Unit 2 Structures and Properties of Matter

Solutions to Practice Problems in
Chapter 3 Atomic Models and Properties of Atoms

Determining Quantum Numbers
(Student textbook page 179)

1. For the quantum number $n = 3$, what values of l are allowed, what values of m_l are possible, and how many orbitals are there?

What Is Required?
You need to determine the allowed values of l and m_l when the value of n is given.

What Is Given?
You are given the value of the principle quantum number, $n = 3$.

Plan Your Strategy	Act on Your Strategy
Use the rules that give the relationships among the quantum numbers: Allowed values for l are 0 to $(n - 1)$.	When $n = 3$, l can be 0, 1, or 2.
Allowed values for m_l are all integers from $-l$ to $+l$.	When $l = 0$, $m_l = 0$ (one orbital). When $l = 1$, m_l can be -1, 0, or +1. When $l = 2$, m_l can be -2, -1, 0, 1, or 2.
To find the number of orbitals, count the total number of possible values for combinations of l and m_l.	There are 9 orbitals.

Check Your Solution
The total number of orbitals for any given value for n is n^2.
Thus, when $n = 3$, the number of orbitals is $3^2 = 9$.

2. If $n = 5$ and $l = 2$, what orbital type is this, what are the possible values for m_l, and how many orbitals are there?

What Is Required?
You need to determine the type of orbital (*s*, *p*, *d*, *f*) when n and l are given.
You need to determine the possible values of m_l when n and l are given.
You need to determine the number of orbitals.

What Is Given?
You are given the values for the principle quantum number and the orbital shape quantum numbers: $n = 5$ and $l = 2$

Plan Your Strategy	Act on Your Strategy
The value of l determines the shape. When $l = 0$, the orbital shape is *s*. When $l = 1$, the orbital shape is *p*. When $l = 2$, the orbital shape is *d*. When $l = 3$, the orbital shape is *f*. The value of n is written before the shape symbol.	When $n = 5$ and $l = 2$, the orbital is a 5*d* orbital.
The value for l also determines the allowed values for m_l, according to the rule that m_l can be any integer from $-l$ to $+l$.	Since $l = 2$, m_l can be -2, -1, 0, +1, or +2.
The number of orbitals is the number of allowed values for m_l. Count the number of values of m_l.	There are five allowed m_l values, so there are five orbitals.

Check Your Solution
For any given value of l there are $2l + 1$ allowed values of m_l and thus $2l + 1$ orbitals.
When $l = 2$, the number of allowed m_l values $= 2l + 1 = 2(2) + 1 = 5$. This is in agreement with the answers determined.

3. What are the n, l, and possible m_l values for the following orbital types?

a. 2*s*
b. 3*p*
c. 5*d*
d. 4*f*

What Is Required?
You need to determine the values of n and l and all of the possible values for m_l for the given orbital types.

What Is Given?
The orbital types 2*s*, 3*p*, 5*d*, and 4*f* are given.

Plan Your Strategy	Act on Your Strategy
The number is the n value and the letter is the symbol for the l value.	**a.** 2*s*: $n = 2$ **b.** 3*p*: $n = 3$ **c.** 5*d*: $n = 5$ **d.** 4*f*: $n = 4$
Use the rules for writing orbital types. The $l = 0$ orbital has the symbol *s*. The $l = 1$ orbital has the symbol *p*. The $l = 2$ orbital has the symbol *d*. The $l = 3$ orbital has the symbol *f*.	**a.** 2*s*: $l = 0$ **b.** 3*p*: $l = 1$ **c.** 5*d*: $l = 2$ **d.** 4*f*: $l = 3$
Use the relationship between l and m_l to determine the possible values for m_l. Allowed m_l values are all integers from $-l$ to $+l$.	**a.** 2*s*: $m_l = 0$ **b.** 3*p*: $m_l = -1, 0, +1$ **c.** 5*d*: $m_l = -2, -1, 0, +1, +2$ **d.** 4*f*: $m_l = -3, -2, -1, 0, +1, +2, +3$

Check Your Solution
The results are in agreement with all of the rules.

4. What orbital type can be described by the following sets of quantum numbers?
a. $n = 2$, $l = 0$, $m_l = 0$
b. $n = 5$, $l = 3$, $m_l = -2$

What Is Required?
You need to determine all possible orbital types for the given principal, orbital-shape, and magnetic quantum numbers.

What Is Given?
You are given the combinations of quantum numbers:
a. $n = 2$, $l = 0$, $m_l = 0$
b. $n = 5$, $l = 3$, $m_l = -2$

Plan Your Strategy	Act on Your Strategy
Write the value of n and then write the symbol for the value of l. The values of m_l are not written into the symbols for orbital types because they only affect the orientation in space, not the shape of the orbitals.	**a.** 2*s* **b.** 5*f* The values of m_l are not written into the symbols for orbital types because they only affect the orientation in space, not the shape of the orbitals.

Check Your Solution
The symbols for the orbital types agree with the given values of the quantum numbers.

5. How many orbitals are associated with each of the following types?

a. 1*s*
b. 5*f*
c. 4*f*
d. 2*p*

What Is Required?

You need to determine the number of orbitals associated with the given orbital types.

What Is Given?

You are given the orbital types:
a. 1*s*
b. 5*f*
c. 4*f*
d. 2*p*

Plan Your Strategy	Act on Your Strategy
Determine the value for n. The first symbol is the n value.	**a.** 1*s*: $n = 1$ **b.** 5*f*: $n = 5$ **c.** 4*f*: $n = 4$ **d.** 2*p*: $n = 2$
Determine the value for l. Allowed l values are 0 to (n - 1).	**a.** 1*s*: $n = 1$; $l = 0$ **b.** 5*f*: $n = 5$; $l = 3$ **c.** 4*f*: $n = 4$; $l = 3$ **d.** 2*p*: $n = 2$; $l = 1$
Determine the value for m_l. Allowed m_l values are all integers from $-l$ to $+l$.	**a.**1*s*: $n = 1$; $l = 0$; $m_l = 0$ **b.** 5*f*: $n = 5$; $l = 3$; m_l = -3, -2, -1, 0, +1, +2, +3 **c.** 4*f*: $n = 4$; $l = 3$; m_l = -3, -2, -1, 0, +1, +2, +3 **d.** 2*p*: $n = 2$; $l = 1$; m_l = -1, 0, +1
Count the number values for m_l to determine the total number of orbitals.	**a.** 1 possible orbital **b.** 7 possible orbitals **c.**7 possible orbitals **d.** 3 possible orbitals

Check Your Solution

The number of orbitals for a given l value is $2l + 1$.
a. $l = 0$: $2(0) + 1 = 1$
b. $l = 3$: $2(3) + 1 = 7$
c. $l = 3$: $2(3) + 1 = 7$
d. $l = 1$: $2(1) + 1 = 3$
These values are all in agreement with the answers.

6. What sets of quantum numbers are possible for a 4*d* orbital? List them.

What Is Required?
You need to list all of the possible sets of quantum numbers for a given orbital type.

What Is Given?
The given orbital type is 4*d*.

Plan Your Strategy	Act on Your Strategy
Assign values for n and l based on the orbital type symbol 4*d*.	For orbital type 4*d*, $n = 4$ and $l = 2$.
List the possible values for m_l based on the value for l, given that allowed m_l values are all integers from $-l$ to $+l$.	$n = 4$, $l = 2$, and $m_l = -2$ $n = 4$, $l = 2$, and $m_l = -1$ $n = 4$, $l = 2$, and $m_l = 0$ $n = 4$, $l = 2$, and $m_l = +1$ $n = 4$, $l = 2$, and $m_l = +2$

Check Your Solution
The values of n and l agree with the orbital type 4*d*. For any l value, the number of possible orbitals is $2l + 1$. Thus, for $l = 2$, there would be $2(2) + 1 = 5$ orbital types. This agrees with the number of orbital types.

7. What is one possible value for the missing number in each of the following sets?
a. $n = 3$, $l = 1$, $m_l = ?$
b. $n = 2$, $l = ?$, $m_l = -3$

What Is Required?
You need to fill numbers into the two sets of quantum numbers that will create allowed sets of quantum numbers.

What Is Given?
You are given the partial sets of quantum numbers:
a. $n = 3$, $l = 1$, $m_l = ?$
b. $n = 2$, $l = ?$, $m_l = -3$

Plan Your Strategy	Act on Your Strategy
a. Determine the allowed values for m_l when $l = 1$ and choose one.	**a.** When $l = 1$, m_l can be -1, 0, or +1. Choose -1. The set of quantum numbers is: $n = 3$, $l = 1$, $m_l = -1$ Sample: $n = 3$, $l = 1$, $m_l = -1$
b. Determine the allowed values for l when $n = 2$ and $m_l = -3$ and choose one.	**b.** When $n = 2$, l must be 0 or 1. When $m_l = -3$, l must be 3 or greater. Therefore, there is no allowed number for l, given that $n = 2$ and $m_l = -3$. The values $n = 2$ and $m_l = -3$ are not compatable. Therefore, there is no solution.

Check Your Solution
a. The set of quantum numbers are allowed because when $n = 3$, l can be 0, 1, or 2.So, $l = 1$ is allowed. The allowed values for m_l are $-l$ to $+l$. So, the value of $m_l = -1$ is allowed.
b. When $n = 2$, l can be 0 or 1 because the allowed values for l are 0 to $(n - 1)$. Allowed m_l values are all integers from -l to +l, which means that for $l = 0$, m_l must be 0. For $l = 1$, m_l must be -1, 0, or +1. Therefore, m_l cannot be -3 making $n = 2$ incompatible with $m_l = -3$.

8. Write two possible sets of quantum numbers for a 6*p* orbital.

What Is Required?
You need to determine two sets of quantum numbers that correspond to the given orbital type.

What Is Given?
You are given the orbital type 6*p*.

Plan Your Strategy	Act on Your Strategy
Determine possible combinations of n and l that correspond to the orbital type 6*p*.	For orbital type 6*d*, $n = 6$ and $l = 1$.
Determine the possible values for m_l that will correspond to the values for n and l that you determine.	The value of m_l can be any integer from -3 to +3. The possible sets of quantum numbers are: $n = 6$, $l = 1$, $m_l = -1$; $n = 6$, $l = 1$, $m_l = 0$; $n = 6$, $l = 1$, $m_l = +1$;
Write at least two sets of these quantum numbers.	Sample answer: $n = 6$, $l = 1$, $m_l = -1$; $n = 6$, $l = 1$, $m_l = 0$; $n = 6$, $l = 1$, $m_l = +1$

Check Your Solution
All of the possible sets of answers are in agreement with the rules for determining allowed values for m_l given l.

9. The following sets of quantum numbers are not allowed. Identify the problem and change one number to give an allowed set.
a. $n = 1$, $l = 2$, $m_l = -2$
b. $n = 4$, $l = 1$, $m_l = -2$

What Is Required?
For each set of quantum numbers, determine which values do not obey the rules. Change one of the numbers in each set so that the set of numbers will be allowed.

What Is Given?
You are given the sets of quantum numbers:
a. $n = 1$, $l = 2$, $m_l = -2$
b. $n = 4$, $l = 1$, $m_l = -2$

a. $n = 1$, $l = 2$, $m_l = -2$

Plan Your Strategy	Act on Your Strategy
Analyze the given sets of quantum numbers in light of the rules for allowed values of l given n, and allowed values for m_l given l. Determine the values that do not conform to the rules.	Given that $n = 1$, the value of l must be 0 since it has to be a positive integer less than 1. However, if $l = 0$, m_l cannot be -2 since m_l has to be between $-l$ and $+l$. So, if $l = 0$, m_l must be 0.
Change one value to create an allowed set of quantum numbers.	The only way to change one value is by setting $n = 3$ or higher. Sample answer: $n = 3$, $l = 2$, $m_l = -2$

b. $n = 4$, $l = 1$, $m_l = -2$

Plan Your Strategy	Act on Your Strategy
Analyze the given sets of quantum numbers in light of the rules for allowed values for l given n and allowed values for m_l given l. Determine the values that do not conform to the rules.	The values of $n = 4$ and $l = 1$ are allowed but m_l cannot be -2 since the magnitude of m_l cannot be greater than l.
Change one value to create an allowed set of quantum numbers.	You could change m_l to -1, 0, or +1 to create an allowed set of quantum numbers. Or, you could change l to 2 or 3, each of which is compatible with $n = 4$, and you would not have to change m_l. Sample answers: $n = 4$, $l = 2$, $m_l = -2$; $n = 4$, $l = 3$, $m_l = -2$; $n = 4$, $l = 1$, $m_l = -1$; $n = 4$, $l = 1$, $m_l = 0$ $n = 4$, $l = 1$, $m_l = +1$

Check Your Solution
All of the examples are in agreement with the rules that describe allowed sets of quantum numbers.

10. Label each of the following sets of quantum numbers as *allowed* or *not allowed*. Identify the problem for each of the *not allowed* sets.

a. $n = 3$, $l = 2$, $m_l = 0$
b. $n = 1$, $l = 1$, $m_l = -1$
c. $n = 0$, $l = 0$, $m_l = 0$
d. $n = 5$, $l = 1$, $m_l = 3$

What Is Required?
Determine whether each of four sets of quantum numbers is allowed or not allowed.

What Is Given?
You are given the four sets of quantum numbers:
a. $n = 3$, $l = 2$, $m_l = 0$
b. $n = 1$, $l = 1$, $m_l = -1$
c. $n = 0$, $l = 0$, $m_l = 0$
d. $n = 5$, $l = 1$, $m_l = 3$

a. $n = 3$, $l = 2$, $m_l = 0$

Plan Your Strategy	Act on Your Strategy
Analyze the set of quantum numbers against the rules for allowed quantum numbers.	allowed
If the set is not allowed, explain why it does not obey the rules.	

b. $n = 1$, $l = 1$, $m_l = -1$

Plan Your Strategy	Act on Your Strategy
Analyze the set of quantum numbers against the rules for allowed quantum numbers.	not allowed
If the set is not allowed, explain why it does not obey the rules.	The value of l must be a positive integer that is less than n. Therefore, when $n = 1$, l cannot be 1. In addition, if $l = 0$, then m_l must also be zero.

c. $n = 0$, $l = 0$, $m_l = 0$

Plan Your Strategy	Act on Your Strategy
Analyze the set of quantum numbers against the rules for allowed quantum numbers.	not allowed
If the set is not allowed, explain why it does not obey the rules.	The principle quantum number must be a positive integer greater than 0. Thus, n cannot be 0.

d. $n = 5$, $l = 1$, $m_l = 3$

Plan Your Strategy	Act on Your Strategy
Analyze the set of quantum numbers against the rules for allowed quantum numbers.	not allowed
If the set is not allowed, explain why it does not obey the rules.	The value for m_l must be between $-l$ and $+l$. So, when $l = 1$, m_l cannot be 3. It must be -1, 0, or +1.

Check Your Solution

The answers are all in agreement with the rules.

Inferring the Characteristics of an Element (Student textbook pages 188-9)

11. Write complete and condensed electron configurations for yttrium, Y.

What Is Required?
You need to determine the electron configuration for the given element and then write it in complete and in condensed form.

What Is Given?
You are given the element yttrium, Y.

Plan Your Strategy	Act on Your Strategy
Determine the atomic number of yttrium which will give you the number of electrons in the neutral atom.	The atomic number of yttrium is 39. Therefore, the neutral atom has 39 electrons.
Follow the guidelines for filling orbitals that are given on page 184 of the student textbook.	Use the order of orbitals shown on page 182 to determine the order of filling energy levels $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < \ldots$
For each occupied orbital, write the electron configuration in the form $n(\text{symbol for } l)^{\text{number of electrons in orbital}}$. The complete electron configuration includes symbols for all occupied orbitals. Write the symbols, keeping track of the number of electrons. Stop at 39.	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^1$
The condensed electron configuration includes the symbol for the nearest, smaller noble gas in square brackets and then the symbols for occupied orbitals that contain valence electrons. The first eight orbital notations are identical to the noble gas, krypton, Kr.	$[Kr]5s^2 4d^1$.

Check Your Solution
If you add the values in the exponents of the complete electron configuration you get 39, which is correct for yttrium. The first eight symbols all have filled orbitals, which is consistent with noble gases.

12. Write complete and condensed electron configurations for lead, Pb.

What Is Required?
You need to determine the electron configuration for the given element and then write it in complete and condensed forms.

What Is Given?
You are given the element lead, Pb.

Plan Your Strategy	Act on Your Strategy
Determine the atomic number of lead which will give you the number of electrons in the neutral atom.	The atomic number of lead is 82 therefore the neutral atom has 82 electrons.
Follow the guidelines for filling orbitals which are given on page 184 of the student textbook.	($1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < ...$)
For each occupied orbital, write the electron configuration in the form $n(\text{symbol for } l)^{\text{number of electrons in orbital}}$ The complete electron configuration includes symbols for all occupied orbitals. Write the symbols, keeping track of the number of electrons. Stop at 82.	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$
The condensed electron configuration includes the symbol for the nearest, smaller noble gas in square brackets and then the symbols for occupied orbitals that contain valence electrons. The first eleven orbital notations are identical to the noble gas, xenon, Xe.	[Xe] $6s^2 4f^{14} 5d^{10} 6p^2$

Check Your Solution
If you add the values in the exponents of the complete electron configuration you get 82, which is correct for lead. The first eleven symbols all have filled orbitals, which is consistent with noble gases.

13. What elements have the valence electron configuration that is given by ns^2?

What Is Required?
Determine which elements have a given electron configuration.

What Is Given?
You are given the electron configuration ns^2.

Plan Your Strategy	Act on Your Strategy
Analyze the electron configuration to determine which features are similar and which can vary.	The principle quantum number *n* can be any value. The highest orbital within an energy level is the *s* orbital. Therefore, the elements must be in the s block of the periodic table. The *s* orbital has 2 electrons, which fills the orbital. The elements must be in the second column of the *s* block elements.
Use this information to narrow the range of eligible elements.	All of the elements in Group 2 of the periodic table

Check Your Solution
All elements in Group 2 have two valence electrons. Therefore, they all include s^2 in their electron configuration.

14. What elements have the valence electron configuration that is given by $ns^2(n - 1)d^3$?

What Is Required?
Determine which elements have a given electron configuration.

What Is Given?
You are given the electron configuration $ns^2(n - 1)d^3$.

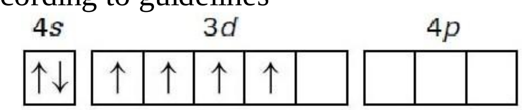
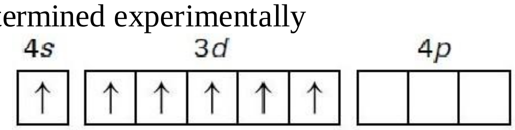
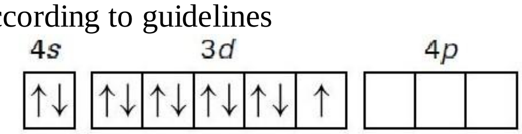
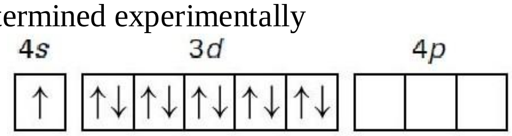
Plan Your Strategy	Act on Your Strategy
Analyze the electron configuration to determine which features are similar and which can vary.	The principle quantum number *n* can have any value. The elements must have valence electrons with a filled *s* orbital. There must be three electrons is the *d* orbital that has a principle quantum number that is one less than the principle quantum number of the valence electrons in the *s* orbital.
Use this information to narrow the range of eligible elements.	The elements will be in the *d* block. Elements in only Period 4 and greater have electrons in *d* orbitals. Therefore, *n* must have a value of 4 or higher. In Periods 4 and above, the *d* orbitals, with a principle quantum number that is one less than that of the *s* orbital, begin to fill after the *s* orbital is filled. Therefore, the elements must have 5 electrons (2 in the *s* orbital and 3 in the *d* orbital). Elements in Group 5 have five valence electrons. Therefore, the elements that fit the electron configuration, $ns^2(n - 1)d^3$, are in Group 5 of Periods 4 and above. These elements are vanadium, V; niobium, Nb; tantalum, Ta; and dubnium, Db.

Check Your Solution
All elements in the *d* block of Group 5 have five valence electrons. Therefore, they all include d^3 in their electron configuration.

15. What are the two exceptions to the guidelines for filling orbitals in Period 4? Draw the partial orbital diagrams you would expect for them, based on the aufbau principle. Then draw partial orbital diagrams that represent their actual electron configurations. Finally, explain why the discrepancy arises.

What Is Required?
You need to find elements which are exceptions to the guidelines for filling orbitals and explain why they do not fill orbitals according to the guidelines. Draw orbital diagrams of the theoretical configurations and the correct configurations.

What Is Given?
You are given Period 4 and told that there are two exceptions to the guidelines.

Plan Your Strategy	Act on Your Strategy
These data are experimental so you must find a source of electron configurations and determine those that have unfilled orbitals that theoretically are at lower energy levels than others that have begun to fill.	Find a periodic table that contains electron configurations or inspect Table 3.5 on page 186 of the student text to find the exceptions to the guidelines. You will find that the elements chromium, Cr, and Cu, copper, have 4*s* orbitals with only one electron even though there are electrons in the 3*d* orbitals that should be at a higher energy.
To draw orbital diagrams, draw boxes for 4*s*, 3*d*, and 4*p* orbitals and add the correct number of electrons, first according to the guidelines and then the correct diagrams.	Cr according to guidelines  Cr determined experimentally  Cu according to guidelines  Cu determined experimentally 

Plan Your Strategy	Act on Your Strategy
To explain why these configurations exist, you must determine a reason why these exceptions might have overall lower total stability than those for which the orbitals fill according to the general sequence for energy of orbitals.	Comparison of the theoretical and experimental orbital diagrams for chromium shows that the 3*d* orbitals are exactly half filled, whereas in the theoretical diagram, there is no such pattern. Therefore, when all of the 3*d* orbitals are half filled, the configuration is more stable than the theoretical configuration. Comparison of the theoretical and experimental orbital diagrams for copper shows that the 3*d* orbitals in the experimental diagram are completely filled, whereas, in the theoretical diagram they are not. Therefore, completely filling the 3*d* orbitals must confer stability on the electron configuration more than does filling the *s* orbital. In summary, when the 3*d* orbitals are either all half-filled or all completely filled, the configuration becomes stable, even if the *s* orbital is not filled.

Check Your Solution

Cr and Cu are the only elements in Period 4 of the *d* block elements that have one electron in the *s* orbital and at least one or more electrons in the *d* orbital of the valence electrons.

16. The condensed electron configuration for strontium is [Kr]$5s^2$. Without using a periodic table, identify the group number to which strontium belongs. Show your reasoning.

What Is Required?
You must determine the group number of an element from its electron configuration without consulting a periodic table.

What Is Given?
You are given the electron configuration for strontium, [Kr]$5s^2$.

Plan Your Strategy	Act on Your Strategy
For all *s*, *p*, and *d* block elements, the number of valence electrons corresponds to the group number. The exponents in the orbital symbols show the number of electrons that occupy the orbital.	Since there are two valence electrons in strontium, it must be in Group 2.

Check Your Solution
A periodic table confirms that strontium is in Group 2.

17. Identify the following elements and write condensed electron configurations for atoms of each element.

a. The *d*-block element in Period 4 with 10 valence electrons.
b. The element in Period 6 with 3 valence electrons.

What Is Required?
You must identify elements given information about the period and the number of valence electrons. Once the elements are identified, you must write the condensed electron configuration.

What Is Given?
You are given the data:
a. The *d*-block element in Period 4 with 10 valence electrons.
b. The element in Period 6 with 3 valence electrons.

a. The *d*-block element in Period 4 with 10 valence electrons.

Plan Your Strategy	Act on Your Strategy
For all elements other than *f*-block elements, the number of valence electrons corresponds to the group number. The principle quantum number of the *s* orbital corresponds to the period number.	Since there are no *f*-block elements below Period 6, you can use the number of valence electrons to determine the group number.
Determine the block and period numbers and, using a periodic table, find the element.	The element is in Group 10. When you look for the element in Group 10 of Period 4, you will find that it is nickel.
When you know the period, find the noble gas in the previous period to start the condensed electron configuration.	The noble gas that starts the condensed electron configuration is at the end of group 3: argon, Ar.
Complete the electron configuration by filling orbitals with the valence electrons according to the guidelines.	Two of the ten valence electrons fill the 4*s* orbital leaving 8 for the *d* orbitals. Therefore, the condensed electron configuration for nickel is $[Ar]4s^23d^8$.

b. The element in Period 6 with 3 valence electrons.

Plan Your Strategy	Act on Your Strategy
For all elements other than *f*-block elements, the number of valence electrons corresponds to the group number. The principle quantum number of the *s* orbital corresponds to the period number.	Since the element is in Period 6, it can have *f*-block electrons. The element has 3 valence electrons, indicating that the 6*s* orbital is filled.

Plan Your Strategy	Act on Your Strategy
Determine the block and period numbers and, using a periodic table, find the element.	Analysis of the blocks in the skeleton periodic table in Figure 3.20 on page 187 of the student textbook shows that the third electron in Periods 6 and 7 enter the (n - 1)d orbital, before starting to fill the f orbitals.
When you know the period, find the noble gas in the previous period to start the condensed electron configuration.	When you look for the element in Period 6 that has three valence electrons, you will find that it is lanthanum, La. The noble gas at the end of Period 5 is xenon, Xe.
Complete the electron configuration by filling orbitals with the valence electrons according to the guidelines.	The condensed electron configuration for lanthanum is [Xe]$6s^2 5d^1$.

Check Your Solution

Consulting a periodic table that includes electron configurations will confirm the electron configurations.

18. The condensed electron configuration for titanium is $[Ar]4s^23d^2$. Without a periodic table, identify the period number to which titanium belongs. Show your reasoning.

What Is Required?
You must identify the period number of an element, given its condensed electron configuration.

What Is Given?
You are given the condensed electron configuration of titanium which is $[Ar]4s^23d^2$.

Plan Your Strategy	Act on Your Strategy
The principle quantum of the *s* electrons in the condensed electron configuration is the same as the period number. The principle quantum number of the *s* electrons in the condensed electron configuration of titanium is 4.	Titanium is a Period 4 element.

Check Your Solution
The noble gas at the beginning of the condensed electron configuration is argon which is at the end of Period 3. This is in agreement with the idea that titanium is in Period 4.

19. The condensed electron configuration for arsenic is $[Ar]4s^23d^{10}4p^3$. Without using a periodic table, identify the orbital block in the periodic table to which arsenic belongs. Show your reasoning.

What Is Required?
Given a condensed electron configuration and without using a periodic table, you must identify the orbital block to which the element belongs.

What Is Given?
You are given the condensed electron configuration of arsenic, $[Ar]4s^23d^{10}4p^3$.

Plan Your Strategy	Act on Your Strategy
For elements other than the exceptions to the rules for filling orbitals, the only partially filled orbital of an element identifies the block to which it belongs.	The *p* orbital is the highest energy orbital that is only partially filled. Therefore, arsenic is a *p*-block element.

Check Your Solution
Consulting a periodic table that indicates blocks confirms that arsenic is a *p*-block element.

20. Without a periodic table, and based on the condensed electron configuration given below, identify the group number, period number, and orbital block to which each of the following elements belongs. Show your reasoning.
a. francium, $[Rn]7s^1$
b. tungsten, $[Xe]6s^2 4f^{14} 5d^4$
c. antimony, $[Kr]5s^2 4d^{10} 5p^3$

What Is Required?
Without using a periodic table, you must identify the group number, period number, and the orbital block to which three elements belong.

What Is Given?
You are given the elements and electron configurations:
a. francium, $[Rn]7s^1$
b. tungsten, $[Xe]6s^2 4f^{14} 5d^4$
c. antimony, $[Kr]5s^2 4d^{10} 5p^3$

a. francium, $[Rn]7s^1$

Plan Your Strategy	Act on Your Strategy
The quantum number in front of the *s* orbital of the valence electrons indicates the period number.	Since the number 7 is in front of the *s* orbital symbol, francium is in Period 7.
The total number of electrons in the *s*, *p*, and *d* orbitals of the valence electrons corresponds to the group number.	There is a total of one valence electron, so francium is in Group 1.
For elements other than the exceptions to the rules for filling orbitals, the only partially filled orbital of an element identifies the block to which it belongs.	The only partially filled orbital is an *s* orbital. Therefore, francium is an *s*-block element.

b. tungsten, $[Xe]6s^2 4f^{14} 5d^4$

Plan Your Strategy	[**Act on Your Strategy**
The quantum number in front of the *s* orbital of the valence electrons indicates the period number.	Since the number 6 is in front of the *s* orbital symbol for the valence electrons, tungsten is in Period 6.
The total number of electrons in the *s*, *p*, and *d* orbitals of the valence electrons corresponds to the group number.	There is a total of six valence electrons ($2s + 4p$) in orbitals other than *f* orbitals. So, tungsten is in Group 6.
For elements other than the exceptions to the rules for filling orbitals, the only partially filled orbital of an element identifies the block to which it belongs.	Since the only partially filled orbital in the valence shell is a *d* orbital, francium is a *d*-block element.

c. antimony, $[Kr]5s^2 4d^{10} 5p^3$

Plan Your Strategy	Act on Your Strategy
The quantum number in front of the *s* orbital of the valence electrons indicates the period number.	Since the number 5 is in front of the *s* orbital symbol for the valence electrons, antimony is in Period 5.
The total number of electrons in the *s*, *p*, and *d* orbitals of the valence electrons corresponds to the group number.	There is a total of 15 valence electrons ($2s + 10d + 3p$), none of which are *f* electrons. So, antimony is in Group 15.
For elements other than the exceptions to the rules for filling orbitals, the only partially filled orbital of an element identifies the block to which it belongs.	Since the only partially filled orbital valence shell is a *p* orbital, antimony is a *p*-block element.

Check Your Solution

Consulting a periodic table that includes orbital blocks confirms the answers.

Solutions to Practice Problems in Chapter 4 Chemical Bonding and Properties of Matter

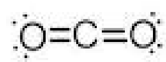
Drawing Lewis Structures of Molecules
(Student textbook page 232)

1. Draw the Lewis structure for a molecule of carbon dioxide, CO_2(g).

What Is Required?
You need to draw the Lewis structure for carbon dioxide, CO_2 (g).

What Is Given?
The chemical formula tells you that the molecule consists of one carbon atom and two oxygen atoms.

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is carbon. Therefore, it is the central atom.
Draw a skeleton structure of the carbon atom with one single bond between it and each of the two oxygen atoms.	O-C-O
Determine the total number of valence electrons in all the atoms of the molecule. The carbon atom has four valence electrons, and the oxygen atoms each have six valence electrons.	$V = \left(1\ \cancel{\text{C atom}} \times \dfrac{4e^-}{\cancel{\text{C atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \dfrac{6e^-}{\cancel{\text{O atom}}}\right)$ $= 4e^- + 12e^-$ $= 16e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{C atom}} \times \dfrac{8e^-}{\cancel{\text{C atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \dfrac{8e^-}{\cancel{\text{O atoms}}}\right)$ $= 8e^- + 16e^-$ $= 24e^-$

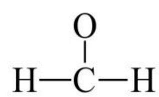
Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 24e^- - 16e^-$ $= 8e^-$ $\text{bonds} = \frac{S}{2} = \frac{8}{2} = 4 \text{ covalent bonds}$
Draw the structure with any necessary double or triple bonds. To create four covalent bonds, put a double bond between the C atom and each of the O atoms.	O = C = O
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 16e^- - 8e^-$ $= 8e^-$
Complete the Lewis structure. Because there are 8 non-bonding electrons, there are 4 lone pairs to place. Put 2 lone pairs on each O atom.	

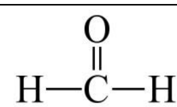
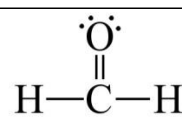
Check Your Solution
Each atom has eight valence electrons which achieves a noble gas configuration. All of the electrons have been accounted for.

2. Draw the Lewis structure for a molecule of formaldehyde (methanal), $CH_2O(g)$.

What Is Required?
You need to draw the Lewis structure for formaldehyde (methanal), $CH_2O(g)$.

What Is Given?
The chemical formula tells you that there is one carbon atom, two hydrogen atoms, and one oxygen atom.

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is carbon therefore it is the central atom.
Draw a skeleton structure of the carbon atom with one single bond between it and each of the other atoms.	
Determine the total number of valence electrons in all the atoms of the molecule. The carbon atom has four valence electrons, the hydrogen atoms each have one valence electron, and the oxygen atom has six valence electrons.	$V = \left(1\ \cancel{\text{C atom}} \times \frac{4e^-}{\cancel{\text{C atom}}}\right) + \left(1\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right) + \left(2\ \cancel{\text{H atoms}} \times \frac{1e^-}{\cancel{\text{H atom}}}\right)$ $= 4e^- + 6e^- + 2e^-$ $= 12e^-$

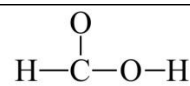
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{C atom}} \times \frac{8e^-}{\cancel{\text{C atom}}}\right) + \left(1\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atom}}}\right) + \left(2\ \cancel{\text{H atoms}} \times \frac{2e^-}{\cancel{\text{H atom}}}\right)$ $= 8e^- + 8e^- + 4e^-$ $= 20e^-$
Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 20e^- - 12e^-$ $= 8e^-$ $\text{bonds} = \frac{S}{2} = \frac{8}{2} = 4$ covalent bonds
Draw the structure with any necessary double or triple bonds. Hydrogen cannot form double bonds because the total number of electrons in its valence shell must be two. Therefore, put a double bond between the C atom and the O atom.	
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 12e^- - 8e^-$ $= 4e^-$
Complete the Lewis structure. Because there are 4 non-bonding electrons, there are 2 lone pairs to place. Put both lone pairs on the O atom.	

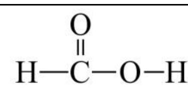
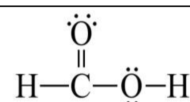
Check Your Solution
Each atom has a filled outer shell of electrons which achieves a noble gas configuration. The hydrogen atoms have only two electrons but that is the noble gas configuration of helium. All of the electrons have been accounted for.

3. Formic acid, or methanoic acid, HCOOH(ℓ), is found naturally in ant venom. Synthesized formic acid is used for its antibacterial and preservative properties in a variety of animal feeds, among other uses. Given that carbon is the central atom of the molecule, with hydrogen, oxygen, and hydroxyl groups, draw its Lewis structure.

What Is Required?
You need to draw the Lewis structure for a molecule of methanoic acid, HCOOH(ℓ).

What Is Given?
The chemical formula tells you that the molecule consists of one carbon atom, two hydrogen atoms, and two oxygen atoms.
You are also told that the carbon atom is the central atom and is bonded to a hydrogen atom, an oxygen atom, and a hydroxyl group.

Plan Your Strategy	Act on Your Strategy
Draw a skeleton structure of the carbon atom with one single bond between it and each of the other atoms or groups as described.	
Determine the total number of valence electrons in all the atoms of the molecule. The carbon atom has four valence electrons, the hydrogen atoms each have one valence electron, and the oxygen atoms have six valence electrons each.	$V = \left(1\ \cancel{\text{C atom}} \times \frac{4e^-}{\cancel{\text{C atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right) + \left(2\ \cancel{\text{H atoms}} \times \frac{1e^-}{\cancel{\text{H atom}}}\right)$ $= 4e^- + 12e^- + 2e^-$ $= 18e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{C atom}} \times \frac{8e^-}{\cancel{\text{C atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atoms}}}\right) + \left(2\ \cancel{\text{H atoms}} \times \frac{2e^-}{\cancel{\text{H atoms}}}\right)$ $= 8e^- + 16e^- + 4e^-$ $= 28e^-$

Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 28e^- - 18e^-$ $= 10e^-$ $\text{bonds} = \frac{S}{2} = \frac{10}{2} = 5$ covalent bonds
Draw the structure with any necessary double or triple bonds. H atoms cannot form double bonds. To create five covalent bonds, put a double bond between the C atom and the O atom that is not part of the hydroxyl group. Put a single bond between the C atom and one H atom, and between the O atom and the H atom of the hydroxyl.	
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 18e^- - 10e^-$ $= 8e^-$
Complete the Lewis structure. Because there are 8 non-bonding electrons, there are 4 lone pairs to place. Put two lone pairs on each O atom.	

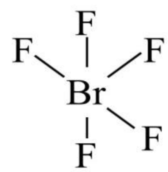
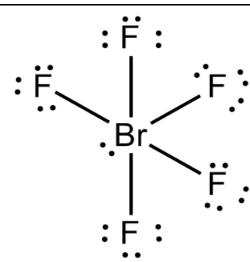
Check Your Solution

Each atom has a filled outer shell of electrons which achieves a noble gas configuration. The hydrogen atoms have only two electrons, but that is the noble gas configuration of helium. All of the electrons have been accounted for.

4. Draw a Lewis structure for a molecule of bromine pentafluoride, $BrF_5(\ell)$. The central bromine molecule has an expanded valence.

What Is Required?
You need to draw a Lewis structure for bromine pentafluoride.

What Is Given?
The chemical formula tells you that the molecule consists of one bromine atom and five fluorine atoms. You are told that the bromine is the central atom and that it has an expanded valence.

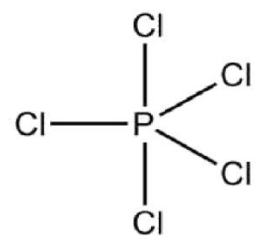
Plan Your Strategy	Act on Your Strategy
Draw a skeleton structure given that the bromine is the central atom.	
Analyze the atoms to see if there is any obvious reason whether single or double bonds should form. Draw the bonds accordingly.	Fluorine has seven valence electrons and therefore forms single bonds by sharing its one unpaired electron with the atom to which it is bonded. Assume that all of the bonds are single bonds. This leaves each fluorine atom with three lone pairs. There is no change in the skeleton structure.
The problem statement did not include anything about co-ordinate covalent bonds so assume there are none, meaning that each atom contributes one electron to each bond. Determine the number of bromine's valence electrons that were used for bonds and how many were not used. Consider the unused valence electrons as forming lone pairs.	Bromine has seven valence electrons and five are used to form single bonds with the five fluorine atoms. The remaining two electrons, therefore, form a lone pair.
Complete the Lewis structure. Add lone pairs to give all of the non-central atoms a complete octet and to give bromine the correct number of valence electrons.	

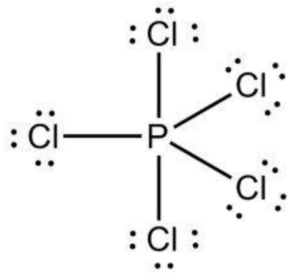
Check Your Solution
Each fluorine atom has an octet of electrons. The bromine atom has twelve valence electrons and thus has an expanded valence. All valence electrons of all of the atoms are accounted for.

5. Draw a Lewis structure for a molecule of phosphorus pentachloride, $PCl_5(s)$.

What Is Required?
You need to draw a Lewis structure for phosphorus pentachloride, $PCl_5(s)$.

What Is Given?
The chemical formula tells you that the molecule consists of one phosphorus atom and five chlorine atoms.

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is phosphorus. Therefore, it is the central atom.
Analyze the remaining atoms to see if any groups might form instead of all of the atoms bonding to the central atom.	The remaining atoms are all chlorine atoms. They typically form one single bond with other atoms and thus are not likely to form bonds to each other which would then bond to the central atom.
Draw a skeleton structure with the least electronegative atom as the central atom and attach the remaining atoms according to your reasoning in the previous step.	
Analyze the skeleton structure to determine whether the central atom must have an expanded valence shell.	Phosphorus is bonded to five atoms, giving it ten valence electrons. It must have an expanded valence.
Analyze the atoms to see if there is any obvious reason whether single or double bonds should form. Draw the bonds accordingly.	Chlorine typically forms one single bond by sharing its only unpaired electron. There should not be any double bonds.
The problem statement did not include anything about co-ordinate covalent bonds so assume there are none, meaning that each atom contributes one electron to the bond. Determine the number of valence electrons on the central atom that were used for bonds and how many were not used. Consider the unused valence electrons as forming lone pairs.	Phosphorous has five valence electrons and has used all of them to form single bonds. There are no lone pairs.

Complete the Lewis structure. Add three lone pairs to each of the chlorine atoms to give them a complete octet.	

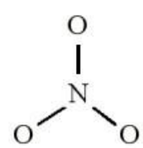
Check Your Solution

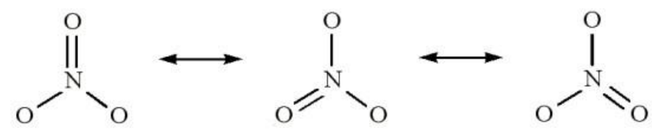
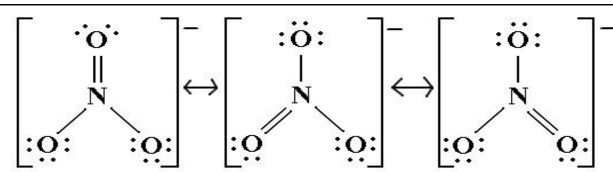
Each chlorine atom has an octet of electrons. The phosphorus atom has ten valence electrons and thus has an expanded valence. All valence electrons of all of the atoms are accounted for.

6. Draw three resonance structures for the $NO_3^-(aq)$ ion.

What Is Required?
You need to draw three resonance structures for a nitrate ion, $NO_3^-(aq)$.

What Is Given?
The chemical formula shows that the nitrate ion has one nitrogen atom, three oxygen atoms, and a charge of 1-.

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is nitrogen therefore it is the central atom.
Draw a skeleton structure of the nitrogen atom with one single bond between it and each of the other atoms.	
Determine the total number of valence electrons in all the atoms of the molecule. The nitrogen atom has five valence electrons, and the oxygen atoms each have six valence electrons. Add one electron to account for the charge.	$V = \left(1\ \cancel{\text{N atom}} \times \frac{5e^-}{\cancel{\text{N atom}}}\right) + \left(3\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right) + 1e^-$ $= 5e^- + 18e^- + 1e^-$ $= 24e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{N atom}} \times \frac{8e^-}{\cancel{\text{N atom}}}\right) + \left(3\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atoms}}}\right)$ $= 8e^- + 24e^-$ $= 32e^-$
Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 32e^- - 24e^-$ $= 8e^-$ $\text{bonds} = \frac{S}{2} = \frac{8}{2} = 4$

Draw the structure with any necessary double or triple bonds. There are three covalent bonds in the skeleton structure and there must be four bonds in the ion. However, it is not possible to distinguish among the oxygen atoms. This is in agreement with the existence of three resonance structures. Any of the three oxygen atoms could be double bonded to the central nitrogen atom at any given time.	
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 24e^{-} - 8e^{-}$ $= 16e^{-}$
Complete the Lewis structure. Because there are 16 non-bonding electrons, there are 8 lone pairs. Place two lone pairs on the double bonded oxygen atom and three long pairs on each of the two single bonded oxygen atoms. Because the structure is an ion with one negative charge, add brackets to the structures and negative signs outside the brackets.	

Check Your Solution

Each atom has eight valence electrons which achieves a noble gas configuration. All of the valence electrons have been accounted for.

7. Draw two resonance structures for a molecule of sulfur dioxide, $SO_2(g)$. An expanded valence structure has been shown to better agree with experimental observations than resonance structures. Draw the Lewis structure for $SO_2(g)$ with an expanded valence.

What Is Required?
You need to draw two resonance structures and one expanded valence structure for sulfur dioxide, $SO_2(g)$.

What Is Given?
The chemical formula shows that the molecule consists of one sulfur atom and two oxygen atoms.
You are told that there could be resonance structrures.
You are told that there could be an expanced valence.

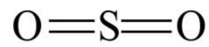
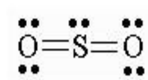
Resonance Structure

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is sulfur therefore it is the central atom.
Draw a skeleton structure with the sulfur atom as the central atom and with one single bond between it and each of the oxygen atoms.	O—S—O
Determine the total number of valence electrons in all the atoms of the molecule. The sulfur atom has six valence electrons, and the oxygen atoms each have six valence electrons.	$V = \left(1\ \cancel{\text{S atom}} \times \frac{6e^-}{\cancel{\text{S atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right)$ $= 6e^- + 12e^-$ $= 18e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{S atom}} \times \frac{8e^-}{\cancel{\text{S atom}}}\right) + \left(2\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atoms}}}\right)$ $= 8e^- + 16e^-$ $= 24e^-$

Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 24e^- - 18e^-$ $= 6e^-$ $\text{bonds} = \frac{S}{2} = \frac{6}{2} = 3$
Draw the structure with any necessary double or triple bonds. There are two covalent bonds in the skeleton structure and there must be three bonds in the molecule. However, it is not possible to distinguish between the oxygen atoms. This is in agreement with the existence of two resonance structures. Either of the oxygen atoms could be double bonded to the central sulfur atom at any given time.	$O{=}S{-}O \longleftrightarrow O{-}S{=}O$
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 18e^- - 6e^-$ $= 12e^-$
Complete the Lewis structure. Since there are 12 non-bonding electrons, there must be 6 lone pairs. Place two lone pairs on the double bonded oxygen atom and three long pairs on each of the single bonded oxygen atoms.	

Expanded Valence

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is sulfur therefore it is the central atom.
Draw a skeleton structure with the sulfur atom as the central atom and with one single bond between it and each of the oxygen atoms.	$O{-}S{-}O$
The problem statement did not include anything about co-ordinate covalent bonds so assume there are none, meaning that each atom contributes one electron to a bond. In addition, the structure cannot be a resonance	

structure. Analyze the atoms to see if there is any obvious reason whether single or double bonds should form. Draw the bonds accordingly. Sulfur has six valence electrons. If two were used to form bonds (one with each of the oxygen atoms), there would be four remaining valence electrons that would form lone pairs. However, the sulfur atom would not have an expanded valence. As well, if the S-O bonds were single bonds, the two oxygen atoms would only have seven valence electrons. Because this cannot be a resonance structure, there cannot be one single and one double bond. Therefore, to give the oxygen atoms a complete octet of electrons, both bonds must be double bonds.	
Determine the number of valence electrons on the central atom that were used for bonds and how many were not used. Consider the unused valence electrons as forming lone pairs.	Four of the six valence electrons of the sulfur atom are used for covalent bonds, thus there are two remaining that will form a lone pair. This gives sulfur ten valence electrons which is an expanded valence. The oxygen atoms use two of their six valence electrons for covalent bonds and thus have four remaining as lone pairs.
Complete the Lewis structure. Add lone pairs to give all of the non-central atoms a complete octet and add a lone pair to the central atom to give it the correct number of valence electrons.	

Check Your Solution

Resonance structure:

Each atom has eight valence electrons which achieves a noble gas configuration. All of the valence electrons have been accounted for.

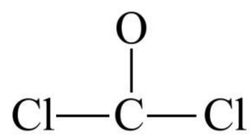
Expanded valence:

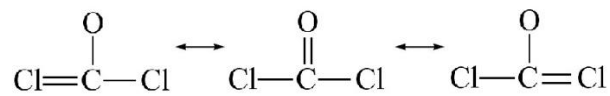
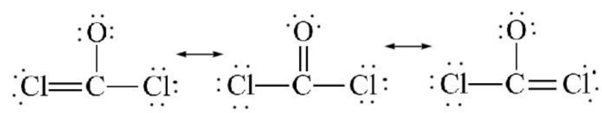
Each oxygen atom has an octet of electrons. The sulfur atom has ten valence electrons and thus has an expanded valence. All valence electrons of all of the atoms are accounted for.

8. Draw three resonance structures for phosgene, $CCl_2O(g)$, a toxic gas.

What Is Required?
You need to draw three resonance structures for phosgene, $CCl_2O(g)$.

What Is Given?
The chemical formula shows that the molecule consists of one carbon atom, two chlorine atoms, and one oxygen atom.

Plan Your Strategy	Act on Your Strategy
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is carbon. Therefore, carbon is the central atom.
Draw a skeleton structure with the carbon atom as the central atom and with one single bond between it and each of the other atoms.	
Determine the total number of valence electrons in all the atoms of the molecule. The carbon atom has four valence electrons, the oxygen atom has six valence electrons, and the chlorine atoms each have seven valence electrons.	$V = \left(1\ \cancel{\text{C atom}} \times \frac{4e^-}{\cancel{\text{C atom}}}\right) + \left(1\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right) + \left(2\ \cancel{\text{Cl atoms}} \times \frac{7e^-}{\cancel{\text{Cl atom}}}\right)$ $= 4e^- + 6e^- + 14e^-$ $= 24e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{C atom}} \times \frac{8e^-}{\cancel{\text{C atom}}}\right) + \left(1\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atom}}}\right) + \left(2\ \cancel{\text{Cl atoms}} \times \frac{8e^-}{\cancel{\text{Cl atom}}}\right)$ $= 8e^- + 8e^- + 16e^-$ $= 32e^-$
Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 32e^- - 24e^-$ $= 8e^-$ $\text{bonds} = \frac{S}{2} = \frac{8}{2} = 4$
Draw the structure with any necessary	

double or triple bonds. There are three covalent bonds in the skeleton structure and there must be four bonds in the ion. However, it is not possible to distinguish between the chlorine atoms. As well, there could be a double bond between the carbon atom and the oxygen atom. This is in agreement with the existence of three resonance structures. Any of the three non-central atoms could be double bonded to the central carbon atom at any given time.	
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$\text{NB} = \text{V} - \text{S}$ $= 24\text{e}^- - 8\text{e}^-$ $= 16\text{e}^-$
Complete the Lewis structure. Because there are 16 non-bonding electrons, there are 8 lone pairs. Place two lone pairs on the double bonded atom and three long pairs on each of the two single bonded atoms.	

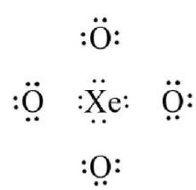
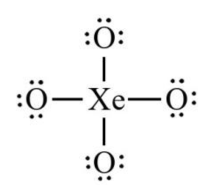
Check Your Solution

Each atom has eight valence electrons, which achieves a noble gas configuration. All of the valence electrons have been accounted for.

9. Compounds of Group 18 elements do not form in nature, but chemists can synthesize such compounds in the lab. One example is xenon tetroxide, XeO_4(g), which contains co-ordinate covalent bonds. Draw the Lewis structure for a molecule of xenon tetroxide.

What Is Required?
You need to draw the Lewis structure for xenon tetroxide, XeO_4(g).

What Is Given?
The chemical formula shows that the molecule consists of one xenon atom and four oxygen atoms.
You are told that xenon tetroxide has co-ordinate covalent bonds.

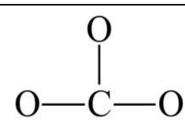
Plan Your Strategy	Act on Your Strategy
Determine the number of valence electrons on each of the atoms.	The xenon atom has eight valence electrons and each of the oxygen atoms has six valence electrons.
Knowing that there are co-ordinate covalent bonds, analyze the atoms and decide which are more likely to donate pairs of electrons to form single bonds and end up with a complete octet of electrons. Decide which atoms could accept a pair of electrons and end up with a complete octet of electrons.	The xenon atom could donate any or all of its paired electrons to form co-ordinate covalent bonds and end up with an octet of electrons. It could not accept pairs of electrons from other atoms because that would result in more than an octet of valence electrons. The oxygen atoms could donate any or all of their electron pairs to form co-ordinate covalent bonds and each could accept one pair of electrons from another atom in the form of a co-ordinate covalent bond and end up with an octet of electrons.
Draw the Lewis structure in accordance with your analysis. Each oxygen atom could accept a pair of electrons from the xenon atom to form co-ordinate covalent bonds and all atoms would have an octet of electrons.	Separate atoms:  Lewis structure: 

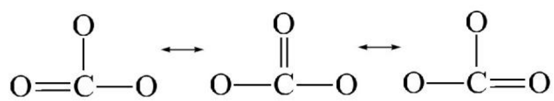
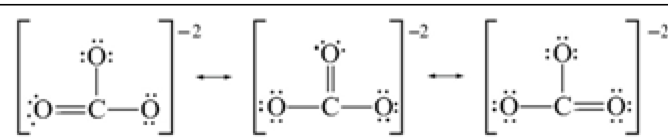
Check Your Solution
All four bonds are co-ordinate covalent bonds and each atom has an octet of electrons.

10. Draw a Lewis structure or structures for the carbonate ion. Consider the possibility of co-ordinate covalent bonds and/or resonance structures.

What Is Required?
You need to draw a Lewis structure or structures for the carbonate ion.

What Is Given?
You know that the compound is a carbonate ion and that there might be co-ordinate covalent bonds or resonance structures.

Plan Your Strategy	Act on Your Strategy
Write the formula for the carbonate ion.	The formula for the carbonate ion is CO_3^{2-}.
Identify the least electronegative atom and make it the central atom in the structure.	The least electronegative atom is the carbon. Therefore, it is the central atom.
Draw a skeleton structure of the central atom with one single bond between it and each of the other atoms.	
Determine the total number of valence electrons in all the atoms of the molecule. The carbon atom has four valence electrons, and each oxygen atom has six valence electrons. The ion has two negative charges, so two electrons must be added to account for the charge.	$V = \left(1\ \cancel{\text{C atom}} \times \frac{4e^-}{\cancel{\text{C atom}}}\right) + \left(3\ \cancel{\text{O atoms}} \times \frac{6e^-}{\cancel{\text{O atom}}}\right) + 2e^-$ $= 4e^- + 18e^- + 2e^-$ $= 24e^-$
Determine the total number of electrons needed for each atom to achieve noble gas configuration.	$T = \left(1\ \cancel{\text{C atom}} \times \frac{8e^-}{\cancel{\text{C atom}}}\right) + \left(3\ \cancel{\text{O atoms}} \times \frac{8e^-}{\cancel{\text{O atoms}}}\right)$ $= 8e^- + 24e^-$ $= 32e^-$

Determine the number of shared electrons and the resulting number of bonds.	$S = T - V$ $= 32e^- - 24e^-$ $= 8e^-$ bonds $= \frac{S}{2} = \frac{8}{2} = 4$
Draw the structure with any necessary double or triple bonds. There are three covalent bonds in the skeleton structure and there must be four bonds in the ion. However, it is not possible to distinguish among the oxygen atoms. This is in agreement with the possibility of resonance structures. Any of the three oxygen atoms could be double bonded to the central carbon atom at any given time. The structures with one double bond added are:	
Determine the number of non-bonding electrons and add them as lone pairs to satisfy the octet rule for all atoms.	$NB = V - S$ $= 24e^- - 8e^-$ $= 16e^-$
Complete the Lewis structure. Because there are 16 non-bonding electrons, there are 8 lone pairs. Place two lone pairs on the double bonded oxygen atom and three long pairs on each of the two single bonded oxygen atoms. Because the structure is an ion with two negative charges, add brackets to the structures and add the exponent, minus two, outside the brackets.	
The above results led to resonance structures. The central atom, a carbon atom, cannot have an expanded valence because it is in Period 2 and Period 2 atoms have no *d* orbitals available. That is, for n = 2, l = 0 or 1, which are the s and p	

orbitals. The energy levels of *d* orbitals with $n = 3$ are so much higher than orbitals with $n = 2$ that no hybrid orbitals could be stable.	

Check Your Solution

Each atom has eight valence electrons which achieves a noble gas configuration. All of the valence electrons have been accounted for.

Predicting the Shape of a Molecule
(Student textbook page 236)

11. What molecular shape is represented by each of the following VSEPR notations?
a. AX_3
b. AX_5E

What Is Required?
You need to assign a molecular shape that is represented by two VSEPR notations.

What Is Given?
You are given the VSEPR notations:
a. AX_3
b. AX_5E

a. AX_3

Plan Your Strategy	Act on Your Strategy
Determine what types of groups and how many of each type are present in the molecule.	There is a central atom and three surrounding atoms but no lone pairs in the molecule.
Apply the principles of electron repulsion to the groups to determine the shape.	Bonding pairs repel equally so the three surrounding atoms will be as far from each other as possible. This would place them at the corners of a triangle with the central atom in the centre. This molecular shape is trigonal planar.

b. AX_5E

Plan Your Strategy	Act on Your Strategy
Determine what types of groups and how many of each group are present.	There is a central atom, five surrounding atoms, and one lone pairs in the molecule.
Apply the principles of electron repulsion to the groups to determine the shape.	The six electron groupings will repel one another but the lone pair will repel more than the others. The electron groupings will be approximately at the corners of an octahedron. The lone pair will not be part of the molecular shape leaving four of the surrounding atoms at the corners of a square and the fifth at the centre and above the square. The central atom will be in the centre of the square. This molecular shape is square pyramidal.

Check Your Solution
A comparison with the figures in Figure 4.31 on page 234 of the student textbook confirms the results.

12. What is the total number of valence electrons surrounding each of the atoms in the following compounds?

a. $CF_4(g)$

b. $NO_3^-(s)$

What Is Required?

You need to determine the total number of valence electrons associated with each of the atoms in two compounds.

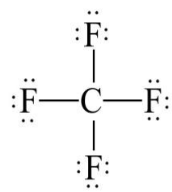
What Is Given?

You are given the compounds:

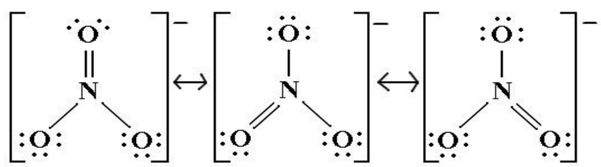
a. $CF_4(g)$

b. $NO_3^-(s)$

a. $CF_4(g)$

Plan Your Strategy	Act on Your Strategy
Draw a Lewis structure of the compound.	
Count the number of electrons surrounding each atom.	There are eight valence electrons around each of the atoms, giving a total of 32 valence electrons in the compound.

b. $NO_3^-(s)$

Plan Your Strategy	Act on Your Strategy
Draw a Lewis structure of the compound.	
Count the number of electrons surrounding each atom.	In all of the three resonance structures, there are eight valence electrons around each of the atoms giving a total of 24 valence electrons in the compound.

Check Your Solution

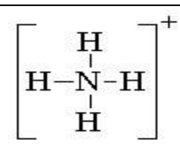
Each atom in both compounds has an octet of valence electrons which creates stability.

13. How many lone pairs and electron groups are associated with the central atom of the following molecules and ions? Remember that double and triple bonds count as only one electron group for the purpose of predicting molecular shape.
a. NH_4^+(s)
b. HCN(g)
c. XeF_4(s)
d. $PbCl_2$(vaporized)

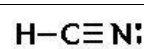
What Is Required?
You need to determine the number of electron groups around the central atom of four compounds.

What Is Given?
You are given the chemical formulas:
a. NH_4^+(s)
b. HCN(g)
c. XeF_4(s)
d. $PbCl_2$(vaporized)

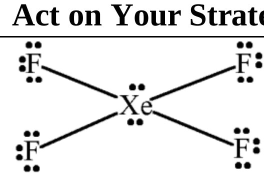
a. NH_4^+(s)

Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure.	
Count the number of each type of electron group around the central atom.	4 BP, 0 LP

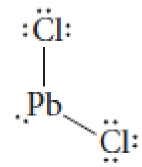
b. HCN(g)

Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure.	
Count the number of each type of electron group around the central atom.	2 BP, 0 LP

c. XeF_4(s)

Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure. (This strucure cannot be obtained by basic steps and requires research. The Xe hybridization is sp^3s^2.)	
Count the number of each type of electron group around the central atom.	4 BP, 2 LP

d. $PbCl_2$(vaporized)

Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure. (This strucure cannot be obtained by basic steps and requires research.)	
Count the number of each type of electron group around the central atom.	2 BP, 1 LP

Check Your Solution
The rules have been followed and the answers are reasonable.

14. Use VSEPR theory to predict the shape of O_3(g).

What Is Required?
You need to predict the shape of ozone, O_3(g).

What Is Given?
The chemical formula shows that the molecule consists of three oxygen atoms.

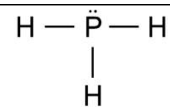
Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure based on the principles for drawing Lewis structures.	
Count the electron groups around the central atom and account for any charge if the compound is an ion.	In either of the resonance forms, there are three electron groups around the central oxygen atom. There is no charge.
Find the name of the electron group arrangement.	This electron group arrangement is called trigonal planar.
Predict the molecular shape based on the location of the bonding pairs and lone pairs in the electron groups.	There are two bonding pairs and one lone pair. This arrangement forms the molecular shape called <u>bent</u>. The bond angle is expected to be less than 120º.

Check Your Solution
This arrangement fits the VSEPR notation AX_2E which represents the bent molecular shape.

15. Use VSEPR theory to predict the shape of $PH_3(g)$.

What Is Required?
You need to predict the molecular shape of phosphorus trihydride, $PH_3(g)$.

What Is Given?
The chemical formula shows that the molecule consists of one phosphorus atom and three hydrogen atoms.

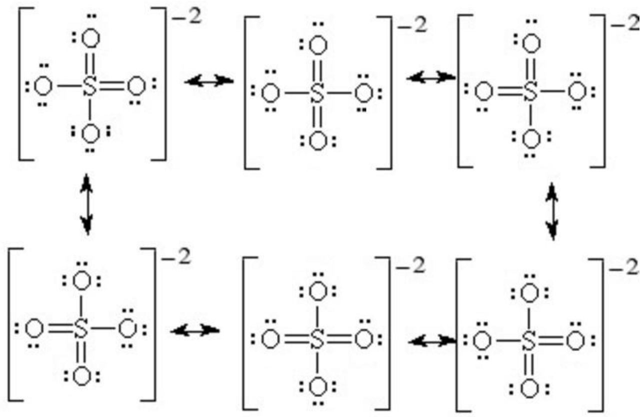
Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure based on the principles for drawing Lewis structures.	
Count the electron groups around the central atom and account for any charge if the compound is an ion.	There are four electron groups around the central phosphorus atom. There is no charge.
Find the name of the electron group arrangement.	This electron group arrangement is called tetrahedral.
Predict the molecular shape based on the location of the bonding pairs and lone pairs in the electron groups.	There are three bonding pairs and one lone pair. This arrangement forms the molecular shape called trigonal pyramidal. The bond angle is expected to be less than 109.5º.

Check Your Solution
This arrangement fits the VSEPR notation AX_3E which represents the trigonal pyramidal molecular shape.

16. Use VSEPR theory to predict the shape of $SO_4^{2-}(s)$.

What Is Required?
You need to predict the shape of the sulfate ion.

What Is Given?
You are given the chemical formula, $SO_4^{2-}(s)$.

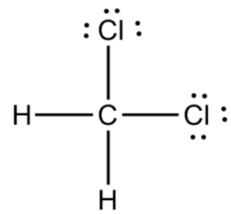
Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure based on the principles for drawing Lewis structures. There are six resonance structures for the sulfate ion. They all contain an expanded valence.	
Count the electron groups around the central atom and account for any charge if the compound is an ion.	In each of the resonance structures, there are four electron groups around the central sulfur atom. The ion has a charge of 2-.
Find the name of the electron group arrangement.	This electron group arrangement is called tetrahedral.
Predict the molecular shape based on the location of the bonding pairs and lone pairs in the electron groups.	There are four bonding pairs and no lone pairs. This arrangement forms the molecular shape called <u>tetrahedral</u>. The bond angle is expected to be 109.5º.

Check Your Solution
This arrangement fits the VSEPR notation AX_4 which represents the tetrahedral molecular shape.

17. Use VSEPR theory to predict the shape and bond angles of $CH_2Cl_2(\ell)$.

What Is Required?
You need to predict the shape of dichloromethane.

What Is Given?
You are given the formula, $CH_2Cl_2(\ell)$.

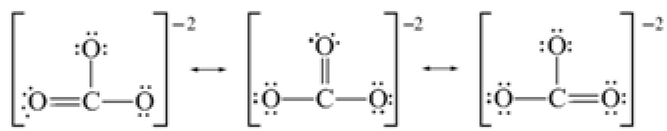
Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure based on the principles for drawing Lewis structures.	
Count the electron groups around the central atom and account for any charge if the compound is an ion.	There are four electron groups around the central carbon atom. There is no charge.
Find the name of the electron group arrangement.	This electron group arrangement is called tetrahedral.
Predict the molecular shape based on the location of the bonding pairs and lone pairs in the electron groups.	There are four bonding pairs and no lone pairs. This arrangement forms the molecular shape called tetrahedral. The bond angle is expected to be 109.5º.

Check Your Solution
This arrangement fits the VSEPR notation AX_4 which represents the tetrahedral molecular shape.

18. Use VSEPR theory to predict the shape and bond angles of CO_3^{2-}(s).

What Is Required?
You need to predict the shape of the carbonate ion.

What is given?
The chemical formula shows that the carbonate ion has one carbon atom, three oxygen atoms, and a charge of 2-.

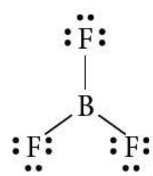
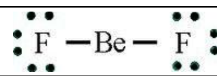
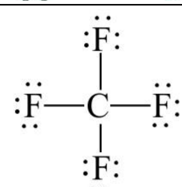
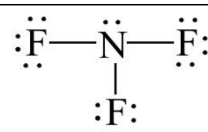
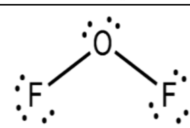
Plan Your Strategy	Act on Your Strategy
Draw the Lewis structure based on the principles for drawing Lewis structures.	
Count the electron groups around the central atom and account for any charge if the compound is an ion.	There are three electron groups around the central carbon atom. There is a charge of 2-.
Find the name of the electron group arrangement.	This electron group arrangement is called trigonal planar.
Predict the molecular shape based on the location of the bonding pairs and lone pairs in the electron groups.	There are three bonding pairs and no lone pairs. This arrangement fits the trigonal planar molecular shape. The bond angle is expected to be 109.5º.

Check Your Solution
This arrangement fits the VSEPR notation AX_3 which represents the trigonal planar molecular shape.

19. Consider the following fluorine-containing molecules to have the general formula XF_n. Arrange them in order of increasing F-X-F bond angles: BF_3(g), BeF_2(g), CF_4(g), NF_3(g), OF_2(g).

What Is Required?
You need to arrange the fluorine containing compounds in order of increasing F-X-F angles, where X is the central atom in each of the compounds.

What Is Given?
You are given the chemical formulas:
BF_3(g), BeF_2(g), CF_4(g), NF_3(g), OF_2(g)

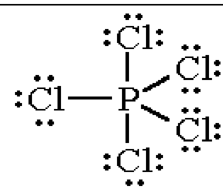
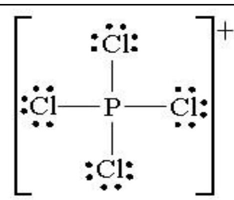
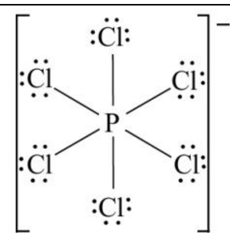
Plan Your Strategy	Act on Your Strategy
Draw Lewis structures of each of the compounds. BF_3(g)	
BeF_2(g)	
CF_4(g)	
NF_3(g)	
OF_2(g)	
Determine the number of bonding pairs and lone pairs in each of the compounds.	BF_3(g): 3 BP, 0 LP, bond angle = 120º BeF_2(g): 2 BP, 0 LP, bond angle =180º CF_4(g): 4 BP, 0 LP, bond angle = 109.5º NF_3(g): 3 BP, 1 LP, bond angle $<$109.5º OF_2(g): 2 BP, 2 LP, bond angle $<<$109.5º
Apply VSEPR theory to predict the deviation of the bond angle from the standard for the standard for the general electron group arrangement.	Order of increasing bond angle: OF_2, NF_3, CF_4, BF_3, BeF_2

Check Your Solution
For the first three compounds in the sequence, the total number of electron groupings increases from 2 to 3 to 4. You would expect that as the number of groups increases, the bond angles would decrease. The last three compounds all have four electrons groups but the number of lone pairs increases from 0 to 1 to 2. Because the lone pairs repel the bonding pairs more than other bonding pairs do, you would expect that as the number of lone pairs increases, the bond angles would decrease. The answer is in agreement with these expectations.

20. Phosphorus pentachloride, used in synthetic chemistry to add chlorine to molecules, exists as molecules of PCl_5(g) in the gas phase. As a solid, though, it exists as alternating ions PCl_4^+(s) and PCl_6^-(s). Predict the shapes of all three species: PCl_5(g), PCl_4^+(s), and PCl_6^-(s).

What Is Required?
You need to predict the molecular shape of three phosphorus- and chlorine-containing compounds.

What Is Given?
You are given the chemical formulas, PCl_5(g), PCl_4^+(s), and PCl_6^-(s).

Plan Your Strategy	Act on Your Strategy
Draw the Lewis structures. PCl_5(g)	
PCl_4^+(s)	
PCl_6^-(s)	
Based on the rules of VSEPR, predict the shapes of the compounds.	PCl_5(g) has the VESPR notation AX_5 and thus its shape is <u>trigonal bipyamidal</u>. PCl_4^+(s) has the VESPR notation AX_4 and thus its shape is <u>tetrahedral</u>. PCl_6^-(s) has the VESPR notation AX_6 and thus its shape is <u>octahedral</u>.

Check Your Solution
None of the compounds have any lone pairs. Thus, the molecular shape is the same as the electron group shape, in agreement with the results.