AP Chapter 8 Study Questions

True/False
*Indicate whether the statement is true or false.*

1. Atoms surrounded by eight valence electrons tend to lose electrons.

2. The greater the lattice energy, the greater the charges on the participatory ions and the smaller their radii.

3. Most transition metals do not form ions with a noble gas configuration.

4. When a metal gains an electron, the process is endothermic.

5. Electron affinity is a measure of how strongly an atom can attract additional electrons.

6. As electronegativity difference increases, bond length will decrease.

7. In some molecules and polyatomic ions, the sum of the valence electrons is odd and as a result the octet rule fails.

8. Bond enthalpy can be positive or negative.

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

9. There are _______ paired and _______ unpaired electrons in the Lewis symbol for a phosphorus atom.
   a. 4, 2  b. 2, 4  c. 2, 3  d. 4, 3  e. 0, 3

10. In the Lewis symbol for a fluorine atom, there are _______ paired and _______ unpaired electrons.
    a. 4, 2  b. 4,1  c. 2, 5  d. 6, 1  e. 0, 5

11. Based on the octet rule, magnesium most likely forms a _______ ion.
    a. Mg²⁺  b. Mg⁰  c. Mg⁶⁺  d. Mg⁶⁺  e. Mg⁺

12. Based on the octet rule, phosphorus most likely forms a _______ ion.
    a. P³⁺  b. P⁰  c. P⁵⁺  d. P⁵⁺  e. P⁺

13. Based on the octet rule, iodine most likely forms an _______ ion.
    a. I²⁺  b. I⁴⁺  c. I⁺  d. I⁺  e. I⁻

14. There are _______ unpaired electrons in the Lewis symbol for an oxygen atom.
    a. 0  b. 1  c. 2  d. 4  e. 3

15. How many unpaired electrons are there in the Lewis structures of a N³⁻ ion?
    a. 0  b. 1  c. 2  d. 3  e. This cannot be predicted.

16. How many unpaired electrons are there in an O²⁻ ion?
    a. 0  b. 1  c. 2  d. 3  e. This cannot be predicted.

17. The electron configuration of the phosphide ion (P³⁻) is _______.
    a. [Ne]3s²  b. [Ne]3s²3p¹  c. [Ne]3s²3p³  d. [Ne]3p²  e. [Ne]3s²3p⁶
18. The halogens, alkali metals, and alkaline earth metals have __________ valence electrons, respectively.
   a. 7, 4, and 6  b. 1, 5, and 7  c. 8, 2, and 3  d. 7, 1, and 2  e. 2, 7, and 4

19. The only noble gas without eight valence electrons is __________.
   a. Ar  b. Ne  c. He  d. Kr  e. All noble gases have eight valence electrons.

20. Which of the following would have to lose two electrons in order to achieve a noble gas electron configuration?
   O Sr Na Se Br

21. Which of the following would have to gain two electrons in order to achieve a noble gas electron configuration?
   O Sr Na Se Br

22. For a given arrangement of ions, the lattice energy increases as ionic radius ________ and as ionic charge ________.
   a. decreases, increases  b. increases, decreases  c. increases, increases  d. decreases, decreases  e. This cannot be predicted.

23. The electron configuration of the S$^{2-}$ ion is
   a. [Ar]3s$^2$3p$^6$  b. [Ar]3s$^2$3p$^2$  c. [Ne]3s$^2$3p$^2$  
   d. [Ne]3s$^2$3p$^6$  e. [Kr]3s$^2$2p$^6$

24. The principal quantum number of the electrons that are lost when tungsten forms a cation is ________.
   a. 6  b. 5  c. 4  d. 3  e. 2

25. Which of the following species has the electron configuration [Ar]3d$^6$?
   a. Mn$^{2+}$  b. Cr$^{3+}$  c. V$^{3+}$  d. Fe$^{3+}$  e. K$^+$

26. What is the electron configuration for the Co$^{2+}$ ion?
   a. [Ar]4s$^2$3d$^6$  b. [Ar]4s$^2$3d$^7$  c. [Ar]4s$^2$3d$^5$
   d. [Ar]4s$^2$3d$^9$  e. [Ne]3s$^2$3p$^{10}$

27. What is the electron configuration for the Fe$^{2+}$ ion?
   a. [Ar]4s$^2$3d$^6$  b. [Ar]4s$^2$3d$^4$  c. [Ar]4s$^2$3d$^8$
   d. [Ar]4s$^2$3d$^8$  e. [Ar]4s$^2$3d$^2$

28. The formula of palladium(IV) sulfide is ________.
   a. Pd$_2$S$_4$  b. PdS$_4$  c. Pd$_4$S  d. PdS$_2$  e. Pd$_2$S$_2$

29. Elements from opposite sides of the periodic table tend to form ________.
   a. covalent compounds  b. ionic compounds  c. compounds that are gaseous at room temperature  d. homonuclear diatomic compounds  e. covalent compounds that are gaseous at room temperature

30. Determining lattice energy from Born-Haber cycle data requires the use of ________.
   a. the octet rule  b. Coulomb's law  c. Periodic law  d. Hess's law  e. Avogadro's number

31. How many single covalent bonds must a silicon atom form to have a complete octet in its valence shell?
   a. 3  b. 4  c. 1  d. 2  e. 0

32. A ________ covalent bond between the same two atoms is the longest.
   a. single  b. double  c. triple  d. They are all the same length  e. strong

33. How many hydrogen atoms must bond to silicon to give it an octet of valence electrons?
   a. 1  b. 2  c. 3  d. 4  e. 5

34. A double bond consists of ________ pairs of electrons shared between two atoms.
   a. 1  b. 2  c. 3  d. 4  e. 6

35. What is the maximum number of double bonds that a hydrogen atom can form?
   a. 0  b. 1  c. 2  d. 3  e. 4
36. What is the maximum number of double bonds that a carbon atom can form?
   a. 4   b. 1   c. 0   d. 2   e. 3

37. In the molecule below, which atom has the largest partial negative charge __________?

   ![Molecule Diagram]


38. The ability of an atom in a molecule to attract electrons is best quantified by the __________.
   a. paramagnetism   b. diamagnetism   c. electronegativity   d. electron change-to-mass ratio   e. first ionization potential

39. Given the electronegativities below, which covalent single bond is most polar?

<table>
<thead>
<tr>
<th>Element</th>
<th>H</th>
<th>C</th>
<th>N</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electronegativity</td>
<td>2.1</td>
<td>2.5</td>
<td>3.0</td>
<td>3.5</td>
</tr>
</tbody>
</table>


40. Electronegativity __________ from left to right within a period and __________ from top to bottom within a group.
   a. decreases, increases   b. increases, increases   c. increases, decreases   d. stays the same, increases   e. increases, stays the same

41. A nonpolar bond will form between two __________ atoms of __________ electronegativity.
   a. different, opposite   b. identical, different   c. different, different   d. similar, different   e. identical, equal

42. The ion ICl₄⁻ has __________ valence electrons.
   a. 34   b. 35   c. 36   d. 28   e. 8

43. The ion NO₃⁻ has __________ valence electrons.
   a. 15   b. 14   c. 16   d. 10   e. 12

44. The Lewis structure of AsH₃ shows __________ nonbonding electron pair(s) on As.
   a. 0   b. 1   c. 2   d. 3   e. This cannot be determined from the data given.

45. The Lewis structure of PF₃ shows that the central phosphorus atom has __________ nonbonding and __________ bonding electron pairs.
   a. 2, 2   b. 1, 3   c. 3, 1   d. 1, 2   e. 3, 3

46. The Lewis structure of HCN (H bonded to C) shows that __________ has __________ nonbonding electron pairs.

47. The formal charge on carbon in the molecule below is __________.

   ![Carbon Molecule Diagram]

   a. 0   b. +1   c. +2   d. +3   e. -1

48. The formal charge on nitrogen in NO₃⁻ is __________.

   ![Nitrogen Molecule Diagram]

   a. -1   b. 0   c. +1   d. +2   e. -2

49. The formal charge on sulfur in SO₄²⁻ is __________, where the Lewis structure of the ion is:

   ![Sulfur Molecule Diagram]

   a. -2   b. 0   c. +2   d. +4   e. -4
50. In the Lewis structure of ClF, the formal charge on Cl is ________ and the formal charge on F is ________.
   a. -1, -1  b. 0, 0  c. 0, -1  d. +1, -1  e. -1, +1

51. In the resonance form of ozone shown below, the formal charge on the central oxygen atom is ________.
   a. 0  b. +1  c. -1  d. +2  e. -2

52. How many equivalent resonance forms can be drawn for CO$_3^{2-}$ (carbon is the central atom)?
   a. 1  b. 2  c. 3  d. 4  e. 0

53. How many equivalent resonance forms can be drawn for SO$_3^{2-}$ without expanding octet on the sulfur atom (sulfur is the central atom)?
   a. 0  b. 2  c. 3  d. 4  e. 1

54. How many equivalent resonance structures can be drawn for the molecule of SO$_3$ without having to violate the octet rule on the sulfur atom?
   a. 5  b. 2  c. 1  d. 4  e. 3

55. How many different types of resonance structures can be drawn for the ion SO$_3^{2-}$ where all atoms satisfy the octet rule?
   a. 1  b. 2  c. 3  d. 4  e. 5

56. Using the table of average bond energies below, the $\Delta H$ for the reaction is ________ kJ.

   Bond: C≡C  C=C  H-I  C-I  C-H
   D (kJ/mol): 839 348 299 240 413

   a. +160  b. -160  c. -217  d. -63  e. +63

57. Using the table of average bond energies below, the $\Delta H$ for the reaction is ________ kJ.

   \[ H - C≡C−H (g) + H−I (g) \rightarrow H_2C≡HI (g) \]

   Bond: C≡C  C=C  H-I  C-I  C-H
   D (kJ/mol): 839 614 299 240 413

   a. +506  b. -931  c. -506  d. -129  e. +129
58. Using the table of average bond energies below, the $\Delta H$ for the reaction is _______ kJ.

\[
\text{C≡O (g) + 2H}_2\text{(g)} \rightarrow \text{H}_3\text{C–O–H (g)}
\]

<table>
<thead>
<tr>
<th>Bond</th>
<th>C-O</th>
<th>C=O</th>
<th>C≡O</th>
<th>C-H</th>
<th>H-H</th>
<th>O-H</th>
</tr>
</thead>
<tbody>
<tr>
<td>$D$ (kJ/mol)</td>
<td>358</td>
<td>799</td>
<td>1072</td>
<td>413</td>
<td>436</td>
<td>463</td>
</tr>
</tbody>
</table>

a. +276  b. -276  c. +735  d. -735  e. -116

59. Using the table of bond dissociation energies, the $\Delta H$ for the following gas-phase reaction is _______ kJ.

\[
\text{H} = \text{C} = \text{C} + \text{H–Cl} \rightarrow \text{H} = \text{C} = \text{C–Cl}
\]

<table>
<thead>
<tr>
<th>Bond</th>
<th>$D$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C=C</td>
<td>348</td>
</tr>
<tr>
<td>C–C</td>
<td>614</td>
</tr>
<tr>
<td>C–H</td>
<td>413</td>
</tr>
<tr>
<td>H–Cl</td>
<td>431</td>
</tr>
<tr>
<td>C–Cl</td>
<td>328</td>
</tr>
</tbody>
</table>

a. -44  b. 38  c. 304  d. 2134  e. -38

60. Using the table of bond dissociation energies, the $\Delta H$ for the following gas-phase reaction is _______ kJ.

\[
\text{H} = \text{C} = \text{C} + \text{H–Br} \rightarrow \text{H} = \text{C} = \text{C–Br}
\]

<table>
<thead>
<tr>
<th>Bond</th>
<th>$D$ (kJ/mol)</th>
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<tr>
<td>C=C</td>
<td>348</td>
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</tr>
<tr>
<td>C–H</td>
<td>413</td>
</tr>
<tr>
<td>H–Br</td>
<td>366</td>
</tr>
<tr>
<td>C–Br</td>
<td>276</td>
</tr>
</tbody>
</table>

a. 291  b. 2017  c. -57  d. -356  e. -291
61. Using the table of bond dissociation energies, the \( \Delta H \) for the following reaction is \__________\ kJ.

\[
2\text{HCl (g)} + \text{F}_2 (g) \rightarrow 2\text{HF (g)} + \text{Cl}_2 (g)
\]

<table>
<thead>
<tr>
<th>Bond</th>
<th>( D ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–Cl</td>
<td>431</td>
</tr>
<tr>
<td>F–F</td>
<td>155</td>
</tr>
<tr>
<td>H–F</td>
<td>567</td>
</tr>
<tr>
<td>Cl–Cl</td>
<td>242</td>
</tr>
</tbody>
</table>

a. -359   b. -223   c. 359   d. 223   e. 208

62. Which ion below has a noble gas electron configuration?
   a. \( \text{Li}^{2+} \) b. \( \text{Be}^{2+} \) c. \( \text{B}^{2+} \) d. \( \text{C}^{2+} \) e. \( \text{N}^{-} \)

63. Of the ions below, only \__________\ has a noble gas electron configuration.
   a. \( \text{S}^{3-} \) b. \( \text{O}^{2+} \) c. \( \text{I}^{+} \) d. \( \text{K}^{-} \) e. \( \text{Cl}^{-} \)

64. Which of the following has eight valence electrons?
   a. \( \text{Ti}^{4+} \) b. \( \text{Kr} \) c. \( \text{Cl}^{-} \) d. \( \text{Na}^{+} \) e. all of the above

65. Which of the following does not have eight valence electrons?
   a. \( \text{Ca}^{+} \) b. \( \text{Rb}^{+} \) c. \( \text{Xe} \) d. \( \text{Br}^{-} \) e. All of the above have eight valence electrons.

66. The chloride of which of the following metals should have the greatest lattice energy?
   a. potassium b. rubidium c. sodium d. lithium e. cesium

67. Lattice energy is \__________\.
   a. the energy required to convert a mole of ionic solid into its constituent ions in the gas phase
   b. the energy given off when gaseous ions combine to form one mole of an ionic solid
   c. the energy required to produce one mole of an ionic compound from its constituent elements in their standard states
   d. the sum of ionization energies of the components in an ionic solid
   e. the sum of electron affinities of the components in an ionic solid

68. In ionic bond formation, the lattice energy of ions \__________\ as the magnitude of the ion charges \__________\ and the radii \__________\.
   a. increases, decrease, increase   b. increases, increase, increase   c. decreases, increase, increase
   d. increases, increase, decrease   e. increases, decrease, decrease

69. The diagram below is the Born-Huber cycle for the formation of crystalline potassium fluoride.

Which energy change corresponds to the electron affinity of fluorine?
   a. 2   b. 5   c. 4   d. 1   e. 6
70. The diagram below is the Born-Huber cycle for the formation of crystalline potassium fluoride.

Which energy change corresponds to the first ionization energy of potassium?

a. 2 b. 5 c. 4 d. 3 e. 6

71. The electron configuration [Kr]4d\(^{10}\) represents ________.

a. Sr\(^{2+}\) b. Sn\(^{2+}\) c. Te\(^{2+}\) d. Ag\(^{+}\) e. Rb\(^{+}\)

72. Fe\(^{2+}\) ions are represented by ________.

a. [Ar]3d\(^3\) b. [Ar]3d\(^4\) c. [Ar]3d\(^6\) d. [Ar]3d\(^{10}\)4s\(^1\) e. [Ar]3d\(^3\)

73. Using the Born-Haber cycle, the \(\Delta H\)\(^f\) of KBr is equal to ________.

a. \(\Delta H\)\(^f\)[K (g)] + \(\Delta H\)\(^f\)[Br (g)] + \(I\)\(_1\)(K) + E(Br) + \(\Delta H\)\(_{lattice}\) b. \(\Delta H\)\(^f\)[K (g)] - \(\Delta H\)\(^f\)[Br (g)] - \(I\)\(_1\)(K) - E(Br) - \(\Delta H\)\(_{lattice}\) c. \(\Delta H\)\(^f\)[K (g)] - \(\Delta H\)\(^f\)[Br (g)] + \(I\)\(_1\)(K) - E(Br) + \(\Delta H\)\(_{lattice}\) d. \(\Delta H\)\(^f\)[K (g)] + \(\Delta H\)\(^f\)[Br (g)] - \(I\)\(_1\) - E(Br) + \(\Delta H\)\(_{lattice}\) e. \(\Delta H\)\(^f\)[K (g)] + \(\Delta H\)\(^f\)[Br (g)] + \(I\)\(_1\)(K) + E(Br) - \(\Delta H\)\(_{lattice}\)

74. The type of compound that is most likely to contain a covalent bond is ________.

a. one that is composed of a metal from the far left of the periodic table and a nonmetal from the far right of the periodic table b. a solid metal c. one that is composed of only nonmetals d. held together by the electrostatic forces between oppositely charged ions e. There is no general rule to predict covalency in bonds.

75. In which of the molecules below is the carbon-carbon distance the shortest?

a. H\(_2\)C=CH\(_2\) b. H-C≡C-H c. H\(_2\)C-CH\(_3\) d. H\(_2\)C=C-CH\(_2\) e. H\(_2\)C-CH\(_2\)-CH\(_3\)

76. Of the atoms below, ________ is the most electronegative.


77. Of the atoms below, ________ is the most electronegative.

a. Si b. Cl c. Rb d. Ca e. S

78. Of the atoms below, ________ is the least electronegative.

a. Rb b. F c. Si d. Cl e. Ca

79. Which of the elements below has the largest electronegativity?

a. Si b. Mg c. P d. S e. Na

80. Of the molecules below, the bond in ________ is the most polar.

a. HBr b. HI c. HCl d. HF e. H\(_2\)

81. Of the bonds below, ________ is the least polar.


82. Which of the following has the bonds correctly arranged in order of increasing polarity?

a. Be\(_{2}\), Mg\(_{2}\), N\(_2\), O\(_2\) b. O\(_2\), N\(_2\), Be\(_{2}\), Mg\(_{2}\) c. O\(_2\), Be\(_{2}\), Mg\(_{2}\), N\(_2\) d. N\(_2\), Be\(_{2}\), Mg\(_{2}\), O\(_2\) e. Mg\(_{2}\), Be\(_{2}\), N\(_2\), O\(_2\)
83. Which two bonds are most similar in polarity?
   a. O-F and Cl-F  b. B-F and Cl-F  c. Al-Cl and I-Br  d. I-Br and Si-Cl  e. Cl-Cl and Be-Cl

84. The bond length in an HI molecule is 1.61 Å and the measured dipole moment is 0.44 D. What is the magnitude (in units of e) of the negative charge on I in HI?
   (1 debye = 3.34 × 10⁻³⁰ coulomb-meters; e=1.6 × 10⁻¹⁹ coulombs)
   a. 1.6 × 10⁻¹⁹ b. 0.057 c. 9.1 d. 1 e. 0.22

85. Which of the following names is/are correct for the compound TiO₂?
   a. titanium dioxide and titanium (IV) oxide  b. titanium (IV) dioxide  c. titanium oxide  d. titanium oxide and titanium (IV) dioxide  e. titanium (II) oxide

86. Which of the following names is/are correct for the compound SnCl₄?
   a. tin (II) chloride and tin (IV) chloride  b. tin tetrachloride and tin (IV) chloride  c. tin (IV) tetrachloride  d. tin chloride  e. tin chloride and tin (II) tetrachloride

87. The Lewis structure of N₂H₂ shows ________.
   a. a nitrogen-nitrogen triple bond  b. a nitrogen-nitrogen single bond  c. each nitrogen has one nonbonding electron pair  d. each nitrogen has two nonbonding electron pairs  e. each hydrogen has one nonbonding electron pair

88. There are ________ valence electrons in the Lewis structure of CH₃CH₂Cl.
   a. 14 b. 12 c. 18 d. 20 e. 10

89. In the Lewis symbol for a sulfur atom, there are ________ paired and ________ unpaired electrons.
   a. 2, 2  b. 4, 2  c. 2, 4  d. 0, 6  e. 5, 1

90. The Lewis structure of the CO₃²⁻ ion is ________.
   a.  
   b.  
   c.  
   d.  
   e.  
   f.  
   g.  
   h.  
   i.  
   j.  
   k.  
   l.  
   m.  
   n.  
   o.  
   p.  
   q.  
   r.  
   s.  
   t.  
   u.  
   v.  
   w.  
   x.  
   y.  
   z.  
   A.  
   B.  
   C.  
   D.  
   E.  
   F.  
   G.  
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   P.  
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   R.  
   S.  
   T.  
   U.  
   V.  
   W.  
   X.  
   Y.  
   Z.
91. In the nitrite ion (NO$_2^-$), __________.
   a. both bonds are single bonds   b. both bonds are double bonds   c. one bond is a double bond and the other is a single bond   d. both bonds are the same   e. there are 20 valence electrons

92. Resonance structures differ by __________.
   a. number and placement of electrons   b. number of electrons only   c. placement of atoms only   d. number of atoms only   e. placement of electrons only

93. The oxidation number of phosphorus in PF$_3$ is __________.
   a. -2 b. +1 c. +3 d. +2 e. -3

94. To convert from one resonance structure to another, __________.
   a. only atoms can be moved   b. electrons and atoms can both be moved   c. only electrons can be moved   d. neither electrons nor atoms can be moved   e. electrons must be added

95. For resonance forms of a molecule or ion, __________.
   a. one always corresponds to the observed structure   b. all the resonance structures are observed in various proportions   c. the observed structure is an average of the resonance forms   d. the same atoms need not be bonded to each other in all resonance forms   e. there cannot be more than two resonance structures for a given species

For the questions that follow, consider the BEST Lewis structures of the following oxyanions:

(i) NO$_2^-$ (ii) NO$_3^-$ (iii) SO$_3^{2-}$ (iv) SO$_4^{2-}$ (v) BrO$_3^-$

96. There can be four equivalent best resonance structures of __________.
   a. (i) b. (ii) c. (iii) d. (iv) e. (v)

97. In which of the ions do all X-O bonds (X indicates the central atom) have the same length?
   a. none b. all c. (i) and (ii) d. (iii) and (v) e. (iii), (iv), and (v)

98. Of the following, __________ cannot accommodate more than an octet of electrons.
   a. P b. As c. O d. S e. I

99. A valid Lewis structure of __________ cannot be drawn without violating the octet rule.
   a. NF$_3$ b. IF$_3$ c. PF$_3$ d. SbF$_3$ e. SO$_4^{2-}$

100. Based on the octet rule, boron will most likely form a __________ ion.
    a. B$^3-$ b. B$^{1+}$ c. B$^{3+}$ d. B$^{2+}$ e. B$^2$-

101. Which of the following does not have eight valence electrons?
    a. Cl$^-$ b. Xe c. Ti$^{4+}$ d. Rb$^{+1}$ e. Sr$^{+1}$

102. A valid Lewis structure of __________ cannot be drawn without violating the octet rule.
    a. PO$_3^{3-}$ b. SiF$_4$ c. CF$_4$ d. SeF$_4$ e. NF$_3$

103. The central atom in __________ does not violate the octet rule.
    a. SF$_4$ b. KrF$_2$ c. CF$_4$ d. XeF$_4$ e. ICl$_4^-$

104. The central atom in __________ violates the octet rule.
    a. NH$_3$ b. SeF$_2$ c. BF$_3$ d. AsF$_3$ e. CF$_4$

105. A valid Lewis structure of __________ cannot be drawn without violating the octet rule.
    a. ClF$_3$ b. PCl$_3$ c. SO$_3$ d. CCl$_4$ e. CO$_2$

106. A valid Lewis structure of __________ cannot be drawn without violating the octet rule.
    a. Ni$_3$ b. SO$_2$ c. ICl$_5$ d. SiF$_4$ e. CO$_2$

107. A valid Lewis structure of __________ cannot be drawn without violating the octet rule.
    a. NF$_3$ b. BeH$_2$ c. SO$_2$ d. CF$_4$ e. SO$_3^{2-}$
108. Why don't we draw double bonds between the Be atom and the Cl atoms in BeCl₂?
   a. That would give positive formal charges to the chlorine atoms and a negative formal charge to the beryllium atom.  
   b. There aren't enough electrons.  
   c. That would result in more than eight electrons around beryllium.  
   d. That would result in more than eight electrons around each chlorine atom.  
   e. That would result in the formal charges not adding up to zero.

109. Which atom can accommodate an octet of electrons, but doesn't necessarily have to accommodate an octet?

110. Bond enthalpy is ________.
   a. always positive  b. always negative  
   c. sometimes positive, sometimes negative  
   d. always zero  e. unpredictable

111. Given that the average bond energies for C-H and C-Br bonds are 413 and 276 kJ/mol, respectively, the heat of atomization of bromoform (CHBr₃) is ________ kJ/mol.
   a. 1241  b. 689  c. -689  d. 1378  e. -1378

112. Of the bonds C-N, C=N, and C≡N, the C-N bond is ________.
   a. strongest/shortest  b. strongest/longest  
   c. weakest/shortest  d. weakest/longest  
   e. intermediate in both strength and length

Completion
Complete each statement.

118. Using the noble gas shorthand notation, write the electron configuration for Fe³⁺.

119. Which halogen, bromine or iodine, will form the more polar bond with phosphorus?

120. Draw the Lewis structure of ICl₂⁺.

113. As the number of covalent bonds between two atoms increases, the distance between the atoms ________ and the strength of the bond between them ________.
   a. increases, increases  b. decreases, decreases  
   c. increases, decreases  d. decreases, increases  
   e. is unpredictable

114. Of the possible bonds between carbon atoms (single, double, and triple), ________.
   a. a triple bond is longer than a single bond  b. a double bond is stronger than a triple bond  
   c. a single bond is stronger than a triple bond  d. a double bond is longer than a single bond  
   e. a single bond is stronger than a double bond

115. Most explosives are compounds that decompose rapidly to produce ________ products and a great deal of ________.
   a. gaseous, gases  b. liquid, heat  c. soluble, heat  
   d. solid, gas  e. gaseous, heat

116. Dynamite consists of nitroglycerine mixed with ________.
   a. potassium nitrate  b. damp KOH  c. TNT  
   d. diatomaceous earth or cellulose  e. solid carbon

117. Dynamite ________.
   a. was invented by Alfred Nobel  b. is made of nitroglycerine and an absorbent such as diatomaceous earth  
   c. is a much safer explosive than pure nitroglycerine  d. is an explosive  
   e. all of the above

121. Benzene is an ________ compound with ________ equivalent Lewis structures.

122. In a reaction, if the bonds in the reactants are stronger than the bonds in the product, the reaction is ________.
123. In compounds of ________ and ________, the octet rule is violated due to the presence of fewer than eight valence electrons.

124. Polyatomic ions with an odd number of electrons will ________ the octet rule.

125. The strength of a covalent bond is measured by its ________.

126. To produce maximum heat, an explosive compound should have ________ chemical bonds and decompose to molecule with ________ bonds.

Essay

127. The electron configuration that corresponds to the Lewis symbol, :\(\text{Cl}^-\), is ________.

128. Write the balanced chemical equation for the reaction for which \(\Delta H^\circ_{\text{rxn}}\) is the lattice energy for potassium bromide.

129. Give the electron configuration of Cu\(^{2+}\).

130. Alternative but equivalent Lewis structures are called ________.

131. Calculate the bond energy of C-F given that the heat of atomization of CHFClBr is 1502 kJ/mol, and that the bond energies of C-H, C-Br, and C-Cl are 413, 276, and 328 kJ/mol, respectively.

132. The reaction below is used to produce methanol:

\[
\text{CO (g)} + 2 \text{H}_2 (g) \rightarrow \text{CH}_3\text{OH (l)} \quad \Delta H_{\text{rxn}} = -128 \text{ kJ}
\]

(a) Calculate the C-H bond energy given the following data:

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<td>O-H</td>
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(b) The tabulated value of the (C-H) bond energy is 413 kJ/mol. Explain why there is a difference between the number you have calculated in (a) and the tabulated value.
# AP Chapter 8 Study Questions
## Answer Section

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<td>1</td>
<td>Sec. 8.7</td>
</tr>
<tr>
<td>108.</td>
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<td>2</td>
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</tr>
<tr>
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<td>1</td>
<td>Sec. 8.7</td>
</tr>
<tr>
<td>110.</td>
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<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
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<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
<td>112.</td>
<td>D</td>
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<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
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<td>1</td>
<td>1</td>
<td>Sec. 8.8</td>
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<tr>
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<td>D</td>
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<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
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<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
<td>116.</td>
<td>D</td>
<td>1</td>
<td>1</td>
<td>Sec. 8.8</td>
</tr>
<tr>
<td>117.</td>
<td>E</td>
<td>1</td>
<td>1</td>
<td>Sec. 8.8</td>
</tr>
</tbody>
</table>

**COMPLETION**

118. ANS: [Ar]3d²

    PTS: 1   DIF: 2   REF: Sec. 8.2

119. ANS: bromine

    PTS: 1   DIF: 1   REF: Sec. 8.4
120. ANS:  

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\end{array}
\]

PTS: 1  DIF: 1  REF: Sec. 8.5

121. ANS: aromatic, two

PTS: 1  DIF: 3  REF: Sec. 8.6

122. ANS: endothermic

PTS: 1  DIF: 1  REF: Sec. 8.7

123. ANS: boron and beryllium

PTS: 1  DIF: 2  REF: Sec. 8.7

124. ANS: fail or violate

PTS: 1  DIF: 2  REF: Sec. 8.7

125. ANS: bond enthalpy

PTS: 1  DIF: 1  REF: Sec. 8.8

126. ANS: weak, strong

PTS: 1  DIF: 4  REF: Sec. 8.8

ESSAY

127. ANS:

\[\text{[Ne]} 3s^23p^5\]

PTS: 1  DIF: 2  REF: Sec. 8.1

128. ANS:

\[\text{KBr (s)} \rightarrow \text{K}^+ (g) + \text{Br}^- (g)\]

PTS: 1  DIF: 1  REF: Sec. 8.2

129. ANS:

\[\text{[Ar]} 3d^9\]

PTS: 1  DIF: 2  REF: Sec. 8.2

130. ANS: resonance structures

PTS: 1  DIF: 1  REF: Sec. 8.6
131. ANS:
\[ \Delta H_{\text{atomization}} = [D(C-H) + D(C-F) + D(C-Cl) + D(C-Br)] \]
\[ D(C-F) = \Delta H_{\text{atomization}} - [D(C-H) + D(C-Cl) + D(C-Br)] \]

\[ = [1502 - (413 + 276 + 328)] \text{kJ/mol} \]

\[ = 485 \text{kJ/mol} \]

PTS: 1 DIF: 1 REF: Sec. 8.8

132. ANS:
(a) \[ \Delta H_{\text{rxn}} = D(C=O) + 2 D(H-H) - [3 D(C-H) + D(C-O) + D(O-H)] \]
3 \[ D(C-H) = -\Delta H_{\text{rxn}} + D(C=O) + 2 D(H-H) - D(C-O) - D(O-H) \]

\[ D(C-H) = (128 + 1072 + 2(436) - 358 - 463)/3 = 417 \text{kJ/mol} \]

(b) Tabulated values, like those in Table 8.4, are averaged from many bond energies measured for C-H bonds in many different molecules.

PTS: 1 DIF: 2 REF: Sec. 8.8

133. ANS:
\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]
\[ \Delta H_{\text{combustion}} = -[(4 \text{ mol C - H})(D_{\text{C-H}}) + (2 \text{ mol O - O})(D_{\text{O-O}})] - \]
\[ [(2 \text{ mol C = O})(D_{\text{C=O}}) - (4 \text{ mol O - H})(D_{\text{O-H}})] \]

\[ = [(4 \times 413 + 2 \times 495) - (2 \times 799 + 4 \times 463)] \text{kJ} \]

\[ \Delta H_{\text{combustion}} = -808 \text{kJ} \]

PTS: 1 DIF: 2 REF: Sec. 8.8