Bonding - Review

Covalent Bonding in Molecular Compounds

1. How is the bonding in O_2 different than in MgO? Give more of an explanation than just the name for the type of bonding in each.
   O_2 is held together by a covalent bond (electrons shared) and MgO is held together by an ionic bond (transfer of electrons)

2. Give the name or chemical formula for each of the following substances and in each case predict whether the bonding is better described as ionic or covalent:

a) iron(II) fluoride – FeF_2 - ionic
b) chromium(III) oxide – Cr_2O_3 - ionic
c) selenium dibromide – SeB_{r2} - covalent
d) SF_6 – sulphur hexafluoride - covalent
e) $CoCl_2$ – cobalt chloride - ionic
f) $ScBr_3$ – scandium bromide - ionic
g) nitrogen dioxide – NO_2 - covalent
h) hydrogen sulfide - H_2S - covalent

3. What happens when a covalent substance dissolves? When it melts?
In each case, intermolecular forces are broken, but molecules remain whole. When a covalent compound dissolves, molecules are solvated (surrounded by solvent molecules)

4. Why do covalent compounds (in general) have lower melting and boiling points than ionic compounds?

   Because intermolecular forces/interactions are weaker than ionic or covalent bonds. To melt/boil a covalent compound you only need to break intermolecular forces, but to melt/boil an ionic compound, ionic bonds must be broken which takes a lot of energy

Lewis Structures

5. What is a Lewis structure and how does it differ from a Lewis dot symbol?
   Lewis structure is for a molecule, but Lewis dot symbol is only for a single element

6. Construct Lewis structures for F_2, O_2, and N_2 in which each atom achieves an octet of electrons. Explain how the covalent bonds differ for each molecule. How many pi and sigma bonds are present in each molecule?
   F_2 – single covalent bond, O_2 – double covalent bond, N_2 – triple covalent bond

7. Draw Lewis Structures for each of the following molecules and show the number of valence electrons in () below each chemical formula:

a) NCl_5	e) PF_3	i) OF_2
b) N_2O_4	f) CO_3^{2-}	j) O_3
c) CO	g) $C_2H_4(OH)_2$	k) BF_3
d) C_2FH_3	h) ClF_5	l) SO_2
m) HCN	o) H_2CO	q) HBr
n) CHC_{l3}	p) SeF_2	r) NH_4^+

s) PO_4^{3-}
t) H_3O^+
u) NO_2^-
v) NO_3^-

8. Draw Lewis structures for the following ions: NO_3^{-1}, PO_4^{-3}, and SO_4^{-2}. Explain why each of these ions is stable with the appropriate charge.

9. Urea, a substance produced in protein metabolism, has a central C atom bonded to an O and two N atoms, NH_2CONH_2. Draw the Lewis structure for urea.

10. Three different Lewis structures are possible for N_2O. Draw all three. Which one is the best description of the bonding in this molecule? Why?

Molecular Shapes and VSEPR

11. Why is it important to understand and be able to predict the shape of molecules?
Molecular shapes affect the properties of the substances (polarity, melting and boiling point, solubility, etc) and help us to explain these properties.

12. What factors do you consider in determining the molecular shape of a molecule?
Number of bonding electron pairs and number of lone pairs

13. What is the VSEPR model for determining molecular shape? What are the limitations of this model?
Valence shell electron pair repulsion. This model predicts shape by explaining that electron pairs repel each other and try to get as far away from each other as possible. In reality, it does not account for all of the factors influencing the shape of a molecule, and makes no predictions about bond length

14. Define:
   a) bonding pair – pair of electrons involved in a bond
   b) nonbonding pair – pair of electrons on central atom that are not part of a bond
   c) electron geometry – arrangement of all electron pairs around the central atom
   d) molecular geometry – arrangement of all atoms around the central atom

15. Describe the effect of nonbonding electrons on bond angles. Where do the nonbonding electrons have to be to affect molecular shape?
Non bonding electrons generally have the effect of decreasing the bond angle between the bonding electron pairs. To affect molecular shape, the non-bonding electrons must be on the central atom

16. The Lewis Structure for which of the following molecules would have two non-bonded pairs of electrons on the central atom?

a) H_2S b) NH_3 c) CH_4 d) HCN e) CO_2

17. For the following descriptions of molecular shape sketch the geometric object on which it is based and sketch a ball and stick drawing of the shape:

a) trigonal planar
b) tetrahedral
c) trigonal pyramidal
d) bent
e) trigonal bipyramidal
f) octahedral
g) seesaw
h) t-shaped
i) square pyramidal
j) square planar

18. Rank the following VSEPR molecular shapes in order of decreasing bond angle: tetrahedral, linear, pyramidal, and triangular. *Linear>trig. Planar (triangular) > tetrahedral> trig. pyramid*

19. Use VSEPR theory to predict the shape of each molecule below and predict whether each molecule is polar or non-polar. Identify each molecule's shape by name, indicate the bond angles and the type of hybridization around the central atom(s)

a) PCl_3
b) N_2O_4
c) NF_3
d) SO_2
e) $C_2H_4(OH)_2$
f) OF_2
g) SO_3
h) BH_3
i) I_3^{-1}
j) NO
k) XeO_2F_2
l) IF_3
m) ClF_5
n) Hydrogen phosphide
o) Hydrobromic acid
p) Sulfuric acid
q) Phosphoric acid
r) Nitric acid
s) Xenon hexafluoride
t) H_2O_2
u) $HClO_4$

20. Describe the difference between a polar covalent bond and a non-polar covalent bond.
In a polar covalent bond, electrons are unequally distributed, while in a non-polar bond, electron sharing is equal

21. Which of the following bonds are polar and which is the more electronegative atom in each polar bond? (use the values listed on your periodic table to answer this question)
a) C - Cl
b) C - O
c) P - F
d) N - Br
e) Si - Si

22. Arrange the bonds in each of the following sets in order of increasing polarity:
a) C - F, O - F, Be – F *Be-F>C-F>O-F*
b) N - Br, P - Br, O – Br P-Br>N-Br>O-Br
c) C - S, B - P, N - O

23. How does a polar molecule differ from a non-polar one?
Polar molecule has one region of the molecule with a partial positive charge and another region of the molecule with a partial negative charge. In a non-polar molecule, no charges are apparent.

Metallic Bonding
24. What is a metallic bond?
A sea of positive metal nuclei held together by their attraction for the delocalized electrons in between them

25. Explain how the high electrical conductivity of metals is explained by the idea of metallic bonds.
Because delocalized electrons can move, they are able to carry charges and hence conduct electricity

26. If you were given a strange unknown solid what are some ways you could determine if it was an ionic compound or a metal?
Metals would conduct electricity in the solid state while ionic compounds would not

<u>Covalent bond characteristics</u>
27. What is a covalent bond? Electrons are shared How does it differ from an ionic bond? In an ionic bond, electrons are transferred

28. What is a single covalent bond? Sharing of two electrons Why does it form? To fulfill the octet requirements of both atoms involved by sharing two electrons, when two orbitals pointing towards each other overlap.

29. Why do multiple bonds (double and triple bonds) form? *When 4 or 6 electrons are shared.* Explain the formation of multiple bonds in terms of orbital overlap. *Multiple bonds occur when pi bonds form which occur when parallel p orbitals overlap*

30. What is the difference between single, double, and triple bonds in terms of bond strength? Bond length? *Single bonds are weaker than double bonds which are weaker than triple bonds. The pattern for length is that single bonds are longest and triple bonds are the shortest*

31. What are the similarities and differences between the bonding in diamond and graphite? *Both are network covalent solids, but diamond has a 3D structure while graphite has a 2D structure with delocalized electrons in between the covalently bonded layers.* How are these differences in bonding reflected in the properties of the materials? *Because graphite has delocalized electrons it can conduct electricity. Because of the 3D network of covalent bonds in diamond, it is very strong and hard, with a very high melting point.*

<u>Characteristics of Ionic compounds/bonds</u>

32. What happens to an ionic compound when it dissolves? Ions dissociate and are surrounded by solvent molecules When it melts? Ionic bonds are broken and ions move around independently

33. Why do ionic compounds conduct electricity when dissolved or melted? *Ions are mobile and can carry charges. Mobile charge carriers are a requirement for electricty to flow*

34. What holds ionic crystals together? *electrostatic attraction between cations and anions*

35. Ionic compounds are generally hard, brittle, and have high melting and boiling points. How can these properties be explained by their bonding? *ionic compounds are not made of individual molecules but of a lattice of ions that are all bonded to their neighbouring counterions. To break the ions apart requires a lot of energy because the electrostatic attraction between the anions and cations is very strong.*

36. How could you predict that a bond between two elements would be ionic? *If one is a metal and one is a non-metal or if their electronegativity values are very different*

Intermolecular Forces

37. What are intermolecular forces? *Forces acting between different molecules*

38. What type of intermolecular forces are present in the following compounds?
    a. Water *h - bonds*
    b. Ammonia *h-bonds*
    c. Iodine *London/dispersion forces*
    d. Ethanol (CH_3CH_2OH) h -bonds
    e. Hydrogen chloride dipole-dipole interactions
    f. Hydrogen fluoride h - bonds
    g. Carbon dioxide London/dispersion forces
    h. Carbon monoxide dipole-dipole forces
    i. Hexane ($CH_3CH_2CH_2CH_2CH_2CH_3$) London/dispersion forces

39. Why do ionic compounds dissolve in polar solvents but not in non-polar solvents? *Polar solvents are attracted to the anions and cations in an ionic solid and can solvate the ions effectively. Non-polar solvents are not attracted to the charged particles and they cannot overcome the ionic bonds holding the crystal together*

40. Why does water have a much higher boiling point than H_2S? h-bonds

Orbital hybridization

41. What are orbitals?
Regions where there is a high probability of finding electrons

42. How are hybrid orbitals created?
Combining the mathematical functions that describe 2 or more orbitals together and coming up with a new function that describes a new shape of orbital

43. What type of hybridization exists in the central atom(s) in each of the following molecules?
    a) CO_2 - sp
    b) CH_4 – sp3
    c) NH_3 – sp3
    d) CH_2O – sp2
    e) HCN - sp
    f) PCl_5 – sp3d

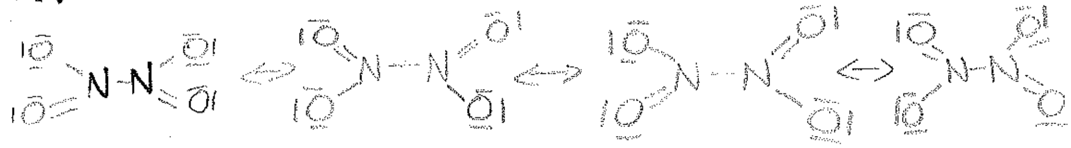
44. How is a sigma bond formed?
Overlap of two orbitals that are pointed towards each other

45. What is a pi bond? How is it formed?
Bond formed by overlap of two parallel p orbitals.

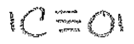
46. How many sigma and pi bonds are there in each of the following molecules?
    a) C_2H_2 – 3 sigma, 2 pi
    b) C_2H_4 – 5 sigma, 1 pi
    c) N_2H_2 – 5 sigma
    d) CO_2 – 2 sigma, 2 pi

7. a) NCl_5
$5 + 5 \times 7 = 40$
omit

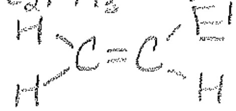
b) N_2O_4
$2(5) + 4(6) = 34$



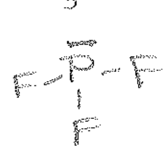
c) $CO \rightarrow 10e^-$



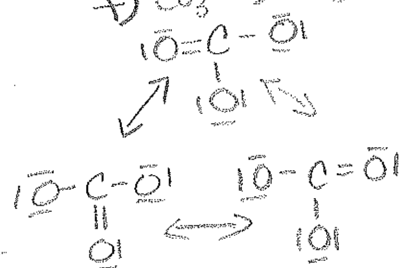
d) $C_2FH_3 \rightarrow 18e^-$



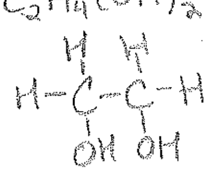
e) $PF_3 \rightarrow 26e^-$



f) $CO_3^{2-} \rightarrow 24e^-$



g) $C_2H_4(OH)_2 \rightarrow 26e^-$



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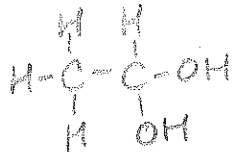
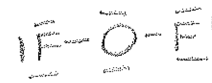


h) $ClF_5 \rightarrow 42e^-$

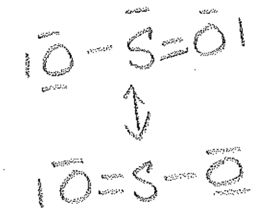
i) $OF_2 \rightarrow 20e^-$



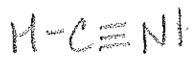
j) $O_3 \rightarrow 18e^-$

k) $BF_3 \rightarrow 24e^-$

l) $SO_2 \rightarrow 18e^-$

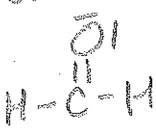


m) $HCN \rightarrow 10e^-$

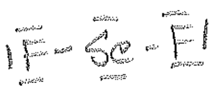


n) $CHCl_3 \rightarrow 26e^-$

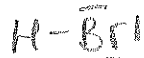
o) $H_2CO \rightarrow 12e^-$



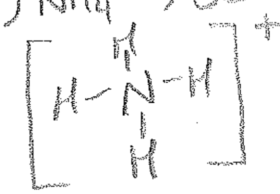
p) $SeF_2 \rightarrow 20e^-$



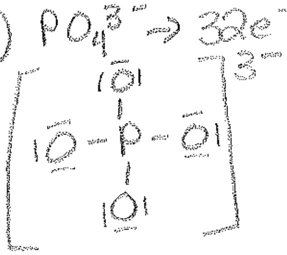
q) $HBr \rightarrow 8e^-$



r) $NH_4^+ \rightarrow 8e^-$

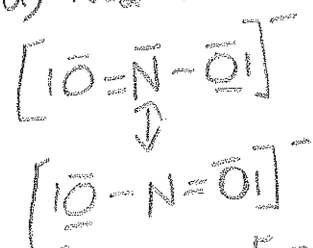


s) $PO_4^{3-} \rightarrow 32e^-$

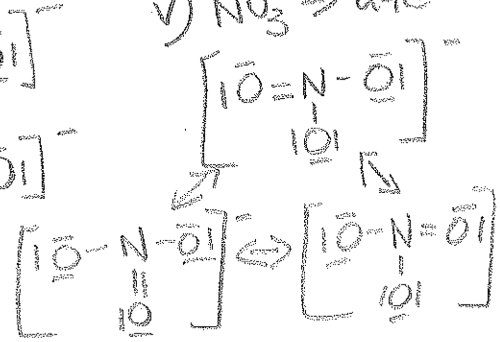


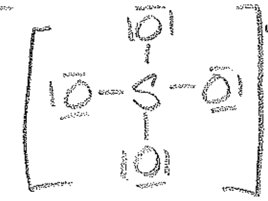
t) $H_3O^+ \rightarrow 8e^-$

u) $NO_2^- \rightarrow 18e^-$



v) $NO_3^- \rightarrow 24e^-$



8. NO_3^- - see #7, PO_4^{3-} - see #7, SO_4^{2-} 

- with less negative charge, these polyatomic ions could not achieve full octets for all atoms

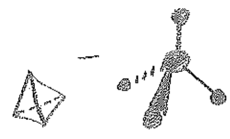
9. 

10. $N_2O \rightarrow 16e^-$

$|\overline{N}=O=\overline{N}|$ OR $|N\equiv O-\underline{\overline{N}}|$

OR $|\underline{\overline{N}}-O\equiv N|$ best picture is probably some blend of all three

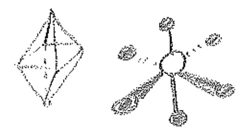
17. a)  - triangle

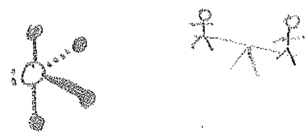
b) 

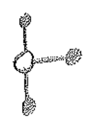
c) 

d) 

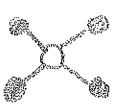
e) 

f) 

g) 

h) 

i) 

j) 

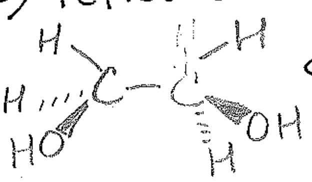
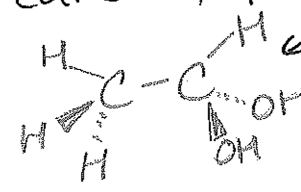
19. a) trigonal pyramid, polar, sp^3, 107.5°

b) trig planar around each N, 120°, sp^2, non-polar

c) trig pyramid, polar, sp^3, 107.5°

d) bent/v-shaped, polar, sp^2, ~120°

e) tetrahedral around each carbon, sp^3, 109.5°

 ← non-polar OR  ← polar

f) bent/v-shaped, polar, sp^3, ~104°

g) trig. planar, sp^2, non-polar, 120°

h) trig. planar, sp^2, non-polar, 120°

i) linear, non polar, 120°, sp

j) linear, polar - oops, omit, odd # e^-'s

k) see-saw, polar, sp^3d, complex bond angles

l) t-shaped, polar, sp^3d, 90° or 180°

m) square pyramid, polar, sp^3d^2, 90°

n) trig. pyramid, polar, sp^3, 107.5 (H_3P)

o) linear, polar, Br - sp^3

p) SO_4^{2-} only → tetrahedral, non-polar, sp^3, 109.5°

q) PO_4^{3-} only ⇒ tetrahedral, non-polar, sp^3, 109.5°

r) NO_3^- only → trig. planar, non-polar, sp^2, 120°

s) too many e^-'s on central atom to name the shape

t) bent around each O, polar, sp^3, ~104°

u) ClO_4^- only ⇒ tetrahedral, non-polar, sp^3, 109.5°

21. a) $\underset{\delta+}{C} \rightarrow \underset{\delta-}{Cl}$ c) $\underset{\delta+}{P} \rightarrow \underset{\delta-}{F}$ b) $\underset{\delta+}{C} \rightarrow \underset{\delta-}{O}$ d) N — Br non-polar e) non polar

22. a) Be-F > C-F > O-F b) $\underset{\delta+}{P} \rightarrow \underset{\delta-}{Br}$ > $\underset{\delta-}{O} \leftarrow \underset{\delta+}{Br}$ > N - Br non polar

c) $\underset{\delta+}{N} \rightarrow \underset{\delta-}{O}$ > $\underset{\delta+}{B} - \underset{\delta-}{P}$ > C - S non-polar