# 7.E: Chemical Bonding and Molecular Geometry (Exercises)

## 7.E.1: 7.1: Ionic Bonding

### 7.E.1.1: Q7.1.1

Does a cation gain protons to form a positive charge or does it lose electrons?

### 7.E.1.2: S7.1.1

The protons in the nucleus do not change during normal chemical reactions. Only the outer electrons move. Positive charges form when electrons are lost.

### 7.E.1.3: Q7.1.2

Iron(III) sulfate [$Fe_2(SO_4)_3$] is composed of $Fe^{3+}$ and $SO_4^{2-}$ ions. Explain why a sample of iron(III) sulfate is uncharged.

### 7.E.1.4: Q7.1.3

Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: P, I, Mg, Cl, In, Cs, O, Pb, Co?

### 7.E.1.5: S7.1.3

P, I, Cl, and O would form anions because they are nonmetals. Mg, In, Cs, Pb, and Co would form cations because they are metals.

### 7.E.1.6: Q7.1.4

Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: Br, Ca, Na, N, F, Al, Sn, S, Cd?

### 7.E.1.7: Q7.1.5

Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

a. P
b. Mg
c. Al
d. O
e. Cl
f. Cs

### 7.E.1.8: S7.1.5

$P^{3-}$; $Mg^{2+}$; $Al^{3+}$; $O^{2-}$; $Cl^{-}$; $Cs^{+}$

### 7.E.1.9: Q7.1.6

Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

1. I
2. Sr
3. K
4. N
5. S
6. In

### 7.E.1.10: S7.1.6

1. $I^-$
2. $Sr^{2+}$
3. $K^+$
4. $N^{3-}$
5. $S^{2-}$

6. $In^{3+}$

### 7.E.1.11: Q7.1.7

Write the electron configuration for each of the following ions:

1. $As^{3-}$
2. $I^{-}$
3. $Be^{2+}$
4. $Cd^{2+}$
5. $O^{2-}$
6. $Ga^{3+}$
7. $Li^{+}$
8. (h) $N^{3-}$
9. (i) $Sn^{2+}$
10. (j) $Co^{2+}$
11. (k) $Fe^{2+}$
12. (l) $As^{3+}$

### 7.E.1.12: S7.1.7

[Ar]$4s^2 3d^{10} 4p^6$; [Kr]$4d^{10} 5s^2 5p^6$ $1s^2$ [Kr]$4d^{10}$; [He]$2s^2 2p^6$; [Ar]$3d^{10}$; $1s^2$ (h) [He]$2s^2 2p^6$ (i) [Kr]$4d^{10} 5s^2$ (j) [Ar]$3d^7$ (k) [Ar]$3d^6$, (l) [Ar]$3d^{10} 4s^2$

### 7.E.1.13: Q7.1.8

Write the electron configuration for the monatomic ions formed from the following elements (which form the greatest concentration of monatomic ions in seawater):

1. Cl
2. Na
3. Mg
4. Ca
5. K
6. Br
7. Sr
8. (h) F

### 7.E.1.14: Q7.1.9

Write out the full electron configuration for each of the following atoms and for the monatomic ion found in binary ionic compounds containing the element:

a. Al
b. Br
c. Sr
d. Li
e. As
f. S

### 7.E.1.15: S7.1.9

$1s^2 2s^2 2p^6 3s^2 3p^1$; $Al^{3+}$: $1s^2 2s^2 2p^6$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$;

$Sr^{2+}$: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; $1s^2 2s^1$;

$Li^{+}$: $1s^2$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; $1s^2 2s^2 2p^6 3s^2 3p^4$; $1s^2 2s^2 2p^6 3s^2 3p^6$

### 7.E.1.16: Q7.1.10

From the labels of several commercial products, prepare a list of six ionic compounds in the products. For each compound, write the formula. (You may need to look up some formulas in a suitable reference.)

## 7.E.2: 7.3: Covalent Bonding

Why is it incorrect to speak of a molecule of solid NaCl?

NaCl consists of discrete ions arranged in a crystal lattice, not covalently bonded molecules.

What information can you use to predict whether a bond between two atoms is covalent or ionic?

Predict which of the following compounds are ionic and which are covalent, based on the location of their constituent atoms in the periodic table:

1. $Cl_2CO$
2. MnO
3. $NCl_3$
4. $CoBr_2$
5. $K_2S$
6. CO
7. $CaF_2$
8. (h) HI
9. (i) CaO
10. (j) IBr
11. (k) $CO_2$

ionic: (b), (d), (e), (g), and (i); covalent: (a), (c), (f), (h), (j), and (k)

Explain the difference between a nonpolar covalent bond, a polar covalent bond, and an ionic bond.

From its position in the periodic table, determine which atom in each pair is more electronegative:

1. Br or Cl
2. N or O
3. S or O
4. P or S
5. Si or N
6. Ba or P
7. N or K

Cl; O; O; S; N; P; N

From its position in the periodic table, determine which atom in each pair is more electronegative:

1. N or P
2. N or Ge
3. S or F
4. Cl or S
5. H or C
6. Se or P
7. C or Si

From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:

1. C, F, H, N, O
2. Br, Cl, F, H, I
3. F, H, O, P, S
4. Al, H, Na, O, P
5. Ba, H, N, O, As

H, C, N, O, F; H, I, Br, Cl, F; H, P, S, O, F; Na, Al, H, P, O; Ba, H, As, N, O

From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:

1. As, H, N, P, Sb
2. Cl, H, P, S, Si

3. Br, Cl, Ge, H, Sr
4. Ca, H, K, N, Si
5. Cl, Cs, Ge, H, Sr

Which atoms can bond to sulfur so as to produce a positive partial charge on the sulfur atom?

N, O, F, and Cl

Which is the most polar bond?

1. C–C
2. C–H
3. N–H
4. O–H
5. Se–H

Identify the more polar bond in each of the following pairs of bonds:

1. HF or HCl
2. NO or CO
3. SH or OH
4. PCl or SCl
5. CH or NH
6. SO or PO
7. CN or NN

HF; CO; OH; PCl; NH; PO; CN

Which of the following molecules or ions contain polar bonds?

1. $O_3$
2. $S_8$
3. $O_2^{2-}$
4. $NO_3^-$
5. $CO_2$
6. $H_2S$
7. $BH_4^-$

## 7.E.3: 7.4: Lewis Symbols and Structures

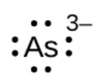
### 7.E.3.1: Q7.4.1

Write the Lewis symbols for each of the following ions:

a. $As^{3-}$
b. $I^-$
c. $Be^{2+}$
d. $O^{2-}$
e. $Ga^{3+}$
f. $Li^+$
g. $N^{3-}$

### 7.E.3.2: S7.4.1

eight electrons:

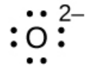



eight electrons:

no electrons

$Be^{2+}$;

eight electrons:



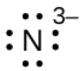
no electrons

$Ga^{3+}$;

no electrons

$Li^{+}$;

eight electrons:



### 7.E.3.3: Q7.4.2

Many monatomic ions are found in seawater, including the ions formed from the following list of elements. Write the Lewis symbols for the monatomic ions formed from the following elements:

a. Cl
b. Na
c. Mg
d. Ca
e. K
f. Br
g. Sr
h. F

### 7.E.3.4: Q7.4.3

Write the Lewis symbols of the ions in each of the following ionic compounds and the Lewis symbols of the atom from which they are formed:

1. MgS
2. $Al_2O_3$
3. $GaCl_3$
4. $K_2O$
5. $Li_3N$
6. KF

(a)

Two Lewis structures are shown. The left shows the symbol M g with a superscripted two positive sign while the right shows the symbol S surrounded by eight dots and a superscripted two negative sign. ;

(b)

Two Lewis structures are shown. The left shows the symbol A l with a superscripted three positive sign while the right shows the symbol O surrounded by eight dots and a superscripted two negative sign. ;

(c)

Two Lewis structures are shown. The left shows the symbol G a with a superscripted three positive sign while the right shows the symbol C l surrounded by eight dots and a superscripted negative sign. ;

(d)

Two Lewis structures are shown. The left shows the symbol K with a superscripted positive sign while the right shows the symbol O surrounded by eight dots and a superscripted two negative sign. ;

(e)

Two Lewis structures are shown. The left shows the symbol L i with a superscripted positive sign while the right shows the symbol N surrounded by eight dots and a superscripted three negative sign. ;

(f)

Two Lewis structures are shown. The left shows the symbol K with a superscripted positive sign while the right shows the symbol F surrounded by eight dots and a superscripted negative sign.

In the Lewis structures listed here, M and X represent various elements in the third period of the periodic table. Write the formula of each compound using the chemical symbols of each element:

(a)

Two Lewis structures are shown side-by-side, each surrounded by brackets. The left structure shows the symbol M with a superscripted two positive sign. The right shows the symbol X surrounded by four lone pairs of electrons with a superscripted two negative sign outside of the brackets.

(b)

Two Lewis structures are shown side-by-side, each surrounded by brackets. The left structure shows the symbol M with a superscripted three positive sign. The right structure shows the symbol X surrounded by four lone pairs of electrons with a superscripted negative sign and a subscripted three both outside of the brackets.

(c)

Two Lewis structures are shown side-by-side, each surrounded by brackets. The left structure shows the symbol M with a superscripted positive sign and a subscripted two outside of the brackets. The right structure shows the symbol X surrounded by four lone pairs of electrons with a superscripted two negative sign outside of the brackets.

(d)

Two Lewis structures are shown side-by-side, each surrounded by brackets. The left structure shows the symbol M with a superscripted three positive sign and a subscripted two outside of the brackets. The right structure shows the symbol X surrounded by four lone pairs of electrons with a superscripted two negative sign and subscripted three both outside of the brackets.

Write the Lewis structure for the diatomic molecule $P_2$, an unstable form of phosphorus found in high-temperature phosphorus vapor.

A Lewis diagram shows two phosphorus atoms triple bonded together each with one lone electron pair.

Write Lewis structures for the following:

a. $H_2$
b. HBr
c. $PCl_3$
d. $SF_2$
e. $H_2CCH_2$
f. HNNH
g. $H_2CNH$
h. (h) $NO^-$
i. (i) $N_2$
j. (j) CO
k. (k) $CN^-$

Write Lewis structures for the following:

a. $O_2$
b. $H_2CO$
c. $AsF_3$
d. ClNO
e. $SiCl_4$
f. $H_3O^+$
g. $NH_4^+$
h. (h) $BF_4^-$
i. (i) HCCH
j. (j) ClCN
k. (k) $C_2^{2+}$

(a)

A Lewis structure shows two oxygen atoms double bonded together, and each has two lone pairs of electrons.

In this case, the Lewis structure is inadequate to depict the fact that experimental studies have shown two unpaired electrons in each oxygen molecule.

(b)

A Lewis structure shows a carbon atom that is single bonded to two hydrogen atoms and double bonded to an oxygen atom. The oxygen atom has two lone pairs of electrons. ;

(c)

A Lewis structure shows an arsenic atom single bonded to three fluorine atoms. Each fluorine atom has a lone pair of electrons. ;

(d)

A Lewis structure shows a nitrogen atom with a lone pair of electrons single bonded to a chlorine atom that has three lone pairs of electrons. The nitrogen is also double bonded to an oxygen which has two lone pairs of electrons.

;

(e)

A Lewis structure shows a silicon atom that is single bonded to four chlorine atoms. Each chlorine atom has three lone pairs of electrons. ;

(f)

A Lewis structure shows an oxygen atom with a lone pair of electrons single bonded to three hydrogen atoms. The structure is surrounded by brackets with a superscripted positive sign. ;

(g)

A Lewis structure shows a nitrogen atom single bonded to four hydrogen atoms. The structure is surrounded by brackets with a superscripted positive sign. ;

(h)

A Lewis structure shows a boron atom single bonded to four fluorine atoms. Each fluorine atom has three lone pairs of electrons. The structure is surrounded by brackets with a superscripted negative sign. ;

(i)

A Lewis structure shows two carbon atoms that are triple bonded together. Each carbon is also single bonded to a hydrogen atom. ;

(j)

A Lewis structure shows a carbon atom that is triple bonded to a nitrogen atom that has one lone pair of electrons. The carbon is also single bonded to a chlorine atom that has three lone pairs of electrons. ;

(k)

A Lewis structure shows two carbon atoms joined with a triple bond. A superscripted 2 positive sign lies to the right of the second carbon.

Write Lewis structures for the following:

1. $ClF_3$
2. $PCl_5$
3. $BF_3$
4. $PF_6^-$

Write Lewis structures for the following:

1. $SeF_6$
2. $XeF_4$
3. $SeCl_3^+$
4. $Cl_2BBCl_2$ (contains a B–B bond)

$SeF_6$:

A Lewis structure shows a selenium atom single bonded to six fluorine atoms, each with three lone pairs of electrons. ;

$XeF_4$:

A Lewis structure shows a xenon atom with two lone pairs of electrons. It is single bonded to four fluorine atoms each with three lone pairs of electrons. ;

$SeCl_3^+$:

A Lewis structure shows a selenium atom with one lone pair of electrons single bonded to three chlorine atoms each with three lone pairs of electrons. The whole structure is surrounded by brackets. ;

$Cl_2BBCl_2$:

A Lewis structure shows two boron atoms that are single bonded together. Each is also single bonded to two chlorine atoms that both have three lone pairs of electrons.

Write Lewis structures for:

1. $PO_4^{3-}$
2. $ICl_4^-$
3. $SO_3^{2-}$
4. HONO

Correct the following statement: “The bonds in solid $PbCl_2$ are ionic; the bond in a HCl molecule is covalent. Thus, all of the valence electrons in $PbCl_2$ are located on the $Cl^-$ ions, and all of the valence electrons in a HCl molecule are shared between the H and Cl atoms.”

Two valence electrons per Pb atom are transferred to Cl atoms; the resulting $Pb^{2+}$ ion has a $6s^2$ valence shell configuration. Two of the valence electrons in the HCl molecule are shared, and the other six are located on the Cl atom as lone pairs of electrons.

Write Lewis structures for the following molecules or ions:

1. $SbH_3$
2. $XeF_2$
3. $Se_8$ (a cyclic molecule with a ring of eight Se atoms)

Methanol, $H_3COH$, is used as the fuel in some race cars. Ethanol, $C_2H_5OH$, is used extensively as motor fuel in Brazil. Both methanol and ethanol produce $CO_2$ and $H_2O$ when they burn. Write the chemical equations for these combustion reactions using Lewis structures instead of chemical formulas.

Two reactions are shown using Lewis structures. The top reaction shows a carbon atom, single bonded to three hydrogen atoms and single bonded to an oxygen atom with two lone pairs of electrons. The oxygen atom is also bonded to a hydrogen atom. This is followed by a plus sign and the number one point five, followed by two oxygen atoms bonded together with a double bond and each with two lone pairs of electrons. A right-facing arrow leads to a carbon atom that is double bonded to two oxygen atoms, each of which has two lone pairs of electrons. This structure is followed by a plus sign, a number two, and a structure made up of an oxygen with two lone pairs of electrons single bonded to two hydrogen atoms. The bottom reaction shows a carbon atom, single bonded to three hydrogen atoms and single bonded to another carbon atom. The second carbon atom is single bonded to two hydrogen atoms and one oxygen atom with two lone pairs of electrons. The oxygen atom is also bonded to a hydrogen atom. This is followed by a plus sign and the number three, followed by two oxygen atoms bonded together with a double bond. Each oxygen atom has two lone pairs of electrons. A right-facing arrow leads to a number two and a carbon atom that is double bonded to two oxygen atoms, each of which has two lone pairs of electrons. This structure is followed by a plus sign, a number three, and a structure made up of an oxygen with two lone pairs of electrons single bonded to two hydrogen atoms.

Many planets in our solar system contain organic chemicals including methane ($CH_4$) and traces of ethylene ($C_2H_4$), ethane ($C_2H_6$), propyne ($H_3CCCH$), and diacetylene (HCCCCH). Write the Lewis structures for each of these molecules.

Carbon tetrachloride was formerly used in fire extinguishers for electrical fires. It is no longer used for this purpose because of the formation of the toxic gas phosgene, $Cl_2CO$. Write the Lewis structures for carbon tetrachloride and phosgene.

Two Lewis structures are shown. The left depicts a carbon atom single bonded to four chlorine atoms, each with three lone pairs of electrons. The right shows a carbon atom double bonded to an oxygen atom that has two lone pairs of electrons. The carbon atom is also single bonded to two chlorine atoms, each of which has three lone pairs of electrons.

Identify the atoms that correspond to each of the following electron configurations. Then, write the Lewis symbol for the common ion formed from each atom:

1. $1s^22s^22p^5$
2. $1s^22s^22p^63s^2$
3. $1s^22s^22p^63s^23p^64s^23d^{10}$
4. $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$
5. $1s^22s^22p^63s^23p^64s^23d^{10}4p^1$

The arrangement of atoms in several biologically important molecules is given here. Complete the Lewis structures of these molecules by adding multiple bonds and lone pairs. Do not add any more atoms.

the amino acid serine:

A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to a hydrogen atom and two other carbon atoms. One of these carbon atoms is single bonded to two hydrogen atoms and an oxygen atom. The oxygen atom is bonded to a hydrogen atom. The other carbon atom is single bonded to two oxygen atoms, one of which is bonded to a hydrogen atom.

urea:

A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to an oxygen atom and another nitrogen atom. That nitrogen atom is then single bonded to two hydrogen atoms.

pyruvic acid:

A Lewis structure is shown. A carbon atom is single bonded to three hydrogen atoms and another carbon atom. The second carbon atom is single bonded to an oxygen atom and a third carbon atom. This carbon is then single bonded to two oxygen atoms, one of which is single bonded to a hydrogen atom.

uracil:

A Lewis hexagonal ring structure is shown. From the top of the ring (moving clockwise), three carbon atoms, one nitrogen atom, a carbon atom, and a nitrogen atom are single bonded to each another. The top carbon atom is single bonded to an oxygen atom. The second and third carbons and the nitrogen atom are each single bonded to a hydrogen atom. The next carbon atom is single bonded to an oxygen atom, and the last nitrogen atom is single bonded to a hydrogen atom.

carbonic acid:

A Lewis structure is shown. A carbon atom is single bonded to three oxygen atoms. Two of those oxygen atoms are each single bonded to a hydrogen atom.

(a)

A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to a hydrogen atom and two other carbon atoms. One of these carbon atoms is single bonded to two hydrogen atoms and an oxygen atom. The oxygen atom is bonded to a hydrogen atom. The other carbon is single bonded to two oxygen atoms, one of which is bonded to a hydrogen atom. The oxygen atoms have two lone pairs of electron dots, and the nitrogen atom has one lone pair of electron dots.

;

(b)

A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to an oxygen atom and one nitrogen atom. That nitrogen atom is then single bonded to two hydrogen atoms. The oxygen atom has two lone pairs of electron dots, and the nitrogen atoms have one lone pair of electron dots each.

;

(c)

A Lewis structure is shown. A carbon atom is single bonded to three hydrogen atoms and a carbon atom. The carbon atom is single bonded to an oxygen atom and a third carbon atom. This carbon is then single bonded to two oxygen atoms, one of which is single bonded to a hydrogen atom. Each oxygen atom has two lone pairs of electron dots.

;

(d)

A Lewis hexagonal ring structure is shown. From the top of the ring, three carbon atoms, one nitrogen atom, a carbon atom and a nitrogen atom are single bonded to one another. The top carbon is single bonded to an oxygen, the second and third carbons and the nitrogen atom are each single bonded to a hydrogen atom. The next carbon is single bonded to an oxygen atom and the last nitrogen is single bonded to a hydrogen atom. The oxygen atoms have two lone pairs of electron dots, and the nitrogen atoms have one lone pair of electron dots.

;

(e)

A Lewis structure is shown. A carbon atom is single bonded to three oxygen atoms. Two of those oxygen atoms are each single bonded to a hydrogen atom. Each oxygen atom has two lone pairs of electron dots.

A compound with a molar mass of about 28 g/mol contains 85.7% carbon and 14.3% hydrogen by mass. Write the Lewis structure for a molecule of the compound.

A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen by mass. Write the Lewis structure for a molecule of the compound.

A Lewis structure is shown. A carbon atom is single bonded to three hydrogen atoms and another carbon atom. The second carbon atom is double bonded to another carbon atom and single bonded to a hydrogen atom. The last carbon is single bonded to two hydrogen atoms.

Two arrangements of atoms are possible for a compound with a molar mass of about 45 g/mol that contains 52.2% C, 13.1% H, and 34.7% O by mass. Write the Lewis structures for the two molecules.

How are single, double, and triple bonds similar? How do they differ?

Each bond includes a sharing of electrons between atoms. Two electrons are shared in a single bond; four electrons are shared in a double bond; and six electrons are shared in a triple bond.

## 7.E.4: 7.5: Formal Charges and Resonance

Write resonance forms that describe the distribution of electrons in each of these molecules or ions.

1. selenium dioxide, OSeO
2. nitrate ion, $NO_3^-$
3. nitric acid, $HNO_3$ (N is bonded to an OH group and two O atoms)
4. benzene, $C_6H_6$:

A Lewis structure shows a hexagonal ring composed of six carbon atoms. They form single bonds to each another and single bonds to one hydrogen atom each.

the formate ion:

A Lewis structure shows a carbon atom single bonded to two oxygen atoms and a hydrogen atom. The structure is surrounded by brackets and there is a superscripted negative sign.

Write resonance forms that describe the distribution of electrons in each of these molecules or ions.

1. sulfur dioxide, $SO_2$
2. carbonate ion, $CO_3^{2-}$

3. hydrogen carbonate ion, $HCO_3^-$ (C is bonded to an OH group and two O atoms)
4. pyridine:

A Lewis structure depicts a hexagonal ring composed of five carbon atoms and one nitrogen atom. Each carbon atom is single bonded to a hydrogen atom.

the allyl ion:

A Lewis structure shows a carbon atom single bonded to two hydrogen atoms and a second carbon atom. The second carbon atom is single bonded to a hydrogen atom and a third carbon atom. The third carbon atom is single bonded to two hydrogen atoms. The whole structure is surrounded by brackets, and there is a superscripted negative sign.

(a)

Two Lewis structures are shown with a double-headed arrow in between. The left structure shows a sulfur atom with a lone pair of electrons single bonded to the left to an oxygen atom with three lone pairs of electrons. The sulfur atom is also double bonded on the right to an oxygen atom with two lone pairs of electrons. The right structure depicts the same atoms, but this time the double bond is between the left oxygen and the sulfur atom. The lone pairs of electrons have also shifted to account for the change of bond types. The sulfur atom in the right structures, also has a third electron dot below it.

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(b)

Three Lewis structures are shown, with double-headed arrows in between, each surrounded by brackets and a superscripted two negative sign. The left structure depicts a carbon atom bonded to three oxygen atoms. It is single bonded to two of these oxygen atoms, each of which has three lone pairs of electrons, and double bonded to the third, which has two lone pairs of electrons. The double bond is located between the bottom oxygen and the carbon. The central and right structures are the same as the first, but the position of the double bonded oxygen has moved to the left oxygen in the right structure while the central structure only has single bonds. The lone pairs of electrons change to correspond with the bonds as well.

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(c)

Two Lewis structures are shown, with a double-headed arrow in between, each surrounded by brackets and a superscripted negative sign. The left structure depicts a carbon atom bonded to three oxygen atoms. It is single bonded to one of these oxygen atoms, which has three lone pairs of electrons, and double bonded to the other two, which have two lone pairs of electrons. One of the double bonded oxygen atoms also has a single bond to a hydrogen atom. The right structure is the same as the first, but there is only one double bonded oxygen. The oxygen with the single bonded hydrogen now has a single bond to the carbon atom. The lone pairs of electrons have also changed to correspond with the bonds.

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(d)

Two Lewis structures are shown with a double-headed arrow in between. The left structure depicts a hexagonal ring composed of five carbon atoms, each single bonded to a hydrogen atom, and one nitrogen atom that has a lone pair of electrons. The ring has alternating single and double bonds. The right structure is the same as the first, but each double bond has rotated to a new position.

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(e)

Two Lewis structures are shown with a double-headed arrow in between. The left structure shows a carbon atom single bonded to two hydrogen atoms and a second carbon atom. The second carbon atom is single bonded to a hydrogen atom and double bonded to a third carbon atom. The third carbon atom is single bonded to two hydrogen atoms. The whole structure is surrounded by brackets and a superscripted negative sign. The right structure shows a carbon atom single bonded to two hydrogen atoms and double bonded to a second carbon atom. The second carbon atom is single bonded to a hydrogen atom and a third carbon atom. The third carbon atom is single bonded to two hydrogen atoms. The whole structure is surrounded by brackets and a superscripted negative sign.

Write the resonance forms of ozone, $O_3$, the component of the upper atmosphere that protects the Earth from ultraviolet radiation.

Sodium nitrite, which has been used to preserve bacon and other meats, is an ionic compound. Write the resonance forms of the nitrite ion, $NO_2^-$.

Two pairs of Lewis structures are shown with a double-headed arrow in between each pair. The left structure of the first pair shows a nitrogen atom with one lone pair of electrons single bonded to an oxygen atom with three lone pairs of electrons. It is also double bonded to an oxygen with two lone pairs of electrons. The right image of this pair depicts the mirror image of the left. Both images are surrounded by brackets and a superscripted negative sign. They are labeled, "For N O subscript two superscript negative sign." The left structure of the second pair shows an oxygen atom with one lone pair of electrons single bonded to an oxygen atom with three lone pairs of electrons. It is also double bonded to an oxygen atom with two lone pairs of electrons. The right structure appears as a mirror image of the left. These structures are labeled, "For O subscript three."

In terms of the bonds present, explain why acetic acid, $CH_3CO_2H$, contains two distinct types of carbon-oxygen bonds, whereas the acetate ion, formed by loss of a hydrogen ion from acetic acid, only contains one type of carbon-oxygen bond. The skeleton structures of these species are shown:

Two Lewis structures are shown with a double headed arrow in between. The left structure shows a carbon atom single bonded to three hydrogen atoms and a second carbon atom. The second carbon is single bonded to two oxygen atoms. One of the oxygen atoms is single bonded to a hydrogen atom. The right structure, surrounded by brackets and with a superscripted negative sign, depicts a carbon atom single bonded to three hydrogen atoms and a second carbon atom. The second carbon atom is single bonded to two oxygen atoms.

Write the Lewis structures for the following, and include resonance structures where appropriate. Indicate which has the strongest carbon-oxygen bond.

1. $CO_2$
2. CO

(a)

This structure shows a carbon atom double bonded to two oxygen atoms, each of which has two lone pairs of electrons.

(b)

The right structure of this pair shows a carbon atom with one lone pair of electrons triple bonded to an oxygen with one lone pair of electrons.

CO has the strongest carbon-oxygen bond because there is a triple bond joining C and O. $CO_2$ has double bonds.

Toothpastes containing sodium hydrogen carbonate (sodium bicarbonate) and hydrogen peroxide are widely used. Write Lewis structures for the hydrogen carbonate ion and hydrogen peroxide molecule, with resonance forms where appropriate.

Determine the formal charge of each element in the following:

1. HCl
2. $CF_4$
3. $PCl_3$
4. $PF_5$

H: 0, Cl: 0; C: 0, F: 0; P: 0, Cl 0; P: 0, F: 0

Determine the formal charge of each element in the following:

1. $H_3O^+$
2. $SO_4^{2-}$
3. $NH_3$
4. $O_2^{2-}$
5. $H_2O_2$

Calculate the formal charge of chlorine in the molecules $Cl_2$, $BeCl_2$, and $ClF_5$.

Cl in $Cl_2$: 0; Cl in $BeCl_2$: 0; Cl in $ClF_5$: 0

Calculate the formal charge of each element in the following compounds and ions:

1. $F_2CO$
2. $NO^-$
3. $BF_4^-$
4. $SnCl_3^-$
5. $H_2CCH_2$
6. $ClF_3$
7. $SeF_6$
8. (h) $PO_4^{3-}$

Draw all possible resonance structures for each of these compounds. Determine the formal charge on each atom in each of the resonance structures:

1. $O_3$
2. $SO_2$
3. $NO_2^-$
4. $NO_3^-$

Two Lewis structures are shown with a double-headed arrow in between. The left structure shows an oxygen atom with one lone pair of electrons single bonded to an oxygen atom with three lone pairs of electrons. It is also double bonded to an oxygen atom with two lone pairs of electrons. The symbols and numbers below this structure read, “( 0 ), ( positive 1 ), ( negative 1 ).” The phrase, “Formal charge,” and a right-facing arrow lie to the left of this structure. The right structure appears as a mirror image of the left and the symbols and numbers below this structure read, “( negative 1 ), ( positive 1 ), ( 0 ).”

;

(b)

Two Lewis structures are shown, with a double-headed arrow in between. The left structure shows a sulfur atom with one lone pair of electrons single bonded to an oxygen atom with three lone pairs of electrons. The sulfur atom also double bonded to an oxygen atom with two lone pairs of electrons. The symbols and numbers below this structure read, “( negative 1 ), ( positive 1 ), ( 0 ).” The right structure appears as a mirror image of the left and the symbols and numbers below this structure read, “( 0 ), ( positive 1 ), ( negative 1 ).”

;

(c)

[Two Lewis structures are shown, with brackets surrounding each with a superscripted negative sign and a double ended arrow in between. The left structure shows a nitrogen atom with one lone pair of electrons single bonded to an oxygen atom with three lone pairs of electrons and double bonded to an oxygen atom with two lone pairs of electrons. The symbols and numbers below this structure read “open parenthesis, 0, close parenthesis, open parenthesis, 0, close parenthesis, open parenthesis, negative 1, close parenthesis. The right structure appears as a mirror image of the left and the symbols and numbers below this structure read “open parenthesis, negative 1, close parenthesis, open parenthesis, 0, close parenthesis, open parenthesis, 0, close parenthesis.]

;

(d)

[Three Lewis structures are shown, with brackets surrounding each with a superscripted negative sign and a double ended arrow in between. The left structure shows a nitrogen atom single bonded to two oxygen atoms, each with three lone pairs of electrons and double bonded to an oxygen atom with two lone pairs of electrons. The single bonded oxygen atoms are labeled, from the top of the structure and going clockwise, "open parenthesis, negative 1, close parenthesis, open parenthesis, positive 1, close parenthesis". The symbols and numbers below this structure read "open parenthesis, 0, close parenthesis, open parenthesis, negative 1, close parenthesis. The middle structure shows a nitrogen atom single bonded to two oxygen atoms, each with three lone pairs of electrons, one of which is labeled "open parenthesis, positive 1, close parenthesis" and double bonded to an oxygen atom with two lone pairs of electrons labeled "open parenthesis, 0, close parenthesis". The symbols and numbers below this structure read "open parenthesis, negative 1, close parenthesis, open parenthesis, negative 1, close parenthesis. The right structure shows a nitrogen atom single bonded to two oxygen atoms, each with three lone pairs of electrons and double bonded to an oxygen atom with two lone pairs of electrons. One of the single bonded oxygen atoms is labeled, "open parenthesis, negative 1, close parenthesis while the double bonded oxygen is labeled, "open parenthesis, positive 1, close parenthesis". The symbols and numbers below this structure read "open parenthesis, negative 1, close parenthesis" and "open parenthesis, 0, close parenthesis".]

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in nitrosyl chloride: ClNO or ClON?

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in hypochlorous acid: HOCl or OClH?

HOCl

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in sulfur dioxide: OSO or SOO?

Draw the structure of hydroxylamine, $H_3NO$, and assign formal charges; look up the structure. Is the actual structure consistent with the formal charges?

The structure that gives zero formal charges is consistent with the actual structure:

A Lewis structure shows a nitrogen atom with one lone pair of electrons single bonded to two hydrogen atoms and an oxygen atom which has two lone pairs of electrons. The oxygen atom is single bonded to a hydrogen atom.

Iodine forms a series of fluorides (listed here). Write Lewis structures for each of the four compounds and determine the formal charge of the iodine atom in each molecule:

1. IF
2. $IF_3$
3. $IF_5$
4. $IF_7$

Write the Lewis structure and chemical formula of the compound with a molar mass of about 70 g/mol that contains 19.7% nitrogen and 80.3% fluorine by mass, and determine the formal charge of the atoms in this compound.

$NF_3$;

A Lewis structure shows a nitrogen atom with one lone pair of electrons single bonded to three fluorine atoms, each with three lone pairs of electrons.

Which of the following structures would we expect for nitrous acid? Determine the formal charges:

Two Lewis structures are shown, with the word "or" in between. The left structure shows a nitrogen atom single bonded to an oxygen atom with three lone pairs of electrons. It is also single bonded to a hydrogen atom and double bonded to an oxygen atom with two lone pairs of electrons. The right structure shows a hydrogen atom single bonded to an oxygen atom with two lone pairs of electrons. The oxygen atom is single bonded to a nitrogen atom which is double bonded to an oxygen atom with two lone pairs of electrons.

Sulfuric acid is the industrial chemical produced in greatest quantity worldwide. About 90 billion pounds are produced each year in the United States alone. Write the Lewis structure for sulfuric acid, $H_2SO_4$, which has two oxygen atoms and two OH groups bonded to the sulfur.

A Lewis structure shows a hydrogen atom single bonded to an oxygen atom with two lone pairs of electrons. The oxygen atom is single bonded to a sulfur atom. The sulfur atom is double bonded to two oxygen atoms, each of which have three lone pairs of electrons, and single bonded to an oxygen atom with two lone pairs of electrons. This oxygen atom is single bonded to a hydrogen atom.

## 7.E.5: 7.6: Strengths of Ionic and Covalent Bonds

Which bond in each of the following pairs of bonds is the strongest?

1. C–C or $\mathrm{C=C}$
2. C–N or $\mathrm{C\equiv N}$
3. $\mathrm{C\equiv O}$ or $\mathrm{C=O}$
4. H–F or H–Cl
5. C–H or O–H
6. C–N or C–O

Using the bond energies in Table, determine the approximate enthalpy change for each of the following reactions:

1. $\mathrm{H_2}(g)+\mathrm{Br_2}(g)\longrightarrow 2\,\mathrm{HBr}(g)$
2. $\mathrm{CH_4}(g)+\mathrm{I_2}(g)\longrightarrow \mathrm{CH_3I}(g)+\mathrm{HI}(g)$

3. (c) $C_2H_4(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 2\,H_2O(g)$

1. −114 kJ;
2. 30 kJ;
3. (c) −1055 kJ

Using the bond energies in Table, determine the approximate enthalpy change for each of the following reactions:

1. $Cl_2(g) + 3\,F_2(g) \longrightarrow 2\,ClF_3(g)$
2. $H_2C = CH_2(g) + H_2(g) \longrightarrow H_3CCH_3(g)$
3. (c) $2\,C_2H_6(g) + 7\,O_2(g) \longrightarrow 4\,CO_2(g) + 6\,H_2O(g)$

When a molecule can form two different structures, the structure with the stronger bonds is usually the more stable form. Use bond energies to predict the correct structure of the hydroxylamine molecule:

Two Lewis structures are shows with the word "or" written in between them. The left structure shows a nitrogen atom with one lone pair of electrons single bonded to two hydrogen atoms. It is also bonded to an oxygen atom with two lone pairs of electrons. The oxygen atom is bonded to a hydrogen atom. The right structure shows a nitrogen atom single bonded to three hydrogen atoms and an oxygen atom with three lone pairs of electrons.

The greater bond energy is in the figure on the left. It is the more stable form.

How does the bond energy of HCldiffer from the standard enthalpy of formation of HCl($g$)?

Using the standard enthalpy of formation data in Appendix G, show how the standard enthalpy of formation of HCl($g$) can be used to determine the bond energy.

$$\begin{array}{ll}
HCl(g) \longrightarrow \frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g) & \Delta H_1^\circ = -\Delta H^\circ_{\mathrm{f}[HCl(g)]} \\
\frac{1}{2}H_2(g) \longrightarrow H(g) & \Delta H_2^\circ = \Delta H^\circ_{\mathrm{f}[H(g)]} \\
\frac{1}{2}Cl_2(g) \longrightarrow Cl(g) & \Delta H_3^\circ = \Delta H^\circ_{\mathrm{f}[Cl(g)]} \\
\hline
HCl(g) \longrightarrow H(g) + Cl(g) & \Delta H^\circ_{298} = \Delta H_1^\circ + \Delta H_2^\circ + \Delta H_3^\circ
\end{array}$$

$$D_{HCl} = \Delta H^\circ_{298} = \Delta H^\circ_{\mathrm{f}[HCl(g)]} + \Delta H^\circ_{\mathrm{f}[H(g)]} + \Delta H^\circ_{\mathrm{f}[Cl(g)]} \tag{7.E.1}$$

$$= -(-92.307\text{ kJ}) + 217.97\text{ kJ} + 121.3\text{ kJ} \tag{7.E.2}$$

$$= 431.6\text{ kJ} \tag{7.E.3}$$

Using the standard enthalpy of formation data in Appendix G, calculate the bond energy of the carbon-sulfur double bond in $CS_2$.

Using the standard enthalpy of formation data in Appendix G, determine which bond is stronger: the S–F bond in $SF_4(g)$ or in $SF_6(g)$?

The S–F bond in $SF_4$ is stronger.

Using the standard enthalpy of formation data in Appendix G, determine which bond is stronger: the P–Cl bond in $PCl_3(g)$ or in $PCl_5(g)$?

Complete the following Lewis structure by adding bonds (not atoms), and then indicate the longest bond:

A Lewis structure is shown that is missing its bonds. It shows a horizontal row of six carbon atoms, equally spaced. Three hydrogen atoms are drawn around the first carbon, two around the second, one above the fifth, and two by the sixth.

A Lewis structure is shown. A carbon atom that is single bonded to three hydrogen atoms is bonded to a second carbon atom. The second carbon atom is single bonded to two hydrogen atoms. The second carbon atom is single bonded to a third carbon atom that is triple bonded to a fourth carbon atom and single bonded to a fifth carbon atom. The fifth carbon atom is single bonded to a hydrogen atom and double bonded to a sixth carbon atom that is single bonded to two hydrogen atoms.

The C–C single bonds are longest.

Use the bond energy to calculate an approximate value of Δ$H$ for the following reaction. Which is the more stable form of $FNO_2$?

Two Lewis structures are shown with a right-facing arrow in between. The left structure shows a nitrogen atom double bonded to an oxygen atom with two lone pairs of electrons. It is also bonded to a fluorine atom and another oxygen atom, each with three lone pairs of electrons. The right structure shows an oxygen atom with two lone pairs of electrons double bonded to a nitrogen atom with one lone pair of electrons. This nitrogen atom is single bonded to an oxygen with two lone pairs of electrons. The oxygen atom is single bonded to a fluorine atom with three lone pairs of electrons.

Use principles of atomic structure to answer each of the following:[1]

1. The radius of the Ca atom is 197 pm; the radius of the $Ca^{2+}$ ion is 99 pm. Account for the difference.
2. The lattice energy of CaO($s$) is –3460 kJ/mol; the lattice energy of $K_2O$ is –2240 kJ/mol. Account for the difference.

3. (c) Given these ionization values, explain the difference between Ca and K with regard to their first and second ionization energies.

| Element | First Ionization Energy (kJ/mol) | Second Ionization Energy (kJ/mol) |
|---|---|---|
| K | 419 | 3050 |
| Ca | 590 | 1140 |

The first ionization energy of Mg is 738 kJ/mol and that of Al is 578 kJ/mol. Account for this difference.

When two electrons are removed from the valence shell, the Ca radius loses the outermost energy level and reverts to the lower *n* = 3 level, which is much smaller in radius. The +2 charge on calcium pulls the oxygen much closer compared with K, thereby increasing the lattice energy relative to a less charged ion. (c) Removal of the 4*s* electron in Ca requires more energy than removal of the 4*s* electron in K because of the stronger attraction of the nucleus and the extra energy required to break the pairing of the electrons. The second ionization energy for K requires that an electron be removed from a lower energy level, where the attraction is much stronger from the nucleus for the electron. In addition, energy is required to unpair two electrons in a full orbital. For Ca, the second ionization potential requires removing only a lone electron in the exposed outer energy level. In Al, the removed electron is relatively unprotected and unpaired in a *p* orbital. The higher energy for Mg mainly reflects the unpairing of the 2*s* electron.

The lattice energy of LiF is 1023 kJ/mol, and the Li–F distance is 200.8 pm. NaF crystallizes in the same structure as LiF but with a Na–F distance of 231 pm. Which of the following values most closely approximates the lattice energy of NaF: 510, 890, 1023, 1175, or 4090 kJ/mol? Explain your choice.

For which of the following substances is the least energy required to convert one mole of the solid into separate ions?

1. MgO
2. SrO
3. (c) KF
4. CsF
5. $MgF_2$

(d)

The reaction of a metal, M, with a halogen, $X_2$, proceeds by an exothermic reaction as indicated by this equation: $\mathrm{M}(s)+\mathrm{X}_2(g)\longrightarrow \mathrm{MX}_2(s)$ . For each of the following, indicate which option will make the reaction more exothermic. Explain your answers.

1. a large radius vs. a small radius for $M^{+2}$
2. a high ionization energy vs. a low ionization energy for M
3. (c) an increasing bond energy for the halogen
4. a decreasing electron affinity for the halogen
5. an increasing size of the anion formed by the halogen

The lattice energy of LiF is 1023 kJ/mol, and the Li–F distance is 201 pm. MgO crystallizes in the same structure as LiF but with a Mg–O distance of 205 pm. Which of the following values most closely approximates the lattice energy of MgO: 256 kJ/mol, 512 kJ/mol, 1023 kJ/mol, 2046 kJ/mol, or 4008 kJ/mol? Explain your choice.

4008 kJ/mol; both ions in MgO have twice the charge of the ions in LiF; the bond length is very similar and both have the same structure; a quadrupling of the energy is expected based on the equation for lattice energy

Which compound in each of the following pairs has the larger lattice energy? Note: $Mg^{2+}$ and $Li^{+}$ have similar radii; $O^{2-}$ and $F^{-}$ have similar radii. Explain your choices.

1. MgO or MgSe
2. LiF or MgO
3. (c) $Li_2O$ or LiCl
4. $Li_2Se$ or MgO

Which compound in each of the following pairs has the larger lattice energy? Note: $Ba^{2+}$ and

$K^+$ have similar radii; $S^{2-}$ and $Cl^-$ have similar radii. Explain your choices.

1. $K_2O$ or $Na_2O$
2. $K_2S$ or BaS
3. (c) KCl or BaS
4. BaS or $BaCl_2$

$Na_2O$; $Na^+$ has a smaller radius than $K^+$; BaS; Ba has a larger charge than K; (c) BaS; Ba and S have larger charges; BaS; S has a larger charge

Which of the following compounds requires the most energy to convert one mole of the solid into separate ions?

1. MgO
2. SrO
3. (c) KF
4. CsF
5. $MgF_2$

Which of the following compounds requires the most energy to convert one mole of the solid into separate ions?

1. $K_2S$
2. $K_2O$
3. (c) CaS
4. $Cs_2S$
5. CaO

(e)

The lattice energy of KF is 794 kJ/mol, and the interionic distance is 269 pm. The Na–F

distance in NaF, which has the same structure as KF, is 231 pm. Which of the following values is the closest approximation of the lattice energy of NaF: 682 kJ/mol, 794 kJ/mol, 924 kJ/mol, 1588 kJ/mol, or 3175 kJ/mol? Explain your answer.

## 7.E.6: 7.7: Molecular Structure and Polarity

Explain why the HOH molecule is bent, whereas the HBeH molecule is linear.

The placement of the two sets of unpaired electrons in water forces the bonds to assume a tetrahedral arrangement, and the resulting HOH molecule is bent. The HBeH molecule (in which Be has only two electrons to bond with the two electrons from the hydrogens) must have the electron pairs as far from one another as possible and is therefore linear.

What feature of a Lewis structure can be used to tell if a molecule's (or ion's) electron-pair geometry and molecular structure will be identical?

Explain the difference between electron-pair geometry and molecular structure.

Space must be provided for each pair of electrons whether they are in a bond or are present as lone pairs. Electron-pair geometry considers the placement of all electrons. Molecular structure considers only the bonding-pair geometry.

Why is the H–N–H angle in $NH_3$ smaller than the H–C–H bond angle in $CH_4$? Why is the H–N–H angle in $NH_4^+$ identical to the H–C–H bond angle in $CH_4$?

Explain how a molecule that contains polar bonds can be nonpolar.

As long as the polar bonds are compensated (for example. two identical atoms are found directly across the central atom from one another), the molecule can be nonpolar.

As a general rule, $MX_n$ molecules (where M represents a central atom and X represents terminal atoms; n = 2 – 5) are polar if there is one or more lone pairs of electrons on M. $NH_3$ (M = N, X = H, n = 3) is an example. There are two molecular structures with lone pairs that are exceptions to this rule. What are they?

Predict the electron pair geometry and the molecular structure of each of the following molecules or ions:

1. $SF_6$
2. $PCl_5$

3. (c) $BeH_2$
4. $CH_3^+$

1. Both the electron geometry and the molecular structure are octahedral.
2. Both the electron geometry and the molecular structure are trigonal bipyramid.
3. (c) Both the electron geometry and the molecular structure are linear.
4. Both the electron geometry and the molecular structure are trigonal planar.

Identify the electron pair geometry and the molecular structure of each of the following molecules or ions:

1. $IF_6^+$
2. $CF_4$
3. (c) $BF_3$
4. $SiF_5^-$
5. $BeCl_2$

What are the electron-pair geometry and the molecular structure of each of the following molecules or ions?

1. $ClF_5$
2. $ClO_2^-$
3. (c) $TeCl_4^{2-}$
4. $PCl_3$
5. $SeF_4$
6. $PH_2^-$

electron-pair geometry: octahedral, molecular structure: square pyramidal; electron-pair geometry: tetrahedral, molecular structure: bent; (c) electron-pair geometry: octahedral, molecular structure: square planar; electron-pair geometry: tetrahedral, molecular structure: trigonal pyramidal; electron-pair geometry: trigonal bypyramidal, molecular structure: seesaw; electron-pair geometry: tetrahedral, molecular structure: bent (109°)

Predict the electron pair geometry and the molecular structure of each of the following ions:

1. $H_3O^+$
2. $PCl_4^-$
3. (c) $SnCl_3^-$
4. $BrCl_4^-$
5. $ICl_3$
6. $XeF_4$
7. (g) $SF_2$

Identify the electron pair geometry and the molecular structure of each of the following molecules:

1. ClNO (N is the central atom)
2. $CS_2$
3. (c) $Cl_2CO$ (C is the central atom)
4. $Cl_2SO$ (S is the central atom)
5. $SO_2F_2$ (S is the central atom)
6. $XeO_2F_2$ (Xe is the central atom)
7. (g) $ClOF_2^+$ (Cl is the central atom)

electron-pair geometry: trigonal planar, molecular structure: bent (120°); electron-pair geometry: linear, molecular structure: linear; (c) electron-pair geometry: trigonal planar, molecular structure: trigonal planar; electron-pair geometry: tetrahedral, molecular structure: trigonal pyramidal; electron-pair geometry: tetrahedral, molecular structure: tetrahedral; electron-pair geometry: trigonal bipyramidal, molecular structure: seesaw; (g) electron-pair geometry: tetrahedral, molecular structure: trigonal pyramidal

Predict the electron pair geometry and the molecular structure of each of the following:

1. $IOF_5$ (I is the central atom)
2. $POCl_3$ (P is the central atom)
3. (c) $Cl_2SeO$ (Se is the central atom)

4. $ClSO^+$ (S is the central atom)
5. $F_2SO$ (S is the central atom)
6. $NO_2^-$
7. (g) $SiO_4^{4-}$

Which of the following molecules and ions contain polar bonds? Which of these molecules and ions have dipole moments?

1. $ClF_5$
2. $ClO_2^-$
3. (c) $TeCl_4^{2-}$
4. $PCl_3$
5. $SeF_4$
6. $PH_2^-$
7. (g) $XeF_2$

All of these molecules and ions contain polar bonds. Only $ClF_5$, $ClO_2^-$, $PCl_3$, $SeF_4$, and $PH_2^-$ have dipole moments.

Which of the molecules and ions in Exercise contain polar bonds? Which of these molecules and ions have dipole moments?

1. $H_3O^+$
2. $PCl_4^-$
3. (c) $SnCl_3^-$
4. $BrCl_4^-$
5. $ICl_3$
6. $XeF_4$
7. (g) $SF_2$

Which of the following molecules have dipole moments?

1. $CS_2$
2. $SeS_2$
3. (c) $CCl_2F_2$
4. $PCl_3$ (P is the central atom)
5. ClNO (N is the central atom)

$SeS_2$, $CCl_2F_2$, $PCl_3$, and ClNO all have dipole moments.

Identify the molecules with a dipole moment:

1. $SF_4$
2. $CF_4$
3. (c) $Cl_2CCBr_2$
4. $CH_3Cl$
5. $H_2CO$

The molecule $XF_3$ has a dipole moment. Is X boron or phosphorus?

P

The molecule $XCl_2$ has a dipole moment. Is X beryllium or sulfur?

Is the $Cl_2BBCl_2$ molecule polar or nonpolar?

nonpolar

There are three possible structures for $PCl_2F_3$ with phosphorus as the central atom. Draw them and discuss how measurements of dipole moments could help distinguish among them.

Describe the molecular structure around the indicated atom or atoms:

1. the sulfur atom in sulfuric acid, $H_2SO_4$ [$(HO)_2SO_2$]
2. the chlorine atom in chloric acid, $HClO_3$ [$HOClO_2$]
3. (c) the oxygen atom in hydrogen peroxide, HOOH

4. the nitrogen atom in nitric acid, $HNO_3$ [$HONO_2$]
5. the oxygen atom in the OH group in nitric acid, $HNO_3$ [$HONO_2$]
6. the central oxygen atom in the ozone molecule, $O_3$
7. (g) each of the carbon atoms in propyne, $CH_3CCH$
8. (h) the carbon atom in Freon, $CCl_2F_2$
9. (i) each of the carbon atoms in allene, $H_2CCCH_2$

tetrahedral; trigonal pyramidal; (c) bent (109°); trigonal planar; bent (109°); bent (109°); (g) $CH_3CCH$ tetrahedral, $CH_3CCH$ linear; (h) tetrahedral; (i) $H_2CCCH_2$ linear; $H_2CCCH_2$ trigonal planar

Draw the Lewis structures and predict the shape of each compound or ion:

1. $CO_2$
2. $NO_2^-$
3. (c) $SO_3$
4. $SO_3^{2-}$

A molecule with the formula $AB_2$, in which A and B represent different atoms, could have one of three different shapes. Sketch and name the three different shapes that this molecule might have. Give an example of a molecule or ion for each shape.

Three Lewis diagrams are shown. The first diagram shows the letter A single bonded to the left and right to the letter B. An example, “C O subscript 2,” and the term, “linear,” are written beside this diagram. The second diagram shows the letter A with two lone pairs of electrons, single bonded to the left and lower right to the letter B. An example, “H subscript 2 O,” and the term, “bent with an approximately 109 degree angle,” are written beside this diagram. The third diagram shows the letter A with one lone electron pair, single bonded to the left and lower right to the letter B. An example, “S O subscript 2,” and the term, “bent with an approximately 120 degree angle,” are written beside this diagram.

A molecule with the formula $AB_3$, in which A and B represent different atoms, could have one of three different shapes. Sketch and name the three different shapes that this molecule might have. Give an example of a molecule or ion that has each shape.

Draw the Lewis electron dot structures for these molecules, including resonance structures where appropriate:

1. $CS_3^{2-}$
2. $CS_2$
3. (c) CS

predict the molecular shapes for $CS_3^{2-}$ and $CS_2$ and explain how you arrived at your predictions

(a)

The figure shows three Lewis structures that are each surrounded by brackets and have a superscripted 2 negative sign. They are written with a double-headed arrow in between each diagram. The first of this trio has a carbon atom single bonded to two sulfur atoms, each of which has thee lone pairs of electrons, and double bonded to a third sulfur atom with two lone pairs of electrons. The second and third diagrams have the same atoms present, but each time the double bond moves between a different carbon and sulfur pair. The lone electron pairs also shift to correspond with the bond changes.

;

(b)

The Lewis structure shows a carbon atom double bonded to two sulfur atoms, each of which has two lone pairs of electrons.;

(c)

This diagram shows a carbon with one lone electron pair triple bonded to a sulfur with one lone electron pair.;

$CS_3^{2-}$ includes three regions of electron density (all are bonds with no lone pairs); the shape is trigonal planar; $CS_2$ has only two regions of electron density (all bonds with no lone pairs); the shape is linear

What is the molecular structure of the stable form of $FNO_2$? (N is the central atom.)

A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen. What is its molecular structure?

The Lewis structure is made from three units, but the atoms must be rearranged:

A Lewis structure is shown in which a carbon atom is single bonded to three hydrogen atoms and a second carbon atom. The second carbon is single bonded to a hydrogen atom and double bonded to a third carbon atom which is single bonded to two hydrogen atoms.

Use the simulation to perform the following exercises for a two-atom molecule:

1. Adjust the electronegativity value so the bond dipole is pointing toward B. Then determine what the electronegativity values must be to switch the dipole so that it points toward A.
2. With a partial positive charge on A, turn on the electric field and describe what happens.
3. (c) With a small partial negative charge on A, turn on the electric field and describe what happens.

4. Reset all, and then with a large partial negative charge on A, turn on the electric field and describe what happens.

Use the simulation to perform the following exercises for a real molecule. You may need to rotate the molecules in three dimensions to see certain dipoles.

1. Sketch the bond dipoles and molecular dipole (if any) for $O_3$. Explain your observations.
2. Look at the bond dipoles for $NH_3$. Use these dipoles to predict whether N or H is more electronegative.
3. (c) Predict whether there should be a molecular dipole for $NH_3$ and, if so, in which direction it will point. Check the molecular dipole box to test your hypothesis.

The molecular dipole points away from the hydrogen atoms.

Use the Molecule Shape simulator to build a molecule. Starting with the central atom, click on the double bond to add one double bond. Then add one single bond and one lone pair. Rotate the molecule to observe the complete geometry. Name the electron group geometry and molecular structure and predict the bond angle. Then click the check boxes at the bottom and right of the simulator to check your answers.

Use the Molecule Shape simulator to explore real molecules. On the Real Molecules tab, select $H_2O$. Switch between the "real" and "model" modes. Explain the difference observed.

The structures are very similar. In the model mode, each electron group occupies the same amount of space, so the bond angle is shown as 109.5°. In the "real" mode, the lone pairs are larger, causing the hydrogens to be compressed. This leads to the smaller angle of 104.5°.

Use the Molecule Shape simulator to explore real molecules. On the Real Molecules tab, select "model" mode and $S_2O$. What is the model bond angle? Explain whether the "real" bond angle should be larger or smaller than the ideal model angle.

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