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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).

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

B. Dipole moment (Section 2.2).

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

C. Formal charge (Section 2.3).

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

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