

Chapter 15 and Chapter 16 Exam Multiple Choice

Multiple Choice

Identify the letter of 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. The electrons involved in the formation of a chemical bond are called
 - a. dipoles.
 - b. s electrons.
 - c. Lewis electrons.
 - d. valence electrons.
3. The electrostatic attraction between positively charged nuclei and negatively charged electrons permits two atoms to be held together by a(n)
 - a. chemical bond.
 - b. London force.
 - c. neutron.
 - d. ion.
4. Atoms are _____ when they are combined.
 - a. more stable
 - b. less stable
 - c. not bound together
 - d. at a high potential energy
5. Atoms naturally move
 - a. toward high potential energy.
 - b. toward low potential energy.
 - c. toward less stability.
 - d. away from each other.
6. The chemical bond formed when two atoms share electrons is called a(n)
 - a. ionic bond.
 - b. orbital bond.
 - c. Lewis structure.
 - d. covalent bond.
7. If two covalently bonded atoms are identical, the bond is
 - a. nonpolar covalent.
 - b. polar covalent.
 - c. nonionic.
 - d. coordinate covalent.
8. When atoms share electrons, the electrical attraction of an atom for the electrons is called the atom's
 - a. electron affinity.
 - b. electronegativity.
 - c. resonance.
 - d. hybridization.
9. If the atoms that share electrons have an unequal attraction for the electrons, the bond is called
 - a. nonpolar.
 - b. polar.
 - c. ionic.
 - d. dipolar.
10. The electrostatic attraction between _____ forms an ionic bond.
 - a. ions
 - b. dipoles
 - c. electrons
 - d. orbitals
11. Most chemical bonds are
 - a. purely ionic.
 - b. purely covalent.
 - c. partly ionic and partly covalent.
 - d. metallic.
12. 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.
13. The greater the electronegativity difference between two bonded atoms, the greater the percentage of
 - a. ionic character.
 - b. covalent character.
 - c. metallic character.
 - d. electron sharing.

14. Bonds that are between 5% and 50% ionic are considered
- ionic.
 - pure covalent.
 - polar covalent.
 - nonpolar covalent.
15. Bonds that are more than 50% ionic are considered
- polyatomic.
 - polar covalent.
 - ionic.
 - nonpolar covalent.
16. A bond that is less than 5% ionic is considered
- polar covalent.
 - ionic.
 - nonpolar covalent.
 - metallic.
17. The pair of elements that forms a bond with the least ionic character is
- Na and Cl.
 - H and Cl.
 - O and Cl.
 - Br and Cl.
18. The B—F bond in BF_3 (electronegativity for B is 2.0; electronegativity for F is 4.0) is
- polar covalent.
 - ionic.
 - nonpolar covalent.
 - pure covalent.
19. The percentage ionic character for the C—F bond (electronegativity for C is 2.5; electronegativity for F is 4.0) in CF_4 is
- 9%.
 - 15%.
 - 38%.
 - 62%.
20. The percentage ionic character and the type of bond in Br_2 (electronegativity for Br is 2.8) is
- 0%; nonpolar covalent.
 - 100%; polar covalent.
 - 0%; pure ionic.
 - 100%; pure ionic.
21. The percentage ionic character and the type of bond in LiCl (electronegativity for Li is 1.0; electronegativity for Cl is 3.0) is
- 50%; ionic.
 - 50%; polar covalent.
 - 67%; polar covalent.
 - 67%; ionic.
22. In which of these compounds is the bond between the atoms NOT a nonpolar covalent bond?
- Cl_2
 - H_2
 - HCl
 - O_2
23. Which of the following is NOT an example of a molecular formula?
- H_2O
 - B
 - NH_3
 - O_2
24. When a covalent bond forms, as the distance between two atoms decreases, the potential energy
- increases.
 - decreases.
 - remains constant.
 - becomes zero.
25. A covalent bond forms when the attraction between two atoms is balanced by repulsion and the potential energy is
- at a maximum.
 - zero.
 - at a minimum.
 - equal to the kinetic energy.
26. Bond length is
- the separation for which potential energy is at a minimum.
 - the separation for which kinetic energy is at a maximum.
 - the separation for which potential energy is at a maximum.
 - one-half the diameter of the electron cloud.
27. In a molecule of fluorine, the two shared electrons give each fluorine atom ____ electron(s) in the outer energy level.
- 1
 - 2
 - 8
 - 32

28. When the octet rule is satisfied, the outermost _____ are filled.
- d* and *f* orbitals
 - s* and *p* orbitals
 - s* and *d* orbitals
 - d* and *p* orbitals
29. In drawing a Lewis structure, the central atom is the
- atom with the greatest mass.
 - atom with the highest atomic number.
 - atom with the fewest electrons.
 - least electronegative atom.
30. To draw a Lewis structure, it is NOT necessary to know
- bond energies.
 - the types of atoms in the molecule.
 - the number of valence electrons for each atom.
 - the number of atoms in the molecule.
31. If, after drawing a Lewis structure, too many valence electrons have been used, the molecule probably contains
- too many atoms.
 - one or more multiple covalent bonds.
 - too many unshared pairs of electrons.
 - an ionic bond.
32. Multiple covalent bonds may occur in atoms that contain carbon, nitrogen, or
- chlorine.
 - hydrogen.
 - oxygen.
 - helium.
33. The substance whose Lewis structure shows three covalent bonds is
- H₂O.
 - CH₂Cl₂.
 - NH₃.
 - CCl₄.
34. How many double bonds are in the Lewis structure for hydrogen fluoride, HF?
- none
 - one
 - two
 - three
35. How many extra electrons are in the Lewis structure of the phosphate ion, PO₄³⁻?
- 0
 - 2
 - 3
 - 4
36. How many electrons must be shown in the Lewis structure of the hydroxide ion, OH⁻?
- 1
 - 8
 - 9
 - 10
-



37. What is the Lewis structure for hydrogen chloride, HCl?
- A
 - B
 - C
 - D
-

- b. positive ions to form.
 - c. these pairs to be separated as far as possible.
 - d. light to reflect.
49. According to VSEPR theory, the structure of the ammonia molecule, NH_3 , is
- a. linear.
 - b. bent.
 - c. pyramidal.
 - d. tetrahedral.
50. Use VSEPR theory to predict the shape of the hydrogen chloride molecule, HCl .
- a. tetrahedral
 - b. linear
 - c. bent
 - d. trigonal planar
51. Use VSEPR theory to predict the shape of the carbon tetraiodide molecule, CI_4 .
- a. tetrahedral
 - b. linear
 - c. bent
 - d. trigonal planar
52. Use VSEPR theory to predict the shape of the chlorate ion, ClO_3^- .
- a. trigonal planar
 - b. octahedral
 - c. trigonal pyramidal
 - d. bent
53. Use VSEPR theory to predict the shape of the hydrogen sulfide molecule, H_2S .
- a. tetrahedral
 - b. linear
 - c. bent
 - d. octahedral
54. Use VSEPR theory to predict the shape of carbon dioxide, CO_2 .
- a. tetrahedral
 - b. linear
 - c. bent
 - d. octahedral
55. Dipole-dipole forces are considered the most important forces in polar substances because the London dispersion forces
- a. act only in nonpolar substances.
 - b. are usually much weaker than the dipole-dipole forces.
 - c. are too unpredictable.
 - d. act only in solids.
56. The strength of London dispersion forces between molecules depends on
- a. only the number of electrons in the molecule.
 - b. only the number of protons in the molecule.
 - c. both the number of electrons in the molecule and the mass of the molecule.
 - d. both the number of electrons and the number of neutrons in the molecule.
57. The strong forces of attraction between the positive and negative regions of molecules are called
- a. dipole-dipole forces.
 - b. London forces.
 - c. lattice forces.
 - d. orbital forces.
58. The intermolecular attraction between a hydrogen atom bonded to a strongly electronegative atom and the unshared pair of electrons on another strongly electronegative atom is called
- a. electron affinity.
 - b. covalent bonding.
 - c. hydrogen bonding.
 - d. electronegativity.
59. The weak intermolecular forces resulting from instantaneous and induced dipoles are called
- a. London dispersion forces.
 - b. dipole-dipole forces.
 - c. hydrogen forces.
 - d. polar covalent bonding.
60. Compared with molecular bonds, the strength of intermolecular forces is
- a. weaker.
 - b. stronger.
 - c. about the same.
 - d. too variable to compare.
61. The equal but opposite charges present in the two regions of a polar molecule create a(n)
- a. electron sea.
 - b. dipole.
 - c. crystal lattice.
 - d. ionic bond.

62. That the boiling point of water (H_2O) is higher than the boiling point of hydrogen sulfide (H_2S) is partially explained by
- London forces.
 - covalent bonding.
 - ionic bonding.
 - hydrogen bonding.
63. The following molecules contain polar bonds. The only nonpolar molecule is
- HCl .
 - H_2O .
 - CO_2 .
 - NH_3 .
64. A polar molecule contains
- ions.
 - a region of positive charge and a region of negative charge.
 - only London forces.
 - no bonds.
65. A molecule of hydrogen chloride is polar because
- it is composed of ions.
 - it is magnetic.
 - it contains metallic bonds.
 - the chlorine attracts the shared electrons more strongly than does the hydrogen atom.
66. When a polar molecule attracts the electron in a nonpolar molecule,
- a dipole is induced.
 - a crystal lattice forms.
 - an ionic bond forms.
 - a Lewis structure forms.
67. Iodine monochloride (ICl) has a higher boiling point than bromine (Br_2) partly because iodine monochloride is a(n)
- nonpolar molecule.
 - ion.
 - crystal.
 - polar molecule.

Problems

- Draw a Lewis structure for the oxalate ion, $\text{C}_2\text{O}_4^{2-}$.
- Draw a Lewis structure for the ammonium ion, NH_4^+ .
- Draw a Lewis structure for the nitrate ion, NO_3^- .
- Draw a Lewis structure for the sulfate ion, SO_4^{2-} .
- Draw a Lewis structure for the phosphate ion, PO_4^{3-} .
- For the following chemical reaction, determine if it is EXOTHERMIC (gives off energy) or ENDOTHERMIC (requires energy for the reaction to occur). Use the table below to determine bond dissociation energies. **5 points**



- Balance the equation. Remember to be watchful of the coefficients as you do your math in the next steps.
- Draw the structural formulas. Determine which bonds get broken and which get formed.
- Then, using the table for the bond dissociation energies, show all your arithmetic and calculations.

D. State whether the reaction is exothermic or endothermic using the numbers you have calculated.

Bond Energies Table

<i>Bond</i>	<i>Bond Energy kJ/mole</i>	<i>Bond</i>	<i>Bond Energy kJ/mole</i>
H-H	436	N=N	945
C-H	414	H-F	570
C-C	347	H-Cl	431
C=C	619	H-Br	364
C≡C	812	H-I	298
C-O	335	F-F	158
C=O	707	Cl-Cl	243
O-O	138	Br-Br	194
O=O	498	I-I	153
O-H	463	C-Cl	326
N-H	389	C-Br	276
N-N	159		

**Chapter 15 and Chapter 16 Exam Multiple Choice
Answer Section**

MULTIPLE CHOICE

1. C
2. D
3. A
4. A
5. B
6. D
7. A
8. B
9. B
10. A
11. C
12. A
13. A
14. C
15. C
16. C
17. D
18. B
19. C
20. A
21. D
22. C
23. B
24. B
25. C
26. A
27. C
28. B
29. D
30. A
31. B
32. C
33. C
34. A
35. C
36. B
37. D
38. B
39. B
40. A
41. C

42. A
 43. B
 44. D
 45. D
 46. D
 47. B
 48. C
 49. C
 50. B
 51. A
 52. C
 53. C
 54. B
 55. B
 56. C
 57. A
 58. C
 59. A
 60. A
 61. B
 62. D
 63. C
 64. B
 65. D
 66. A
 67. D

PROBLEMS

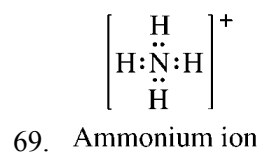
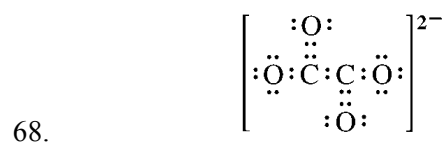
2 C atoms with 4 electrons $\Rightarrow 2 \times 4 = 8$

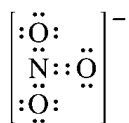
4 O atoms with 6 electrons $\Rightarrow 4 \times 6 = 24$

$8 + 24 = 32$ valence electrons

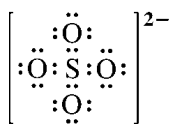
2- charge $\Rightarrow 2$ extra electrons

$32 + 2 = 34$ electrons total

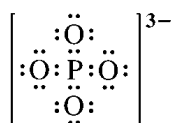




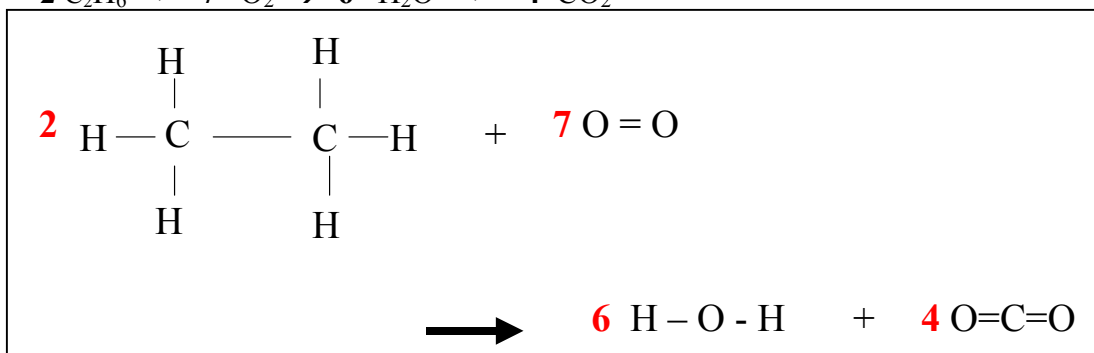
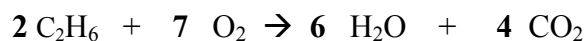
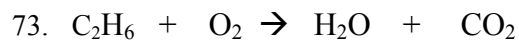
70. Nitrate ion



71. Sulfate ion



72. Phosphate ion



2 (6) 414 kJ for C-H bonds broken

2 (1) 347 kJ for C-C bonds broken

7(1) 498 kJ for O=O bonds broken

9148 kJ used up to break bonds

compared to.....

6(2) 463 for O-H bonds formed

4(2) 707 for C=O bonds formed

11212 kJ released when bonds are formed

So..... 11212 kJ – 9148 kJ = **2064 kJ given off: EXOTHERMIC REACTION**

VSEPR structures (structures NOT drawn here)

BBr_3 3 pair of electrons, trigonal planar shape

H_2O 4 pair of electrons, angular shape

XeBr_4 6 pair of electrons, square planar shape

PCl_5 5 pair of electrons, trigonal bipyramidal shape

SiCl_4 4 pair of electrons, tetrahedral shape

NF_3 4 pair of electrons, trigonal pyramidal shape

XeF_5^{+1} 6 pair of electrons, square pyramidal shape

CO_2 2 pair of electrons, linear shape (remember that Oxygen is NOT a halogen!)