NAME DO NOT REMOVE From DESK DATE HOUR

Ch. 6 Review Worksheet - Chemical Bonding

In the space provided,write the letter of the term or phrase that best completes each statement or best answers each question.

D **1.** In a crystal of an ionic compound, each cation is surrounded by a number of
(a) molecules. **(c)** dipoles.
(b) positive ions. **(d)** negative ions.

B **2.** The lattice energy of compound A is greater in magnitude than that of compound B. What can be concluded from this fact?
(a) Compound A is not an ionic compound.
(b) It will be more difficult to break the bonds in compound A than those in compound B.
(c) Compound B has larger crystals than compound A.
(d) Compound A has larger crystals than compound B.

B B **3.** The salts NaCl(*s*) and CaCl2(*s*)
a. are good conductors of electricity. **c.** are positively charged.
b. are held together by ionic bonds. **d.** Both (a) and (b)

A **4.** Compared with solid ionic compounds, solid molecular compounds generally
a. have lower melting points. **c.** are more brittle.
b. are harder. **d.** conduct electricity as liquids.

D **5.** Because strong attractive forces hold the layers in an ionic crystal in relatively fixed positions, ionic compounds
a. are hard. **c.** are brittle.
b. are not electrical conductors as solids. **d.** All of the above

B **6.** In metals, the valence electrons are considered to be
(a) attached to particular positive ions. **(c)** immobile.
(b) shared by all surrounding atoms. **(d)** involved in covalent bonds.

D **7.** The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their
(a) chemical bonds. **(c)** enthalpies of vaporization.
(b) London forces. **(d)** polarity.

D **8.** Mobile electrons in the metallic bond are responsible for
(a) luster. **(c)** electrical conductivity.
(b) thermal conductivity. **(d)** All of the above.

D **9.** Chemical bonding in metals is
a. the same as ionic bonding. **c.** a combination of ionic and covalent bonding.
b. the same as covalent bonding. **d.** different from ionic or covalent bonding.

C **10.** Which of the following properties is *not* explained by metallic bonding?
a. electrical conductivity **c.** brittleness
b. thermal conductivity **d.** ductility

A **11.** Metals are malleable because when struck, one plane of metal atoms
a. can slide past another plane without breaking bonds.
b. cannot easily move out of the way.
c. moves in a way that maximizes the repulsive forces within the metal.
d. bonds to the plane directly beneath it.

C **12.** Which of these is responsible for the good electrical conductivity of metals?
a. the arrangement of metal atoms in separate layers
b. the high density of metals atoms in the crystal lattice
c. the ability of electrons to move freely about the crystal structure
d. the fact that metal atoms contain many orbitals separated by very small energy

B 13. The arrangement of valence electrons in a metallic bond is best described as
a. fixed positions in a lattice.
b. a sea of free-moving electrons.
c. concentrated electron density around specific atoms.
d. electron pairs existing in multiple bonds.

Answer the following questions in the space provided.

14. How does the behavior of electrons in metals contribute to the metal's ability to conduct electricity and heat?

the electrons in metals are free to move around inbetween the two or more atoms (metals)

15. What type of energy best represents the strength of an ionic bond?

Bond energy

16. Arrange the ionic bonds in the table below in order of increasing strength from weakest to strongest.

most negative

	Ionic bond	**Lattice energy (kJ/mol)**
2	NaCl	-787
4	CaO	-3384
1	KCl	-715
5	MgO	-3760
3	LiCl	-861

17. Use the electronegativity values shown in **Figure 20,** on page 161 of the text, to determine whether each of the following bonds is nonpolar covalent, polar covalent, or ionic.

a. H—F ionic
b. Na—Cl ionic
c. H—O polar covalent
d. H—H non-polar covalent
e. H—C non-polar covalent/polar covalent
f. H—N polar covalent

18. How is a hydrogen bond different from an ionic or covalent bond?

hydrogen bonds are intermolecular (between 2 separate molecules)
ionic/covalent bonds are intramolecular (with the same molecule)

19. H_2S and H_2O have similar structures and their central atoms belong to the same group. Yet H_2S is a gas at room temperature and H_2O is a liquid. Use bonding principles to explain why this is.

H_2S has polar covalent bonds which are weaker than the bonds in H_2O, because of their electronegativities.

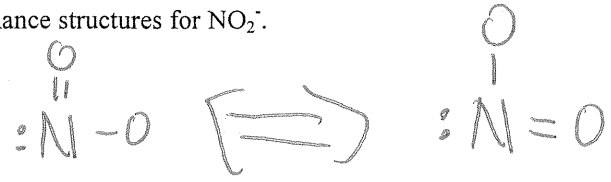
20. In what way is a polar-covalent bond similar to an ionic bond?

There is an unequal sharing of electrons

21. Explain the electrical conductivity of melted and of aqueous solutions of ionic compounds, using the characteristics of ionic bonds.

ions are free to move around

22. Draw resonance structures for NO_2^-.



40. Complete the following table:

	Metals	Ionic Compounds	Molecular Compounds
Components	metals atoms	metal & nonmetal + & - ions	2 or more nonmetals molecules
Overall charge	neutral	neutral	neutral
Conductive in the solid state	Yes	no only in water & when molten	never
Melting point	middle low to high?	highest	lowest
Hardness	soft to hard	hardest	softest
Malleable	yes	no	no
Ductile	yes	no	no
How bonds are formed	attraction between + metal ions & - electrons	transfer of e⁻	sharing of e⁻

24 41. How does bond energy relate to bond length? Which molecule has a greater bond energy, C=C or C-C? Which molecule has a greater bond length? C-C

shorter length

shorter bond length = higher bond energy

25 42. List the intermolecular forces in order of decreasing strength.

hydrogen bonding, dipole-dipole, dispersion

26 43. Explain how hydrogen bonding causes NH_3 to have a higher boiling point than CH_4.

NH_3 has H bonding, CH_4 does not; stronger forces of attraction requires more energy to break

27 44. Which intermolecular force would be present in a molecule of HF? Explain.

Ex: HBr

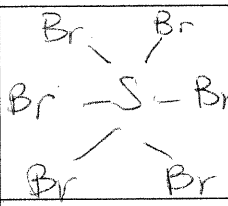
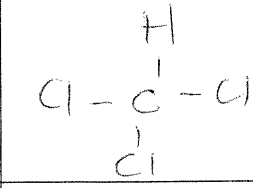
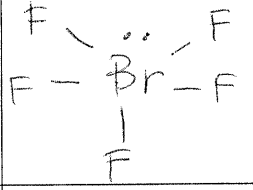
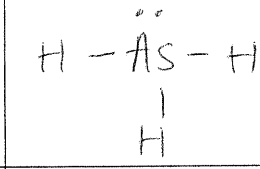
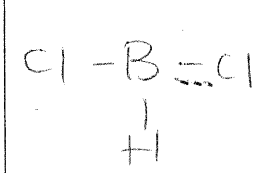
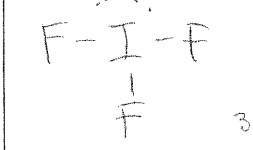
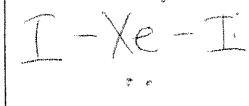
H bonding

Chapter 6 Study Guide

1. List the characteristics of a ionic, covalent and metallic bonds.
2. Be able to explain the properties of ionic compounds, molecular compounds, and metals.
3. Be able to draw structural formulas.
4. Explain resonance and draw structures of them.
5. Identify the shapes of covalent compounds using VSEPR theory and be able to identify the bond angles of each shape.
6. Categorize the degree of polarity of a bond using E.N. values.
7. Be able to determine whether a molecule is polar or nonpolar based on its shape.
8. Describe and name the weak attractive forces that hold molecules together.
9. Know how bond length and bond energy are related. Know how these terms relate to multiple bonds. #41

4. Complete the chart below.

45.

Compound	Dot Diagram	Structural Formula	Shape	Bond Angle	Polar or Nonpolar?
SBr_6 6+7(6) 48			octahedral	90°	nonpolar
$CHCl_3$			tetrahedral	109.5	polar
BrF_5 42			square pyramidal	<90°	polar
AsH_3 5+3=8			trigonal pyramidal	<109.5	non polar
$BHCl_2$			trigonal planar	120°	polar
H_2Se			bent	<109.5°	polar
CS_2			linear	180°	nonpolar
IF_3 28		 3/2	t-shaped	90° & <120°	polar
XeI_2 8+14=22		 2/2	linear	180°	non polar