PRACTICE QUIZ
Mr. Scott

Select the best answer.

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)

A chemical bond results from the mutual attraction of the nuclei of atoms and

a) electrons. b) protons.
c) neutrons. d) dipoles.

3)

The electrons involved in the formation of a chemical bond are called

a) dipoles. b) *s* electrons.
c) Lewis electrons. d) valence electrons.

4)

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.

5)

In a chemical bond, the link between atoms results from the attraction between electrons and

a) Lewis structures. b) nuclei.
c) van der Waals forces. d) isotopes.

6)

As independent particles, atoms are

a) at relatively high potential energy. b) at relatively low potential energy.
c) very stable. d) part of a chemical bond.

7)

Atoms are __________ when they are combined.

a) more stable
b) less stable
c) not bound together
d) at a high potential energy

8)

Atoms naturally move

a) toward high potential energy.
b) toward low potential energy.
c) toward less stability.
d) away from each other.

9)

As atoms bond with each other, they

a) increase their potential energy, thus creating less-stable arrangements of matter.
b) decrease their potential energy, thus creating less-stable arrangements of matter.
c) increase their potential energy, thus creating more-stable arrangements of matter.
d) decrease their potential energy, thus creating more-stable arrangements of matter.

10)

A chemical bond resulting from the electrostatic attraction between positive and negative ions is called a(n)

a) covalent bond.
b) ionic bond.
c) charged bond.
d) dipole bond.

11)

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.

12)

If two covalently bonded atoms are identical, the bond is

a) nonpolar covalent.
b) polar covalent.
c) nonionic.
d) coordinate covalent.

13)

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.

14)

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.

15)

The electrostatic attraction between __________ forms an ionic bond.

a) ions
b) dipoles
c) electrons
d) orbitals

16)

A covalent bond results when __________ are shared.

a) ions
b) Lewis structures
c) electrons
d) dipoles

17)

Most chemical bonds are

a) purely ionic.
b) purely covalent.
c) partly ionic and partly covalent.
d) metallic.

18)

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.

19)

Purely ionic bonds do not occur because the atom that gives up an electron in such a

a) is not an ion.

b) still has some attraction for the electron.

c) has a negative charge.

d) shares the electron equally with the other atom in the bond.

20)

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.

21)

Bonds that are between 5% and 50% ionic are considered

a) ionic.

b) pure covalent.

c) polar covalent.

d) nonpolar covalent.

22)

Bonds that are more than 50% ionic are considered

a) polyatomic.

b) polar covalent.

c) ionic.

d) nonpolar covalent.

23)

A bond that is less than 5% ionic is considered

a) polar covalent.
b) ionic.
c) nonpolar covalent.
d) metallic.

24)

The pair of elements that forms a bond with the least ionic character is

a) Na and Cl.
b) H and Cl.
c) O and Cl.
d) Br and Cl.

25)

The B—F bond in BF_3 (electronegativity for B is 2.0; electronegativity for F is 4.0) is

a) polar covalent.
b) ionic.
c) nonpolar covalent.
d) pure covalent.

26)

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

a) 9%.
b) 15%.
c) 38%.
d) 62%.

27)

The percentage ionic character and the type of bond in Br_2 (electronegativity for Br is 2.8) is

a) 0%; nonpolar covalent.
b) 100%; polar covalent.
c) 0%; pure ionic.
d) 100%; pure ionic.

28)

The percentage ionic character and the type of bond in LiCl (electronegativity for Li is 1.0; electronegativity for Cl is 3.0) is

a) 50%; ionic.
b) 50%; polar covalent.
c) 67%; polar covalent.
d) 67%; ionic.

29)

In which of these compounds is the bond between the atoms NOT a nonpolar covalent bond?

a) Cl_2

b) H_2

c) HCl

d) O_2

30)

A neutral group of atoms held together by covalent bonds is a

a) molecular formula.

b) chemical formula.

c) compound.

d) molecule.

31)

A molecule is a

a) negatively charged group of atoms held together by covalent bonds.

b) positively charged group of atoms held together by covalent bonds.

c) neutral group of atoms held together by covalent bonds.

d) neutral group of atoms held together by ionic bonds.

32)

All of the following are true statements about a molecule EXCEPT

a) it is capable of existing on its own.

b) it may consist of two or more atoms of the same type.

c) it must be part of a molecular compound.

d) it may consist of two or more atoms of different types.

33)

A __________ shows the types and numbers of atoms joined in a single molecule of a molecular compound.

a) molecular formula

b) chemical formula

c) covalent bond

d) ionic bond

34)

Which of the following is NOT an example of a molecular formula?

a) H_2O

b) B

c) NH_3

d) O_2

35)

When a covalent bond forms, as the distance between two atoms decreases, the potential energy

a) increases.

b) decreases.

c) remains constant.

d) becomes zero.

36)

A covalent bond forms when the attraction between two atoms is balanced by repulsion and the potential energy is

a) at a maximum.

b) zero.

c) at a minimum.

d) equal to the kinetic energy.

37)

Bond length is

a) the separation for which potential energy is at a minimum.

b) the separation for which kinetic energy is at a maximum.

c) the separation for which potential energy is at a maximum.

d) one-half the diameter of the electron cloud.

38)

Bond energy is the energy

a) required to break a chemical bond.

b) released when a chemical bond breaks.

c) required to form a chemical bond.

d) absorbed when a chemical bond forms.

39)

When two atoms are separated and do not attract each other, the potential energy is

a) at a minimum.
b) zero.
c) at a maximum.
d) equal to the sum of the potential energy of each atom divided by two.

40)

The energy released when a covalent bond forms is the difference between zero and the

a) maximum potential energy.
b) kinetic energy of the atom.
c) minimum potential energy.
d) bond length expressed in nanometers.

41)

In a molecule of fluorine, the two shared electrons give each fluorine atom ________ electron(s) in the outer energy level.

a) 1
b) 2
c) 8
d) 32

42)

In many compounds, atoms of main-group elements form bonds so that the number of electrons in the outermost energy levels of each atom is

a) 2.
b) 6.
c) 8.
d) 10.

43)

The octet rule states that chemical compounds tend to form so that each atom has an octet of electrons in

a) its highest occupied energy level.
b) the 1*s* orbital.
c) its *d* orbitals.
d) its *p* orbitals.

44)

An octet is equal to

a) 2.
b) 4.
c) 5.
d) 8.

45)

What principle states that atoms tend to form compounds so that each atom can have eight electrons in its outermost energy level?

a) rule of eights
b) Avogadro principle
c) configuration rule
d) octet rule

46)

The electron configuration of nitrogen is $1s^2\ 2s^2\ 2p^3$. How many more electrons does nitrogen need to satisfy the octet rule?

a) 1
b) 3
c) 5
d) 8

47)

The elements of the __________ group satisfy the octet rule without forming compounds.

a) main
b) noble gas
c) alkali metal
d) alkaline-earth metal

48)

When the octet rule is satisfied, the outermost __________ are filled.

a) *d* and *f* orbitals
b) *s* and *p* orbitals
c) *s* and *d* orbitals
d) *d* and *p* orbitals

49)

In drawing a Lewis structure, each nonmetal atom except hydrogen should be surrounded by

a) 2 electrons.
b) 4 electrons.
c) 8 electrons.
d) 10 electrons.

50)

In drawing a Lewis structure, the central atom is the

a) atom with the greatest mass.
b) atom with the highest atomic number.
c) atom with the fewest electrons.
d) least electronegative atom.

51)

To draw a Lewis structure, one must know the

a) number of valence electrons in each atom.

b) atomic mass of each atom.

c) bond length of each atom.

d) ionization energy of each atom.

52)

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) determine the total number of valence electrons in each atom.

d) determine the electronegativity of each atom.

53)

To draw a Lewis structure, it is NOT necessary to know

a) bond energies.

b) the types of atoms in the molecule.

c) the number of valence electrons for each atom.

d) the number of atoms in the molecule.

54)

The Lewis structure for the ammonium ion has

a) 2 valence electrons.

b) 4 valence electrons.

c) 8 valence electrons.

d) 12 valence electrons.

55)

If, after drawing a Lewis structure, too many valence electrons have been used, the molecule probably contains

a) too many atoms.

b) one or more multiple covalent bonds.

c) too many unshared pairs of electrons.

d) an ionic bond.

56)

Multiple covalent bonds may occur in atoms that contain carbon, nitrogen, or

a) chlorine.
b) hydrogen.
c) oxygen.
d) helium.

57)

The substance whose Lewis structure shows three covalent bonds is

a) H_2O.
b) CH_2Cl_2.
c) NH_3.
d) CCl_4.

58)

How many double bonds are in the Lewis structure for hydrogen fluoride, HF?

a) none
b) one
c) two
d) three

59)

How many extra electrons are in the Lewis structure of the phosphate ion, PO_4^{3-}?

a) 0
b) 2
c) 3
d) 4

60)

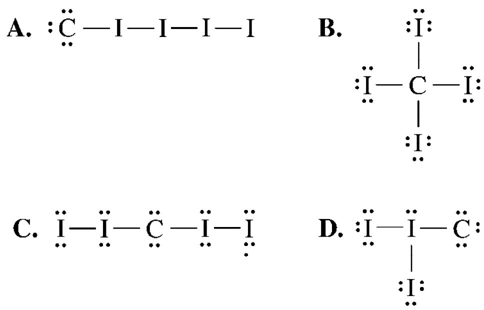
How many electrons must be shown in the Lewis structure of the hydroxide ion, OH^-?

a) 1
b) 8
c) 9
d) 10

61)

What is the Lewis structure for hydrogen chloride, HCl?

a) A
b) B
c) C
d) D



62)

What is the Lewis structure for carbon tetraiodide, which contains one carbon atom and four iodine atoms?

a) A
b) B
c) C
d) D

63)

Bonding in molecules or ions that cannot be correctly represented by a single Lewis structure is

a) covalent bonding.
b) resonance.
c) single bonding.
d) double bonding.

64)

Resonance structures are also called

a) dichotomous structures.
b) Lewis structures.
c) molecular structures.
d) resonance hybrids.

65)

To indicate resonance, a __________ is placed between a molecule's resonance structures.

a) double-headed arrow
b) single-headed arrow
c) series of dots
d) Lewis structure

66)

Chemists once believed 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) Lewis structure.
d) single-double bonding.

67)

The chemical formula for an ionic compound represents the

a) number of atoms in each molecule.

b) number of ions in each molecule.

c) simplest ratio of the combined ions that balances total charges.

d) total number of ions in the crystal lattice.

68)

A shorthand representation of the composition of a substance using atomic symbols and numerical subscripts is called a(n)

a) Lewis structure.

b) chemical formula.

c) polyatomic ion.

d) multiple bond.

69)

A formula that shows the types and numbers of atoms combined in a single molecule is called a(n)

a) molecular formula.

b) ionic formula.

c) Lewis structure.

d) covalent formula.

70)

A formula unit of an ionic compound

a) is an independent unit that can be isolated and studied.

b) is the simplest ratio of ions that balances total charge.

c) describes the crystal lattice.

d) all of the above

71)

The chemical formula for water, a covalent compound, is H_2O. This formula is an example of a(n)

a) formula unit.

b) Lewis structure.

c) ionic formula.

d) molecular formula.

72)

An ionic compound is not represented by a molecular formula because an ionic compound

a) does not contain bonds.
b) does not have charged particles.
c) lacks molecules.
d) always has a positive charge.

73)

In the NaCl crystal, each Na^+ and Cl^- ion has clustered around it __________ of the oppositely charged ions.

a) 1
b) 2
c) 4
d) 6

74)

In an ionic compound, the orderly arrangement of ions in a crystal is the state of

a) maximum potential energy.
b) minimum potential energy.
c) average potential energy.
d) zero potential energy.

75)

The ions in an ionic compound are organized into a(n)

a) molecule.
b) Lewis structure.
c) polyatomic ion.
d) crystal.

76)

In a crystal of an ionic compound, each cation is surrounded by

a) molecules.
b) positive ions.
c) dipoles.
d) anions.

77)

In a crystal, the valence electrons of adjacent ions

a) repel each other.
b) attract each other.
c) neutralize each other.
d) have no effect on each other.

78)

The simplest repeating unit of a crystal is known as a(n)

a) unit cell. b) crystal lattice.

c) ionic compound. d) salt.

79)

The energy released when 1 mol of an ionic crystalline compound is formed from gaseous ions is called the

a) bond energy. b) potential energy.

c) lattice energy. d) energy of crystallization.

80)

The lattice energy is a measure of the

a) strength of an ionic bond. b) strength of a metallic bond.

c) strength of a covalent bond. d) number of ions in a crystal.

81)

Lattice energy is the energy released in the formation of a(n)

a) Lewis structure. b) polar covalent compound.

c) nonpolar covalent compound. d) ionic compound.

82)

Compared with energies of neutral atoms, a crystal lattice has

a) higher potential energy. b) lower potential energy.

c) equal potential energy. d) less stability.

83)

If the lattice energy of compound A is greater than that of compound B,

a) compound A is not an ionic compound. b) the bonds in compound A are stronger than the bonds in compound B.

c) compound B is probably a gas. d) compound A has larger crystals than compound B.

84)

Which of the following is NOT a property of an ionic compound?

a) low boiling point

b) brittleness

c) hardness

d) molten compound conducts electricity

85)

Compared with ionic compounds, molecular compounds

a) have higher boiling points.

b) are brittle.

c) have lower melting points.

d) are harder.

86)

The forces of attraction between molecules in a molecular compound are

a) stronger than the forces of ionic bonding.

b) weaker than the forces of ionic bonding.

c) approximately equal to the forces of ionic bonding.

d) zero.

87)

A compound that vaporizes readily at room temperature is most likely to be a(n)

a) molecular compound.

b) ionic compound.

c) metal.

d) brittle compound.

88)

Because the particles in ionic compounds are more strongly attracted than in molecular compounds, the melting points of ionic compounds are

a) equal for all ionic compounds.

b) lower than melting points of molecular compounds.

c) higher than melting points of molecular compounds.

d) approximately equal to room temperature.

89)

Ionic compounds are brittle because the strong attractive forces

a) allow the layers to shift easily.

b) cause the compound to vaporize easily.

c) keep the surface dull.

d) hold the layers in relatively fixed positions.

90)

The properties of both ionic and molecular compounds are related to the

a) lattice energies of the compounds.

b) strengths of attraction between the particles in the compounds.

c) number of covalent bonds each contains.

d) mobile electrons that they contain.

91)

A chemical bond formed by the attraction between positive ions and surrounding mobile electrons is a(n)

a) nonpolar covalent bond.

b) ionic bond.

c) polar covalent bond.

d) metallic bond.

92)

Compared with nonmetals, the number of valence electrons in metals is generally

a) smaller.

b) greater.

c) about the same.

d) almost triple that of nonmetals.

93)

In metals, the valence electrons

a) are attached to particular positive ions.

b) are shared by all of the atoms.

c) are immobile.

d) form covalent bonds.

94)

In the electron-sea model of a metallic bond,

a) electrons are stationary.
b) electrons are bonded to particular positive ions.
c) some electrons are valence electrons and some are not.
d) mobile electrons are shared by all the atoms.

95)

A metallic bond forms when positive ions attract

a) stationary electrons.
b) nonvalence electrons.
c) cations.
d) mobile electrons.

96)

In metallic bonds, the mobile electrons surrounding the positive ions are called a(n)

a) Lewis structure.
b) electron sea.
c) electron cloud.
d) dipole.

97)

The electron-sea model of bonding represents

a) covalent bonding.
b) metallic bonding.
c) ionic bonding.
d) hydrogen bonding.

98)

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 heat.
d) change light to electricity.

99)

The shiny appearance of a metal is most closely related to the metal's

a) electron sea.
b) covalent bonds.
c) brittle crystalline structure.
d) positive ions.

100)

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.

101)

The electron-sea model of the metallic bond helps to explain why metals are

a) brittle.
b) dull.
c) shiny.
d) malleable.

102)

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

103)

Metals are malleable because the metallic bonding

a) holds the layers of ions in rigid positions.
b) does not produce ions.
c) allows one plane of ions to slide past another.
d) is easily broken.

104)

Which best explains the observation that metals are malleable and ionic crystals are

a) their chemical bonds
b) their London forces
c) their heats of vaporization
d) their polarity

105)

Because metallic bonds permit one plane of ions to slide past another without breaking bonds, metals are

a) brittle.
b) malleable.
c) nonreflective.
d) poor conductors of electricity.

106)

Malleability and ductility are characteristic of substances with

a) covalent bonds.
b) ionic bonds.
c) Lewis structures.
d) metallic bonds.

107)

Compared with metals, ionic crystals are

a) ductile.
b) malleable.
c) brittle.
d) lustrous.

108)

Planes of ions can slide past one another without breaking bonds in substances with

a) metallic bonds.
b) ionic bonds.
c) covalent bonds.
d) no bonds.

109)

Shifting the layers of an ionic crystal causes the crystal to

a) be drawn into a wire.
b) shatter.
c) become metallic.
d) freeze.

110)

The model for predicting the shape of a molecule that is based on the repulsion of electrons for each other is called

a) hybridization.
b) Lewis structure.
c) London force model.
d) VSEPR theory.

111)

According to VSEPR theory, an AB_2 molecule is

a) trigonal planar.
b) tetrahedral.
c) linear.
d) octahedral.

112)

VSEPR theory is a model for predicting

a) the strength of metallic bonds.

b) the shape of molecules.

c) lattice energy values.

d) ionization energy.

113)

The concept that electrostatic repulsion between electron pairs surrounding an atom causes these pairs to be separated as far as possible is the foundation of

a) VSEPR theory.

b) the hybridization model.

c) the electron-sea model.

d) Lewis theory.

114)

According to VSEPR theory, the electrostatic repulsion between electron pairs surrounding an atom causes

a) an electron sea to form.

b) positive ions to form.

c) these pairs to be separated as far as possible.

d) light to reflect.

115)

According to VSEPR theory, the shape of an AB_3 molecule is

a) trigonal planar.

b) tetrahedral.

c) linear.

d) bent.

116)

According to VSEPR theory, the structure of the ammonia molecule, NH_3, is

a) linear.

b) bent.

c) pyramidal.

d) tetrahedral.

117)

Use VSEPR theory to predict the shape of the hydrogen chloride molecule, HCl.

a) tetrahedral

b) linear

c) bent

d) trigonal planar

118)

Use VSEPR theory to predict the shape of the magnesium hydride molecule, MgH_2.

a) tetrahedral
b) linear
c) bent
d) octahedral

119)

Use VSEPR theory to predict the shape of the carbon tetraiodide molecule, CI_4.

a) tetrahedral
b) linear
c) bent
d) trigonal planar

120)

Use VSEPR theory to predict the shape of the chlorate ion, ClO_3^-.

a) trigonal planar
b) octahedral
c) trigonal pyramidal
d) bent

121)

Use VSEPR theory to predict the shape of the hydrogen sulfide molecule, H_2S.

a) tetrahedral
b) linear
c) bent
d) octahedral

122)

Use VSEPR theory to predict the shape of carbon dioxide, CO_2.

a) tetrahedral
b) linear
c) bent
d) octahedral

123)

The hybridized orbitals responsible for the bent shape of the water molecule are

a) $1s^2\ 2s^2$.
b) ps^1.
c) sp^3.
d) $2s^2\ sp^2$.

124)

The hybridized orbitals responsible for the shape of the CH_4 molecule are

a) $1s^1\ 1p^3$.
b) sp^2.
c) $2s^2\ 2p^2$.
d) sp^3.

125)

The mixing of two or more atomic orbitals of similar energies on the same atom to produce new orbitals of equal energies is called

a) VSEPR theory.
b) malleability.
c) hybridization.
d) dipole-dipole interaction.

126)

Orbitals of equal energy produced by the combination of two or more orbitals in the same atom are called

a) *q*-orbitals.
b) hybrid orbitals.
c) Lewis orbitals.
d) *n*-orbitals.

127)

Which hybrid orbitals help explain how methane bonds?

a) sp^3 orbitals
b) *sp* orbitals
c) pd^3 orbitals
d) df^3 orbitals

128)

Hybridization helps explain molecular bonding when the valence electrons on the uncombined atoms

a) total more than 8.
b) are in orbitals with different shapes.
c) total less than 3.
d) are not available for bonding.

129)

Four hybrid sp^3 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.

130)

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.

131)

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.

132)

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.

133)

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.

134)

The weak intermolecular forces resulting from instantaneous and induced dipoles are

a) London dispersion forces. b) dipole-dipole forces.

c) hydrogen forces. d) polar covalent bonding.

135)

Compared with molecular bonds, the strength of intermolecular forces is

a) weaker. b) stronger.

c) about the same. d) too variable to compare.

136)

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.

137)

That the boiling point of water (H_2O) is higher than the boiling point of hydrogen sulfide (H_2S) is partially explained by

a) London forces. b) covalent bonding.

c) ionic bonding. d) hydrogen bonding.

138)

The following molecules contain polar bonds. The only polar molecule is

a) CCl_4. b) CO_2.

c) NH_3. d) CH_4.

139)

The following molecules contain polar bonds. The only nonpolar molecule is

a) HCl. b) H_2O.

c) CO_2. d) NH_3.

140)

A polar molecule contains

a) ions.

b) a region of positive charge and a region of negative charge.

c) only London forces.

d) no bonds.

141)

A molecule of hydrogen chloride is polar because

a) it is composed of ions.

b) it is magnetic.

c) it contains metallic bonds.

d) the chlorine attracts the shared electrons more strongly than does the hydrogen atom.

142)

When a polar molecule attracts the electron in a nonpolar molecule,

a) a dipole is induced.

b) a crystal lattice forms.

c) an ionic bond forms.

d) a Lewis structure forms.

143)

Iodine monochloride (ICl) has a higher boiling point than bromine (Br_2) partly because iodine monochloride is a(n)

a) nonpolar molecule.

b) ion.

c) crystal.

d) polar molecule.

Read each question or statement, and write your response in the space provided.

144)

Differentiate between an ionic compound and a molecular compound.

145)

Why do most atoms form chemical bonds?

146)

Explain why scientists use resonance structures to represent some molecules.

147)

Explain what is meant by the lattice energy of BeF_2.

Read each question or statement, and write your response in the space provided.

148)

List the six basic steps used in drawing Lewis structures.

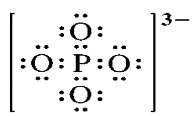
Solve the following problems. Show your answer and your work.

149)

Draw a Lewis structure for the sulfate ion, SO_4^{2-}.

150)

Draw a Lewis structure for the phosphate ion, PO_4^{3-}.



Phosphate ion

151)

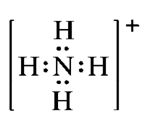
Draw a Lewis structure for the nitrate ion, NO_3^{-}.

152)

Draw a Lewis structure for the oxalate ion, $C_2O_4^{2-}$.

153)

Draw a Lewis structure for the ammonium ion, NH_4^+.



Ammonium ion

ANSWER KEY

1) c
2) a
3) d
4) a
5) b
6) a
7) a
8) b
9) d
10) b
11) d
12) a
13) b
14) b
15) a
16) c
17) c
18) a
19) b
20) a
21) c
22) c
23) c
24) d
25) b
26) c
27) a
28) d
29) c
30) d
31) c
32) c
33) a
34) b
35) b
36) c
37) a

38) a
39) b
40) c
41) c
42) c
43) a
44) d
45) d
46) b
47) b
48) b
49) c
50) d
51) a
52) c
53) a
54) c
55) b
56) c
57) c
58) a
59) c
60) b
61) d
62) b
63) b
64) d
65) a
66) b
67) c
68) b
69) a
70) b
71) d
72) c
73) d

74) b
75) d
76) d
77) a
78) a
79) c
80) a
81) d
82) b
83) b
84) a
85) c
86) b
87) a
88) c
89) d
90) b
91) d
92) a
93) b
94) d
95) d
96) b
97) b
98) b
99) a
100) d
101) c
102) b
103) c
104) a
105) b
106) d
107) c
108) a
109) b

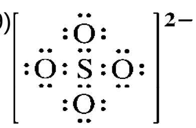
110) d
111) c
112) b
113) a
114) c
115) a
116) c
117) b
118) b
119) a
120) c
121) c
122) b
123) c
124) d
125) c
126) b
127) a
128) b
129) d
130) b
131) c
132) a
133) c
134) a
135) a
136) b
137) d
138) c
139) c
140) b
141) d
142) a
143) d
144) Atoms in a molecular compound share electrons to achieve stability. Atoms in an ionic compound gain or lose electrons to form ions.

145) 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.

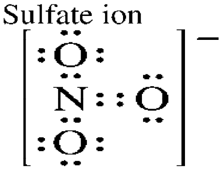
146) Resonance structures picture the bonding in molecules that cannot be correctly pictured with a single Lewis structure.

147) The lattice energy of BeF_2 is the energy released when 1 mol of solid BeF_2 is formed from beryllium ions and fluoride ions.

148) 1. Determine the type and number of atoms in the molecule.
2. Write the electron-dot notation for each type of atom in the molecule.
3. Determine the total number of valence electrons in the atoms to be combined.
4. Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. If not, the least-electronegative atom is central.
5. Add unshared pairs of electrons so that each hydrogen atom shares a pair of electrons and each other nonmetal is surrounded by eight electrons.
6. Count the electrons in the structure to check that the number of valence electrons used equals the number available.

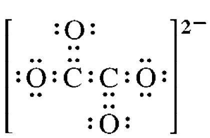
149) 

Sulfate ion

151) 

Nitrate ion

152) 2 C atoms with 4 electrons $\Longrightarrow 2 \times 4 = 8$
4 O atoms with 6 electrons $\Longrightarrow 4 \times 6 = 24$
$8 + 24 = 32$ valence electrons
$2-$ charge $\Longrightarrow$ 2 extra electrons
$32 + 2 = 34$ electrons total



153)