

Chem 115 POGIL Worksheet - Week 11 - Solutions
Chemical Bonds

Key Questions

1. Using the electronegativity values in the table above, calculate $\delta\chi$ for the pairs of elements forming bonds in each of the following species, and describe the type of bonding as either pure covalent, polar covalent, or ionic: H_2 , HCl , PbBr_2 , NaCl , NCl_3 .

H_2	$\delta\chi = 2.1 - 2.1 = 0$	pure covalent
HCl	$\delta\chi = 2.1 - 3.0 = 0.9$	polar covalent
PbBr_2	$\delta\chi = 1.8 - 2.8 = 1.0$	polar covalent
NBr_3	$\delta\chi = 3.0 - 2.8 = 0.2$	polar covalent (slightly)
NaCl	$\delta\chi = 0.9 - 3.0 = 2.1$	ionic

2. Calculate $\delta\chi$ for the bonds in $\text{HF}(g)$ and $\text{NCl}_3(l)$ and describe them.

$$\text{HF} \quad \delta\chi = |2.1 - 4.0| = 1.9 \quad \text{very polar covalent}$$

Although $\delta\chi$ suggests an ionic bond, HF is a gas and exists as a molecular species. Its bonds are very polar covalent, but HF does not contain discrete H^+ and F^- ions like a typical ionic compound.

$$\text{NCl}_3 \quad \delta\chi = |3.0 - 3.0| = 0 \quad \text{virtually pure covalent}$$

The zero value of difference in electronegativity suggests that there is very little polarity to the N–Cl bonds in NCl_3 . But because N and Cl are two different elements, there is actually some difference in their abilities to attract electrons in the bond. Pauling electronegativity values are simply too imprecise to indicate the polarity.

3. Explain the following trends in lattice energy.

- a. $\text{KF} < \text{CaO} < \text{ScN}$

The ion charges are +1/–1 for KCl , +2/–2 for CaO , and +3/–3 for ScN . Lattice energy rises sharply with increasing ion charge.

- b. $\text{NaCl} > \text{RbBr} > \text{CsBr}$

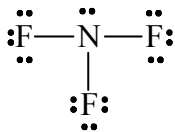
All of these are +1/–1 ion pairings, so size is the dominating factor. As ion size increases, the separation between oppositely charged ions increases, causing the forces of attraction to diminish.

- c. $\text{MgO} > \text{MgCl}_2$

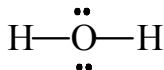
The Cl^- ion is larger and has a lesser charge than the O^{2-} ion. Both factors result in MgCl_2 having a lower lattice energy than MgO .

4. Following the steps given above, draw Lewis dot structures for the following molecular species: NF_3 , H_2O , NH_4^+ , SO_4^{2-} , H_2CO , C_2H_2 .

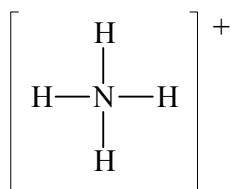
$$\text{NF}_3 \quad \text{N} + 3\text{F} = 5 + (3)(7) = 26$$



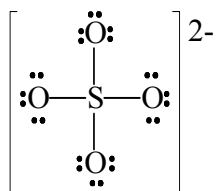
$$\text{H}_2\text{O} \quad \text{O} + 2\text{H} = 6 + (2)(1) = 8$$



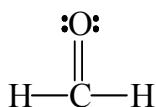
$$\text{NH}_4^+ \quad \text{N} + 4\text{H} + (+) = 5 + (4)(1) - 1$$



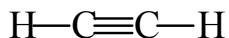
$$\text{SO}_4^{2-} \quad \text{S} + 4\text{O} + (2-) = 6 + 4(6) + 2 = 32$$



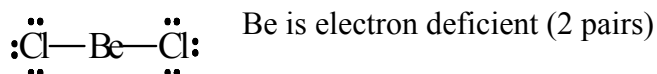
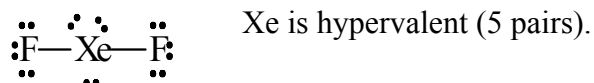
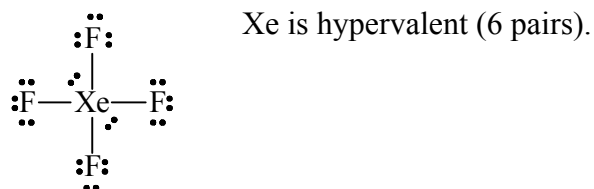
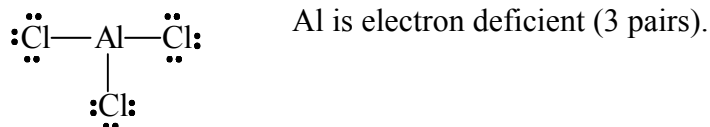
$$\text{H}_2\text{CO} \quad 2\text{H} + \text{C} + \text{O} = (2)(1) + 4 + 6 = 12$$



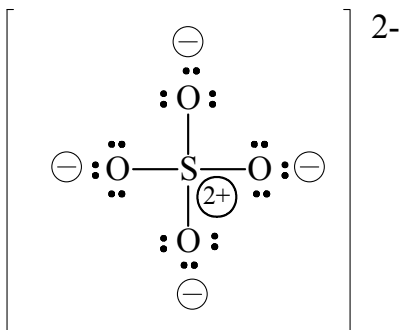
$$\text{C}_2\text{H}_2 \quad 2\text{C} + 2\text{H} = (2)(4) + (2)(1) = 10$$



5. Following the steps above, draw Lewis dot structures for the following molecular species, which may be electron deficient or hypervalent: AlCl_3 , XeF_4 , XeF_2 , BeCl_2

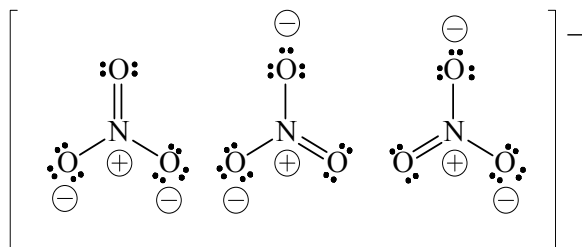


6. Show the formal charges on all atoms in the sulfate ion, SO_4^{2-} , whose Lewis dot structure you developed in Key Question 4.



7. The nitrate ion, NO_3^- , can be represented by a series of three equivalent resonance forms. Show the resonance forms of NO_3^- , including formal charges, and estimate the bond order of the N–O bonds in this ion.

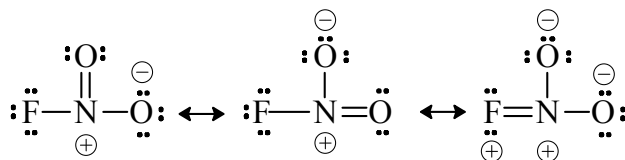
$$\text{NO}_3^- \quad \text{N} + 3\text{O} + (-) = 5 + (3)(6) + 1 = 24$$



As with ozone, the single and double bonds in these structures do not really exist in the nitrate ion. All N–O bonds are equivalent. The three N–O bonds in any one of the three equivalent resonance forms involve four pairs of electrons. Therefore the approximate bond order for a N–O bond in the nitrate ion is $4/3 = 1\frac{1}{3}$.

8. FNO_2 can be represented by three non-equivalent resonance forms. Draw the three canonical forms of FNO_2 and assess their relative contributions to the resonance hybrid description of the molecule. What is the approximate N–O bond order?

$$\text{FNO}_2 \quad \text{F} + \text{N} + 2\text{O} = 7 + 5 + (2)(6) = 24$$



The third form is insignificant (adjacent like formal charges). In the remaining forms, the two N–O bonds are approximately equivalent. The two N–O bonds use 3 pairs, so the bond order of each is about $3/2 = 1\frac{1}{2}$.