Bonding Practice Test

Multiple Choice
Identify the choice that best completes the statement or answers the question.

____ 1. A mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together is called a(n)
   a. dipole. 
   b. Lewis structure. 
   c. chemical bond. 
   d. London force.

____ 2. When atoms share electrons, the electrical attraction of an atom for the shared electrons is called the atom's
   a. electron affinity. 
   b. electronegativity. 
   c. resonance. 
   d. hybridization.

____ 3. What are shared in a covalent bond?
   a. ions 
   b. Lewis structures 
   c. electrons 
   d. dipoles

____ 4. Most chemical bonds are
   a. purely ionic. 
   b. purely covalent. 
   c. partly ionic and partly covalent. 
   d. metallic.

____ 5. Nonpolar covalent bonds are not common because
   a. one atom usually attracts electrons more strongly than the other. 
   b. ions always form when atoms join. 
   c. the electrons usually remain equally distant from both atoms. 
   d. dipoles are rare in nature.

____ 6. The greater the electronegativity difference between two bonded atoms, the greater the percentage of ____ in the bond.
   a. ionic character 
   b. covalent character 
   c. metallic character 
   d. electron sharing

____ 7. In the three molecules, F₂, HF, and N₂, what atom would have a partial negative charge?
   a. oxygen 
   b. hydrogen 
   c. chlorine 
   d. fluorine

____ 8. After drawing a Lewis structure, one should
   a. determine the number of each type of atom in the molecule. 
   b. add unshared pairs of electrons around nonmetal atoms. 
   c. confirm that the total number of valence electrons used equals the number available. 
   d. determine the electronegativity of each atom.

____ 9. Multiple covalent bonds may occur in atoms that contain carbon, nitrogen, or
   a. chlorine. 
   b. hydrogen. 
   c. oxygen. 
   d. helium.

____ 10. Bonding in molecules or ions that cannot be correctly represented by a single Lewis structure is
   a. polyatomic. 
   b. resonance. 
   c. single bonding. 
   d. double bonding.

____ 11. Chemists once theorized that a molecule that contains a single bond and a double bond split its time existing as one of these two structures. This effect became known as
   a. alternation. 
   b. resonance. 
   c. multiple bonding. 
   d. single-double bonding.

____ 12. The Lewis structure for the ammonium ion, NH₄⁺, has
   a. nonpolar covalent bond. 
   b. ionic bond. 
   c. polar covalent bond. 
   d. metallic bond.

____ 13. How many extra electrons are in the Lewis structure of the phosphate ion, PO₄³⁻?
   a. 0 
   b. 2 
   c. 3 
   d. 4

____ 14. How many electrons must be shown in the Lewis structure of the hydroxide ion, OH⁻?
   a. 1 
   b. 8 
   c. 9 
   d. 10

____ 15. Compared with nonmetals, the number of valence electrons in metals is generally
   a. smaller. 
   b. greater. 
   c. about the same. 
   d. almost triple.

____ 16. To appear shiny, a material must be able to
   a. form crystals. 
   b. absorb and re-emit light of many wavelengths. 
   c. absorb light and change it all to energy as heat. 
   d. change light to electricity.

____ 17. The shiny appearance of a metal is most closely related to the metal's
   a. highly mobile valence electrons. 
   b. covalent bonds. 
   c. brittle crystalline structure. 
   d. positive ions.

____ 18. As light strikes the surface of a metal, the electrons in the electron sea
a. allow the light to pass through.
b. become attached to particular positive ions.
c. fall to lower energy levels.
d. absorb and re-emit the light.

19. If a material can be shaped or extended by physical pressure, such as hammering, which property does the material have?
   a. conductivity  
   b. malleability  
   c. ductility  
   d. luster

20. Metals are malleable because the metallic bonding
   a. holds the layers of ions in rigid positions.
   b. maximizes the repulsive forces within the metal.
   c. allows one plane of ions to slide past another.
   d. is easily broken.

21. Use VSEPR theory to predict the shape of the carbon tetraiodide molecule, ClI₄.
   a. tetrahedral  
   b. linear  
   c. bent  
   d. trigonal-planar

22. Use VSEPR theory to predict the shape of the chlorate ion, ClO₃⁻.
   a. trigonal-planar  
   b. octahedral  
   c. trigonal-pyramidal  
   d. bent

23. Use VSEPR theory to predict the shape of carbon dioxide, CO₂.
   a. tetrahedral  
   b. linear  
   c. bent  
   d. octahedral

24. Which hybrid orbitals help explain the bonding in methane, CH₄?
   a. sp³ orbitals  
   b. sp orbitals  
   c. pd³ orbitals  
   d. df³ orbitals

25. Four hybrid sp³ orbitals are formed from
   a. two s orbitals and two p orbitals.
   b. an s orbital and a p orbital.
   c. three s orbitals and one p orbital.
   d. one s orbital and three p orbitals.

26. The reason the boiling point of water (H₂O) is higher than the boiling point of hydrogen sulfide (H₂S) is partially explained by
   a. London forces.  
   b. covalent bonding.  
   c. ionic bonding.  
   d. hydrogen bonding.

27. The following molecules contain polar bonds. The only polar molecule is
   a. CCl₄  
   b. CO₂  
   c. NH₃  
   d. CH₄.

Short Answer

28. Why do most atoms form chemical bonds?
29. Explain why scientists use resonance structures to represent some molecules.
30. Differentiate between an ionic compound and a molecular compound.
31. Explain why metals are good conductors of electricity.

Problem

32. Draw a Lewis structure for carbon disulfide, CS₂.
33. Draw a Lewis structure for the nitrate ion, NO₃⁻. Use VSEPR theory to predict its molecular geometry.
Bonding Practice Test
Answer Section

MULTIPLE CHOICE

1. ANS: C
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 1

2. ANS: B
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 3

3. ANS: C
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 3

4. ANS: C
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 4

5. ANS: A
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 4

6. ANS: A
   PTS: 1
   DIF: I
   REF: 1
   OBJ: 4

7. ANS: C
   PTS: 1
   DIF: III
   REF: 1
   OBJ: 5

8. ANS: C
   PTS: 1
   DIF: I
   REF: 2
   OBJ: 4

9. ANS: C
   PTS: 1
   DIF: II
   REF: 2
   OBJ: 5

10. ANS: B
    PTS: 1
    DIF: I
    REF: 2
    OBJ: 5

11. ANS: B
    PTS: 1
    DIF: I
    REF: 2
    OBJ: 5

12. ANS: C
    PTS: 1
    DIF: I
    REF: 3
    OBJ: 4

13. ANS: C
    PTS: 1
    DIF: III
    REF: 3
    OBJ: 4

14. ANS: B
    PTS: 1
    DIF: III
    REF: 3
    OBJ: 4

15. ANS: A
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 1

16. ANS: B
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 2

17. ANS: A
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 2

18. ANS: D
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 2

19. ANS: B
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 3

20. ANS: C
    PTS: 1
    DIF: I
    REF: 4
    OBJ: 3

21. ANS: A
    PTS: 1
    DIF: III
    REF: 5
    OBJ: 2

22. ANS: C
    PTS: 1
    DIF: III
    REF: 5
    OBJ: 2

23. ANS: B
    PTS: 1
    DIF: III
    REF: 5
    OBJ: 2

24. ANS: A
    PTS: 1
    DIF: I
    REF: 5
    OBJ: 3

25. ANS: D
    PTS: 1
    DIF: II
    REF: 5
    OBJ: 3

26. ANS: D
    PTS: 1
    DIF: II
    REF: 5
    OBJ: 4

27. ANS: C
    PTS: 1
    DIF: III
    REF: 5
    OBJ: 5

SHORT ANSWER

28. ANS:
Atoms form chemical bonds to establish a more-stable arrangement. As independent particles, they are at high potential energy. By bonding, they decrease their potential energy, thus becoming more stable.

29. ANS:
Resonance structures represent the bonding in molecules that cannot be adequately represented with a single Lewis structure.

30. ANS:
Atoms in a molecular compound share electrons to achieve stability. Atoms in an ionic compound gain or lose electrons to form ions, which combine so that the number of positive and negative charges is equal.

31. ANS:
The valence electrons in a metal's structure are delocalized, so they can move freely and carry an electric charge throughout the metal.

PROBLEM

32. ANS:
1 C atom with 4 valence electrons ⇒ 1 × 4 = 4
2 S atoms with 6 valence electrons ⇒ 2 × 6 = 12
4 + 12 = 16 valence electrons

\[ \text{\ :\text{S}} \quad \text{C} \quad \text{\ :\text{S}} \]

33. ANS:

\[
\begin{bmatrix}
\vdots \\
\text{\ :\text{O}} \\
\text{\ :\text{N}} \\
\text{\ :\text{O}} \\
\vdots
\end{bmatrix}^{\text{\text{-}}}
\]

Nitrate ion

trigonal-planar