CH 221 Chapter Four Concept Guide

1. Lewis Structures

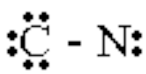
Problem
Draw the Lewis Dot Structure for cyanide ion, CN^-.

Solution
Step 1. Add valence electrons.

1 C at 4 electrons	=	4 electrons
1 N at 5 electrons	=	5 electrons
-1 charge	=	+ 1 electron
Total	=	10 electrons

Step 2. Place a bond between C and N to represent the sharing of 2 electrons. 8 electrons remain.
C-N

Step 3. Place lone pairs of the remaining 8 electrons around the outside atoms.



Step 4. In Step 3, nitrogen has only 4 valence electrons. Multiple bonds are needed to satisfy the octet rule. Two pairs of electrons are moved from C to form bonds between C and N.
:C≡N:

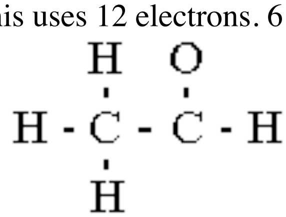
2. Lewis Structures

Problem
Draw the Lewis Dot Structure for acetaldehyde, CH_3CHO. Note: the right-hand H atom is bonded to the right-hand C atom.

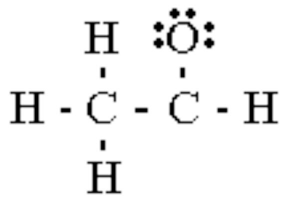
Solution
Step 1. Add valence electrons.

2 C at 4 electrons each	=	8 electrons
4 H at 1 electron each	=	4 electrons
1 O at 6 electrons	=	6 electrons
Total	=	18 electrons

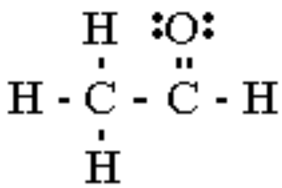
Step 2. Write the structure of CH_3CHO and place a bond between atoms to represent the sharing of 2 electrons. This uses 12 electrons. 6 electrons remain.



Step 3. The structure in Step 2 shows one of two carbon atoms and the oxygen atom without octets. Place lone pairs using the remaining 6 electrons around the oxygen atom.



Step 4. Multiple bonds are necessary to give carbon an octet, so a lone pair is moved from the oxygen atom to form a second bond between C and O.



3. Resonance Structures

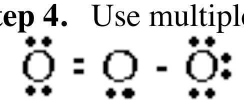
Problem
Draw all resonance structures for O_3.

Solution
Step 1. Add valence electrons.
3 O at 6 electrons each = 18 electrons

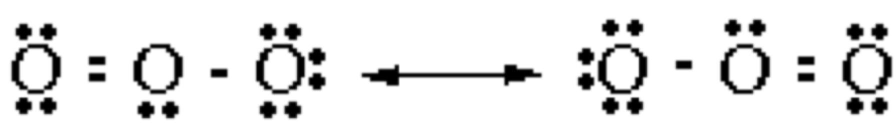
Step 2. Place a bond between atoms. This uses 4 electrons. 14 electrons remain.
O-O-O

Step 3. Place lone pairs of the remaining 14 electrons around the oxygen atoms, starting with the terminal atoms. After each terminal O atom has 8 electrons, 2 electrons remain and are placed on the central atom.

Step 4. Use multiple bonds to obtain octets around each O.



Step 5. Electron delocalization leads to two resonance structures for O_3.



4. Lewis Structures

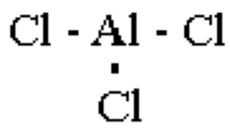
Problem
Draw the Lewis Dot Structure for $AlCl_3$.

Solution

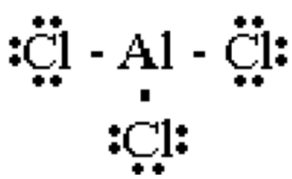
Step 1. Add valence electrons.

1 Al at 3 electrons	=	3 electrons
3 Cl at 7 electrons each	=	21 electrons
Total	=	24 electrons

Step 2. Place a bond between the atoms to represent the sharing of 2 electrons. 18 electrons remain.



Step 3. Place lone pairs of the remaining 18 electrons around the outside, terminal atoms. Each chlorine atom should have an octet of electrons. Neither Al nor Cl can form multiple bonds, so even though Al has only 6 electrons, this is the final Lewis structure.

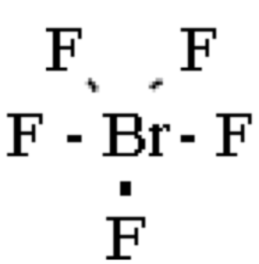


5. Lewis Structures

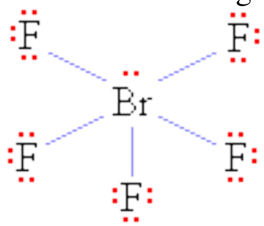
Problem
Draw the Lewis Dot Structure for BrF_5.
Solution:
Step 1. Add valence electrons.

1 Br at 7 electrons	=	7 electrons
5 F at 7 electrons each	=	35 electrons
Total	=	42 electrons

Step 2. Place a bond between Br and F atoms to represent the sharing of 2 electrons. 32 electrons remain.



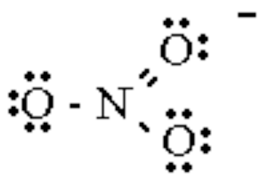
Step 3. Place lone pairs of the remaining 32 electrons around the 5 fluorine atoms. Each F atom has an octet of electrons. The remaining electrons are placed on the central atom.



6. Oxidation Numbers and Formal Charge

Problem

Determine the oxidation number and formal charge for N in the nitrate ion, NO_3^-.



Element	Electronegativity Values
H	2.1
B	2.0
C	2.5
N	3.0
O	3.5
F	4.0
Cl	3.0

Solution

Step 1. According to oxidation number rules, all valence electrons are considered to be held by the more electronegative atom. Therefore, each O atom has an oxidation number of –2. There are three oxygens for a total of –6. The charge on the ion is –1. Solving for x, where x is the oxidation number for nitrogen:

$$x + (3)(-2) = -1$$
$$x = +5$$

Nitrogen has an oxidation number of +5.

Step 2. The formal charge is based on the assumption that electrons are shared equally between covalently bonded atoms.

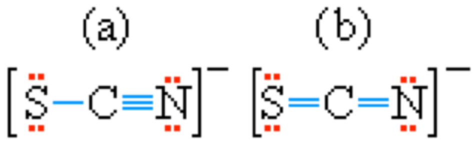
$$\text{Formal charge} = (\text{group number in periodic table}) - (\text{number of lone electrons}) - \tfrac{1}{2}(\text{number of bonding electrons})$$

For nitrogen, formal charge = 5 – 0 – 4 = +1. In this case, N has a formal charge of +1.

7. Resonance Structures

Question

Which is the better resonance structure for thiocyanate ion, SCN^-?



Solution

Step 1. Determine the formal charges on each element in both resonance structures.

$$\text{Formal charge} = (\text{group number in periodic table}) - (\text{number of lone electrons}) - \tfrac{1}{2}(\text{number of bonding electrons})$$

Formal Charge

	Structure (a)	Structure (b)
Sulfur	+1	0
Carbon	0	0
Nitrogen	–2	-1

Step 2. The structure that has the fewest number of atoms with a formal charge is the better structure. Consequently, structure (b) is the correct one.

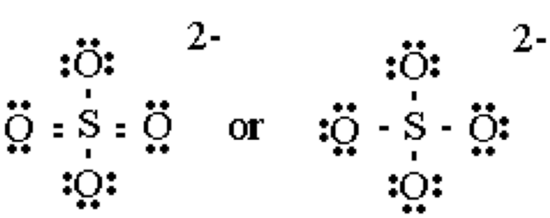
8. Using VSEPR: Predicting Geometries

Question

Using the VSEPR theory, what is the geometry of SO_4^{2-}?

Solution

Step 1. Draw the Lewis Dot Structure. There is more than one.



Step 2. Both structures have 4 bonded atoms and no lone pairs on the central S atom. Therefore, both the electron pair geometry and the molecular geometry are tetrahedral.
(Note: the VSEPR theory applies to both ions and molecules.) If the central S had one or more lone pairs, the electron pair and molecular geometries would differ.

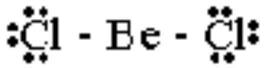
9. Using VSEPR: Predicting Geometries

Question

Using the VSEPR theory, what is the geometry of $BeCl_2$?

Solution

Step 1. Draw the Lewis Dot Structure.



Step 2. Neither Be nor Cl is able to form multiple bonds, so the central atom Be remains electron deficient. There are no lone pairs on the central atom, Be. The electron pair geometry is linear. This molecule is linear.

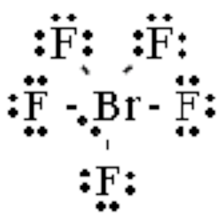
10. Using VSEPR: Geometry and Bond Angles

Problem

Based on the VSEPR theory, predict the electron-pair geometry for BrF_5 and F-Br -F bond angles.

Solution

Step 1. Draw the Lewis structure of the molecule.



Step 2. Determine the number of bonded pairs and lone pairs of electrons around the central atom. In this molecule, there are 5 bonded atoms and 1 lone pair around Br. There are 6 structurally significant electron pairs.

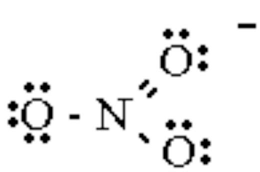
Step 3. The electron pair geometry is octahedral. The ideal molecular geometry is square pyramidal.

Step 4. The ideal bond angles in an octahedral structure are 90 degrees. The angles in BrF_5 should be compressed slightly from the ideal 90 degrees angle due to the lone pair of electrons on Br.

11. Bond Properties and Resonance

Problem

NO_3^- has three resonance structures and a bond order of 1.3. Predict the bond length for N-O.



N-O single bond length = 136 pm
N-O double bond length = 115 pm
N-O triple bond length = 108 pm

Solution

Bond length depends partly on bond order. An N-O bond order of 1.3 suggests that the N-O bond length is a value between that for a single and double N-O bond. The bond length is 122 pm, which is in fact the length of a single N-O bond added to one third the difference between single and double N-O bond lengths.
Note: in order to use the $\Delta H_{formation}$ data, all reactants must be in the gas phase.

12. Bond Polarity

Problem

Arrange the following covalent bonds in order of increasing polarity:

O-H I-Br C-F P-H S-Cl.

Electronegativity Values	
H	2.1
B	2.0
P	2.1
I	2.5
C	2.5
S	2.5
Br	2.8
N	3.0
O	3.5
F	4.0
Cl	3.0

Approach

Consider the differences in electronegativity values for each bonded pair of atoms.

Solution

The polarity of a bond increases with increasing difference in electronegativity of the bonded atoms. The covalent bonds, therefore, have the following order of increasing polarity with the electronegativity differences shown in parentheses:

→ Increasing electronegativity →

P-H (0.0) < I-Br (0.3) < S-Cl (0.5) < O-H (1.4) < C-F (1.5)

13. Molecular Polarity

Question

Is NF_3 polar or nonpolar?

Solution

NF_3 has the same pyramidal structure as NH_3. Fluorine is more electronegative than N, thus each bond is polar. The NF_3 molecule is asymmetrical and polar.

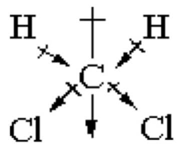
14. Molecular Polarity

Question

Is CH_2Cl_2 polar or nonpolar?

Solution:

In CH_2Cl_2, chlorine atoms are the most electronegative, followed by carbon atoms, then hydrogen atoms.

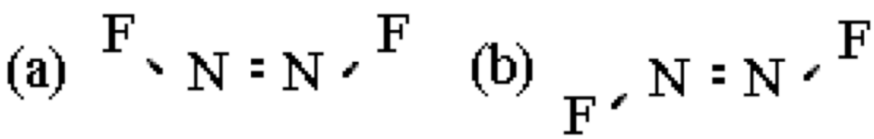


There is a net movement of electron density away from H atoms and toward Cl atoms. The asymmetric arrangement of the atoms, and the negative end of the bond, is toward the two Cl atoms. The positive end toward the 2 H atoms makes CH_2Cl_2 a polar molecule.

15. Molecular Polarity

Problem

There are two different molecules with the formula N_2F_2.   Is either molecule polar?



Solution

Fluorine atoms are more electronegative than nitrogen atoms. The negative ends of the bonds are toward the two fluorine atoms and the positive ends are toward the two nitrogen atoms. In molecule (a), the two dipoles do not cancel each other, thus this molecule is polar. In molecule (b), however, the two dipoles are opposite in direction and do cancel each other, making this molecule appear to be nonpolar. However, both nitrogens have a lone pair (not shown), making both molecules polar (although molecule (b) would be less polar than molecule (a).)