Contents

1 General Chemistry Review

Structural Theory of Matter

- **Constitutional isomers**: same molecular formula, but different in the way the atoms are connect, i.e. their constitution is different.
- Each element forms a predictable number of bonds, from one to four.
- \ch{x-x} single: -, double: =, triple: +. e.g. $CH_3{-}CH_3$, $CH_2{=}CH_2$, $CH{\equiv}CH$

Electrons, Bonds, and Lewis Structures

- **Covalent bond**: two atoms sharing a pair of electrons.
- The lowest energy (most stable) state of two atoms is determined both by bond length and bond strength.
- **Lewis structures**: drawings that show free electrons.
- Valence electrons are determined by the group, 1A–8A, of the periodic table.
- **Lone pair**: unshared, or nonbonding, electrons.

Identifying Formal Charges

- **Formal charge**: any atom that does not exhibit the appropriate number of valance electrons.
- Less than expected results in positive charge.
- More than expected results in negative charge.

Induction and Polar Covalent Bonds

- Bonds are classified into three categories: covalent, polar covalent, 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.
- If the difference in electronegativity is less than 0.5, then the electrons are considered equally shared and this covalent.
- If the difference in electronegativity is between 0.5 and 1.7, then the electrons are not equally shared and thus a polar covalent bond.

- **Induction**: the withdrawl of electrons towards to more electronegative atom. δ^+ represnets partial positive charged gained when electrons are pulled away, while δ^- represnets the partial negative charge pulled closer.
- If the difference in electronegativity is greater than 1.7 then the electrons are not shared and results in an ionic bond which is just a result of the force between two oppositely charged ions.

Atomic Orbitals

- **Atomic orbital (AO)**: *s(1), p(3), d(5), f(7)*.
- Locations where ψ is zero is called **nodes**.
- The more nodes that an orbital has, the greater it's energy.
- **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.

Valence Bond Theory

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Molecular Orbital Theory

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Hybridized Atomic Orbitals

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Molecular Geometry

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Dipole Moments and Molecular Polarity

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Intermolecular Forces and Physical Properties

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Solubility

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