Solutions to suggested end-of-chapter problems for Ch. 8

Question numbers correspond to the 9th edition. The number in parentheses corresponds to the 10th edition.

25 (29). CO_2 , 4 + 2(6) = 16 valence electrons

The formal charges are shown above the atoms in the three Lewis structures. The best Lewis structure for CO_2 from a formal charge standpoint is the first structure having each oxygen double bonded to carbon. This structure has a formal charge of zero on all atoms (which is preferred). The other two resonance structures have nonzero formal charges on the oxygens, making them less reasonable. For CO_2 , we usually ignore the last two resonance structures and think of the first structure as the true Lewis structure for CO_2 .

27 (31). Using the periodic table, the general trend for electronegativity is:

- (1) Increase as we go from left to right across a period
- (2) Decrease as we go down a group

Using these trends, the expected orders are:

a.	C < N -	<0	b.	Se <	< S < 0	Cl	c.	Sn •	< Ge < Si	d.	Tl <	< Ge < S
28 (32).	a. Rb	< K < Na		b.	Ga <	B < 0		c.	Br < Cl < F		d.	S < O < F

- **29 (33).** The most polar bond will have the greatest difference in electronegativity between the two atoms. From positions in the periodic table, we would predict:
 - a. Ge–F b. P–Cl c. S–F d. Ti–Cl
- **30** (**34**). a. Sn-H b. Tl-Br c. Si-O d. O-F
- **33** (**37**). Use the electronegativity trend to predict the partial negative end and the partial positive end of the bond dipole (if there is one). To do this, you need to remember that H has electronegativity between B and C and identical to P. Answers b, d, and e are incorrect. For d (Br₂), the bond between two Br atoms will be a pure covalent bond, where there is equal sharing of the bonding electrons, and no dipole moment exists. For b and e, the bond polarities are reversed. In Cl–I, the more electronegative Cl atom will be the partial negative end of the bond dipole, with I having the partial positive end. In O–P, the more electronegative oxygen will be the partial negative end of the bond dipole, we used arrows to indicate the bond dipole. The arrow always points to the partial negative end of a bond dipole (which always is the most electronegative atom in the bond).

$$Cl - I = 0 - P$$

35 (39). Bonding between a metal and a nonmetal is generally ionic. Bonding between two nonmetals is covalent, and in general, the bonding between two different nonmetals is usually polar covalent. When two different nonmetals have very similar electronegativities, the bonding is pure covalent or just covalent.

a.	ionic	b.	covalent	с.	polar covalent
----	-------	----	----------	----	----------------

d. ionic e. polar covalent f. covalent

42 (**46**). Te^{2-} : [Kr] $5s^{2}4d^{10}5p^{6}$; Cl⁻: [Ne] $3s^{2}3p^{6}$; Sr²⁺: [Ar] $4s^{2}3d^{10}4p^{6}$; Li⁺: $1s^{2}$

44 (48). a.	Al ³⁺ and Cl ⁻ ; AlCl ₃ , aluminum chloride	b.	Na^+ and O^{2-} ; Na_2O , sodium oxide
с.	Sr^{2+} and F^- ; SrF_2 , strontium fluoride	d.	Ca ²⁺ and S ²⁻ ; CaS, calcium sulfide

46 (50). a. Sr^{2+} : [Ar] $4s^23d^{10}4p^6$; Cs⁺: [Kr] $5s^24d^{10}5p^6$; In⁺: [Kr] $5s^24d^{10}$; Pb²⁺: [Xe] $6s^24f^{14}5d^{10}$

b. P^{3-} and S^{2-} : [Ne]3s²3p⁶; Br⁻: [Ar]4s²3d¹⁰4p⁶

- **48 (52).** a. Cs_2S is composed of Cs^+ and S^{2-} . Cs^+ has the same electron configuration as Xe, and S^{2-} has the same configuration as Ar.
 - b. SrF_2 ; Sr^{2+} has the Kr electron configuration, and F^- has the Ne configuration.
 - c. Ca_3N_2 ; Ca^{2+} has the Ar electron configuration, and N^{3-} has the Ne configuration.
 - d. AlBr₃; Al³⁺ has the Ne electron configuration, and Br⁻ has the Kr configuration.

54 (58). a. $V > V^{2+} > V^{3+} > V^{5+}$ b. $Cs^+ > Rb^+ > K^+ > Na^+$ c. $Te^{2-} > I^- > Cs^+ > Ba^{2+}$ d. $P^{3-} > P^- > P$ e. $Te^{2-} > Se^{2-} > S^{2-} > O^{2-}$

56 (60). a. LiF; Li⁺ is smaller than Cs⁺. b. NaBr; Br⁻ is smaller than Γ .

c. BaO; O^{2-} has a greater charge than Cl⁻. d. CaSO₄; Ca²⁺ has a greater charge than Na⁺.

e. K_2O ; O^{2-} has a greater charge than F⁻. f. Li₂O; both ions are smaller in Li₂O.

58 (62).	$Mg(s) \rightarrow Mg(g)$	$\Delta H = 150. \text{ kJ}$	(sublimation)
	$Mg(g) \rightarrow Mg^+(g) + e^-$	$\Delta H = 735 \text{ kJ}$	(IE_1)
	$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	$\Delta H = 1445 \text{ kJ}$	(IE_2)
	$F_2(g) \rightarrow 2 F(g)$	$\Delta H = 154 \text{ kJ}$	(BE)
	$2 \operatorname{F}(g) + 2 \operatorname{e}^{-} \rightarrow 2 \operatorname{F}^{-}(g)$	$\Delta H = 2(-328) \text{ kJ}$	(EA)
	$Mg^{2+}(g) + 2 F(g) \rightarrow MgF_2(s)$	$\Delta H = -2913 \text{ kJ}$	(LE)
	$Mg(s) + F_2(g) \rightarrow MgF_2(s)$	$\Delta H_{f}^{o} = -1085 \text{ kJ/m}$	nol

60 (64).	$Na(g) \rightarrow Na^{+}(g) + e^{-}$ F(g) + e^{-} \rightarrow F^{-}(g)	$\Delta H = IE = 495 \text{ kJ} (Table 7.5)$ $\Delta H = EA = -327.8 \text{ kJ} (Table 7.7)$		
	$Na(g) + F(g) \rightarrow Na^{+}(g) + F^{-}(g)$	$\Delta H = 167 \text{ kJ}$		

The described process is endothermic. What we haven't accounted for is the extremely favorable lattice energy. Here, the lattice energy is a large negative (exothermic) value, making the overall formation of NaF a favorable exothermic process.

61 (65). Use Figure 8.11 as a template for this problem.

$Li(s) \rightarrow Li(g)$	$\Delta H_{sub} = ?$
$Li(g) \rightarrow Li^+(g) + e^-$	$\Delta H = 520. \text{ kJ}$
$1/2 I_2(g) \rightarrow I(g)$	$\Delta H = 151/2 \text{ kJ}$
$I(g) + e^- \rightarrow I^-(g)$	$\Delta H = -295 \text{ kJ}$
$\text{Li}^+(g) + \Gamma(g) \rightarrow \text{LiI}(s)$	$\Delta H = -753 \text{ kJ}$
$\text{Li}(s) + 1/2 \text{ I}_2(g) \rightarrow \text{LiI}(s)$	$\Delta H = -292 \text{ kJ}$

 ΔH_{sub} + 520. + 151/2 - 295 - 753 = -292, ΔH_{sub} = 161 kJ

64 (68). Lattice energy is proportional to the charge of the cation times the charge of the anion Q_1Q_2 .

Compound	Q_1Q_2	Lattice Energy
FeCl ₂	(+2)(-1) = -2	-2631 kJ/mol
FeCl ₃	(+3)(-1) = -3	-5359 kJ/mol
Fe_2O_3	(+3)(-2) = -6	-14,744 kJ/mol

65 (69). a. $H - H + Cl - Cl \rightarrow 2H - Cl$

Bonds broken:

Bonds formed:

1 H–H (432 kJ/mol) 2 H–Cl (427 kJ/mol) 1 Cl–Cl (239 kJ/mol)

$$\Delta H = \Sigma D_{broken} - \Sigma D_{formed}, \ \Delta H = 432 \text{ kJ} + 239 \text{ kJ} - 2(427) \text{ kJ} = -183 \text{ kJ}$$

b. N≡N+3 H−−H → 2 H−−N−−H | H

Bonds broken:	Bonds formed:
1 N≡N (941 kJ/mol)	6 N–H (391 kJ/mol)

3 H–H (432 kJ/mol)

 $\Delta H = 941 \text{ kJ} + 3(432) \text{ kJ} - 6(391) \text{ kJ} = -109 \text{ kJ}$

66 (70). Sometimes some of the bonds remain the same between reactants and products. To save time, only break and form bonds that are involved in the reaction.



Bonds broken:

Bonds formed:

1 C=N (891 kJ/mol) 2 H-H (432 kJ/mol) 2 N-H (305 kJ/mol) 2 N-H (391 kJ/mol)

 $\Delta H = 891 \text{ kJ} + 2(432 \text{ kJ}) - [305 \text{ kJ} + 2(413 \text{ kJ}) + 2(391 \text{ kJ})] = -158 \text{ kJ}$



Bonds broken:

1 N–N (160. kJ/mol) 4 N–H (391 kJ/mol) 2 F–F (154 kJ/mol) Bonds formed:

 mol)
 4 H-F (565 kJ/mol)

 nol)
 1 N \equiv N (941 kJ/mol)

 nol)
 1 N \equiv N (941 kJ/mol)

67 (71).





Bonds broken: 1 C–N (305 kJ/mol) Bonds formed: 1 C–C (347 kJ/mol)

 $\Delta H = \Sigma D_{broken} - \Sigma D_{formed}, \ \Delta H = 305 - 347 = -42 \text{ kJ}$

Note: Sometimes some of the bonds remain the same between reactants and products. To save time, only break and form bonds that are involved in the reaction.

 $\begin{array}{cccc} 68 \ (72). & H & H & O \\ H - C - O - H & + & C \equiv O & \longrightarrow & H - C - C - O - H \\ H & H & H & H \end{array}$

Bonds broken:	Bonds formed:
1 C≡O (1072 kJ/mol)	1 C–C (347 kJ/mol)
1 C–O (358 kJ/mol)	1 C=O (745 kJ/mol)
	1 C–O (358 kJ/mol)

 $\Delta H = 1072 + 358 - [347 + 745 + 358] = -20. \text{ kJ}$



Bonds broken:	Bonds formed:
2 S-H (347 kJ/mol)	4 S-F (327 kJ/mol)
3 F-F (154 kJ/mol)	2 H-F (565 kJ/mol)

$$\Delta H = 2(347) + 3(154) - [4(327) + 2(565)] = -1282 \text{ kJ}$$



 $\Delta H = -549 \text{ kJ} = 614 \text{ kJ} + 154 \text{ kJ} - [347 \text{ kJ} + 2D_{CF}], 2D_{CF} = 970., D_{CF} = 485 \text{ kJ/mol}$

74 (78). Let x = bond energy for A_2 , so 2x = bond energy for AB.

$$\Delta H = -285 \text{ kJ} = x + 432 \text{ kJ} - [2(2x)], 3x = 717, x = 239 \text{ kJ/mol}$$

The bond energy for A_2 is 239 kJ/mol.

78 (82). $NH_3(g) \rightarrow N(g) + 3 H(g) \Delta H^\circ = 3D_{NH} = 472.7 + 3(216.0) - (-46.1) = 1166.8 kJ$

$$D_{\rm NH} = \frac{1166.8 \text{ kJ}}{3 \text{ mol NH bonds}} = 388.93 \text{ kJ/mol} \approx 389 \text{ kJ/mol}$$

 $D_{calc} = 389 \text{ kJ/mol}$ as compared with 391 kJ/mol in Table 8.4. There is good agreement.

80 (84). $1/2 N_2(g) + 1/2 O_2(g) \rightarrow NO(g) \Delta H = 90. kJ$

Bonds broken:

Bonds formed:

1 NO (D_{NO} = NO bond energy) 1/2 N≡N (941 kJ/mol) 1/2 O=O (495 kJ/mol)

$$\Delta H = 90. \text{ kJ} = 1/2(941) + 1/2(495) - (D_{NO}), D_{NO} = 628 \text{ kJ/mol}$$

From this data, the calculated NO bond energy is 628 kJ/mol.

83 (87). Sometimes part of drawing Lewis structures is trial and error. However, the first two steps are always the same. These steps are (1) count the valence electrons available in the molecule/ion, and (2) attach all atoms to each other with single bonds (called the skeletal structure). The atom listed first is usually assumed to be the atom in the middle, called the central atom, and all other atoms in the formula are attached to this atom. But for formulas that begin with H, e.g., H_2O , H_2CO , etc. hydrogen can never be a central atom since this would require H to have more than two electrons. In these compounds, the atom listed second is assumed to be the central atom. Also, if there is only one atom of an element and more than one of other(s) in the molecule, the single-atom element is generally the central atom.

After counting valence electrons and drawing the skeletal structure, the rest is trial and error. We place the remaining electrons around the various atoms in an attempt to satisfy the octet rule (or duet rule for H).

a. CCl_4 has 4 + 4(7) = 32 valence



Skeletal Lewis structure

structure

structure

structure



SeCl₂ has 6 + 2(7) = 20 valence d. ICl has 7 + 7 = 14 valence electrons.



84 (88). a. $POCl_3$ has 5 + 6 + 3(7) = 32 valence electrons.



Skeletal structure

Lewis structure

 SO_4^{2-} has 6 + 4(6) + 2 = 32 valence electrons.



Note: A negatively charged ion will have additional electrons to those that come from the valence shell of the atoms. The magnitude of the negative charge indicates the number of extra electrons to add in.

 XeO_4 , 8 + 4(6) = 32 e⁻ PO_4^{3-} , 5 + 4(6) + 3 = 32 e⁻





 ClO_4^{-} has 7 + 4(6) + 1 = 32 valence electrons



Note: All of these species have the same number of atoms and the same number of valence electrons. They also have the same Lewis structure.

b. NF₃ has 5 + 3(7) = 26 valence electrons. SO₃²⁻, $6 + 3(6) + 2 = 26 e^{-1}$



Note: Species with the same number of atoms and valence electrons have similar Lewis structures.

c. ClO_2^{-} has 7 + 2(6) + 1 = 20 valence

Skeletal structure

Lewis structure

Note: Species with the same number of atoms and valence electrons have similar Lewis structures.

d. Molecules ions that have the same number of valence electrons and the same number of atoms will have similar Lewis structures.

85 (89). BeH₂, 2 + 2(1) = 4 valence electrons

 BH_3 , 3 + 3(1) = 6 valence electrons



87 (91). PF₅, 5 +5(7) = 40 valence electrons SF₄, 6 + 4(7) = 34 e⁻



Row 3 and heavier nonmetals can have more than 8 electrons around them when they have to. Row 3 and heavier elements have empty d orbitals that are close in energy to valence s and p orbitals. These empty d orbitals can accept extra electrons.

For example, P in PF_5 has its five valence electrons in the 3s and 3p orbitals. These s and p orbitals have room for three more electrons, and if it has to, P can use the empty 3d orbitals for any electrons above 8.

90 (94). Ozone: O_3 has 3(6) = 18 valence electrons.

Sulfur dioxide: SO_2 has 6 + 2(6) = 18 valence electrons.

$$\ddot{0}$$
 $=$ \ddot{s} $\ddot{0}$ \cdot \cdot $\ddot{0}$ \ddot{s} $=$ $\ddot{0}$

Sulfur trioxide: SO_3 has 6 + 3(6) = 24 valence electrons.



94 (98). CO_3^{2-} has 4 + 3(6) + 2 = 24 valence electrons.

Three resonance structures can be drawn for $\text{CO}_3^{2^-}$. The actual structure for $\text{CO}_3^{2^-}$ is an average of these three resonance structures. That is, the three C–O bond lengths are all equivalent, with a length somewhere between a single and a double bond. The actual bond length of 136 pm is consistent with this resonance view of $\text{CO}_3^{2^-}$.

98 (102).
$$H_2NOH (14 e^{-})$$

 $H \longrightarrow H$
 $H \longrightarrow H$
Single bond between N and O
 $N_2O (16 e^{-}): \dot{N} = N = \ddot{O} \iff N \equiv N - \ddot{O}: \iff \dot{N} - N \equiv O:$
Average of a double bond between N and O
 $NO^+ (10 e^{-}): [:N \equiv O:]^+$ Triple bond between N and O
 $NO_2^- (18 e^{-}): [\ddot{O} = \ddot{N} - \ddot{O}:]^- \iff [:\ddot{O} - \ddot{N} = \ddot{O}]^-$
Average of 1 1/2 bond between N and O

$$NO_{3} (24 \text{ e}) = \begin{bmatrix} :0: \\ ... \\$$

Average of 1 1/3 bond between N and O

From the Lewis structures, the order from shortest \rightarrow longest N–O bond is:

 $NO^{+} < N_2O < NO_2^{-} < NO_3^{-} < H_2NOH$

101 (105). See Exercise 84 (88 in 10^{th} ed.) for the Lewis structures of POCl₃, SO₄²⁻, ClO₄⁻ and PO₄³⁻. All these compounds/ions have similar Lewis structures to those of SO₂Cl₂ and XeO₄ shown below. Formal charge = [number of valence electrons on free atom] – [number of lone pair electrons on atom + 1/2(number of shared electrons of atom)].

a. POCl₃: P, FC =
$$5 - 1/2(8) = +1$$
 b. SO₄²⁻: S, FC = $6 - 1/2(8) = +2$

c.
$$CIO_4^{-1}$$
: $CI, FC = 7 - 1/2(8) = +3$
e. $SO_2CI_2, 6 + 2(6) + 2(7) = 32 e^{-1}$
f. $XeO_4, 8 + 4(6) = 32 e^{-1}$
g. $CIO_3^{-1}, 7 + 3(6) + 1 = 26 e^{-1}$
h. $NO_4^{3-}, 5 + 4(6) + 3 = 32 e^{-1}$
[$: \bigcirc -N - \bigcirc :$
 $: \bigcirc : -N - \bigcirc :$
 $: \bigcirc : -N - \bigcirc :$
 $: \bigcirc : \odot : -N - \bigcirc :$
 $: \bigcirc : \odot : \end{bmatrix}^{3-1}$
CI, FC = 7 - 2 - 1/2(6) = +2
N, FC = 5 - 1/2(8) = +1

106 (110). The nitrogen-nitrogen bond length of 112 pm is between a double (120 pm) and a triple (110 pm) bond. The nitrogen-oxygen bond length of 119 pm is between a single (147 pm) and a double bond (115 pm). The third resonance structure shown below doesn't appear to be as important as the other two since there is no evidence from bond lengths for a nitrogen-oxygen triple bond or a nitrogen-nitrogen single bond as in the third resonance form. We can adequately describe the structure of N_2O using the resonance forms:

$$\ddot{N} = N = \ddot{O} \iff N \equiv N = \ddot{O}$$

Assigning formal charges for all three resonance forms:

$$N = N = 0 \quad \longleftrightarrow \quad N = N = 0 \quad \longleftrightarrow \quad N = 0$$

$$(\dot{N} =), FC = 5 - 4 - 1/2(4) = -1$$

 $(= N =), FC = 5 - 1/2(8) = +1$, Same for $(= N -)$ and $(- N =)$
 $(\ddot{N} -), FC = 5 - 6 - 1/2(2) = -2$; $(N =), FC = 5 - 2 - 1/2(6) = 0$
 $(= \dot{O}), FC = 6 - 4 - 1/2(4) = 0$; $(= \dot{O}:), FC = 6 - 6 - 1/2(2) = -1$
 $(= O:), FC = 6 - 2 - 1/2(6) = +1$

We should eliminate N-N=O because it has a formal charge of +1 on the most electronegative element (O). This is consistent with the observation that the N-N bond is between a double and triple bond and that the N-O bond is between a single and double bond.

a. PCl_3 has 5 + 3(7) = 26 valence electrons.

b. SCl_2 has 6 + 2(7) = 20valence electrons.



V-shaped; angle is <109.5°.

c. SiF₄ has 4 + 4(7) = 32 valence electrons.

Trigonal pyramid; all angles are <109.5°.



Note: In PCl₃, SCl₂, and SiF₄, there are four pairs of electrons about the central atom in each case in this exercise. All of the structures are based on a tetrahedral geometry, but only SiF_4 has a tetrahedral structure. We consider only the relative positions of the atoms when describing the molecular structure.

115 (121). a. XeCl₂ has 8 + 2(7) = 22 valence electrons.



There are five pairs of electrons about the central Xe atom. The structure will be based on a trigonal bipyramid geometry. The most stable arrangement of the atoms in $XeCl_2$ is a linear molecular structure with a 180° bond angle.

b. ICl_3 has 7 + 3(7) = 28 valence electrons.



T-shaped; the CIICl angles are $\approx 90^{\circ}$. Since the lone pairs will take up more space, the CIICl bond angles will probably be slightly less than 90° .

c. TeF₄ has 6 + 4(7) = 34 valence electrons.

d. PCl_5 has 5 + 5(7) = 40 valence electrons.



See-saw or teeter-totter or distorted tetrahedron

Trigonal bipyramid

All the species in this exercise have five pairs of electrons around the central atom. All the structures are based on a trigonal bipyramid geometry, but only in PCl_5 are all the pairs, bonding pairs. Thus PCl_5 is the only one for which we describe the molecular structure as trigonal bipyramid. Still, we had to begin with the trigonal bipyramid geometry to get to the structures (and bond angles) of the others.

- **118** (124). All have polar bonds; in SiF₄, the individual bond dipoles cancel when summed together, and in PCl₃ and SCl₂, the individual bond dipoles do not cancel. Therefore, SiF₄ has no net dipole moment (is nonpolar), and PCl₃ and SCl₂ have net dipole moments (are polar). For PCl₃, the negative end of the dipole moment is between the more electronegative chlorine atoms, and the positive end is around P. For SCl₂, the negative end is between the more electronegative Cl atoms, and the positive end of the dipole moment is around S.
- **119** (125). All have polar bonds, but only TeF_4 and ICl_3 have dipole moments. The bond dipoles from the five P–Cl bonds in PCl₅ cancel each other when summed together, so PCl₅ has no net dipole moment. The bond dipoles in XeCl₂ also cancel:

Because the bond dipoles from the two Xe–Cl bonds are equal in magnitude but point in opposite directions, they cancel each other, and $XeCl_2$ has no net dipole moment (is

nonpolar). For TeF_4 and ICl_3 , the arrangement of these molecules is such that the individual bond dipoles do *not* all cancel, so each has an overall net dipole moment (is polar).

121 (127). Molecules that have an overall dipole moment are called polar molecules, and molecules that do not have an overall dipole moment are called nonpolar molecules.



V-shaped, polar; OCl_2 is polar because the two O–Cl bond dipoles don't cancel each other. The resulting dipole moment is shown in the drawing.

 $KrF_2, 8 + 2(7) = 22 e^{-1}$:F $\stackrel{\longleftarrow}{\longrightarrow}$ Kr $\stackrel{\longleftarrow}{\longrightarrow}$ F:

Linear, nonpolar; the molecule is nonpolar because the two Kr–F bond dipoles cancel each other.

Linear, nonpolar; Be–H bond dipoles are equal and point in opposite directions. They cancel each other. BeH_2 is nonpolar. V-shaped, polar; the S–O bond dipoles do not cancel, so SO_2 is polar (has a net dipole moment). Only one resonance structure is shown.

Note: All four species contain three atoms. They have different structures because the number of lone pairs of electrons around the central atom are different in each case.

b. SO₃, $6 + 3(6) = 24 e^{-1}$



Trigonal planar, nonpolar; bond dipoles cancel. Only one resonance structure is shown.

NF₃, $5 + 3(7) = 26 e^{-7}$



Trigonal pyramid, polar; bond dipoles do not cancel.

IF₃ has 7 + 3(7) = 28 valence electrons.



T-shaped, polar; bond dipoles do not cancel.

Note: Each molecule has the same number of atoms but different structures because of differing numbers of lone pairs around each central atom.

c. CF_4 , $4 + 4(7) = 32 e^-$



Tetrahedral, nonpolar; bond dipoles cancel.



See-saw, polar; bond dipoles do not cancel.

 KrF_4 , 8 + 4(7) = 36 valence electrons



Square planar, nonpolar; bond dipoles cancel.

Note: Again, each molecule has the same number of atoms but different structures because of differing numbers of lone pairs around the central atom.

d. IF₅, $7 + 5(7) = 42 e^{-7}$



Square pyramid, polar; bond dipoles do not cancel.

AsF₅, $5 + 5(7) = 40 e^{-7}$



Trigonal bipyramid, nonpolar; bond dipoles cancel.

Note: Yet again, the molecules have the same number of atoms but different structures because of the presence of differing numbers of lone pairs.