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Chapter 2 – Polar Covalent Bonds; Acids and Bases

Chapter Outline

I. Polar covalent bonds (Sections 2.1–2.3).

A. Electronegativity (Section 2.1).

- Although some bonds are totally ionic and some are totally covalent, most chemical bonds are polar covalent bonds.
 - In these bonds, electrons are attracted to one atom more than to the other atom.
- Bond polarity is due to differences in electronegativity (EN).
 - Elements on the right side of the periodic table are more electronegative than elements on the left side.
 - Carbon has an EN of 2.5.
 - Elements with $EN > 2.5$ are more electronegative than carbon.
 - Elements with $EN < 2.5$ are less electronegative than carbon.
- The difference in EN between two elements can be used to predict the polarity of a bond.
 - If $\Delta EN < 0.4$, a bond is nonpolar covalent.
 - If ΔEN is between 0.4 and 2.0, a bond is polar covalent.
 - If $\Delta EN > 2.0$, a bond is ionic.
 - The symbols δ^+ and δ^- are used to indicate partial charges.
 - A crossed arrow is used to indicate bond polarity.
 - The tail of the arrow is electron-poor, and the head of the arrow is electron-rich.
- Electrostatic potential maps are also used to show electron-rich (red) and electron-poor (blue) regions of molecules.
- An inductive effect is an atom's ability to polarize a bond.

B. Dipole moment (Section 2.2).

- Dipole moment is the measure of a molecule's overall polarity.
- Dipole moment (μ) = $Q \times r$, where Q = charge and r = distance between charges.
 - Dipole moment is measured in debyes (D).
- Dipole moment can be used to measure charge separation.
- Water and ammonia have large values of D; methane and ethane have $D = 0$.

C. Formal charge (Section 2.3).

- Formal charge (FC) indicates electron "ownership" in a molecule.
- $$(FC) = \left[\begin{array}{c} \# \text{ of valence} \\ \text{electrons} \end{array} \right] - \left[\begin{array}{c} \# \text{ of bonding electrons} \\ \uparrow \\ 2 \end{array} \right] - \left[\begin{array}{c} \# \text{ nonbonding} \\ \text{electrons} \end{array} \right]$$

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