces.

 - Thus many covalent solids are liquids or gases at room temperature.

 - Intermolecular forces vary the melting and boiling points:

 - increasing molecular mass

 - the extent of polarity within the bonds of molecules

Solubility: The ease with which the solid because dispersed through a liquid and disperse through liquid. Ionic solids disperse in water readily because the **partial charges of water molecules** are attracted to the **oppositely charged ions in the lattice structure**. When these ions are surrounded with water they are said to be hydrated.

 Polar compounds are soluble in water because they partial charges are attracted to the opposite partial charges of water molecules. Sugar and Ethanol for example form hydrogen bonds with the water.

Non-polar substances are not soluble in water because they have no partial charge.

The rules are reversed when we consider an organic non-polar organic solvent like hexane.

 -Ionic compounds will not dissolve because there is not attraction for any ions in their lattices structures.

 -Polar compounds will not dissolve because their partial charges have no interaction.

 -Non-polar compounds will dissolve because they can interact with the non-polar solvent through van der Waals forces (recall that these occur when electrons create temporary dipoles).

The general rule is *Like Dissolves Like*

 Ionic Compounds are soluble in Polar solvents but not soluble in Non polar solvents

 Polar Compounds are soluble in polar solvents but not in non-polar solvents (solubility in polar solvents increases with increasing polarity, solubility in non-polar solvents increases with decreasing polarity)

 Non Polar Compounds are not soluble in Polar Solvents but are in non-polar solvent

 Electrical Conductivity: Depends on whether it has mobile ions that can carry a charge.

 - Ionic Compounds do not conduct electricity in solid state, but can conduct in an aqueous state.

 - All covalent compounds whether polar or non-polar do not conduct electricity in any state.