1. The chemical bond - attractive force between two atoms or ions that binds them together as a unit.
	1. Occurs due to a redistribution of valence electrons
	2. Bonds form to:
		1. decrease potential energy (PE)
		2. increase stability
2. There are three types of primary bonds:
	1. covalent bond - (prefix "co" means jointly so co-valent means jointly shared valence electrons)
		1. between non-metals
		2. atoms in a covalent bond share electrons to obey the octet rule
		3. Forms between atoms that have similar electronegativities
		4. Compounds formed from covalent bonds are:
			1. a. molecules - small (water and carbon dioxide)
			2. covalent network solids - large (diamonds)
		5. Lewis Dot Structures - are used to show covalent bonds between atoms.
			1. Steps for Lewis Dot Structures:
				1. 1. Figure out how many valence electrons are possessed by each atom in the molecule and arrange them around the atomic symbol with the least number of paired electrons
				2. If there are two different types of atoms, put the least electronegative atom in the center.
				3. Arrange the electrons so that each atom contributes one electron to a single bond between each atom
				4. Count the electrons around each atom and see if they obey octet rule. If so, they are complete.
				5. If octets are not complete, continue to bond lone electrons between atoms (double and triple bonds).
		6. Geometry - VSEPR Theory - Valence Shell Electron Pair Repulsion
			1. Idea that electron domains want to be as far away from each other as possible.
			2. Used to predict 3-D shape of molecule
			3. Common structures include: (DRAW THEM!!!!!!!!!!!!!)
				1. Linear
				2. Trigonal Planar
				3. Tetrahedral
				4. Pyramidal
				5. Bent
		7. Hybridization - (recall the shapes of the orbitals)
			1. A hybrid occurs when two things are combined and the result has characteristics of both
				1. EX: hybrid car (uses gas and electricity)
			2. During chemical bonding, different atomic orbitals undergo hybridization.
			3. Consider methane, CH4 - The carbon atom has four valence electrons with the electron configuration of [He]2s22p2.
			4. You may expect the two unpaired p electrons to bond with other atoms and the two paired s electrons to remain as a lone pair
			5. However, carbon undergoes hybridization, a process in which atomic orbitals mix and form new, identical, hybrid orbitals.
			6. With four hybrid orbitals - carbon can now make 4 bonds
			7. Lone pairs can also occupy hybrid orbitals
				1. dictate the shape of the molecule
				2. Example: Water(Bent)
		8. Valence Bond Theory
			1. Hybridization is a major player in this approach to bonding.
			2. There are two ways orbitals can overlap to form bonds between atoms.
				1. Sigma bonds are characterized by

Head-to-head overlap.

Cylindrical symmetry of electron density about the internuclear axis.

* + - * 1. Pi bonds are characterized by

Side-to-side overlap.

Electron density above and below the internuclear axis

* + - 1. Single bonds are always σ bonds, because σ overlap is greater, resulting in a stronger bond and more energy lowering.
			2. Multiple bonds
				1. In a multiple bond one of the bonds is a σ bond and the rest are π bonds.
				2. In a molecule like formaldehyde an *sp2* orbital on carbon overlaps in σ fashion with the corresponding orbital on the oxygen.
				3. The unhybridized *p* orbitals overlap in π fashion.
				4. In triple bonds, as in acetylene, two *sp* orbitals form a σ bond between the carbons, and two pairs of *p* orbitals overlap in π fashion to form the two π bonds.
			3. Delocalized Electrons: Resonance
				1. When writing Lewis structures for species like the nitrate ion, we draw resonance structures to more accurately reflect the structure of the molecule or ion.
				2. In reality, each of the four atoms in the nitrate ion has a *p* orbital.
				3. The *p* orbitals on all three oxygens overlap with the *p* orbital on the central nitrogen.
				4. This means the π electrons are not localized between the nitrogen and one of the oxygens, but rather are delocalized throughout the ion.
				5. The organic molecule benzene has six σ bonds and a *p* orbital on each carbon atom.
				6. In reality the π electrons in benzene are not localized, but delocalized.
				7. The even distribution of the π electrons in benzene makes the molecule unusually stable.
	1. Ionic Bond - an attraction between two oppositely charged particles (ions)
		1. Usually occurs between metals and non-metals
		2. Occurs between two particles that have different electronegativities - the larger the difference the stronger the ionic bond.
		3. Formation occurs when a metal, who has a low Ionization Energy, releases some its electrons and when a non-metal gains electrons and they achieve stable electron configurations.
		4. Compounds formed from ionic bonds are salts.
		5. Writing Formulas of Ionic Compounds
			1. Chemical Formulas - shows the number of atoms in a compound and must have a neutral charge.
			2. To write an ionic compound's chemical formula:
				1. Find the charges of the ions involved in the salt.
				2. Criss-Cross Rule - the charge of the cation becomes the subscript of the anion and the charge of the anion becomes the subscript of the cation.
				3. Make sure the subscripts are in lowest possible terms
				4. If writing a salt made with a polyatomic ion, all rules are the same but if you need multiple of a polyatomic ion, you must put parenthesis around it.
	2. Metallic bond - forms between metals
		1. Occurs when valence shells of metals overlap each other and allows the valence electrons to flow easily from atom to atom.
		2. The attraction comes from the flow of the "conduction" electrons and the fixed positive nuclei
1. Properties of each bond’s compounds

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | **Covalent-Network** | **Ionic** | **Metallic** | **Molecule** |
| **Melting Point** |  |  |  |  |
| **Solubility** |  |  |  |  |
| **Conductivity of Solid** |  |  |  |  |
| **Hardness** |  |  |  |  |
| **Brittleness** |  |  |  |  |