8

COVALENT BONDING

A. Completion

Use this completion exercise to check your understanding of the concepts and terms introduced in this chapter. Each blank can be completed with a term, short phrase, or number.

When atoms share electrons to gain the _1_ configuration of a noble gas, the bonds formed are _2_. A _3_ pair of valence electrons constitutes a _4_ covalent bond. Sometimes two or three pairs of electrons may be shared to give _5_ covalent bonds. In some cases only one of the atoms in a bond provides the pair of bonding electrons; this is a _6_.

When like atoms are joined by a covalent bond, the bonding electrons are shared _7_, and the bond is _8_. When the atoms in a bond are not the same, the bonding electrons are shared _9_, and the bond is _10_. The degree of polarity of a bond between any two atoms is determined by consulting a table of _11_. Some molecules are _12_ because they contain polar covalent bonds. Bond dissociation energies are affected by bond polarity and the electronegativity of the joined atoms. The attractions between opposite poles of polar molecules constitute _13_. The dipole interaction is one of several weak attractions between molecules. Another weak attractive force is the _14_. These weak forces determine whether a covalent compound will be a solid, liquid, or gas at room temperature.

As a general rule, molecules adjust their three-dimensional shapes so that the valence-electron pairs around a central atom are as far apart as possible. This is the guiding principle in the valence-shell electron-pair repulsion, or _15_ theory of molecular geometries.
B. True-False

Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

16. In a polar covalent bond the more electronegative atom has a slight positive charge.

17. In general, the electronegativity values of nonmetallic elements are greater than the electronegativity values of metallic elements.

18. A molecule with polar bonds must itself be polar.

19. If two or more atoms are covalently bonded together, a molecule of a compound results.

20. A hydrogen bond must involve a hydrogen atom that is covalently bonded to a very electronegative atom.

21. To attain a noble gas electron structure, a nitrogen atom must lose its five valence electrons.

22. The compound $\text{OF}_2$ contains two double covalent bonds.

23. The tendency of carbon to form four bonds to other atoms can be explained by the four $p$ electrons in its outer shell.

24. The modern interpretation of resonance is that electron pairs rapidly flip back and forth between the various electron dot structures.

25. Unshared pairs of electrons affect the shape of molecules.

26. Covalent compounds are network solids.

C. Questions

Answer the following questions in the space provided.

27. Draw electron dot structures for each of the following compounds:
   a. $\text{Br}_2$
   b. $\text{HCN}$
   c. $\text{NH}_4^+$

28. Classify the following compounds as ionic or covalent. If the compound is covalent, indicate the degree of polarity.

   a. $\text{BaCl}_2$
   b. $\text{HF}$
   c. $\text{K}_2\text{S}$
   d. $\text{N}_2$
   e. $\text{NO}_2$
29. Explain how atoms form covalent bonds.

30. Arrange the following intermolecular attractions in order of increasing strength:
dipole interactions, dispersion forces, hydrogen bonds.

31. Draw resonance structures for the following substance: NO.

32. Explain the formation of a coordinate covalent bond between a hydrogen ion and an ammonia molecule.

33. From the list below select the compound that is best described by each statement, and write the formula in the blank.

   \[ \text{H}_2\text{O}, \text{CCl}_4, \text{NH}_3, \text{CO}_2 \]
   a. Which compound represents a linear molecule?  
   b. Which compound best represents a tetrahedral molecule?  
   c. Which compound would show a bent molecular structure?  
   d. Which compound's molecules have a pyramidal shape?

D. Additional Questions

Answer the following questions in the space provided.

34. Draw resonance for NO\textsubscript{2}.

35. Indicate the hybrid orbitals used by each carbon atom in the following compound:

\[
\begin{array}{c}
\text{H}_3\text{C} - \text{C} = \text{C} - \text{C} = \text{C} - \text{CH}_3 \\
\text{H}
\end{array}
\]
Structures of Covalent Compounds

In the last worksheet you learned about the ionic bonds. These are caused by the give-and-take of electrons. Not all bonds involve the gain and loss of electrons, however. Some bonds are formed by the sharing of electrons between two atoms. These are called covalent bonds, and they are the subject of this worksheet.

Example A

Write the electron dot structure for hydrogen fluoride, HF.

Solution The two atoms must share a pair of electrons in order to form a single covalent bond. Hydrogen shares its one unpaired electron with fluorine, giving fluorine eight electrons, like neon, whereas fluorine shares its one unpaired electron with hydrogen, giving hydrogen 2 electrons, like helium.

\[ H: F^- \]

You Try It

1. Draw the electron dot structure for phosphorus trifluoride, PF₃.

Your Solution

Example B

Predict the shape and bond angle for the compound carbon tetrafluoride, CF₄.

Solution The four fluorine atoms are covalently bonded to the central carbon atom. The four shared-pairs of electrons repel each other to the corners of a tetrahedron. All four bond angles are 109.5°.

You Try It

2. Predict the shape and bond angle for phosphorus trifluoride, PF₃.

Your Solution
Example C
Predict the type of hybridized orbitals involved in the compound boron trichloride, BCl₃.

Solution  Boron, the central atom, has an s orbital and two p orbitals involved in the bonding. The type of hybridized orbitals involved are \( sp^2 \) orbitals.

You Try It
3. Predict the hybridized orbitals involved in the bonds in silicon tetrafluoride, SiF₄.

Your Solution

Problems For You To Try
4. Draw the electron dot structure for nitrogen trichloride.

5. Predict the shape and bond angle of oxygen difluoride, OF₂.

6. Draw the electron dot configuration for acetylene, C₂H₂.

7. Predict the shape of the CH₂CF₂ molecule by drawing it. What hybridization is involved in the carbon–carbon bonds?
A. Completion
1. stable electron
2. covalent
3. shared
4. single
5. double or triple
6. coordinate
7. equally
8. nonpolar
9. unequally
10. polar
11. electronegativities
12. polar
13. dipole interactions
14. hydrogen bond
15. VSEPR

B. True-False
16. NT
17. AT
18. ST
19. ST
20. AT
21. NT
22. NT
23. NT
24. NT
25. AT
26. ST

C. Questions
27. a. Br : Br
   b. H : C : : N : 
   c. [H : N : H ]

28. a. ionic
   b. covalent, very polar
   c. ionic
   d. covalent, nonpolar
   e. covalent, moderately polar

29. Atoms form covalent bonds by sharing electrons to attain a noble-gas electron configuration.
30. dispersion forces, dipole interactions, hydrogen bonds

32. The hydrogen ion has lost its valence electron and has a positive charge. The nitrogen of the ammonia molecule has an unshared pair of electrons. The positive hydrogen ion is attracted to the unshared pair of electrons to form a coordinate covalent bond with the nitrogen.

D. Additional Questions
33. a. CO₂
   b. CCl₄
   c. H₂O
   d. NH₃

34.

35. Reading from left to right: sp² sp² sp² sp sp sp².

1. Phosphorus has 15 electrons; 2 fill the 1s orbital, 8 fill the 2s and 2p orbitals, and the remaining 5, located in the 3s and 3p orbitals, are its valence electrons. It needs 8 electrons to fill its outer energy level, so it needs 3 more electrons to fill the 3p orbitals. Fluorine has 9 electrons, 2 in the filled 1s and the other 7 (valence) electrons in the 1s² 2p¹ arrangement. Fluorine needs one more electron to fill its second energy level. The phosphorus and fluorine atoms can fulfill each other's needs by sharing electrons in single covalent bonds. Since each fluorine atom only needs one electron and phosphorus needs three electrons, three fluorine atoms are required to bond with phosphorus.

2. The four valence-electron pairs repel each other, but the unshared pair is held closer to the phosphorus than the three bonding pairs. The unshared pair repels the shared pairs more strongly. Thus the angle between bonds is expected to be slightly smaller than the tetrahedral bond angle of 109.5°. The actual bond angle for NH₃, a similar molecule, is 107°.

3. Silicon's bonds involve the s orbital and all three of the p orbitals. The hybridization in SiF₄ is sp³.

4. Nitrogen has 7 electrons; 2 fill the 1s and the remaining 5, located in the 2s and 2p orbitals (2s²2p³), are its valence electrons. It needs 8 electrons in its outer energy level to achieve a noble gas configuration, so it needs 3 more electrons to fill its second energy level. Chlorine has 17 electrons, 2 in the filled 1s orbital, 8 in the filled 2s and 2p orbitals, and the other 7 (valence) electrons in the 3s²3p⁵ arrangement. Chlorine thus needs one more electron to achieve a noble gas configuration. The nitrogen and chlorine atoms can satisfy each other's needs by sharing electrons in single covalent bonds. Because each chlorine atom needs only one electron and nitrogen needs 3 electrons, three chlorine atoms are required to bond with nitrogen.

5. Oxygen is the central atom in this molecule. It has 6 valence electrons, two of which are bonding electrons. The other 4 electrons are located in 2 unshared pairs. These 2 unshared pairs repel the two bonding pairs and prevent CF₄ from being a linear tetraatomic molecule. The molecule is a bent tetraatomic molecule, with a bond angle of approximately 104.5°. This angle is slightly smaller than the tetrahedral bond angle because the two unshared pairs repel each other more strongly than the two shared pairs.

6. Because carbon makes four single covalent bonds, there is an apparent shortage of atoms with which to bond. This is a clue that a carbon–carbon multiple bond exists in this compound. Each carbon atom has a single electron, shares one electron with one of the two hydrogen atoms, which each need one electron each to fill the first energy level and become like helium in electron configuration. The remaining three electrons for each carbon atom form a triple covalent bond. The electron dot structure is as follows.

7. Because each carbon can make single covalent bonds with four other atoms, there exists in this compound an apparent shortage of atoms with which to bond. This is a clue that CF₄ contains a carbon–carbon multiple bond. The two hydrogen atoms bond with the carbon atom on the left while the two fluorine atoms bond with the carbon atom on the right. This is indicated by the positioning of the hydrogen atoms (H₃C) and the two fluorine atoms (F₃C) in the structural formula given. After the carbon atoms have each bonded with two other atoms, each carbon atom has two electrons left over. These two electrons from each atom form a carbon–carbon double covalent bond. Since this gives H–C–H and F–C–F bond angles of 120°, the hybridization involved in the carbon–carbon bond is sp², just as in simpler boron compounds. The molecule looks very much like the ethene molecule (CH₂).