

chapter 6 review chemical bonding answers

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Introduction

Chapter 6 of most introductory chemistry textbooks is dedicated to the fascinating world of **chemical bonding**. Whether you're a high school student tackling the basics or a university freshman exploring the nuances of molecular structure, the concepts in this chapter form the foundation of all chemical science. In this comprehensive review, we'll walk through the most common questions students face, providing clear, concise answers that illuminate the principles of ionic, covalent, and metallic bonding. By the end of this article, you'll have a solid grasp of how atoms combine, how to predict bond types, and why the shapes and properties of molecules matter in real world chemistry.

Understanding Chapter 6: Chemical Bonding

Key Concepts Covered

- Ionic and covalent bond formation
- Electronegativity and bond polarity
- Lewis structures and octet rule
- VSEPR theory and molecular geometry
- Hybridization and orbital overlap
- Bond order, bond length, and bond strength
- Resonance and hypervalency
- Metallic bonding and crystal lattices

Common Questions and Answers

1. What is the difference between ionic and covalent bonds?

In an **ionic bond**, electrons are transferred from a metal to a nonmetal, creating oppositely charged ions that attract each other. This transfer leads to the formation of a crystal lattice with high melting points. In contrast, a **covalent bond** involves the sharing of one or more pairs of electrons between atoms, typically between nonmetals. Covalent bonds can be nonpolar (equal sharing) or polar (unequal sharing). The key distinction lies in electron transfer versus electron sharing, which dictates the physical properties and reactivity of the resulting compounds.

2. How do you determine bond type using electronegativity?

Electronegativity is the tendency of an atom to attract electrons. The difference in electronegativity values between two bonded atoms guides the bond type:

- $0 \leq \Delta EN < 0.5$: Nonpolar covalent
- $0.5 \leq \Delta EN < 1.7$: Polar covalent
- $\Delta EN \geq 1.7$: Ionic (often with some covalent character)

These thresholds are approximate, but they provide a useful rule of thumb for predicting bond behavior.

3. What is a Lewis structure?

A Lewis structure, or electron dot diagram, represents atoms, bonds, and lone pairs of electrons in a molecule. The

diagram uses dots or lines to depict valence electrons, helping to visualize how atoms share or transfer electrons to achieve stable electron configurations. Constructing a Lewis structure involves the following steps:

1. Count total valence electrons.
2. Choose a central atom.
3. Connect atoms with single bonds.
4. Add lone pairs to satisfy the octet rule.
5. Adjust for formal charges if necessary.

Lewis structures are essential for predicting molecular geometry and reactivity.

4. Explain VSEPR theory.

Valence Shell Electron Pair Repulsion (VSEPR) theory predicts the three dimensional shapes of molecules by assuming that electron pairs around a central atom repel each other and arrange themselves as far apart as possible. The theory classifies electron groups into:

- Bonding pairs
- Lone pairs

By counting these groups, VSEPR assigns a geometry (e.g., linear, trigonal planar, tetrahedral, trigonal bipyramidal, octahedral). This geometry influences a molecule's physical properties, such as polarity and reactivity.

5. How do resonance structures affect bond order?

Resonance occurs when a molecule can be represented by two or more valid Lewis structures. The true electronic structure is a hybrid of all contributing structures, leading to partial bonds that are intermediate between single and double bonds. Bond order is calculated as:

For example, in the nitrate ion (NO_3^-), each N–O bond has a bond order of 1.33, reflecting equal distribution of double bond character among the three bonds.

6. What is hybridization?

Hybridization describes the mixing of atomic orbitals to form new, equivalent hybrid orbitals that dictate molecular geometry. Common types include:

- sp (linear)
- sp^2 (trigonal planar)
- sp^3 (tetrahedral)
- sp^3d (trigonal bipyramidal)
- sp^3d^2 (octahedral)

Hybrid orbitals overlap to form sigma bonds, while unhybridized p orbitals can form pi bonds.

7. Describe ionic vs covalent bonding in real compounds (e.g., NaCl, H₂O).

In **NaCl**, sodium donates an electron to chlorine, forming Na^+ and Cl^- ions that assemble into a crystalline lattice. The bond is largely ionic with some covalent character due to polarizability. In **H₂O**, oxygen shares electrons with two hydrogen atoms, creating polar covalent bonds. The unequal sharing results in a bent molecular shape with a dipole moment. These differences explain why NaCl is a solid salt at room temperature while H₂O is a liquid with a high boiling point.

8. What is the concept of bond polarity?

Bond polarity refers to the unequal distribution of electron density between bonded atoms. If electrons are shared

more strongly by one atom, that atom becomes partially negative (δ^-) and the other partially positive (δ^+). Polarity affects intermolecular forces, solubility, and boiling points. Molecules with polar bonds may exhibit dipole-dipole interactions, whereas nonpolar molecules rely on London dispersion forces.

9. How do you calculate bond energy?

Bond energy, or bond dissociation energy, is the energy required to break one mole of bonds in a gaseous molecule. It can be estimated using thermochemical data or derived from Hess's law calculations. The general relationship is:

High bond energies indicate strong, stable bonds, while low values suggest easier bond cleavage.

10. What are the factors influencing bond strength?

Bond strength depends on:

- Bond length (shorter bonds are stronger)
- Bond order (double > single)
- Atomic size and electronegativity
- Hybridization (sp^3 bonds are stronger than sp^2)
- Resonance stabilization (delocalized electrons lower energy)

Understanding these factors helps predict reactivity and stability of molecules.

11. Explain the difference between single, double, and triple bonds.

A **single bond** consists of one sigma (σ) bond formed by head-on

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