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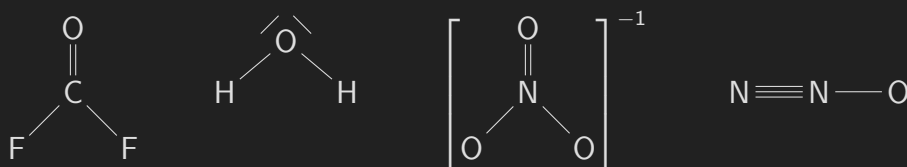
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1 General Chemistry Review

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.
 - **F, O, N, Cl** (Br, I). Most electronegative elements, from left to right, that are often encountered.
- ▷ **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

- ▷ **sp³-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.
- ▷ **sp²-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 expain geometry of compounds bearing a double bond.
 - A double bond if 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. Determin 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 = groups - 1
 4. For groups 5-6: sp³d^x; x = 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 occupy 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 plan.
- ▷ **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, Bipyramid: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.
 - **Ion-dipole:** electrostatic interaction between an ion and a molecule with a dipole.
 - **Hydrogen bonding:** molecules with a hydrogen attached to an F, O, or N.
 - Not actually a bond, just an interaction.

- When hydrogen bonds to a electronegative atom, then the hydrogen will have a δ^+ .
- 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.
- Stronger than dipole-dipole interactions.
- **Fleeting dipole-dipole interactions:**
 - Electrons are considered to be in constant motion, which result in the center of negative charge to vary.
 - **London Dispersion Forces (LDFs):** On average, the dipole moment is zero, though it can experience transient dipole moments, initiating fleeting attraction/repulsion.
 - All atoms and molecules have LDFs.
 - Weakest, but the dominant force in non-polar molecules.
 - Dispersion forces directly related to molar mass.
 - 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? (increases boiling point)
 - Formation of hydrogen bonds? (increase boiling point)
 - Number of electrons. (more electrons, higher boiling point)
 - Number of carbon atoms. (more surface area, higher boiling point)
 - Degree of branching of compound. (more branching, more surface area)

Structural Theory of Matter

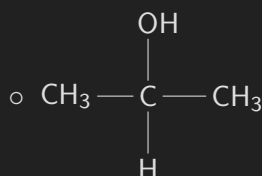
- ▷ **Constitutional isomers:** aka structural isomers; same **chemical formula**, but different in the way the **atoms are connect**, i.e. their constitution is different.
 - Consistent with the octet rule.
 - Each element forms a predictable number of bonds, from one to four.

- 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.*

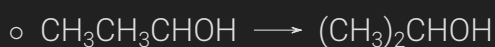
2 Molecular Representations

Types of Molecular Representations

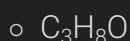
- ▷ **Partially condensed structures:** the C–H bonds are not always drawn, saving space.



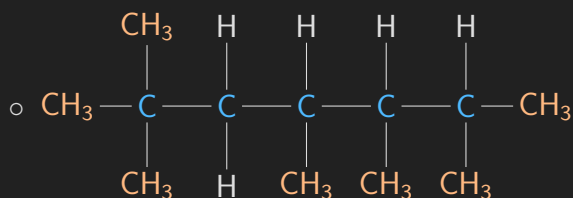
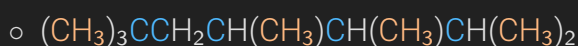
- ▷ **Condensed structures:** single bonds are not drawn and groups of atoms are clustered when possible.



- ▷ **Molecular formula:** simply shows number of each type of atom with no structural information.



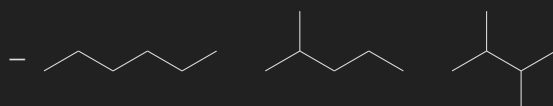
- ▷ Example of converting a condensed structure into a partially condensed structure:



- This shows just one isomer, more partially condensed structures are possible.

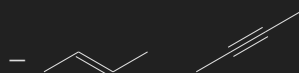
Bond-Line Structures

- **Bond-line structures;** aka skeletal structures; simplify drawing process of chemical structures and are easier to read.
- Each corner or endpoint represents a carbon atom.



- All examples have 6 carbon atoms

- Double bonds are shown with two lines, triple with three.

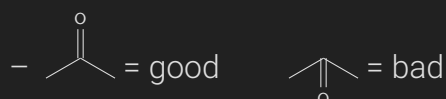


- Triple bonds are drawn linearly due to sp-hybridization

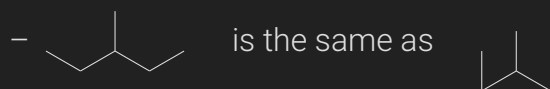
- Hydrogens are not shown; it is assumed that each carbon possesses enough to satisfy octet rule.

Notes on Drawing Bond-Line Structures

- Carbon atoms in a straight chain should be drawn in zigzag format in order to accurately show each carbon.
- Double bonds should be drawn as far apart as possible:



- Direction of a single bond is irrelevant:



- All **heteroatoms** (atoms other than carbon and hydrogen) must be drawn.
 - Hydrogens next to heteroatoms must be shown.
- Carbons cannot have more than four bonds.

Hydrogen Deficiency Index: Degrees of Unsaturation

Excerpt from Chapter 14: Infrared Spectroscopy and Mass Spectrometry

- ▶ **Saturated compounds:** the maximum number of hydrogen atoms possible, relative to number of carbon present.
 - Determining saturation using molecular formula: C_nH_{2n+2} n = carbon atoms
 - Halogens:** takes the place of a hydrogen atom; **add one H** for each halogen.
 - Oxygen:** no affect on saturation; **ignore**.
 - Nitrogen:** needs an extra hydrogen; **subtract one H** for each nitrogen.
- ▶ **Unsaturated compounds:** a compound that contains at least one π bond, resulting fewer than the maximum number of hydrogen atoms.
 - Compounds with rings also result in an unsaturated compound.
 - Degree of unsaturation:** a number that represents **half** the "missing" number of hydrogen atoms when compared to a saturated compound.
- ▶ **Hydrogen deficiency index (HDI):** the measure of degrees of unsaturation.
 - e.g. two degrees of unsaturation results in a HDI of 2.

- Degrees of freedom help represent possible structures, indicating possible double bonds, triple bonds, rings, or various combinations of each.
- Only helpful when molecular formula is known for certainty.
- Formula: $\text{HDI} = \frac{1}{2}(2C + 2 + N - H - X)$
 - X: halogen atoms.

Identifying Functional Groups

FUNCTIONAL GROUP*	CLASSIFICATION	EXAMPLE	CHAPTER	FUNCTIONAL GROUP*	CLASSIFICATION	EXAMPLE	CHAPTER
$\text{R}-\ddot{\text{X}}:$ (X=Cl, Br, or I)	Alkyl halide	 <i>n</i> -Propyl chloride	7	 $\text{R}-\text{C}(=\text{O})-\text{R}$	Ketone	 2-Butanone	19
$\text{R}-\text{C}(\text{R})=\text{C}(\text{R})-\text{R}$	Alkene	 1-Butene	7, 8	 $\text{R}-\text{C}(=\text{O})-\text{H}$	Aldehyde	 Butanal	19
$\text{R}-\text{C}\equiv\text{C}-\text{R}$	Alkyne	 1-Butyne	9	 $\text{R}-\text{C}(=\text{O})-\text{OH}$	Carboxylic acid	 Pentanoic acid	20
$\text{R}-\ddot{\text{O}}\text{H}$	Alcohol	 1-Butanol	12	 $\text{R}-\text{C}(=\text{O})-\text{X}$	Acyl halide	 Acetyl chloride	20
$\text{R}-\ddot{\text{O}}-\text{R}$	Ether	 Diethyl ether	13	 $\text{R}-\text{C}(=\text{O})-\text{O}-\text{C}(=\text{O})-\text{R}$	Anhydride	 Acetic anhydride	20
$\text{R}-\ddot{\text{S}}\text{H}$	Thiol	 1-Butanethiol	13	 $\text{R}-\text{C}(=\text{O})-\text{O}-\text{R}$	Ester	 Ethyl acetate	20
$\text{R}-\ddot{\text{S}}-\text{R}$	Sulfide	 Diethyl sulfide	13	 $\text{R}-\text{C}(=\text{O})-\text{N}(\text{R})_2$	Amide	 Butanamide	20
	Aromatic (or arene)	 Methylbenzene	17, 18	 $\text{R}-\text{N}(\text{R})_2$	Amine	 Diethylamine	22

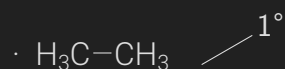
* The "R" refers to the remainder of the compound, usually carbon and hydrogen atoms.

- **Functional group (R):** specific substituents or moieties within molecules that may be responsible for the characteristic chemical reactions.
- **Substituents:** an atom or group of atoms which replaces one or more hydrogen atoms on the parent hydrocarbon chain.
 - **Moiety:** a part of a molecule which is typically found within other molecules and often given a specific name.

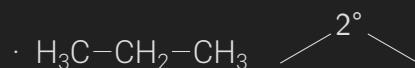
Characterizing Carbon Centers and Functional Groups

Characterizing Carbon Centers

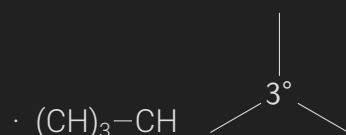
- Primary 1°: a carbon with only one carbon-carbon bond.



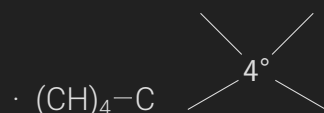
- Secondary 2°: a carbon with two carbon-carbon bonds.



- Tertiary 3°: a carbon with 3 carbon-carbon bonds.



- Quaternary 4°: a carbon with four carbon-carbon bonds.



Characterizing Functional Groups

- Certain functional groups can be characterized as 1°, 2°, or 3°, based on how many carbon bonds are attached to the carbon with the functional group.

Identifying Lone Pairs

- ▷ Formal charges must always be drawn on bond line structures, otherwise the resulting bond line structures would be inferred incorrectly.
- ▷ Lone pairs do not have to be drawn and usually are omitted.
- ▷ The formal charge allows you to determine lone pairs.

- Formula: $FC = V - N - \frac{B}{2}$

- V = valence electrons of element

- N = lone pair electrons

- B = bonded electrons

- Solve for lone pairs: $N = V - FC - \frac{B}{2}$

- ▷ Frequent usage will allow for intuition for lone pairs.

Common Patterns Between Formal Charge and Lone Pairs

o Associated Patterns for Oxygen

- A **negative** (\ominus) charge corresponds with **1 bond** and **3 lone pairs**.
- The absence of charge corresponds with **2 bonds** and **2 lone pairs**.
- A **positive** (\oplus) charge corresponds with **3 bonds** and **1 lone pair**.

o Associated Patterns for Nitrogen

- A **negative** charge corresponds with **2 bonds** and **2 lone pairs**.
- The absence of charge corresponds with **3 bonds** and **1 lone pair**.
- A **positive** charge corresponds with **4 bonds** and **0 lone pairs**.

Resonance

- ▷ **Resonance:** description of bonding in molecules or ions by the combination of multiple contributing structures.
- o **Resonance structures:** each contributing structure of the resonance hybrid.
 - Formal charges are important to include when drawing resonance structures as it clarifies where locations of lone pairs and movement of electrons.
 - Total charge must remain the same between structures.
- ▷ Resonance does not describe any real process, rather it's a method to overcome inadequacy of bond-line drawings.
- ▷ Different from isomerism, which differs in arrangements of atomic nuclei in space, rather than how the electrons are assigned to the depictions.

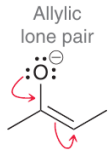

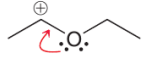


Resonance: Curved Arrows

- o **Curved arrows:** a tool used to help draw resonance structures by **representing electrons as if** they were moving.
 - Somewhat different from curved arrow notation in reactions, which actually represent the flow of electron density.
 - Can help show how to change the formal charge:
 - Formal charges at the **tail** become more **positive**, since it's losing an electron.
 - Formal charges at the **head** more **negative**, since it's gaining an electron.

- **Avoid breaking a single bond.**
 - Structures must have atoms connected in same order, though there are minor exceptions that *will be discussed later*.
 - This rule affects the placement of the **tail** of the arrow, as it represents distribution of previous electrons.
- **Never exceed an octet for second-row elements.**
 - Not a violation to have less than an octet.
 - This rule affects the placement of the **head** of the arrow, as it represents sharing of new electrons.
- Can only be used on adjacent atoms, though the electrons can be pushed multiple times.
- "Legal" moves:
 - π bond \rightarrow lone pair.
 - Lone pair $\rightarrow \pi$ bond.
 - π bond $\rightarrow \pi$ bond.
 - Every resonance structure can be built through a combination of above three moves.

Common Patterns of Resonance Structures

- **Vinylic:** the two carbon atoms bearing the double bond of a carbon-carbon double bond.
- **Allylic:** atoms connected directly to vinylic positions.

<p>Allylic lone pair</p>  <p>Two curved arrows</p>	<p>Allylic carbocation</p>  <p>One curved arrow</p>	<p>Lone pair adjacent to C+</p>  <p>One curved arrow</p>	<p>π bond between two atoms of differing electronegativity</p>  <p>One curved arrow</p>	<p>Conjugated π bonds enclosed in a ring</p>  <p>Three curved arrows</p>
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Resonance Hybrid

- **Resonance hybrid:** represents the *average* of the contributing structures, with bond lengths and partial charges taking on intermediate values.
- No matter how many resonance structures are drawn, they collectively represent one entity.

- Drawn partial bonds and charges to illustrate the delocalization of electrons.

Delocalization

- **Delocalization**: the spreading of electrons between multiple atoms or covalent bonds.
 - **Resonance stabilization**: molecules and ions that are stabilized by the delocalization of electrons.
 - Plays a major role in the outcome of many reactions.
- When a lone pair participates in resonance, it will occupy a p orbital rather than hybridized; important for 3d shapes of proteins.
- **Localized lone pair**: when a lone pair is not allylic to a π bond.
 - Whenever an atom possesses both a π bond and a lone pair, they will not both participate in resonance.
 - Usually π bonds participate first.

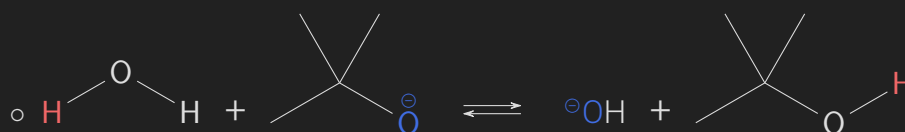
Contributor Significance

- Some resonance structures may resemble the actual molecule more than another, in regards to energy and stability.
- Structures with low potential energy are more stable compared to those of higher values and resemble the actual structure more.
- **Major contributors**: the most stable contributing structures.
- **Minor contributors**: less favorable contributing structures.
- Rules for contributing significance, descending:
 - The greatest number of filled octets.
 - The greatest number of covalent bonds.
 - Minimize formally charged atoms.
 - Separation of unlike and like charges, minimized and maximized respectively.
 - Negative charges placed on the most electronegative atoms, positive charges placed on the less electronegative atoms.
 - Do not deviate substantially from idealized bond lengths and angles.
 - Maintain aromatic substructures locally while avoiding anti-aromatic ones.

3 Acids and Bases

Introduction to Brønsted-Lowry Acids and Bases

- ▷ **Acid:** a proton donor; i.e., a H^+ donor.
- ▷ **Base:** a proton acceptor; i.e., a OH^- (hydroxide ion), which wants a H^+ to form the more stable H_2O .
- ▷ General definition: $\text{acid} + \text{base} \rightleftharpoons \text{conjugate base} + \text{conjugate acid}$
- ▷ Symbolically: $\text{HA} + \text{B} \rightleftharpoons \text{A}^- + \text{HB}^+$
- ▷ Example using bond-line structures:



Lewis Acids and Bases

- ▷ The Lewis definition is more broad than the Brønsted-Lowry definition.
- ▷ Lewis describes acidity in terms of electrons, rather than protons.
- ▷ **Lewis acid:** electron-pair acceptor.
- ▷ **Lewis base:** electron-pair donor.
- ▷ Most reactions are described in terms of Lewis base and acids, since molecules without donatable protons are unable to be described by the Brønsted-Lowry definition.

Nucleophiles and Electrophiles

Excerpt from Chapter 6: Chemical Reactivity and Mechanisms ⇌

- ▷ **Ionic reactions**, aka polar reactions: reactions that involve the participation of ions as reactants, intermediates, or products.
 - Most cases ions act as intermediates.
 - Radical reactions and pericyclic reactions are also major categories, but are typically not discussed in undergraduate courses.
 - Ionic reactions occur when one reactant has a site of high electron density and the other reactant has a site of low electron density.
- ▷ **Nucleophiles:** an electron rich atom that is capable of donating a pair of electrons.
 - Nucleophiles are Lewis bases.

- Any atom that possesses a localized lone pair can be nucleophilic.
- π bonds can also function as nucleophiles due to their region of space having high electron density.
- **Polarizability**: the ability of an atom to distribute its electron density unevenly in response to external influences.
 - Correlated with size of the atom, which increases the number electrons that are distant from the nucleus.
- ▷ **Electrophiles**: an electron-deficient atom that is capable of accepting a pair of electrons.
 - Electrophiles are Lewis acids.

Flow of Electron Density: Curved-Arrow Notation

- ▷ All reactions are accomplished via a flow of electron density.
- ▷ Electron density flow is illustrated with curved arrows.
 - **Reaction mechanism**: how the reaction occurs in terms of the motion
 - All ionic mechanisms, regardless of complexity, are combinations of four characteristic patterns of electron flow (*discussed later*).

Notes on Drawing Curved Arrows

- **Tails** must be placed on either a bond or a lone pair.
 - Shows the **source**, i.e., the electron donor (base).
 - Electrons can only be found in lone pairs or bonds, so **never place the tail** of a curved arrow on a **positive charge**.
- **Heads** must be placed so that it shows either the formation of a bond or the formation of a lone pair.
 - Shows the **destination**, i.e., the electron acceptor (acid).
 - Avoid drawing an arrow that violates the octet rule, so never draw an arrow that gives more than four orbitals to a second-row element.

4 Alkanes and Cycloalkanes

Introduction to Alkanes



Nomenclature of Alkanes

