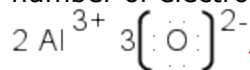


*CH301 Fall 2009 Worksheet 4 Answer Key: A Lovely Worksheet on part of Chapter 2. Great for Exam 1.

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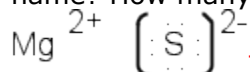
1. Draw Lewis structure for the following ionic compounds and answer the questions posed.

a. A salt formed from aluminum and oxygen. What is the unit formula? What is the total number of electrons exchanged? Which species gives up electrons?



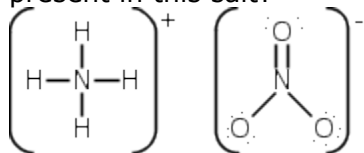
The unit formula is Al_2O_3 . A total of 6 electrons are exchanged and it is the aluminum atoms that give up electrons to form aluminum ions.

b. A salt formed from magnesium and sulfur. What is the unit formula? What would be its name? How many electrons are exchanged?



The unit formula is MgS . Its name would be magnesium sulfide. Two electrons are exchanged in total.

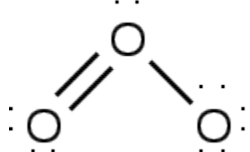
c. A salt formed from ammonium (NH_4^+) and nitrate (NO_3^-). What sort of bonds are present in this salt?



Both covalent bonds (within the polyatomic ions) and ionic bonds (between the polyatomic ions) are present.

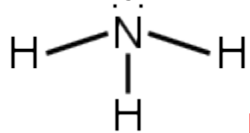
2. Draw Lewis structure for the following species and decide which would exhibit resonance. Name the electronic geometry of each molecule as well.

a. O_3



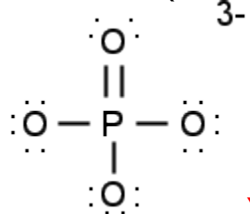
Yes, this would exhibit resonance - the double bond could be drawn between the central O atom and either of the peripheral O atoms. Trigonal planar.

b. NH_3



No, this would not exhibit resonance - there are no double bonds, and therefore no possible alternative structures. Tetrahedral.

c. PO_4^{3-} (see the "formal charge" topic in your course notes to get the correct structure.)



Yes, this would exhibit resonance - the double bond could be drawn between the central P atom and any of the peripheral O atoms. Tetrahedral.

d. HCN



No, this would not exhibit resonance - the triple bond between C and N cannot have its electrons redistributed as H can only form one bond. Linear

3. How many resonance structures would each of the following species exhibit?

a. C₆H₆, aka benzene (hint: the carbons form a ring and the molecule is symmetrical)

2

b. NO₂

2

c. H₂O

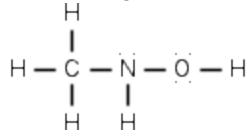
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4. Define resonance in your own words.

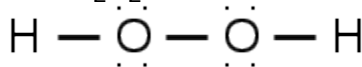
Resonance describes when a single Lewis structure is insufficient to describe the location of all electrons in a molecule. An equivalent statement is that resonance describes molecules for which distinct but degenerate Lewis structures can be drawn.

5. Draw the Lewis structure for each of the following.

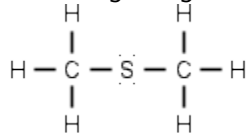
a. CH₃NHOH



d. H₂O₂

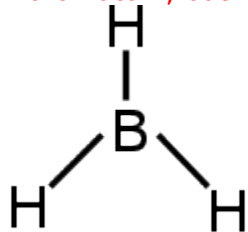


c. CH₃SCH₃



6. Members of which group on the periodic table would be most likely to form stable compounds with fewer than 8 valence electrons? Why? Give an example, draw its Lewis structure.

Members of Group III (group 13 in the IUPAC convention, aka the triels or the earth metals), especially Boron. This is because they have only 3 valence electrons and thus can only easily form 3 bonds. BH₃ (aka borane) would have only 6 valence electrons on the Boron atom, but would nonetheless be stable.



7. Which element always fails to satisfy the octet "rule" and yet is always "happy?"

Hydrogen, when bonded in a molecule, always has exactly 2 electrons, which for it constitutes a filled outer shell and consequently a stable configuration.

8. a. An atom, ion or molecule which has an odd number of valence electrons is called what?

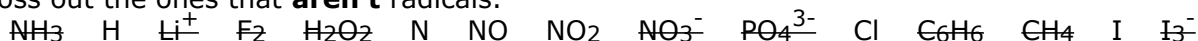
A radical, or sometimes a "free radical."

b. In such a molecule, which atom will most likely have the unpaired electron?

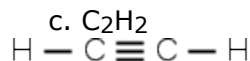
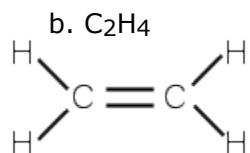
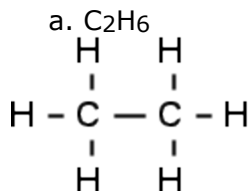
Typically, the least electronegative atom will have the unpaired electron because its more

electronegative counterparts are better able to attract and retain full pairs.

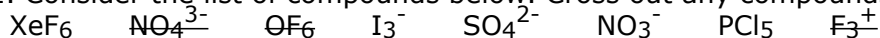
9. Without drawing Lewis structures, consider the atoms, ions and molecules below and cross out the ones that **aren't** radicals.



10. Draw Lewis structures for the compounds below.



11. Consider the list of compounds below. Cross out any compounds that **cannot** exist.



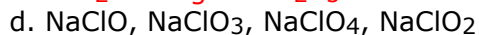
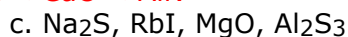
12. What do the molecules that you crossed out in question 11 have in common? Why can't they exist?

What they all have in common is a central atom which is found in the second row of the periodic table and which would have more than 8 valence electrons. This is impossible because hypervalence (aka, "too big" or expanded valence) requires that the atom in question has empty *d* orbitals available in its outermost shell. No member of row 2 will satisfy this requirement because there is no 2*d* subshell.

13. What are the two main factors that determine lattice energy? Which of our periodic trends is most useful here? Which of the two makes the greater contribution to differences in lattice energy observed in different salts?

The two main factors are the charges of the ions involved and the distances separating the ions. We can use the ionic radius trend to make qualitative comparisons about the distances between ions. Because the charge comes in integer increments (i.e. 1, 2, 3, 4 and so on), and the radii of common ions vary relatively little (between 50 and 140 pm typically), charge will have a greater impact on differences in lattice energy.

14. Rank the following sets of salts from least to greatest lattice energy.

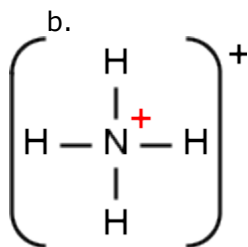
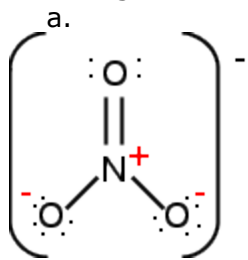


15. Calculate the difference in electronegativity (ΔEN) for the following diatomic species. Approximations are fine.

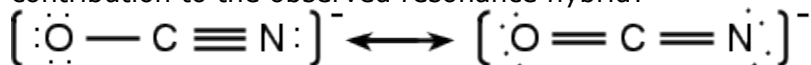
- a. LiH $\Delta EN = 2.1 - 1.0 = 1.1$
- b. BeO $\Delta EN = 3.5 - 1.5 = 2.0$
- c. HF $\Delta EN = 4.0 - 2.1 = 1.9$
- d. BN $\Delta EN = 3.5 - 2.5 = 1.0$
- e. LiF $\Delta EN = 4.0 - 1.0 = 3.0$

16. Rank the following species from lowest to highest ΔEN : IBr, ClI, IF, F₂, BrF
 $F_2 < IBr < ClI < BrF < IF$

17. Assign formal charges to each atom in the following species.



18. Consider the two resonance structures of cyanate below. Which makes the greater contribution to the observed resonance hybrid?



The left resonance structure likely makes the greater contribution to the resonance hybrid. Since the overall charge on cyanate is -1, there must be a negative formal charge somewhere within the molecule. The most likely place for this additional negative charge to go is the most electronegative atom, in this case oxygen.

19. What is the proportionality between...

- a. ΔEN and bond energy?

These are directly proportional. Large ΔEN corresponds to greater bond energy.

- b. ΔEN and bond length?

These are inversely proportional. Large ΔEN corresponds to shorter bond length.

- c. bond energy and bond length?

These are inversely proportional. Strong bonds have short lengths.

20. Rank the following species from (____): IBr, ClI, IF, F₂, BrF

- a. greatest to least bond energy

$IF < BrF < ClI < IBr < F_2$

- b. shortest to longest bond length

$IF < BrF < ClI < IBr < F_2$