Name: Key

Chemical Bonding Study Guide

A. Fill in the Blank

1. The electrons involved in the formation of a chemical bond are called valence electrons.
2. A chemical bond that results from the electrostatic attraction between positive and negative ions is called a(n) ionic bond.
3. If electrons involved in bonding spend most of the time closer to one atom rather than the other, the bond is polar.
4. If a bond's character is more than 50% ionic, then the bond is called a(n) ionic bond.
5. A bond's character is more than 50% ionic if the electronegativity difference between the two atoms is greater than 1.67 (or 1.7)
6. Write the formula for an example of each of the following compounds:
   a. nonpolar covalent compound F_2
   b. polar covalent compound H_2O
   c. ionic compound NaF
7. Electron dot or Lewis dot notation shows only the valence electrons of an atom of a particular element arranged around it as dots.
8. What type of bonding holds a polyatomic ion together?
   covalent (nonmetals)
9. Describe the force that holds two atoms together in an ionic bond.
   Opposite charges attract w/ electrostatic force
10. What type of energy best represents the strength of an ionic bond?
    lattice energy
11. **Molecular polarity** is determined by overall polarities summed as a dipole moment. (# of lone pairs + bonding atoms)

12. According to **VSEPR theory**, the repulsion between the sets of valence-level electrons surrounding an atom causes atoms to be as far from each other as possible

13. What molecular geometry corresponds to *sp* **hybridization**? linear

14. Tetrahedral geometry is a result of what **hybridization**? sp^3

15. A molecule with three **hybrid orbitals** has what type of hybridization? sp^2

~~16.~~ In what direction does current in a **dipole** flow? + to −

17. **Hydrogen bonding** is responsible for the properties of water.

18. **London forces** (dispersion) are the only intermolecular forces that act upon nonpolar (covalent) molecules

19. As atomic or molar masses increase, **London forces** increase.

20. Identify the major assumption of the **VSEPR theory** that is used to predict the shape of atoms. electrons repel

21. In water, two hydrogen atoms are bonded to one oxygen atom. Why isn't water a linear molecule? O has 2 lone pairs, forcing the H's closer.

22. What orbitals combine together to form sp^3 hybrid orbitals around a carbon atom? 2s and all 3 2p orbitals ($2p_x$, $2p_y$, $2p_z$)

23. What two factors determine whether or not a molecule is polar? OR ✗
① Shape ② bonding atoms (electroneg. diff)
① lone pairs
② bonding atoms

24. How are dipole-dipole attractions, London dispersion forces, and hydrogen bonding similar? they are all intermolecular forces

B. True or False

Write *true* or *false* for each statement.

1. The **bond length** is the distance between two bonded atoms at their maximum potential energy. F

2. The energy required to form a chemical bond between two neutral atoms is called **bond energy**. T

3. According to the **octet rule**, chemical compounds tend to form so that each atom has an octet of electrons in its highest energy level. T

4. When the orbitals of the hydrogen atoms in a hydrogen molecule overlap, the **electron density** between the nuclei decreases. T

5. A **polyatomic ion** is a group of covalently bonded atoms. T

6. NaCl is a **molecular compound**. F ionic

7. **Lattice energy** is released when the ions in a crystal lattice are separated from each other. F

8. An **ionic compound** is composed of independent, neutral elements that can be isolated from each other. F

9. The atoms in **molecular compounds** are held together by covalent bonds. T

C. Skill Builder

1. Use the electronegativity values listed in Table 1 to classify the bonds that would form between atoms of each of the following elements listed in Table 2.

Table 1

Element	Electronegativity	Element	Electronegativity
Hydrogen (H)	2.1	Oxygen (O)	3.5
Sodium (Na)	0.9	Carbon (C)	2.5
Calcium (Ca)	1.0	Cobalt (Co)	1.8
Lithium (Li)	1.0	Nitrogen (N)	3.0

Table 2

Bond	Type of Bond	Bond	Type of Bond
H and O	polar	N and O	polar
H and C	polar	Li and Ca	metallic → ignore
Na and Co polar	metallic	N and Ca	ionic
Li and N	ionic	Na and O	ionic
Ca and H	polar	C and O	polar

ignore ← Na and Co

2. Bond energy is related to bond length. Use the data in the tables below to arrange the bonds listed in order of increasing bond length, from shortest bond to longest.

a.

Bond	Bond energy (kJ/mol)	Order
H–F	569	1
H–I	299	4
H–Cl	432	2
H–Br	366	3

b.

Bond	Bond energy (kJ/mol)	Order
C—C	346	3
C≡C	835	1
C=C	612	2

shortest

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

Ionic Bond	Lattice energy (kJ/mol)	Order
NaCl	–787	2
CaO	–3384	4
KCl	–715	1
MgO	–3760	5
LiCl	–861	3

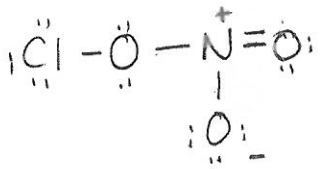
4. Arrange the following types of attractions between atoms in order of increasing strength, with 1 being the weakest and 4 the strongest.

Bonds are always stronger than IMF

3 covalent
4 ionic
} these are bonds

2 dipole-dipole
1 London dispersion
} these are IMF

5. The chlorine nitrate molecule ($ClONO_2$) is believed to be involved in the destruction of ozone in the Antarctic stratosphere. Draw a plausible Lewis structure for this molecule.



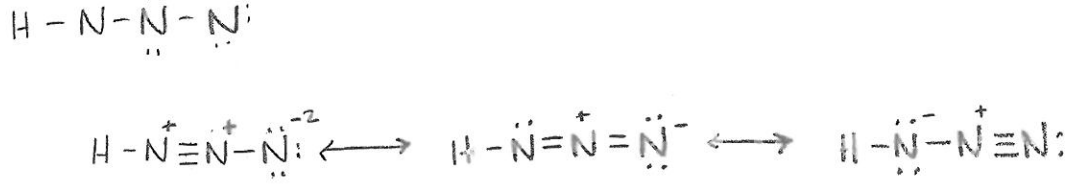
(don't need +/-)

7
6
5
+12
30

6. Draw three resonance structures for hydrazoic acid, HN_3. The atomic arrangement is HNNN. Calculate and show the formal charges on each structure.

*NO formal charges

15
+
16e⁻



7. Imagine that an ionic bond is formed between a cation, A^+ and an anion, B^-. How would the energy of the ionic bond be affected by the following changes? You would use Coulomb's law to solve this problem:
You would use Coulomb's law to answer this question:

$$E = k \frac{Q_{cation} Q_{anion}}{r}$$

a. Doubling the radius of A^+. Will E increase, decrease or stay the same?

↓

b. Tripling the charge on A^+. Will E increase or decrease, and by what factor? (i.e be halved, doubled, etc.)

3× larger

c. Doubling the charges on A^+ and B^-. Will E increase or decrease, and by what factor?

4× larger

d. Decreasing the radii of both A^+ and B^- to half their original values. Will E increase or decrease, and by what factor?

4× larger

10.17
10.21

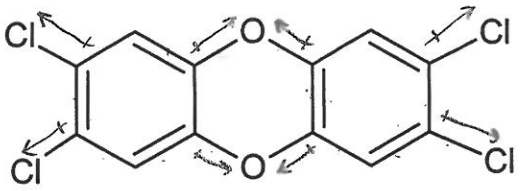
8. List the following molecules in order of increasing dipole moment: H_2O, CBr_4, H_2S, HF, NH_3, CO_2.

$CO_2 < CBr_4 <$ $< HF$

9. What types of hybrid orbitals are in each of the following molecules?

Molecule	Hybrid Orbitals
a. $H_3C–CH_3$	sp^3
b. $CH_3–C\equiv C–CH_2OH$	sp, sp^3
c. $CH_3CH=O$	sp^3, sp^2
d. CH_3COOH	sp^3, sp^2
e. $H_3C–CH=CH_2$	sp^3, sp^2

10. TCDD, or 2,3,7,8-tsetrachlorodibenzo-p-dioxin, is a highly toxic compound. It gained considerable notoriety in 2004 when it was implicated in the murder plot of a Ukranian politician.



a. Describe its molecular geometry and state whether the molecule has a dipole moment or not.

_____

b. How many pi bonds and sigma bonds are there in the molecule?

sigma bonds 20 pi bonds 6

Standardized Test Prep

Circle the letter of the best answer.

1. A large difference in electronegativity between two atoms in a bond results in
   a. polar covalent bonding.
   b. nonpolar covalent bonding.
   c. ionic bonding.
   d. repulsion between the two atoms.

2. A chemical bond between atoms results from the attraction between electrons and
   a. protons
   b. neutrons
   c. isotopes
   d. Lewis structures

3. A covalent bond consists of
   a. a shared electron
   b. a shared electron pair
   c. two different ions
   d. an octet of electrons

4. If two covalently bonded atoms are identical, the bond is identified as
   a. nonpolar
   b. polar covalent
   c. nonionic
   d. dipolar

5. A covalent bond in which there is an unequal attraction for the shared electrons is
   a. nonpolar
   b. polar
   c. ionic
   d. dipolar

6. Atoms with a strong attraction for electrons they share with another atom exhibit
   a. zero electronegativity
   b. low electronegativity
   c. high electronegativity
   d. Lewis electronegativity

7. Bonds that possess between 5% and 50% ionic character are considered to be
   a. ionic
   b. pure covalent
   c. polar covalent
   d. nonpolar covalent

need Bond character for final

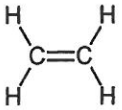
8. The greater the electronegativity difference between two atoms bonded together, the greater the bond's percentage of
   a. ionic character
   b. covalent character
   c. metallic character
   d. electron sharing

9. The molecular formula H_2O indicates a molecule
   a. with two atoms of oxygen bonded with one atom of hydrogen.
   b. with two atoms of hydrogen bonded with one atom of oxygen
   c. that includes only ionic bonds
   d. that is diatomic

10. Noble-gas atoms are able to exist independently in nature because
   a. they are exceptions to the octet rule.
   b. their bond energies are low compared to their bond lengths.
   c. their electron configurations are more stable than those of other atoms.
   d. they share electrons in overlapping orbitals with other noble-gas atoms.

11. The electron-dot notation of hydrogen is H•. This configuration
   a. that a hydrogen atom has only one occupied energy level.
   b. that hydrogen is a noble gas.
   c. that the n = 1 level of a hydrogen atom contains a single electron.
   d. both a and c.

12. Which of the following statements is true?
   a. A pair of dots in electron-dot notation represents an electron pair of a lone atom in a compound.
   b. Only valence electrons of a particular atom are shown in electron-dot notation.
   c. A single bond is formed when a single electron brings two atoms together.
   d. Covalent bond formation involves the electrons closest to the nuclei of the atoms.

13. The Lewis structure for C_2H_4 is  . This Lewis structure shows that

a. the bonds between carbon and hydrogen are double bonds.
b. each hydrogen atom has two valence electrons.
c. the bond between the carbon atoms is a double bond.
d. the bond energies of the carbon-hydrogen bonds are higher than the bond energy of the carbon-carbon bond.

14. Which statement is true of molecular compounds?
a. When dissolved in water, they conduct electricity.
b. They are hard, brittle substances.
c. Their lattice energies are positive rather than negative.
d. Some are completely gaseous at room temperature.

15. Which statement is true of ionic compounds?
a. The distances between the ions vary through out the crystal.
b. All ionic crystals have the same structure.
c. The ions cannot move in the solid state.
d. All ionic compounds are able to dissolve in water.

16. Which statement is true?
a. Calculating lattice energies helps chemists compare bond strengths in ionic compounds.
b. Table salt is a molecular compound.
c. Na_2Cl_2 is a formula unit.
d. Ionic crystals form so that the potential energy of a compound is maximized.

17. Which statement is true?
a. Molecular formulas represent the simplest formulas for compounds.
b. The forces holding ions together in ionic bonds are relatively weak.
c. Intermolecular attractions that hold molecules together are stronger than ionic action.
d. A polyatomic ion with a shortage of electrons has a positive charge.

18. VSEPR theory does *not* explain
a. the shapes of molecules.
b. how to use Lewis structures to predict the shapes of polyatomic ions.
c. the relationship between a molecule's geometry and occupied orbitals.
d. that double and triple bonds are treated as single bonds in describing molecular geometry.

19. The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their
a. chemical bonds.
b. London forces.
c. heats of vaporization.
d. polarity.

20. The lattice energy of compound A is greater 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 in compound B.
c. Compound B is probably a gas.
d. Compound A has larger crystals than compound B.

21. Why are metals able to conduct electricity?

a. The valence electrons that make up a metal are highly mobile.
b. Electrons are bound to individual ions that are held in place in metallic crystal structures.
c. Metals are shiny and malleable
d. The shortage of electrons in metals gives them a positive electrical charge.

22. A blacksmith can shape metal into a horseshoe. This shows that the metal is

a. ductile.
b. a good conductor of electricity.
c. malleable.
d. a liquid at room temperature.

23. In metals, the valence electrons are considered to be

a. attached to particular positive ions.
b. shared by all surrounding atoms.
c. immobile.
d. involved in covalent bonds.

24. A molecule with 2 atoms bonded to the central atom generally has which shape?

a. linear.
b. bent.
c. trigonal-planar.
d. either a or b. (depends on lone pairs)

25. Complete the following sentence: The higher the boiling point of a substance,

a. the weaker the force between particles.
b. the stronger the force between particles.
c. the more likely it is that the substance is nonpolar-covalent.
d. the more likely it is that the substance is polar-covalent.

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

a. molecules.
b. positive ions.
c. dipoles.
d. negative ions.

27. The forces of attraction between molecules in a molecular compound are
- **a.** stronger than the attractive forces in ionic bonding.
- **b.** weaker than the attractive forces in ionic bonding.
- **c.** approximately equal to the attractive forces in ionic bonding.
- **d.** equal to zero

28. Why are metals able to conduct electricity?
- **a.** The valence electrons that make up a metal are highly mobile.
- **b.** Electrons are bound to individual ions that are held in place in metallic crystal structures.
- **c.** Metals are shiny and malleable
- **d.** The shortage of electrons in metals gives them a positive electrical charge.

repeat

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

repeat

30. The fact that metals are malleable and ionic crystals are brittle is best explained in terms of their
- **a.** chemical bonds.
- **b.** London forces.
- **c.** heats of vaporization.
- **d.** polarity.

31. Which is true of an AB_4 molecule?
- **a.** The electron pairs separate to form a tetrahedron.
- **b.** The molecule contains hydrogen.
- **c.** The molecule has ionic bonds.
- **d.** The molecule has both single and double bonds.

32. Which of the following affects molecular geometry?
- **a.** the number of lone pairs of electrons on the atoms.
- **b.** the sizes of the different types of atoms.
- **c.** the hybridization of some of the orbitals of the atoms.
- **d.** all of the above.

33. Complete the following sentence: The higher the boiling point of a substance,
- **a.** the weaker the force between particles.
- **b.** the stronger the force between particles.
- **c.** the more likely it is that the substance is nonpolar-covalent.
- **d.** the more likely it is that the substance is polar-covalent.

34. In general, intermolecular forces are
- **a.** stronger than bonds that join atoms in molecules.
- **b.** weaker than bonds that join atoms in molecules, but stronger than ionic bonds.
- **c.** stronger than bonds that join metal atoms in solid metals.
- **d.** weaker than bonds that join atoms in molecules and ions in ionic compounds.

you can do this one

35. Which statement is true?
   a. London forces explain the high boiling point of water.
   b. Individual bond dipoles are contained only in ionic solutions.
   c. Hydrogen bonding explains the high boiling point of ammonia.
   d. Intermolecular forces are the forces of attraction between atoms.

36. In a crystal of an ionic compound, each cation is surrounded by a number of
   a. molecules.
   b. positive ions.
   c. dipoles.
   d. negative ions.

37. The notation for sodium chloride, NaCl, stands for one
   a. formula unit.
   b. molecule.
   c. crystal.
   d. atom.