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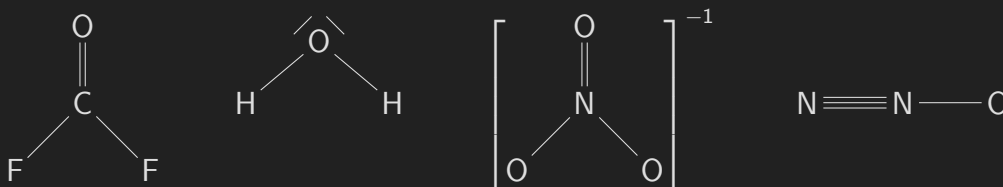
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1 General Chemistry Review

Electrons, Bonds, and Lewis Structures

- ▷ **Covalent bond:** two atoms sharing a pair of electrons.
- ▷ **Octet rule:** *main group elements* that tend to bond in a way that each atom has **eight** electrons in its valence shell.
 - Atoms that do not have eight will share electrons with other elements in order to maintain a stable state.
- ▷ **Main group elements:** sometimes called representative elements, are groups 1, 2 and 13–18 in periodic table.
 - Some elements in group 3 and 12 share properties between transition metals and the main group.
- ▷ The lowest energy (most stable) state of two atoms is determined both by bond length and bond strength.
- ▷ Valence electrons are determined by the group, 1A–8A, of the periodic table.
- ▷ **Lone pair:** unshared, or nonbonding, electrons.
- ▷ **Lewis structures:** 2D model that represents covalent bonds as straight lines and lone pairs as dots.
- ▷ Examples: COF_2 , H_2O , NO_3^- , N_2O :



- ▷ **Resonance structures:** a set of two or more Lewis structures that collectively describe the electronic bonding of a single polyatomic species, including fractional bonds.

Identifying Formal Charges

- ▷ **Formal charge:** any atom that does not exhibit the appropriate number of valence electrons.
- ▷ Determining formal charge:
 - Formula: $FC = V - N - \frac{B}{2}$
 - V = valence electrons of element
 - N = lone pair electrons

- B = bonded electrons
- ▷ **Less** than expected number of valence electrons results in a **positive** charge.
- ▷ **More** than expected number of valence electrons results in a **negative** charge.
- ▷ The lower the **magnitude** of formal charge, the **greater the stability** of the whole molecule.
- ▷ Atoms that are **more electronegative** hold **negative** formal charges better, which results in **greater stability** vs when the negative charge is spread on less electronegative elements in a polyatomic species.
 - The dominant resonance structure will be that of the greatest stability.

Induction and Polar Covalent Bonds

- ▷ Bonds can be classified into three categories: covalent, polar covalent, and ionic.
- ▷ The categories emerge from the electronegativity values of the atoms sharing a bond.
- ▷ **Electronegativity**: a measure of the ability of an atom to attract electrons.
 - Electronegativity generally **increases left to right**, and from the **bottom to top** of the periodic table.
- ▷ **Covalent bond**: when the difference in electronegativity is **less than 0.5**.
- ▷ **Polar covalent bond**: when the difference in electronegativity is **between 0.5 and 1.9**, then the electrons are not equally shared and become polar.
- ▷ **Induction**: the withdrawal of electrons towards a more electronegative atom. δ^+ represents partial positive charge gained when electrons are pulled away, while δ^- represents the partial negative charge pulled closer.
- ▷ **Ionic bond**: when the difference in electronegativity is **greater than 1.9**.
 - Electrons are not shared in this case, and attraction is instead just the result of oppositely charged ions.

Atomic Orbitals

- ▷ **Atomic orbital (AO)**: standing quantum wave (excitation in electron field) around an atom.
 - More energy leads to higher orbital levels.
 - Gives principle quantum number, n , as is associated with distance from nucleus.
 - Orbital levels: s(1 pair), p(3 pairs), d(5 pairs), f(7 pairs).

- Angular momentum quantum number that describes three-dimensional region of space that the electron density occupies.
- Magnetic quantum number describes orientation in space of electron density.
 - $m_l = 0$; s orbital
 - $m_l = -1, 0, 1$; p_x, p_y, p_z orbitals.
- Locations where ψ (quantum wave function) is zero are called **nodes**.
 - The **more nodes** that an orbital has, the **greater** it's energy.
- *Spin*: allows an orbital to contain only two electrons, $\pm\frac{1}{2}$
- ▷ **Degenerate orbitals**: orbitals with the same energy level.
- ▷ Order in which orbitals are filled is determined by three principles:
 - **Aufbau principle**: lowest energy orbital is filled first.
 - **Pauli exclusion principle**: each orbital can accommodate a maximum of two electrons that have opposite spin.
 - **Hund's rule**: electrons are placed in each degenerate orbital before being paired up.
- ▷ Describing the nature of atomic orbital is done with two commonly used theories: *Valence Bond Theory* and *Molecular Orbital Theory*.
- ▷ The commonly used theories give a deeper understanding of covalent bonds, which is essentially just the **overlap of atomic orbitals**.
- ▷ **Constructive/destructive interference**: the result of two waves that approach each other, or overlap.
 - Constructive interference produces a wave with the vector sum of both waves.
 - Destructive interference cancel each other out and produces a node.

Valence Bond Theory

- ▷ **Valence bond theory**: the sharing of electron density between two atoms is a result of the constructive interference of their atomic orbitals.
- ▷ *Bond axis*: the line that can be drawn between two hydrogen atoms.
- ▷ **Sigma bond (σ)**: a particular type of covalent bond that has circular symmetry with respect to the bond axis.
 - All single bonds are σ bonds.

- The strongest type of covalent bond.
- ▷ **Pi bond (π)**: covalent bonds where two lobes of an orbital overlap with two lobes of another atom.
 - Each atomic orbital has zero electron density at a shared nodal plane, passing through the two bonded nuclei.
 - π bonds form double ($\sigma + \pi$) and triple bonds ($\pi + \sigma + \pi$).
 - Individual π bonds are weaker than σ bonds.

Molecular Orbital Theory

- ▷ **Molecular orbital theory (MO)**: uses linear combinations of atomic orbitals to model and explore the consequences of orbital overlap.
 - The newly described orbitals are called **molecular orbitals** according to MO theory.
- ▷ Atomic orbitals refer to an individual atom, while molecular orbitals is associated with an entire molecular.
- ▷ In other words, MO theory states that atomic orbitals cease to exist when they overlap. Instead they are replaced with multiple molecular orbitals which span the entire molecule.
- ▷ Molecular orbitals are more stable (lower energy) since electrons are attracted by both nuclei.
- ▷ When there are **nodes** between the nuclei, then the resulting σ^* orbitals become **antibonding**, as they **destabilize** (increase the energy) of a molecular orbital.
- ▷ Best used to produce a quantitative picture of bonding.
 - Describes strength, order, and polarity of bonds.
 - Allows for the presence of paired or unpaired electrons.
 - Has spectroscopic properties.

Hybridized Atomic Orbitals

- ▷ **sp^3 -hybridized orbitals**: produced by averaging one s orbital and **three p** orbitals.
 - Hybridized orbitals explains to geomtry of methane, which results form the **now four degenerate** orbitals pushing apart to achieve tetrahedral geometry.

- Hybridized orbitals become **unsymmetrical**, producing a larger front lobe that is more efficient than standard p orbitals in the ability to form bonds.
- All bonds in are **σ bonds**, and thus can be individually represented by the overlap of atomic orbitals.
- ▷ **sp^2 -hybridized orbitals**: produced by averaging the s orbital with only **two** of p orbitals.
 - The remaining p orbital is unaffected, and free multiple p orbitals results in a π bond.
 - This is done to explain geometry of compounds bearing a double bond.
 - A double bond is formed from one σ bond and one π bond.
 - Associated with *trigonal planar geometry*.
- ▷ **sp -hybridized orbitals**: produced by averaging of one s orbital and **one** p orbital.
 - Leaves two p orbitals and resulting in two π bonds.
 - A triple bond is formed with the addition of one σ bond due to the overlap of the sp orbitals.
 - Geometry of a triple bond has *linear geometry*.
- ▷ Finding the hybridization of any atom can be done simply:
 1. Look at the central item.
 2. Determine groups (number of bonds, π bonds count as 1, and lone pairs attached) of atom.
 - groups aka regions of electron density.
 3. For groups 1-4: sp^x ; $x = \text{groups} - 1$
 4. For groups 5-6: sp^3d^x ; $x = \text{groups} - 4$
- ▷ Bond Strength and Bond Length:
 - Bond length **decreases** with more bonds.
 - Bond strength **increases** with more bonds.
 - The more **s character**, the **shorter** and **stronger** the bond, and the **larger** the bond angle.
 - *s -character*: contribution of the σ bond in a hybridization.
 - e.g. $sp = 50\%$, $sp^2 = 33\%$, $sp^3 = 25\%$
 - sp - sp bond is the strongest, sp^3 - sp^3 is the weakest.

Molecular Geometry

- ▷ **Valence shell electron pair repulsion (VSEPR) theory:** enables the prediction of molecular geometry due to the presumption that all electron pairs repel each other; resulting in a three-dimensional space that maximizes distance from each other.
- ▷ **Steric number:** the total number of electron pairs in a molecule. Can be bonds or lone pairs.
- ▷ **Tetrahedral geometry:** result of four σ bonds and zero lone pairs.
 - produces a tetrahedron with bond angles of 109.5° .
- ▷ **Trigonal pyramidal geometry:** three σ bonds and one lone pair.
 - The lone pair occupies more space than bonded electron pairs, so the remaining angles are slightly less than a tetrahedral, at 107° .
 - The lone pair sits atop the base forming a pyramid like structure.
- ▷ **Bent geometry:** two σ bonds and two lone pairs.
 - VSEPR predicts the lone pairs to be in two corners of the tetrahedral, producing bond angles of 105° .
 - VSEPR predicts geometry H_2O correctly, but for wrong reasons.
 - The lone pairs in H_2O have different energy levels, suggesting one pair occupies a p orbital with the other in a lower-energy hybridized orbital.
- ▷ VSEPR theory is best used for a first approximation and is mostly accurate for most small molecules.
- ▷ **Trigonal planar geometry:** three electron pairs forming three bond angles of 120° and lie on the same plane.
- ▷ **Linear geometry:** two electron pairs that oppose each other at 180° , forming a linear structure.
- ▷ General method of determining structure:
 1. Count steric number—the total number of electron pairs in a molecule. Can be bonds or lone pairs.
 2. Determine predicted geometrical structure predicted (EDG) by VSEPR using steric number.
 - Octahedral:6, Bipyrmaid:5, Tetrahedral:4, Trigonal:3, Linear:2
 3. Determine impact (the MG) of lone pairs; more lone pairs results in less space between bonded pairs. Shape depends on EDG.

Dipole Moments and Molecular Polarity

- ▷ **Dipole moment (μ):** defined as the amount of partial charge, δ , on either end of the dipole multiplied by the distance separation, d :
 - $\mu = \delta d$
 - μ generally has an order of magnitude of 10^{-18} esu·cm due to general partial charge (esu) and distance (cm) values.
 - 1 debye (D) = 10^{-18} esu·cm
- ▷ **Molecular dipole moment:** the vector sum of the individual dipole moments.
 - Lone pairs have significant effect on the molecular dipole moment.
 - Also called the net dipole moment.

Intermolecular Forces and Physical Properties

- ▷ **Intermolecular forces:** the attractive forces between individual molecules that determined the physical properties of a compound.
- ▷ *Electrostatic:* forces that occur as a result of the attraction between opposite charges.
- ▷ Electrostatic interactions for neutral molecules (no formal charge) are often classified as into the following categories:
 - **Dipole-dipole interaction:** Compounds with net dipole moments.
 - In solid space these interactions either repel or attract each other.
 - In liquid space these interactions tend to attract more often, raising melting/boiling point.
 - **Hydrogen bonding:**
 - Not actually a bond, just an interaction.
 - When hydrogen bonds to a electronegative atom, then the hydrogen will have a δ^+ .
 - **F, O, N, Cl** (Br, I). Most electronegative elements, from left to right, that hydrogen most often bonds too.
 - Hydrogen bonding is strong due to size of hydrogen atom, resulting in very close partial charge interactions.
 - The more hydrogen bonds, the higher the boiling point tends to be.
 - **Fleeting dipole-dipole interactions:**

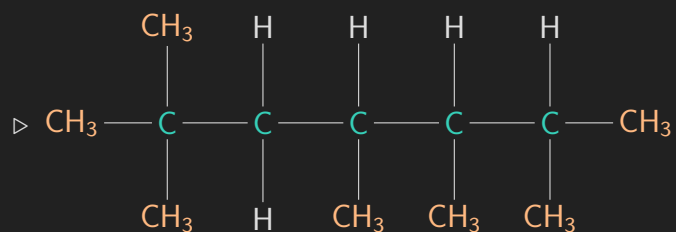
- Electrons are considered to be in constant motion, which result in the center of negative charge to vary.
 - On average, the dipole moment is zero, though can experience transient dipole moments, initiating fleeting attraction/repulsion.
 - Heavier hydrocarbons generally experience a stronger force due to increased surface area, and thus greater chance for non-zero dipole moments, which results in higher boiling points.
 - Branched hydrocarbons generally have decreased surface area, decreasing boiling point relative to others of similar weight.
- ▷ When comparing boiling points of compounds, look for following factors:
- Any dipole-dipole interactions?
 - Formation of hydrogen bonds?
 - Number of carbon atoms. (surface area)
 - Degree of branching of compound. (surface area)

Structural Theory of Matter

- ▷ **Constitutional isomers:** aka structural isomers; same chemical formula, but different in the way the atoms are connect, i.e. their constitution is different.
- Consistent with the octet rule.
 - Each element forms a predictable number of bonds, from one to four.
 - The number of possible constitutional isomers increases as the number of carbon atoms increases
- ▷ **Stereoisomers:** isomers that differ in spatial arrangement of atoms, rather than connectivity.
- **Geometric isomerism:** aka cis–trans; locked into spatial positions due to double bonds or a ring structure.
 - Cis indicates functional groups that are on the same side of the carbon chain.
 - Trans indicates functional groups on opposite sides of the carbon chain.
 - **Enantiomers:** aka optical isomers; mirror images of each other that are non-superposable.
 - Human hands are a macroscopic analogy.
- ▷ More detail will be covered in later sections.

2 Molecular Representations

Molecular Representations



Bond-Line Structures

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