Student Textbook page 171

1. (a) All chemical bonds are electrostatic forces that hold atoms together. An ionic bond is the force of attraction between oppositely charged ions (a cation and an anion). A covalent bond is the result of the balance of the forces of attraction and repulsion that act between the nuclei and electrons of two or more atoms. A polar covalent bond differs from a covalent bond in that one of the nuclei pulls the bonding electrons more strongly than the other nuclei.

(b) Ionic bond:

\[
\begin{array}{c}
\text{Na}^+ \\
\vdots \\
\text{Br}^-
\end{array}
\]

Covalent bond:

\[
\begin{array}{c}
\vdots \\
\text{Br} \\
\vdots
\end{array}
\]

Polar covalent bond:

\[
\begin{array}{c}
\vdots \\
\text{H} \\
\vdots \\
\text{Br}^-
\end{array}
\]

(c) The ionic compound, NaBr, is a white crystalline solid. The covalent compound, Br₂, is an orange liquid. The polar covalent compound, HBr, is a colourless gas. The polar and ionic compounds, NaBr and HBr, are soluble in water. The non-polar covalent compound, Br₂, does not mix with water. The NaBr solution conducts electricity, but the conductivity of the HBr solution is not obvious. The bond energy of NaBr is 743 kJ/mol, which is much larger than the bond energy of HBr (366 kJ/mol), which is again larger than the bond energy of Br₂ (193 kJ/mol).

2. Electronegativities: H = 2.20, N = 3.04, O = 3.44, S = 2.58, Cl = 3.16, Br = 2.96, I = 2.66

Order of decreasing electronegativity difference: O—H (ΔEN = 1.24) > S—O (ΔEN = 0.86) > H—Br (ΔEN = 0.76) > Cl—I (ΔEN = 0.50) > N—Cl (ΔEN = 0.12)

3. The melting point of a metal increases across a period. The melting points of the Group 2 (IIA) metals are higher than the melting points of the Group 1 (IA) metals. The increase in the number of valence electrons from s¹ to s² increases the strength of the metallic bond.

4. In Figure 4.10 on page 172 of the student textbook, when the material is hit with a hammer, the material breaks. The broken part has the same arrangement as the original block. Cleavage is the property that is shown. The type of bonding that accounts for this property is ionic bonding.

5. Tungsten and copper are both metals, and they have metallic properties (conduct electricity, have high melting points and boiling points, and are malleable). Tungsten has four unpaired electrons in the 5d orbitals and is thus a “harder” metal. Copper has no unpaired electrons in the 3d orbitals and is thus “softer” and more ductile. Tungsten has a higher resistance than copper and heats up more quickly when it carries an electric current. Since copper is more flexible than tungsten, it is more useful for wiring. Tungsten, which is tougher and can withstand repeated heating, is used to make filaments in light bulbs.

6. The stability of these compounds in a high-temperature environment needs to be determined. Both of these compounds are ionic. Therefore, the lattice energy is one property to check. The lattice energy of MgCl₂ is 2489 kJ/mol, and the lattice energy of MgO is 3933 kJ/mol. Thus, MgO tends to last longer than MgCl₂ in a high-temperature environment. A second property to check is the melting point. The melting point of MgO is much higher than the melting point of MgCl₂. Therefore, MgO would be the better choice.

7. (a) The strength of iron is due to the strength of its metallic bonds. Iron has two, perhaps three, electrons that are available for metallic bonding, so the bond is very strong. These electrons are also available for bonding with other elements, such as oxygen. Therefore, iron is quite reactive. Iron(III) oxide has a lower energy than the separate atoms, so there is a great tendency for it to form. Once formed, the electrons that used to hold the metal atoms together are held (bond) the iron atoms (ΔEN = 1.83) and oxygen atoms (ΔEN = 3.44). Since ΔEN = 1.61, the bonding between iron and oxygen is polar covalent. The intermolecular forces between the iron(III) oxide particles are weak compared with the metallic bonding in iron metal. The attractive force between the metal and the oxide is even weaker. Therefore, the rust flakes off the metal.

(b) The reasoning used in part (a) also applies to aluminum (ΔEN = 1.61) and oxygen (ΔEN = 3.44), except that the bonding between the aluminum and oxygen is ionic (ΔEN = 1.83). This implies that aluminum oxide (lattice energy = 15 916 kJ/mol) is harder [less likely] to break down than iron(III) oxide (lattice energy = 14 399 kJ/mol). The flakiness of iron(III) oxide could be due to the dipole of the iron(III) molecule not being able to interact with the cations of the iron metal. (The cations of the aluminum metal interact better with the aluminum cations of the oxide.) As well, the oxygen ion is the same size as the aluminum atom. The oxygen ion fits into the metal lattice structure better than it fits into the structure of iron(III) oxide.
1. (a)  
\[ \text{AsH}_3 \text{ has the VSEPR notation AX}_3E. \text{ The molecule is tristrial pyramidal. Since } \Delta EN \text{ for As} \rightarrow \text{H is essentially zero, the molecule is non-polar.} \]

(b)  
\[ \text{CH}_4 \text{CN has the VSEPR notation AX}_4. \text{ The molecule is tetrahedral. Since the CN group is polar, the molecule is polar.} \]

(c)  
\[ \text{Cl}_2 \text{O has the VSEPR notation AX}_1E_2. \text{ The Cl} \rightarrow \text{O bond is polar. Since the molecule is bent, it is polar.} \]

2. All polar molecules must have polar bonds, but not all non-polar molecules have non-polar bonds. If there is symmetry in a molecule such that the dipole of individual bonds cancel each other to give a net zero dipole, the molecule is non-polar. For example, CCl\(_4\) is a non-polar molecule that has four polar bonds.

3. \text{CH}_4 \text{ and CH}_4\text{O} \text{H have the same VSEPR notation: AX}_4. \text{ Both molecules have a tetrahedral electron pair arrangement and a tetrahedral shape. CH}_4 \text{ is non-polar because of the symmetry of the molecule. CH}_4\text{O} \text{H is polar because the polarity of the O} \rightarrow \text{H bond is different from the polarity of the three C} \rightarrow \text{H bonds.} \]

4. The combination reaction can be represented by the following equation.
\[ \text{BF}_3 + \text{NH}_3 \rightarrow \text{F} \text{B} \text{N}_3 \]

5.  
\[ \text{Since } \Delta EN = 3.98 - 2.58 = 1.40, \text{ the } S \rightarrow \text{F bond is polar. The VSEPR notation for the molecule is AX}_4E_2. \text{ The four electron pairs are arranged in a tetrahedral shape, and the shape of the molecule is bent. Thus, the } S \rightarrow \text{F dipoles do not cancel each other, and the molecule is polar.} \]

6.  
\[ \text{The VSEPR notation for the ion is AX}_4E_2. \text{ The electron pairs are arranged in a tetrahedral shape, and the shape of the ion is also tetrahedral.} \]

7. The factors that affect the shape of a molecule are the number of bonding pairs of electrons and the number of non-bonding pairs of electrons around the central atom of the molecule. The factors that affect the polarity of a molecule are the polarity of the individual bonds and the shape of the molecule. The VSEPR theory is based on minimizing the repulsive force between pairs of electrons. If there are two electron pairs around a central atom, these electron pairs are on opposite sides of the central atom. Electron pairs tend to stay as far apart from one another as possible to minimize the repulsion. Thus, the number of bonding and non-bonding pairs around the central atom determines the shape of a molecule. A molecule may be non-polar if the individual bonds are polar and the shape has symmetry. Consider the linear molecule of beryllium and bromine. Each Be–Br bond is polar. Since the dipoles that are associated with the polar bond are pointing in opposite directions, they cancel each other. As a result, the molecule becomes non-polar. Therefore, the shape is very important in determining the polarity of a molecule.
1. The electronegativity difference of 0.76 between H (2.0) and Br (2.96) suggests that the H—Br bond is polar covalent. The intramolecular force is the simultaneous attraction on a pair of electrons by the hydrogen and bromine nuclei. Since bromine has a greater electronegativity, the bromine atom pulls the pair of electrons closer to itself, giving the molecule a permanent dipole. In HBBr3, the positive end of the molecule lines up with the negative end of another molecule, thus setting up an electrostatic attraction (dipole-dipole force) between the two molecules. Therefore, the intermolecular forces in HBBr3 are dipole-dipole and dispersion forces.

2. The boiling point of 150°C indicates that the intermolecular forces in H2O are greater than the intermolecular forces in H2O2 (boiling point 100°C). Both substances can form hydrogen bonds. Water has a total of 10 electrons, while H2O2 has 18. Therefore, the larger H2O2 molecules have greater dispersion forces between them. The stronger dispersion forces in H2O2 most likely account for the difference in the boiling points.

3. (a) non-polar molecular solid  
(b) ionic solid  
(c) metallic solid  
(d) network solid

4. In NH4Cl, the bonding between NH4+ and Cl− is ionic. In the NH4+ ion, three of the H—N bonds are polar covalent bonds and one is a covalent bond.

5. The relatively low melting point of 140°C suggests that the solid is a molecular solid. The solubility of the solid indicates that it is polar. Therefore, the solid is likely a polar covalent solid. To confirm, test the solution of the solid for conductivity. Low or no significant conductivity would confirm a polar covalent solid.

6. The VSEPR notation for the molecule is AX4E. The electron pairs are arranged in a tetrahedral shape, and the shape of the molecule is trigonal pyramidal since the P—Cl and P—F bonds are polar, and the symmetry of the molecule does not allow these dipoles to cancel, the molecule is polar.

7. The VSEPR notation for the molecule is AX4E. The arrangement of the electron pairs about the germanium atom is tetrahedral, and the shape of the molecule is also tetrahedral. The Ge—H bond is essentially non-polar (ΔEN = 0.19). The molecule is symmetrical, so the polarities of the four bonds cancel each other. Therefore, the molecule is non-polar.

8. The electronegativity differences of the four halides decrease in this order: Cs—I (ΔEN = 3.19) > Cs—Br (ΔEN = 2.17) > Cs—Cl (ΔEN = 2.37) > Cs—F. Since the electrostatic force of attraction varies inversely as the square of the distance between the charges, the electrostatic force between a cation and an anion decreases for larger ions. As a result, the force of attraction between the Cs+ ion and the halide ion decreases in the following order: CsF > CsCl > CsBr > CsI. To melt an ionic solid, the bond between the cation and the anion needs to be broken. Therefore, the melting points of the halides, from highest to lowest, are in this order: CsF > CsCl > CsBr > CsI.

9. A canoe often comes into contact with rocks when, for example, going aground at a portage or running rapids. The material for making a canoe has to be strong enough to resist possible punctures as a result of this contact. As well, a canoe has to be light enough to be carried. A material that has the strength of steel and the weight of feathers is ideal. KEVLAR® is used to make canoes because it is produced from very light fibres, which can withstand the impact of a bullet without shredding. KEVLAR® spreads any impact over a wide area, so that no part of the canoe punctures or breaks.

10. Diamond is a network solid. Each carbon atom of diamond is joined to four other carbon atoms in the shape of a tetrahedron, so each diamond is a single molecule.

(a) An electrical insulator is a poor conductor of electricity. It is used to suppress the flow of electricity. Diamond does not conduct electricity because all its electrons are tied up in the covalent bonds between carbon atoms. Thus, diamond would make an excellent insulator. Diamond is not flexible, however, so it cannot be wrapped around wires. Currently, industrial diamonds are very small and are only used for abrasives.

(b) To create an optical fibre, the core of the fibre must have a very high refractive index. The refractive index of glass fibre is 1.512. The refractive index of diamond is 2.417. This optical property of diamond would make it an ideal core fibre. However, no technology has yet been developed to grow or machine diamond into long thin threads. Thus, diamond is not yet a suitable material for making cables for fibre optics.
Chapter 4 Review Answers

Student Textbook pages 209-211

Answers to Knowledge/Understanding Questions

1. An octet in the valence level of atoms is a stable arrangement—an arrangement that corresponds to a minimum total energy for the system. For example, noble gases are chemically unreactive, and each atom of a noble gas element (other than He) has eight valence electrons.

2. All the halogen hydrides are polar. They experience dispersion forces as well as dipole-dipole forces. The dispersion forces are greatest in HI and smallest in HF; because the HI molecule is the largest and contains the most electrons. \( \Delta EN \) is greatest in HF and smallest in HI, so the dipole-dipole forces are greatest in HF and smallest in HI. However, \( \Delta EN \) is small for the halogens, so the difference in the strength of the dipole-dipole forces is not as significant as the difference in the strength of the dispersion forces. Therefore, the boiling point should increase from HF to HI. In HF, though, the relative sizes of the hydrogen atom and fluorine atom allow hydrogen bonding to occur in this molecule. Hydrogen bonding does not occur in the other hydrogen halides. As a result, HF has the strongest intermolecular forces and its boiling point is higher than the boiling points of the other hydrogen halides.

3. (a) CHBr₃
   \[
   \begin{array}{c}
   \vdots \\
   \vdots \\
   \cdots \\
   \end{array}
   \]
   
   H --- C --- Br
   \[
   \begin{array}{c}
   \vdots \\
   \vdots \\
   \cdots \\
   \end{array}
   \]
   
   (b) HS⁻
   \[
   \begin{array}{c}
   \vdots \\
   \vdots \\
   \cdots \\
   \end{array}
   \]

4. An induced dipole exists when a non-polar molecule comes close to an ion or a polar molecule. As soon as the ion or polar molecule moves away, the induced dipole disappears. A permanent dipole is produced by the difference in the electronegativities of the atoms within a molecule. The dipole does not disappear, and its existence is independent of the presence of other molecules.

5. The intermolecular forces of non-polar molecules are dispersion forces. The strength of these forces depends on the number of electrons and the shape and size of the molecule. Oil is a mixture of large organic molecules, which have many car-
bon atoms and a large number of electrons. As a result, the
dispersion forces between the molecules are very strong, even
stronger than the hydrogen bonding in water. Therefore,
cooking oil has a higher boiling point than water.

6. (a) dispersion forces and dipole-dipole forces
   (b) ionic bond
   (c) metallic bond
   (d) dispersion forces

7. Cesium is a metallic solid, and sulfur is a molecular solid.
The other two elements are gases. Thus, cesium has the
highest boiling point, and sulfur has the next highest boil-
ing point. Krypton has more electrons (36) in its atom
than oxygen gas (32). The dispersion forces of krypton are
greater than the dispersion forces of O₂. Thus, the
order of decreasing boiling points is cesium > sulfur >
krypton > oxygen. This order agrees with the measured
boiling points: Cs (944 K), S (717.8 K), Kr (119.9 K),
and O₂ (90.2 K).

8. H₂O and NH₃ have the same number of electrons in their
molecules. ΔEV is greater for O—H (1.24) than for N—
H (0.84). Therefore, the O—H bond is more polar and
forms stronger hydrogen bonds. Thus, hydrogen bonding
is stronger in water than in ammonia.

9. O₂^−

10. The term “crystal” describes something macroscopic—
something we can see. The unit cell of a crystal is micro-
scopic. It cannot be seen with the unaided eye or even with
a microscope. X-ray diffraction is used to make a model of
the crystal, from which the unit cell can be determined.
The unit cell is the structure that, if repeated many times,
makes up the crystal.

11. If an ionic crystal is broken, each of the broken pieces is
still made up of repeating unit cells of the original crystal.
Thus, on a microscopic scale, each broken piece has the
same shape as the original crystal. Glass is an amorphous
solid. If glass is shattered, the silicate tetrahedra do not
arrange themselves identically in the broken pieces.

12. Viscosity is the ability of a liquid to resist flow. In flow-
ing, the molecules must move past one another. A liquid with
low viscosity thus experiences weaker intermolecular forces
between liquid molecules. Surface tension arises due to an
imbalance of forces between molecules at the surface of the
liquid and the forces between the liquid molecules in the
interior. Regardless of the strength of the intermolecular
forces, this net force always exists at the surface of the liq-
uid. Therefore, a liquid with low viscosity can have a high
surface tension.

13. SiO₂ is a network solid, so it has the highest boiling point.
K is a metallic solid. The boiling point of K is therefore rela-
tively lower. C₂H₆ and C₂H₅OH are covalent solids, which
have lower boiling points than SiO₂ and K. C₂H₅OH is
polar, and it may have hydrogen bonding in its solid form.
C₂H₆ is non-polar and will not form hydrogen bonds.
C₂H₅OH probably has a higher boiling point than C₂H₆.
Thus, the order of increasing boiling points is C₂H₅OH <
C₂H₆ < K < SiO₂. Actual boiling point data support
this prediction: C₂H₅OH (351 K), SiO₂ (2603 K), C₂H₅OH
(231 K), K (1059 K).

14. The molecular shapes of CCl₄, CH₃Cl, and CHCl₃ are all
tetrahedral. CCl₄ is non-polar, and CH₃Cl and CHCl₃ are both
polar.

15. Dipole-dipole attractions are intermolecular forces between
two polar molecules. An ionic bond is the attraction between
two ions: one positively charged and the other negatively
charged. Ionic bonds are usually thought of as intramole-
cular forces. In an ionic crystal, however, it is impossible to
differentiate between the intermolecular and intramolecular
forces.

16. (a) The shape of the NOCl molecule is bent.

   ● Cl — N — Cl

   (b) The shape of the AlF₆^3− ion is octahedral.

   [Al — F — F — F — F — F]³−

   (c) The shape of the XeO₃ molecule is trigonal pyramidal.

   ● O — Xe — O

17. (a) Cl₂

   ● Cl — Cl

   (b) BrCl

   O + Br — Cl
18. Substances that have the following VSEPR notations have the same name for the molecular shape and the electron group arrangement: AX₄ (linear), AX₃ (trigonal planar), AX₅ (tetrahedral), AX₆ (trigonal bipyramidal), and AX₇ (octahedral). In other words, a molecule that does not have any lone pairs of electrons has the same name for its molecular shape and electron group arrangement.

19. If the polar bonds in a molecule are arranged around the central atom in such a way that the dipoles can cancel each other, then the molecule is non-polar. Non-polar molecules with polar covalent bonds include binary (two-element) molecules with no electron lone pairs. These molecules have the VSEPR notation AXₙ, where n = 2, 3, 4, 5, and 6. Linear molecules that have the VSEPR notation AX₄F₂ (XeF₂, for example) and square planar molecules that have the VSEPR notation AX₄F₂ (XeF₄, for example) are also non-polar.

Answers to Inquiry Questions

20. 

When the two fluorine atoms are bonded to the central atom along the vertical axis of the trigonal bipyramidal arrangement, the dipoles cancel and the molecule is non-polar. When one fluorine atom and one chlorine atom are bonded to the central atom along the vertical axis, however, the molecule is polar.

21. At temperatures above 4°C, water cools. The molecules move closer together, so the density of the water increases. This increase in density continues until the temperature of the water is 4°C, when the open hexagonal structure (due to hydrogen bonding) stops further contraction in the volume of the water. When the water cools further, below 4°C, the water molecules form more of the open hexagons that are found in ice. The volume begins to increase, so the density begins to decrease.

22. The electronegativity difference between the elements in a compound determines the type of bonding present. ΔEN of NaCl is 2.23, and the compound is ionic. ΔEN of CaCl₂ is 2.16. The compound is also ionic, but it has less ionic character than NaCl. ΔEN of AlCl₃ is 1.55, and the compound is polar covalent. AB⁺ and Cl⁻ ions do not form the lattice structure that is characteristic of ionic solids. The electrostatic force of attraction is greater in NaCl, so NaCl requires the most energy to melt (to break its ionic lattice). Therefore, NaCl has the highest melting point. Since AlCl₃ is more of a covalent compound, the melting point of AlCl₃ is the lower.

23. (a) The decomposition of Au₂S₅ occurs at a lower temperature than the decomposition of Au₂S. Thus, Au₂S₅ has greater ionic character. That is, more energy is required to break the ionic lattice of Au₂S.

(b) The decomposition of the sulfide is the reverse of the formation reaction. Since the decomposition temperature of Au₂S₅ is higher than the decomposition temperature of Au₂S, Au₂S₅ is thermally a more stable compound than Au₂S. Since stability is related to lattice energy, the lattice energy of Au₂S₅ is greater than the lattice energy of Au₂S.

24. The ground state electron configuration of Xe is [Kr] 4d⁰ 5s² 5p⁶. If six fluorine atoms are bonded to Xe, the resulting molecule will have the VSEPR notation AX₆. E in total, there will be one non-bonding pair and six bonding pairs of electrons surrounding the central atom. Thus, the shape is not octahedral. This molecule has the xenon atom in the centre, with a lone pair of electrons directly above it. The six fluorine atoms are arranged at the three points of each of two triangles above and below the xenon atom.

25. From the electronegativity values, the H—F bond (ΔEN = 1.78) is more polar than the O—H bond (ΔEN = 1.24). This suggests that hydrogen fluoride may have a higher boiling point than water. However, hydrogen bonding is much stronger in water than in hydrogen fluoride. The bent shape of the water molecule allows for two O—H bonds to form hydrogen bonds. Each water molecule can be linked to three other molecules, but each hydrogen fluoride molecule can be linked to only two other molecules. As a result, water should have a higher boiling point. From Figure 4.16 on page 193 of the student textbook, the boiling point of water is 373 K and the boiling point of hydrogen fluoride is 293 K.

Answers to Communication Questions

26. The ground state electron configuration of S is [Ne]3s² 3p⁴. To form four bonds, as in SF₄, one electron of S has to be excited to a 3d orbital to give [Ne]3s² 3p⁴ 3d⁴. As a result, the electron arrangement around the central S atom involves four bonding pairs and one non-bonding pair of electrons. The VSEPR notation for this molecule is AX₄E. The electron
- KEVLAR® gloves and vests protect workers in hazardous situations.
- Foam cups are inexpensive and convenient for keeping drinks warm or cold.

Benefits versus risks:
- The benefits of KEVLAR® products outweigh the risks, since the quantity in use remains quite small.
- The benefits of foam cups do not outweigh the risks, since people can use re-usable cups instead.

33. Here are some examples of how intermolecular forces influence the weather:
- If you are standing at the top of a ski hill and snow just begins to fall, you may feel that the temperature is rising. As water sublimes to form snowflakes, hydrogen bonds (intermolecular forces) are formed. The making of bonds produces energy, so you feel that the air temperature is higher.

- On a very hot summer day, if you get out of a swimming pool and lay on a hot concrete patio beside the pool, you may find yourself shivering in the heat. As the water on your body evaporates, the hydrogen bonds in the water need energy to break. The energy is absorbed from your body, so your skin feels cool.

- On a cold winter morning, you may notice frost on the classroom windows. During the previous night, the glass window became very cold. As water vapour passed over the window, it cooled and the water molecules were brought closer together. The hydrogen bonds then formed to give small crystals of ice. The frost "painting" on the window was formed by water vapour subliming to the solid.
Chapter 4

Structures and Properties of Substances

Solutions for Practice Problems
Student Textbook pages 165–166

1. Problem
   Write electron configurations for the following:
   (a) Li⁺
   (b) Ca²⁺
   (c) Br⁻
   (d) O²⁻

   Solution
   First determine the atomic number of the ion, and then add or subtract electrons to match the charge indicated on the ion. Use the aufbau process to write the complete electron configuration of the ion.
   (a) Li⁺: Z = 3, subtract one electron for +1 charge; 1s²
   (b) Ca²⁺: Z = 20, subtract two electrons for +2 charge; 1s²2s²2p⁶3s²3p⁶
   (c) Br⁻: Z = 35, add one electron for charge of −1; 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶
   (d) O²⁻: Z = 8, add two electrons for charge of −2; 1s²2s²2p⁶

   Check Your Solution
   A check of the sequence of orbitals shows that the aufbau process was followed correctly. The number of electrons in each electron configuration matches the difference of Z − (net charge on ion). The answers are correct.

2. Problem
   Draw Lewis structures for the chemical species in question 1.

   Solution
   In a Lewis structure, the nucleus and the inner core electrons are represented by the symbol for the element. The electrons in the valence shell are drawn around the symbol.
   (a) [Li⁺]
   (b) [Ca²⁺]
   (c) [Br⁻]
   (d) [O²⁻]

   Check Your Solution
   Each positive ion shows no electrons. This is expected, since electrons were removed from the valence shell. Each negative ion has 8 valence electrons. This is expected for all ions except helium-like ions.

3. Problem
   Draw orbital diagrams and Lewis structures to show how the following pairs of elements can combine. In each case, write the chemical formula for the product.

   Solution
   In each example, electrons are transferred from the metal to the non-metal so that each ion attains a noble gas electron configuration.
(a) Lewis structure:

\[
\text{Li}^+ \quad \text{S}^2- \quad \rightarrow \quad \text{Li}^+ \quad \text{S}^2-
\]

(b) Lewis structure:

\[
\text{Ca}^{2+} \quad \text{Cl}^- \quad \rightarrow \quad \text{Ca}^{2+} \quad \text{Cl}^-
\]

(c) Lewis structure:

\[
\text{K}^+ \quad \text{Cl}^- \quad \rightarrow \quad \text{K}^+ \quad \text{Cl}^-
\]

(d) Lewis structure:

\[
\text{Na}^+ \quad \text{N}^3- \quad \rightarrow \quad \text{Na}^+ \quad \text{N}^3-
\]

Check Your Solution

In each case the ions have a noble gas electron configuration that takes into account the charge on the ion. The answers are correct.
4. Problem
To which main group on the periodic table does X belong?
(a) MgX
(b) X₂SO₄
(c) X₂O₃
(d) XCO₃

Solution
Since these compounds are all ionic, the zero sum rule applies, and the sum of the charges on the ions must equal zero. Once you know the charge on the ion, it can be related to the characteristic charge on ions for the main group elements.
(a) The charge on Mg is +2. Therefore, (+2) + x = 0, and x = −2. X²⁻ will be in group 16.
(b) The charge on SO₄ is −2. Therefore, 2(x) + (−2) = 0, and x = +1. X⁺ will be in group 1.
(c) The charge on O is −2. Therefore, 2(x) + 3(−2) = 0, and x = +3. X³⁺ will be in group 3.
(d) The charge on CO₃ is −2. Therefore, x + (−2) = 0, and x = +2. X²⁺ will be in group 2.

Check Your Solution
In each case the zero sum rule holds. The charge on each X ion has been determined correctly.

Solutions for Practice Problems
Student Textbook pages 169–170

5. Problem
List the following compounds in order of decreasing bond energy: H — Br, H — I, H — Cl. Use Appendix E to verify your answer.

Solution
Since all bonds are electrical in nature, a greater difference in charge between two atoms indicates a stronger attraction between the atoms, and therefore, a stronger bond. The difference in charge can be estimated using the property of electronegativity. A greater difference in electronegativity between the two atoms in the bond indicates a greater attraction between the atoms. (A greater electronegativity also indicates a greater ionic character of the bond.) For each bond, measure the difference in electronegativity, (ΔEN). The greater ΔEN, the stronger the bond.
H — Cl: ΔEN = 3.16 − 2.20 = 0.96
H — Br: ΔEN = 2.96 − 2.20 = 0.76
H — I: ΔEN = 2.66 − 2.20 = 0.46
Based upon these calculated values for ΔEN, H — Cl has greatest bond energy and H — I has the least bond energy. From Appendix E, the bond energies confirm this conclusion.
H — Cl: bond energy = 432 kJ/mol
H — Br: bond energy = 366 kJ/mol
H — I: bond energy = 298 kJ/mol

Check Your Solution
Since the measured value of bond energy matches the prediction using ΔEN, this is a valid method to estimate bond strength.
6. Problem
Rank the following compounds in sequence from lowest melting point to highest melting point, and give reasons for your decisions: AsBr$_3$, KBr, CaBr$_2$.

**Solution**
When compounds that are composed of a metal ion and a non-metal ion melt, the two types of ions are separated. The greater the attraction between the ions, the more heat energy will be required to melt the compound. The attraction between ions can be estimated by calculating the difference in electronegativity ($\Delta EN$) between the atoms (ions) in each bond.

Based upon this $\Delta EN$, the melting points will be expected to increase from AsBr$_3$ < CaBr$_2$ < KBr.

**Check Your Answer**
The actual melting points of these three compounds can be found in a reference text such as the Handbook of Chemistry and Physics. A check of this reference confirms that the reasoning was correct.

7. Problem
From their position in the periodic table, predict which bond in the following groups is the most polar. Verify your predictions by calculating the $\Delta EN$.

(a) C—H, Si—H, Ge—H
(b) Sn—Br, Sn—I, Sn—F
(c) C—O, C—H, C—N

**Solution**
A greater distance between the positions of two elements in the periodic table indicates a greater expected difference in the atoms’ ability to attract electrons. Therefore, a greater distance between elements in the periodic table indicates that a more polar bond forms between the elements. Using these criteria, the most polar bond in each group is predicted to be:

(a) C—H  (b) Sn—F  (c) C—O

Using the electronegativities given on the periodic table, the following $\Delta EN$ are found:

(a) C—H (0.35), Si—H (0.30), Ge—H (0.19)
(b) Sn—Br (1.00), Sn—I (0.70), Sn—F (2.01)
(c) C—O (0.89), C—H (0.35), C—N (0.49)

Our predictions were accurate.

**Check Your Solution**
Generally, the statement that “the further apart two elements are found in the periodic table, the more polar the bond between them,” is reliable. The use of $\Delta EN$, however, is a more accurate measure of polarity.

8. Problem
Classify the bonding in each of the following as covalent (non-polar), polar covalent, or ionic. Afterwards, rank the polar covalent compounds in order of increasing polarity.

(a) S$_8$  (b) RbCl  (c) PF$_3$
(d) SCl$_2$  (e) F$_2$  (f) SF$_2$
Solution
Calculate the difference in electronegativity, $\Delta EN$, between the elements in each compound. Apply the criteria that for a mostly covalent bond, $\Delta EN$ is < 0.5; for a polar covalent bond, $\Delta EN$ is between 0.5 and 1.7; and for an ionic bond $\Delta EN$ > 1.7.

<table>
<thead>
<tr>
<th></th>
<th>$\Delta EN$</th>
<th>Type of bonding</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td>S₂</td>
<td>0 - 0 = 0</td>
</tr>
<tr>
<td>(b)</td>
<td>RbCl</td>
<td>3.16 - 0.82 = 2.34</td>
</tr>
<tr>
<td>(c)</td>
<td>PF₃</td>
<td>3.98 - 2.19 = 1.79</td>
</tr>
<tr>
<td>(d)</td>
<td>SCl₂</td>
<td>3.16 - 2.58 = 0.58</td>
</tr>
<tr>
<td>(e)</td>
<td>F₂</td>
<td>0 - 0 = 0</td>
</tr>
<tr>
<td>(f)</td>
<td>SF₂</td>
<td>3.98 - 2.58 = 1.40</td>
</tr>
</tbody>
</table>

polar covalent compounds: SCl₂ < SF₂ < PF₃

Check Your Solution
The type of bond predicted correlates with the relative positions of the elements in the periodic table. These results are good indicators of the polarity of the bonds.

Solutions for Practice Problems
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9. Problem
Draw Lewis structures for each of the following molecules or ions:
(a) NH₃  (b) CH₄  (c) CF₄  (d) AsH₃  
(e) BrO⁻  (f) H₂S  (g) H₂O₂  (h) ClNO

Solution
(a) The molecular formula, NH₃, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

Step 1 Nitrogen has the lower electronegativity in the molecule and will be the central atom.

\[
\begin{array}{c}
  \text{H} \quad \text{N} \quad \text{H} \\
  \text{H}
\end{array}
\]

Step 2 Determine the total number of valence electrons.
(1 atom N \times 5 e^-/N) + (3 atoms H \times 1 e^-/H) = 8 e^-
Determine the total number of electrons required for a noble gas configuration.
(1 atom N \times 8 e^-/atom) + (3 atoms N \times 2 e^-/atom) = 14 e^-
To find the number of shared electrons, subtract the first total from the second.
14 e^- - 8 e^- = 6 e^-
Divide the number of shared electrons by two to obtain the number of bonds.
6 e^- / 2 = 3 covalent bonds

Step 3 Determine the number of non-bonding electrons by subtracting the number of shared electrons from the total number of valence electrons:
8 valence e^- - 6 bonding e^- = 2 e^- , or 1 lone pair

\[
\begin{array}{c}
  \text{H} \quad \text{N} \quad \text{H} \\
  \text{H}
\end{array}
\]
(b) The molecular formula, CH₄, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Carbon has the lower electronegativity in the molecule and will be the central atom.

```
   H
  H—C—H
  H
```

**Step 2** Find the total number of valence electrons.

\[
(1 \text{ atom C } \times 4 \text{ e}^-/\text{C}) + (4 \text{ atoms H } \times 1 \text{ e}^-/\text{H}) = 8 \text{ e}^-
\]

Determine the total number of electrons required for a noble gas configuration.

\[
(1 \text{ atom N } \times 8 \text{ e}^-/\text{atom}) + (4 \text{ atoms H } \times 2 \text{ e}^-/\text{atom}) = 16 \text{ e}^-
\]

Determine the total number of electrons used in bonding.

\[
16 \text{ e}^- - 8 \text{ e}^- = 8 \text{ e}^-, \text{ or 4 covalent bonds}
\]

**Step 3** Determine the number of non-bonding electrons.

\[
8 \text{ valence e}^- - 8 \text{ bonding e}^- = 0 \text{ non-bonding e}^-
\]

(c) The molecular formula, CF₄, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Carbon has the lower electronegativity in the molecule and will be the central atom.

```
  F
 F—C—F
 F
```

**Step 2** Calculate the total number of valence electrons.

\[
(1 \text{ atom C } \times 4 \text{ e}^-/\text{C}) + (4 \text{ atoms F } \times 7 \text{ e}^-/\text{F}) = 32 \text{ e}^-
\]

Determine the total number of electrons required for a noble gas configuration.

\[
5 \text{ atoms } \times 8 \text{ e}^-/\text{atom} = 40 \text{ e}^-
\]

Determine the total number of electrons used in bonding.

\[
40 \text{ e}^- - 32 \text{ e}^- = 8 \text{ e}^-, \text{ or 4 covalent bonds}
\]

**Step 3** Calculate the number of non-bonding electrons.

\[
32 \text{ valence e}^- - 8 \text{ bonding e}^- = 24 \text{ e}^-, \text{ or 12 lone pairs}
\]

(d) The molecular formula, AsH₃, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Arsenic has the lower electronegativity in the molecule and will be the central atom.

```
   H
  H—As—H
  H
```
**Step 2** Calculate the total number of valence electrons.

(1 atom As × 5 e⁻/N) + (3 atoms H × 1 e⁻/H) = 8 e⁻

Determine the total number of electrons required for a noble gas configuration.

(1 atom As × 8 e⁻/atom) + (3 atoms H × 2 e⁻/atom) = 14 e⁻

Determine the total number of electrons used in bonding.

14 e⁻ − 8 e⁻ = 6 e⁻, or 3 covalent bonds

**Step 3** Determine the number of non-bonding electrons.

8 valence e⁻ − 6 bonding e⁻ = 2 e⁻, or 1 lone pair

\[ \text{H—As—H} \]

(e) The molecular formula, BrO⁻, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Br → O

**Step 2** Calculate the total number of valence electrons.

(1 atom O × 6 e⁻/O) + (1 atom Br × 7 e⁻/Br) + (1 e⁻ for charge on ion) = 14 e⁻

Determine the total number of electrons required for a noble gas configuration.

2 atoms × 8 e⁻/atom = 16 e⁻

Determine the total number of electrons used in bonding.

16 e⁻ − 14 e⁻ = 2 e⁻, or 1 covalent bond

**Step 3** Determine the number of non-bonding electrons.

14 valence e⁻ − 2 bonding e⁻ = 12 e⁻, or 6 lone pairs

\[ \text{Br—O}^- \]

(f) The molecular formula, H₂S, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Sulfur has the lower electronegativity in the molecule and will be the central atom.

\[ \text{H—S—H} \]

**Step 2** Calculate the total number of valence electrons.

(1 atom S × 6 e⁻/S) + (2 atoms H × 1 e⁻/H) = 8 e⁻

Determine the total number of electrons required for a noble gas configuration.

(1 atom S × 8 e⁻/atom) + (2 atoms H × 2 e⁻/atom) = 12 e⁻

Determine the total number of electrons used in bonding.

12 e⁻ − 8 e⁻ = 4 e⁻, or 2 covalent bonds

**Step 3** Determine the number of non-bonding electrons.

8 valence e⁻ − 4 bonding e⁻ = 4 e⁻, or 2 lone pairs

\[ \text{H—S—H} \]

(g) The molecular formula, H₂O₂, gives the number of each kind of atom. Since there is no single central atom, steps 1–3 are modified.

**Step 1** In this example, the two hydrogen atoms are at either end of the molecule, and the oxygen atoms are bonded.

\[ \text{H—O—O—H} \]

**Step 2** Calculate the total number of valence electrons.

(2 O atoms × 6 e⁻/O) + (2 H atoms × 1 e⁻/H atom) = 14 e⁻

Determine the total number of electrons required for a noble gas configuration.

(2 O atoms × 8 e⁻/O) + (2 H atoms × 2 e⁻/H) = 20 e⁻
Determine the total number of electrons that are used in bonding.

$$20\ e^- - 14\ e^- = 6\ e^-,$$ or 3 covalent bonds

**Step 3** Determine the number of non-bonding electrons.

14 valence e$^- - 6$ bonding e$^- = 8$ non-bonding e$^-$, or 4 lone pairs

Hydrogen has no lone pair, and each oxygen atom has 2 lone pairs.

\[\text{H} \quad \text{O} \quad \text{O} \quad \text{H}\]

(b) The molecular formula, CINO, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Nitrogen is the least electronegative element in the molecule and will be the central atom.

\[\text{Cl} \quad \text{N} \quad \text{O}\]

**Step 2** Calculate the total number of valence electrons.

\[
(1\ \text{atom}\ N \times 5\ e^-/N) + (1\ \text{atom}\ O \times 6\ e^-/O) + \\
(1\ \text{atom}\ Cl \times 7\ e^-/Cl) = 18\ e^-
\]

Determine the total number of electrons required for a noble gas configuration.

3 atoms $\times 8\ e^-/atom = 24\ e^-$

Determine the total number of electrons used in bonding.

$$24\ e^- - 18\ e^- = 6\ e^-,$$ or 3 covalent bonds

**Step 3** Determine the number of non-bonding electrons.

18 valence e$^- - 6$ bonding e$^- = 12\ e^-$, or 6 lone pairs

Add an additional bonding pair between the nitrogen and oxygen atom to satisfy the normal bonding capacity of oxygen.

\[\text{Cl} \quad \text{N} \equiv \text{O}\]

**Check Your Solution**

For each structure, each atom has a noble gas electron configuration in its valence shell. These are acceptable Lewis structures.

**10. Problem**

Draw Lewis structures for each of the following ions. (Note: Consider resonance structures.)

(a) $\text{CO}_3^{2-}$

(b) $\text{NO}^+$

(c) $\text{ClO}_3^-$

(d) $\text{SO}_3^{2-}$

**Solution**

(a) The molecular formula, $\text{CO}_3^{2-}$, gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1** Carbon has the lower electronegativity in the molecule and will be the central atom.

\[
\begin{array}{c}
\text{O} \\
\text{C} \\
\text{O}
\end{array}
\]

**Step 2** Calculate the total number of valence electrons.

\[
(1\ \text{atom}\ C \times 4\ e^-/C) + (3\ \text{atoms}\ O \times 6\ e^-/O) + \\
(2\ e^- \text{for charge on ion}) = 24\ e^-
\]

Determine the total number of electrons required for a noble gas configuration.

4 atoms $\times 8\ e^-/atom = 32\ e^-$
Determine the total number of electrons used in bonding.

\[ 32 \text{ e}^- - 24 \text{ e}^- = 8 \text{ e}^-, \text{ or 4 covalent bonds} \]

**Step 3**

Determine the number of non-bonding electrons.

24 valence \( \text{e}^- \) – 8 bonding \( \text{e}^- \) = 16 \( \text{e}^- \), or 8 lone pairs

More than one Lewis structure can be shown. Resonance structures must exist.

\[ \begin{align*}
\text{O} & \text{C} \text{:} \text{O} \\leftrightarrow \\text{:} \text{O} \text{C} \\leftrightarrow \text{O} \text{:} \text{C} \\leftrightarrow \text{O} \\
\text{O} & \text{C} \text{:} \text{O} \\
\text{O} & \text{C} \text{:} \text{O} \\
\text{O} & \text{C} \text{:} \text{O} \\
\end{align*} \]

(b) **Step 1**

The molecular formula, NO\(^+\), gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

\[ \text{N} \rightarrow \text{O} \]

**Step 2**

Calculate the total number of valence electrons.

\[ (1 \text{ atom } \text{O} \times 6 \text{ e}^-/\text{O}) + (1 \text{ atom N} \times 5 \text{ e}^-/\text{N}) - (1 \text{ e}^- \text{ for charge on ion}) = 10 \text{ e}^- \]

Determine the total number of electrons required for a noble gas configuration.

\[ 2 \text{ atoms} \times 8 \text{ e}^-/\text{atom} = 16 \text{ e}^- \]

Determine the total number of electrons used in bonding.

\[ 16 \text{ e}^- - 10 \text{ e}^- = 6 \text{ e}^-, \text{ or 3 covalent bonds} \]

**Step 3**

Determine the number of non-bonding electrons.

\[ 10 \text{ valence e}^- - 6 \text{ bonding e}^- = 4 \text{ e}^-, \text{ or 2 lone pairs} \]

\[ :\text{N} \equiv \text{O} : \]

(c) The molecular formula, ClO\(_3^−\), gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1**

Chlorine has the lower electronegativity in the molecule and will be the central atom.

\[ \text{O} \]

\[ \text{O} \rightarrow \text{Cl} \rightarrow \text{O} \]

**Step 2**

Calculate the total number of valence electrons.

\[ (1 \text{ atom Cl} \times 7 \text{ e}^-/\text{Cl}) + (3 \text{ atoms } \text{O} \times 6 \text{ e}^-/\text{O}) + (1 \text{ e}^- \text{ for charge on ion}) = 26 \text{ e}^- \]

Determine the total number of electrons required for a noble gas configuration.

\[ 4 \text{ atoms} \times 8 \text{ e}^-/\text{atom} = 32 \text{ e}^- \]

Determine the total number of electrons used in bonding.

\[ 32 \text{ e}^- - 26 \text{ e}^- = 6 \text{ e}^-, \text{ or 3 covalent bonds} \]

**Step 3**

Determine the number of non-bonding electrons.

\[ 26 \text{ valence e}^- - 6 \text{ bonding e}^- = 20 \text{ e}^-, \text{ or 10 lone pairs} \]

\[ \begin{align*}
\text{O} & \text{O} \\
\text{S} & \text{O} \\
\text{O} & \text{O} \\
\text{O} & \text{O} \\
\end{align*} \]

(d) The molecular formula, SO\(_3^{2−}\), gives the number of each kind of atom. Follow steps 1–3 as given in the student textbook.

**Step 1**

Sulfur has the lower electronegativity in the molecule and will be the central atom.

\[ \text{O} \]

\[ \text{O} \rightarrow \text{S} \rightarrow \text{O} \]

**Step 2**

Calculate the total number of valence electrons.

\[ \]
(1 atom S × 6 e−/S) + (3 atoms O × 6 e−/O) +
(2 e− for charge on ion) = 26 e−
Determine the total number of electrons required for a noble gas
configuration.
4 atoms × 8 e−/atom = 32 e−
Determine the total number of electrons used in bonding.
32 e− − 26 e− = 6 e−, or 3 covalent bonds

Step 3 Determine the number of non-bonding electrons.
26 valence e− − 6 bonding e− = 20 e−, or 10 lone pairs

Check Your Solution
For each structure, each atom has a noble gas electron configuration in its valence
shell. These are acceptable Lewis structures.

11. Problem
Dichlorofluoroethane, CH3CFC12, has been proposed as a replacement for
chlorofluorocarbons (CFCs). The presence of hydrogen in CH3CFC12 markedly
reduces the ozone-depleting ability of the compound. Draw a Lewis structure for
this molecule.

Solution

Step 1  There is no central atom in this molecule. The three hydrogen atoms will be
bonded to one carbon, and the fluorine and the two chlorine atoms will be
bonded to the other carbon. The two carbon atoms are bonded to each
other.

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]

Step 2  Calculate the total number of valence electrons.
(2 C atoms × 4 e−/C) + (3 H atoms × 1 e−/H atom) +
(1 F atom × 7 e−/F) + (2 Cl atoms × 7 e−/Cl) = 32 e−
Determine the total number of electrons that are required for a noble
gas configuration.
(2 C atoms × 8 e−/C) + (3 H atoms × 2 e−/H) + (1 F × 8 e−/F) +
(2 Cl × 8 e−/Cl) = 46 e−
Determine the total number of electrons used in bonding.
46 − 32 = 14 e−, or 7 covalent bonds

Step 3  Determine the number of non-bonding electrons.
32 valence e− − 14 bonding e− = 18 non-bonding e−, or 9 lone pairs
Hydrogen and carbon have no lone pairs, and fluorine and the two
chlorine atoms have 3 lone pairs.

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{Cl}
\end{array}
\]

Check Your Solution
Each atom has a noble gas electron configuration in its valence shell. This is an
acceptable Lewis structure.
12. Problem
Draw Lewis structures for each of the following molecules: (Note: Neither of these molecules have a single central atom.)
(a) \( \text{N}_2\text{H}_4 \)
(b) \( \text{N}_2\text{F}_2 \)

Solution
(a) Step 1 Two hydrogen atoms will be bonded to each nitrogen atom. The two nitrogen atoms are bonded to each other.

\[
\begin{array}{c}
\text{H} \\
\text{N}--\text{N} \\
\text{H} \\
\end{array}
\]

Step 2 Calculate the total number of valence electrons.
\((2 \text{ N atoms } \times 5 \text{ e}^-/\text{N}) + (4 \text{ H atoms } \times 1 \text{ e}^-/\text{H atom}) = 14 \text{ e}^-\)
Determine the total number of electrons required for a noble gas configuration.
\((2 \text{ C atoms } \times 8 \text{ e}^-/\text{C}) + (4 \text{ H atoms } \times 2 \text{ e}^-/\text{H}) = 24 \text{ e}^-\)
Determine the total number of electrons used in bonding.
\(24 - 14 = 10 \text{ e}^-, \) or 5 covalent bonds

Step 3 Determine the number of non-bonding electrons.
\(14 \text{ valence } e^- - 10 \text{ bonding } e^- = 4 \text{ non-bonding } e^-\), or 2 lone pairs
Hydrogen has no lone pairs, and each nitrogen atom has one lone pair.

\[
\begin{array}{c}
\text{H} \\
\vdots \text{N}--\text{N}: \\
\text{H} \\
\end{array}
\]

(b) Step 1 There will be one fluorine atom bonded to each nitrogen atom. The nitrogen atoms are bonded to each other.

\[
\begin{array}{c}
\text{F} \\
\text{N}--\text{N}--\text{F}
\end{array}
\]

Step 2 Calculate the total number of valence electrons.
\((2 \text{ N atoms } \times 5 \text{ e}^-/\text{N}) + (2 \text{ F atoms } \times 7 \text{ e}^-/\text{F atom}) = 24 \text{ e}^-\)
Determine the total number of electrons required for a noble gas configuration.
\((2 \text{ N atoms } \times 8 \text{ e}^-/\text{C}) + (2 \text{ F atoms } \times 8 \text{ e}^-/\text{F}) = 32 \text{ e}^-\)
Determine the total number of electrons used in bonding.
\(32 - 24 = 8 \text{ e}^-, \) or 4 covalent bonds

Step 3 Determine the number of non-bonding electrons.
\(24 \text{ valence } e^- - 8 \text{ bonding } e^- = 16 \text{ non-bonding } e^-\), or 8 lone pairs
Each nitrogen atom has one lone pair and each fluorine atom has 3 lone pairs. There will be a double bond between the nitrogen atoms.

\[
\begin{array}{c}
\vdots \text{F}--\text{N}==\text{N}--\text{F}
\end{array}
\]

Check Your Solution
For each structure, each atom has a noble gas electron configuration in its valence shell. These are acceptable Lewis structures.

13. Problem
Although Group 18 (VIIIA) elements are inactive, chemists are able to synthesize compounds of several noble gases, including Xe. Draw a Lewis structure for the \( \text{XeO}_4 \) molecule. Indicate if coordinate covalent bonding is likely a part of the bonding in this molecule.
Solution
The molecular formula, XeO₄, gives the number of each kind of atom.

**Step 1** Xe will be the central atom.

\[
\begin{array}{c}
\text{O} \\
\text{O} - \text{Xe} - \text{O} \\
\text{O}
\end{array}
\]

**Step 2** Calculate the total number of valence electrons.
\[
(1 \text{ atom Xe} \times 8 \text{ e}^-/\text{Xe}) + (4 \text{ atoms O} \times 6 \text{ e}^-/\text{O}) = 32 \text{ e}^-
\]
Determine the total number of electrons required for a noble gas configuration.
\[
5 \text{ atoms} \times 8 \text{ e}^-/\text{atom} = 40 \text{ e}^-
\]
Determine the total number of electrons used in bonding.
\[
40 \text{ e}^- - 32 \text{ e}^- = 8 \text{ e}^-, \text{ or } 4 \text{ covalent bonds}
\]

**Step 3** Determine the number of non-bonding electrons.
32 valence e^- - 8 bonding e^- = 24 e^-, or 12 lone pairs

\[
\begin{array}{c}
\text{:O} \\
\text{:O} - \text{Xe} - \text{O} \\
\text{:O}
\end{array}
\]

All four bonds are co-ordinate covalent.

Check Your Solution
Each atom has a noble gas electron configuration in its valence shell. This is an acceptable Lewis structure.

**Solutions for Practice Problems**

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14. Problem
   Draw Lewis structures for each of the following molecules:
   (a) SF₆
   (b) BrF₅

**Solution**

(a) SF₆
The molecular formula, SF₆, gives the number of each kind of atom.

**Step 1** Sulfur has the lower electronegativity and will be the central atom.

\[
\begin{array}{c}
\text{F} \\
\text{F} - \text{S} - \text{F} \\
\text{F}
\end{array}
\]

**Step 2** Calculate the total number of valence electrons.
\[
(1 \text{ atom S} \times 6 \text{ e}^-/\text{S}) + (6 \text{ atoms F} \times 7 \text{ e}^-/\text{F}) = 48 \text{ e}^-
\]
Determine the total number of electrons required for a noble gas configuration.
\[
7 \text{ atoms} \times 8 \text{ e}^-/\text{atom} = 56 \text{ e}^-
\]
Determine the total number of electrons available in bonding.
\[
56 \text{ e}^- - 48 \text{ e}^- = 8 \text{ e}^-, \text{ or } 4 \text{ covalent bonds}
\]
Since twelve electrons are needed to bond the six fluorine atoms, there are not enough electrons. There must be an expanded octet around the sulfur.
**Step 3** Determine the number of non-bonding electrons.
48 valence e\(^-\) − 12 bonding e\(^-\) = 36 e\(^-\), or 18 lone pairs
When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom.

\[
\begin{align*}
\vdots & : F & : \vdots \\
\vdots & : F & : \vdots \\
: \; S & : \; F & : \\
\vdots & : F & : \vdots \\
\vdots & : F & : \vdots 
\end{align*}
\]

(b) BrF\(_5\)
The molecular formula, BrF\(_5\), gives the number of each kind of atom.

**Step 1** Bromine has the lower electronegativity and will be the central atom.

\[
\begin{align*}
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
: \; S & : \; F & : \\
\end{align*}
\]

**Step 2** Calculate the total number of valence electrons.
(1 atom Br × 7 e\(^-\)/Br) + (5 atoms F × 7 e\(^-\)/F) = 42 e\(^-\)
Determine the total number of electrons required for a noble gas configuration.
6 atoms × 8 e\(^-\)/atom = 48 e\(^-\)
Determine the total number of electrons available in bonding.
46 e\(^-\) − 42 e\(^-\) = 4 e\(^-\), or 2 covalent bonds
Since 10 e\(^-\) are needed to bond the five fluorine atoms, there are not enough electrons. There must be an expanded octet around the bromine.

**Step 3** Determine the number of non-bonding electrons.
42 valence e\(^-\) − 10 bonding e\(^-\) = 32 e\(^-\), or 16 lone pairs
When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any electrons remaining, assign them to the central atom.

\[
\begin{align*}
\vdots & : F & : \vdots \\
\vdots & : F & : \vdots \\
\vdots & : F & : \vdots \\
\vdots & : F & : \vdots \\
: \; F & : \; F & : \\
\end{align*}
\]

**Check Your Solution**
Each fluorine atom has a noble gas electron configuration in its valence shell. Both the sulfur atom and the bromine atom have expanded octets. These are acceptable Lewis structures.

15. **Problem**
   Draw Lewis structures for each of the following molecules.
   (a) XeF\(_4\)
   (b) PF\(_5\)

**Solution**
(a) The molecular formula, XeF\(_4\), gives the number of each kind of atom.

**Step 1** Xenon will be the central atom.

\[
\begin{align*}
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
\vdots & \; F & \vdots \\
: \; Xe & : \; F & : \\
\end{align*}
\]
Step 2 Calculate the total number of valence electrons.
(1 atom Xe × 8 e⁻/Xe) + (4 atoms F × 7 e⁻/F) = 36 e⁻
Determine the total number of electrons required for a noble gas configuration.
5 atoms × 8 e⁻/atom = 40 e⁻
Determine the total number of electrons available in bonding.
40 e⁻ − 36 e⁻ = 4 e⁻, or 2 covalent bonds
Since 8 e⁻ are needed to bond the four fluorine atoms, there are not enough electrons. There must be an expanded octet around the xenon atom.

Step 3 Determine the number of non-bonding electrons.
36 valence e⁻ − 8 bonding e⁻ = 28 e⁻, or 14 lone pairs
When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any electrons remaining, assign them to the central atom.

Check Your Solution
Each fluorine atom has a noble gas electron configuration in its valence shell. The Xe and P atoms have expanded octets. These are acceptable Lewis structures.

16. Problem
How does the arrangement of electrons around the central atom differ in PI₃ and ClI₃? Draw the Lewis structures for these compounds to answer this question.
Solution
The molecular formula, PI₃, gives the number of each kind of atom.

**Step 1** Phosphorus has the lower electronegativity and will be the central atom.

\[
\begin{align*}
 & I \quad P \quad I \\
\end{align*}
\]

**Step 2** Calculate the total number of valence electrons.
\[
(1 \text{ atom } P \times 5 \text{ e}^-/P \text{ atom}) + (3 \text{ atoms } I \times 7 \text{ e}^-/I) = 26 \text{ e}^-
\]
Determine the total number of electrons required for a noble gas configuration.
\[
4 \text{ atoms } \times 8 \text{ e}^-/\text{atom} = 32 \text{ e}^-
\]
Determine the total number of electrons available in bonding.
\[
32 \text{ e}^- - 26 \text{ e}^- = 6 \text{ e}^-, \text{ or } 3 \text{ covalent bonds}
\]

**Step 3** Determine the number of non-bonding electrons.
\[
26 \text{ valence } e^- - 6 \text{ bonding } e^- = 20 \text{ e}^-, \text{ or } 10 \text{ lone pairs}
\]

\[
\begin{align*}
 & : I \\
 & : I - P - I:
\end{align*}
\]

For the second part, the molecular formula, ClI₃, gives the number of each kind of atom.

**Step 1** Chlorine will be the central atom.

\[
\begin{align*}
 & I \quad Cl \quad I
\end{align*}
\]

**Step 2** Calculate the total number of valence electrons.
\[
(1 \text{ atom } Cl \times 7 \text{ e}^-/Cl) + (3 \text{ atoms } I \times 7 \text{ e}^-/I) = 28 \text{ e}^-
\]
Determine the total number of electrons required for a noble gas configuration.
\[
4 \text{ atoms } \times 8 \text{ e}^-/\text{atom} = 32 \text{ e}^-
\]
Determine the total number of electrons available in bonding.
\[
32 \text{ e}^- - 28 \text{ e}^- = 4 \text{ e}^-, \text{ or } 2 \text{ covalent bonds}
\]
Since 6 e⁻ are needed to bond the three iodine atoms, there are not enough electrons. There must be an expanded octet around the chlorine.

**Step 3** Determine the number of non-bonding electrons.
\[
28 \text{ valence } e^- - 6 \text{ bonding } e^- = 22 \text{ e}^- \text{ or } 11 \text{ lone pairs}
\]
When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any electrons remaining, assign them to the central atom.

\[
\begin{align*}
 & : I \\
 & : I - Cl - I:
\end{align*}
\]

The difference in the arrangement of electrons around the central atom in these two molecules is that phosphorus has one lone pair, and chlorine has two lone pairs.

**Check Your Solution**
Each atom in PI₃ has a noble gas electron configuration in its valence shell. In ClI₃, the iodine atoms have eight valence electrons, and the chlorine atom has ten valence electrons, which is an expanded octet. These are acceptable Lewis structures.
17. Problem
Draw a Lewis structure for the ion, \( \text{IF}_4^- \).

Solution
The molecular formula, \( \text{IF}_4^- \), gives the number of each kind of atom.

**Step 1** Iodine has the lower electronegativity and will be the central atom.

\[
\begin{array}{c}
\text{F} \\
\text{I} \\
\text{F}
\end{array}
\]

**Step 2** Calculate the total number of valence electrons.
(1 atom I × 7 e⁻/I) + (4 atoms F × 7 e⁻/F) + (1 e⁻ for the charge on the ion) = 36 e⁻

Determine the total number of electrons required for a noble gas configuration.

5 atoms × 8 e⁻/atom = 40 e⁻

Determine the total number of electrons available in bonding.

40 e⁻ − 36 e⁻ = 4 e⁻, or 2 covalent bonds

Since 8 e⁻ are needed to bond the four fluorine atoms, there are not enough electrons. There must be an expanded octet around the iodine atom.

**Step 3** Determine the number of non-bonding electrons.

36 valence e⁻ − 8 bonding e⁻ = 28 e⁻, or 14 lone pairs

When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom.

Check Your Solution
Each fluorine atom has a noble gas electron configuration in its valence shell. The iodine atom has twelve valence electrons, which is an expanded octet. This is an acceptable Lewis structure.

Solutions for Practice Problems
Student Textbook pages 185–186

18. Problem
Use VSEPR theory to predict the molecular shape of each of the following:

(a) HCN
(b) \( \text{SO}_2 \)
(c) \( \text{SO}_3 \)
(d) \( \text{SO}_4^{2-} \)

Plan Your Strategy
For each molecule or ion, follow the four-step procedure to predict molecular shape. Use Table 4.2 in the textbook to name the molecular shape based upon the electron-group arrangements.
Act on Your Strategy

(a) Step 1  Draw a preliminary Lewis structure for HCN.

\[ \text{Total number of valence electrons:} \]
\[ = (1 \text{ C atom } \times 4 \text{ e}^-/\text{C}) + (1 \text{ H atom } \times 1 \text{ e}^-/\text{H}) + (1 \text{ N atom } \times 5 \text{ e}^-/\text{N}) \]
\[ = 10 \text{ e}^- \]

\[ \text{Total electrons needed for noble gas configuration:} \]
\[ = (1 \text{ C atom } \times 8 \text{ e}^-/\text{C}) + (1 \text{ N atom } \times 8 \text{ e}^-/\text{N}) + (1 \text{ H atom } \times 2 \text{ e}^-/\text{H}) \]
\[ = 18 \text{ e}^- \]

\[ \text{Total electrons used in bonding:} \]
\[ = 18 \text{ e}^- - 10 \text{ e}^- \]
\[ = 8 \text{ e}^-, \text{ or 4 covalent bonds} \]

There must be a triple bond between the carbon and nitrogen atoms.

\[ \text{Number of non-bonding electrons:} \]
\[ = 10 \text{ valence e}^- - 8 \text{ bonding e}^- \]
\[ = 2 \text{ e}^-, \text{ or 1 lone pair} \]

The Lewis structure for HCN is

\[ \text{H} - \text{C} \equiv \text{N} \]

Step 2  The Lewis structure shows two electron groups around the central carbon atom (triple bond counts as one group), both BPs.

Step 3  The geometric arrangement of the two electron groups is linear.

Step 4  The molecular shape for two BPs is linear.

(b) Step 1  Draw a preliminary Lewis structure for SO₂.

\[ \text{Total number of valence electrons:} \]
\[ = (1 \text{ S atom } \times 6 \text{ e}^-/\text{S}) + (2 \text{ O atoms } \times 6 \text{ e}^-/\text{O}) \]
\[ = 18 \text{ e}^- \]

\[ \text{Total electrons needed for noble gas configuration:} \]
\[ = 3 \text{ atoms } \times 8 \text{ e}^-/\text{atom} \]
\[ = 24 \text{ e}^- \]

\[ \text{Total electrons used in bonding:} \]
\[ = 24 \text{ e}^- - 18 \text{ e}^- \]
\[ = 6 \text{ e}^-, \text{ or 3 covalent bonds} \]

\[ \text{Number of non-bonding electrons:} \]
\[ = 18 \text{ valence e}^- - 6 \text{ bonding e}^- \]
\[ = 12 \text{ e}^-, \text{ or 6 lone pair} \]

The Lewis structure for SO₂ will be a resonance hybrid.

\[
\begin{align*}
\text{SO}_2 & \rightarrow \text{SO}_2 \\
\text{O} & \equiv \text{S} \equiv \text{O}
\end{align*}
\]

Step 2  The Lewis structure shows three electron groups around the central sulfur atom (double bond counts as one group), two BPs, and one LP.

Step 3  The geometric arrangement of the three electron groups is trigonal planar.

Step 4  The molecular shape for two BPs and one LP is angular (also called bent, or V-shaped).
(c) Step 1  Draw a preliminary Lewis structure for SO$_3$.

\[
\begin{array}{c}
\text{O} \\
\text{S} \\
\text{O} \quad \text{O}
\end{array}
\]

Total number of valence electrons:
\[= (1 \text{ S atom } \times 6 \text{ e}^-/\text{S}) + (3 \text{ O atoms } \times 6 \text{ e}^-/\text{O}) \]
\[= 24 \text{ e}^-\]

Total electrons needed for noble gas configuration:
\[= 4 \text{ atoms } \times 8 \text{ e}^-/\text{atom} \]
\[= 32 \text{ e}^-\]

Total electrons used in bonding:
\[= 32 \text{ e}^- - 24 \text{ e}^- \]
\[= 8 \text{ e}^-, \text{ or 4 covalent bonds} \]

Number of non-bonding electrons:
\[= 24 \text{ valence e}^- - 8 \text{ bonding e}^- \]
\[= 16 \text{ e}^-, \text{ or 8 lone pairs} \]

The Lewis structure for SO$_3$ will be a resonance hybrid.

Step 2  The Lewis structure shows three electron groups around the central sulfur atom (a double bond counts as one group). All are BPs.  
(Note: In further studies in chemistry, you will learn that sulfur can expand its valence shell. This structure will be presented in a variation of this resonance form.)

Step 3  The geometric arrangement of the three electron groups is trigonal planar.

Step 4  The molecular shape for three BPs is trigonal planar.

(d) Step 1  Draw a preliminary Lewis structure for SO$_4^{2-}$.

\[
\begin{array}{c}
\text{O} \\
\text{O} \quad \text{S} \quad \text{O} \\
\text{O}
\end{array}
\]

Total number of valence electrons:
\[= (1 \text{ S atom } \times 6 \text{ e}^-/\text{S}) + (4 \text{ O atoms } \times 6 \text{ e}^-/\text{O}) + (2 \text{ e}^- \text{ for charge on ion}) \]
\[= 32 \text{ e}^-\]

Total electrons needed for noble gas configuration:
\[= 5 \text{ atoms } \times 8 \text{ e}^-/\text{atom} = 40 \text{ e}^-\]

Total electrons used in bonding:
\[= 40 \text{ e}^- - 32 \text{ e}^- = 8 \text{ e}^-, \text{ or 4 covalent bonds} \]

Number of non-bonding electrons:
\[= 32 \text{ valence e}^- - 8 \text{ bonding e}^- \]
\[= 24 \text{ e}^-, \text{ or 12 lone pairs} \]
The Lewis structure for SO₄²⁻ will be

\[
\begin{array}{c}
\vdots \\
\vdots \\
O \\
\vdots \\
S \\
\vdots \\
O \\
\vdots \\
O \\
\vdots \\
\end{array}
\]

(\textbf{Note:} Other resonance structures have been suggested for this ion.)

\textbf{Step 2}  The Lewis structure shows four electron groups around the central sulfur atom. All four are BP.

\textbf{Step 3}  The geometric arrangement of the four electron groups is tetrahedral.

\textbf{Step 4}  The molecular shape for four BPs is tetrahedral.

\textbf{Check Your Solution}

\(a\) This molecular shape corresponds to the VSEPR notation AX₂.

\(b\) This molecular shape corresponds to the VSEPR notation AX₂E.

\(c\) This molecular shape corresponds to the VSEPR notation AX₃.

\(d\) This molecular shape corresponds to the VSEPR notation AX₄.

\textbf{19. Problem}

Use VSEPR theory to predict the molecular shape for each of the following:

\(a\) CH₂F₂

\(b\) AsCl₅

\(c\) NH₄⁺

\(d\) BF₄⁻

\textbf{Plan Your Strategy}

For each molecule or ion, follow the four-step procedure to predict molecular shape. Use Table 4.2 in the textbook to name the molecular shape based upon the electron-group arrangements.

\(a\) \textbf{Step 1}  Draw a preliminary Lewis structure for CH₂F₂.

\[
\begin{array}{c}
F \\
H \quad C \quad F \\
H
\end{array}
\]

Total number of valence electrons:

\(= (1 \text{ C atom } \times 4 \text{ e⁻/C}) + (2 \text{ H atoms } \times 1 \text{ e⁻/H}) + (2 \text{ F atoms } \times 7 \text{ e⁻/F})
\]

\(= 20 \text{ e⁻}
\)

Total electrons needed for noble gas configuration:

\(= (1 \text{ C atoms } \times 8 \text{ e⁻/C}) + (2 \text{ F atoms } \times 8 \text{ e⁻/F}) + (2 \text{ H atoms } \times 2 \text{ e⁻/H})
\]

\(= 28 \text{ e⁻}
\)

Total electrons used in bonding:

\(= 28 \text{ e⁻} - 20 \text{ e⁻}
\]

\(= 8 \text{ e⁻}, \text{ or 4 covalent bonds}
\)

Number of non-bonding electrons:

\(= 20 \text{ valence e⁻} - 8 \text{ bonding e⁻}
\]

\(= 12 \text{ e⁻}, \text{ or 6 lone pairs}
\)

The Lewis structure for CH₂F₂ will be

\[
\begin{array}{c}
\vdots \\
\vdots \\
F \\
\vdots \\
C \\
\vdots \\
F \\
\vdots \\
\vdots \\
\end{array}
\]

\[
\begin{array}{c}
H \\
\quad C \quad F \\
H
\end{array}
\]
Step 2  The Lewis structure shows four electron groups around the central carbon atom. All four are BPs.

Step 3  The geometric arrangement of the four electron groups is tetrahedral.

Step 4  The molecular shape for four BPs is tetrahedral.

(b) Step 1  Draw a preliminary Lewis structure for AsCl₅

\[
\text{Cl} \quad \text{As} \quad \text{Cl} \\
\text{Cl} \quad \text{Cl} \\
\text{Cl} \\
\text{Cl} \\
\text{Cl}
\]

Total number of valence electrons:
\[= (1 \text{ As atom} \times 5 \text{ e}/\text{As}) + (5 \text{ Cl atoms} \times 7 \text{ e}/\text{Cl}) = 40 \text{ e}^{-}\]

Total electrons needed for noble gas configuration:
\[= 6 \text{ atoms} \times 8 \text{ e}/\text{atom} = 48 \text{ e}^{-}\]

Total electrons used in bonding:
\[= 48 \text{ e}^{-} - 40 \text{ e}^{-} = 8 \text{ e}^{-}, \text{ or 4 covalent bonds} \]

Since 10 e⁻ are needed to bond the five chlorine atoms, there are not enough electrons. There must be an expanded octet around the arsenic.

Number of non-bonding electrons:
\[= 40 \text{ valence e}^{-} - 10 \text{ bonding e}^{-} = 30 \text{ e}^{-}, \text{ or 15 lone pairs} \]

When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom.

The Lewis structure for AsCl₅ will be

\[
\text{:Cl} \quad \text{:Cl} \\
\text{:Cl} \quad \text{:Cl} \\
\text{:Cl} \quad \text{As} \\
\text{:Cl} \quad \text{:Cl}
\]

Step 2  The Lewis structure shows five electron groups around the central arsenic atom. All five are BPs.

Step 3  The geometric arrangement of the five electron groups is trigonal bipyramidal.

Step 4  The molecular shape for five BPs is trigonal bipyramidal.

(c) Step 1  Draw the Lewis structure for NH₄⁺. Since this is a common molecule, with no lone pairs, a full series of steps is not given.

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{N} \\
\text{H} \\
\text{H}
\end{array}
\]

Step 2  The Lewis structure shows four electron groups around the central nitrogen atom. All four are BPs.

Step 3  The geometric arrangement of the four electron groups is tetrahedral.

Step 4  The molecular shape for four BPs is tetrahedral.

(d) Step 1  Draw a preliminary Lewis structure for BF₄⁻.

\[
\text{F} \quad \text{B} \quad \text{F} \\
\text{F} \quad \text{F}
\]
Total number of valence electrons:
\[= (1 \text{ B atom} \times 3 \text{ e}^-/\text{B}) + (4 \text{ F atoms} \times 7 \text{ e}^-/\text{F}) + (1 \text{ e}^- \text{ for charge on ion})\]
\[= 32 \text{ e}^-\]

Total electrons needed for noble gas configuration:
\[= 5 \text{ atoms} \times 8 \text{ e}^-/\text{atom}\]
\[= 40 \text{ e}^-\]

Total electrons used in bonding:
\[= 40 \text{ e}^- - 32 \text{ e}^-\]
\[= 8 \text{ e}^-, \text{ or 4 covalent bonds}\]

Number of non-bonding electrons:
\[= 32 \text{ valence } \text{ e}^- - 8 \text{ bonding } \text{ e}^-\]
\[= 24 \text{ e}^-, \text{ or 12 lone pairs}\]

The Lewis structure for \(\text{BF}_4^-\) will be

\[
\begin{array}{c}
: \text{F} : \\
: \text{F} \quad \text{B} \quad \text{F} : \\
: \text{F}:
\end{array}
\]

**Step 2** The Lewis structure shows four electron groups around the central sulfur atom. All four are BPs.

**Step 3** The geometric arrangement of the four electron groups is tetrahedral.

**Step 4** The molecular shape for four BPs is tetrahedral.

**Check Your Solution**
(a) This molecular shape corresponds to the VSEPR notation \(\text{AX}_4\).
(b) This molecular shape corresponds to the VSEPR notation \(\text{AX}_5\).
(c) This molecular shape corresponds to the VSEPR notation \(\text{AX}_4\).
(d) This molecular shape corresponds to the VSEPR notation \(\text{AX}_4\).

20. Problem
Use VSEPR theory to predict the molecular shapes of \(\text{NO}_2^+\) and \(\text{NO}_2^-\).

**Plan Your Strategy**
For each ion, follow the four-step procedure to predict molecular shape. Use Table 4.2 in the textbook to name the molecular shape based upon the electron-group arrangements.

**Act on Your Strategy**

**Step 1** Draw a preliminary Lewis structure for \(\text{NO}_2^+\).

\[\begin{array}{c}
\text{O} \quad \text{N} \quad \text{O}
\end{array}\]

Total number of valence electrons:
\[= (1 \text{ N atom} \times 5 \text{ e}^-/\text{N}) + (2 \text{ O atoms} \times 6 \text{ e}^-/\text{O}) - (1 \text{ e}^- \text{ for charge on ion})\]
\[= 16 \text{ e}^-\]

Total electrons needed for noble gas configuration:
\[= 3 \text{ atoms} \times 8 \text{ e}^-/\text{atom}\]
\[= 24 \text{ e}^-\]

Total electrons used in bonding:
\[= 24 \text{ e}^- - 16 \text{ e}^-\]
\[= 8 \text{ e}^-, \text{ or 4 covalent bonds}\]

Number of non-bonding electrons:
\[= 16 \text{ valence } \text{ e}^- - 8 \text{ bonding } \text{ e}^-\]
\[= 8 \text{ e}^-, \text{ or 4 lone pairs}\]
The Lewis structure for NO$_2^+$ is

\[
\begin{array}{c}
\text{O} \\
\text{N} \\
\text{O}
\end{array}
\]

**Step 2** The Lewis structure shows two electron groups around the central nitrogen atom (a double bond counts as one group). Both are BPs.

**Step 3** The geometric arrangement of the two electron groups is linear.

**Step 4** The molecular shape for two BPs is linear.

For the second part of the problem,

**Step 1** Draw a preliminary Lewis structure for NO$_2^-$.

\[
\begin{array}{c}
\text{O} \\
\text{N} \\
\text{O}
\end{array}
\]

Total number of valence electrons:

\[
= (1 \text{ N atom} \times 5 \text{ e}^-/\text{N}) + (2 \text{ O atoms} \times 6 \text{ e}^-/\text{O}) + (1 \text{ e}^- \text{ for charge on ion})
\]

\[= 18 \text{ e}^-\]

Total electrons needed for noble gas configuration:

\[= 3 \text{ atoms} \times 8 \text{ e}^-/\text{atom}
\]

\[= 24 \text{ e}^-\]

Total electrons used in bonding:

\[= 24 \text{ e}^- - 18 \text{ e}^-
\]

\[= 6 \text{ e}^-, \text{ or 3 covalent bonds}\]

Number of non-bonding electrons:

\[= 18 \text{ valence e}^- - 6 \text{ bonding e}^-
\]

\[= 12 \text{ e}^-, \text{ or 6 lone pairs}\]

The Lewis structure for NO$_2^-$ is

\[
\begin{array}{c}
\text{:O} \\
\text{N} \\
\text{O}
\end{array}
\]

**Step 2** The Lewis structure shows three electron groups around the central nitrogen atom (a double bond counts as one group). There are 2 BPs and 1 LP.

**Step 3** The geometric arrangement of the three electron groups is trigonal planar.

**Step 4** The molecular shape for 2 BPs and 1 LP is angular or bent.

**Check Your Solution**

These molecular shapes correspond to the VSEPR notation AX$_2$ and AX$_3$E respectively.

21. Problem

Draw Lewis structures for the following molecules and ions, and use VSEPR theory to predict the molecular shape. Indicate the examples in which the central atom has an expanded octet.

(a) XeI$_2$

(b) PF$_6^-$

(c) AsF$_3$

(d) AlF$_4^-$

**Plan Your Strategy**

For each molecule or ion, follow the four-step procedure to predict molecular shape. Use Table 4.2 in the textbook to name the molecular shape based upon the electron-group arrangements.

**Act on Your Strategy**

(a) **Step 1** Draw a preliminary Lewis structure for XeI$_2$.

\[
\begin{array}{c}
\text{I} \\
\text{Xe} \\
\text{I}
\end{array}
\]
Total number of valence electrons:
\[(1 \text{ Xe atom } \times 8 \text{ e}^-/\text{Xe}) + (2 \text{ I atoms } \times 7 \text{ e}^-/\text{I})\] = 22 e\(^-\) 

Total electrons needed for noble gas configuration:
= 3 atoms \(\times 8 \text{ e}^-/\text{atom}\) 
= 24 e\(^-\) 

Total electrons used in bonding:
= 24 e\(^-\) – 22 e\(^-\) 
= 2 e\(^-\), or 1 covalent bond

Since 4 e\(^-\) are needed to bond the two iodine atoms, there are not enough electrons. There must be an expanded octet around the xenon atom.

Number of non-bonding electrons:
= 22 valence e\(^-\) – 4 bonding e\(^-\) 
= 18 e\(^-\), or 9 lone pairs

When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any electrons remaining, assign them to the central atom.

The Lewis structure for XeI\(_2\) will be 
\[
\text{I} \quad \text{Xe} \quad \text{I}
\]

**Step 2** The Lewis structure shows five electron groups around the central xenon atom. There are 2 BPs and 3 LPs.

**Step 3** The geometric arrangement of the five electron groups is trigonal bipyramidal.

**Step 4** The molecular shape for 2 BPs and 3 LPs is linear.

(b) **Step 1** Draw a preliminary Lewis structure for PF\(_6^\text{-}\).

\[
\begin{array}{ccc}
F & P & F \\
F & P & F \\
\end{array}
\]

Total number of valence electrons:
\[(1 \text{ P atom } \times 5 \text{ e}^-/\text{P}) + (6 \text{ F atoms } \times 7 \text{ e}^-/\text{F}) + (1 \text{ e}^-\text{ for charge on ion})\] = 48 e\(^-\) 

Total electrons needed for noble gas configuration:
= 7 atoms \(\times 8 \text{ e}^-/\text{atom}\) 
= 56 e\(^-\) 

Total electrons used in bonding:
= 56 e\(^-\) – 48 e\(^-\) 
= 8 e\(^-\), or 4 covalent bonds

Since 12 e\(^-\) are needed to bond the six fluorine atoms, there are not enough electrons. There must be an expanded octet around the phosphorus.

Number of non-bonding electrons:
= 48 valence e\(^-\) – 12 bonding e\(^-\) 
= 36 e\(^-\), or 18 lone pairs
When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any lone pairs remaining, assign them to the central atom.

The Lewis structure for PF$_6^-$ will be

```
  F : P : F
  |   |   |
  | : | : |
  |   |   |
  F : P : F
```

**Step 2** The Lewis structure shows six electron groups around the central phosphorus atom. All six are BPs.

**Step 3** The geometric arrangement of the six electron groups is octahedral.

**Step 4** The molecular shape for 6 BPs is octahedral.

(c) **Step 1** Draw a preliminary Lewis structure for AsF$_3$.

```
  F — As — F
   \   /  \
    F   \\
```

Total number of valence electrons:

\[ = (1 \text{ As atom} \times 5 \text{ e}^-/\text{As}) + (3 \text{ F atoms} \times 7 \text{ e}^-/\text{F}) \]

\[ = 26 \text{ e}^- \]

Total electrons needed for noble gas configuration:

\[ = 4 \text{ atoms} \times 8 \text{ e}^-/\text{atom} \]

\[ = 32 \text{ e}^- \]

Total electrons used in bonding:

\[ = 32 \text{ e}^- - 26 \text{ e}^- \]

\[ = 6 \text{ e}^-, \text{ or 3 covalent bonds} \]

Number of non-bonding electrons:

\[ = 26 \text{ valence e}^- - 6 \text{ bonding e}^- \]

\[ = 20 \text{ e}^-, \text{ or 10 lone pairs} \]

The Lewis structure for AsF$_3$ will be

```
  : F — As — F :
   \   /  \
    F   \\
```

**Step 2** The Lewis structure shows four electron groups around the central arsenic atom. There are 3 BPs and 1 LP.

**Step 3** The geometric arrangement of the four electron groups is tetrahedral.

**Step 4** The molecular shape for 3 BPs and 1 LP is trigonal pyramidal.

(d) **Step 1** Draw a preliminary Lewis structure for AlF$_4^-$.

```
  F — Al — F
   \   /  \
    F   \\
```

Total number of valence electrons:

\[ = (1 \text{ Al atom} \times 3 \text{ e}^-/\text{Al}) + (4 \text{ F atoms} \times 7 \text{ e}^-/\text{F}) + (1 \text{ e}^- \text{ for charge on ion}) \]

\[ = 32 \text{ e}^- \]

Total electrons needed for noble gas configuration:

\[ = 5 \text{ atoms} \times 8 \text{ e}^-/\text{atom} \]

\[ = 40 \text{ e}^- \]
Total electrons used in bonding:
\[= 40 \text{e}^- - 32 \text{e}^-\]
\[= 8 \text{e}^-, \text{or 4 covalent bonds}\]

Number of non-bonding electrons:
\[= 32 \text{valence e}^- - 8 \text{bonding e}^-\]
\[= 24 \text{e}^-, \text{or 12 lone pairs}\]

The Lewis structure for AlF\(_4^-\) will be

![Lewis structure for AlF\(_4^-\)](image)

**Step 2** The Lewis structure shows four electron groups around the central aluminum atom. All four are BPs.

**Step 3** The geometric arrangement of the four electron groups is tetrahedral.

**Step 4** The molecular shape for 4 BPs is tetrahedral.

**Check Your Solution**

(a) This molecular shape corresponds to the VSEPR notation AX\(_3\)E\(_3\).

(b) This molecular shape corresponds to the VSEPR notation AX\(_5\).

(c) This molecular shape corresponds to the VSEPR notation AX\(_3\)E.

(d) This molecular shape corresponds to the VSEPR notation AX\(_4\).

**22. Problem**

Given the general formula and the shape of the molecule or ion, suggest possible elements that could be the central atom, X, in each of the following:

(a) \(XF_3^+\) (trigonal pyramidal)

(b) \(XF_4^-\) (tetrahedral)

(c) \(XF_3\) (T-shaped)

**Plan Your Strategy**

- Identify the VSEPR notation that will result in the indicated molecular shape.
- Determine the number of BP and LP that are necessary for this molecular shape.
- Determine the number of electrons needed to give this number of BP and LP.
- Calculate the number of electrons contributed by the valence shell of the central atom. This value is equal to the total number of valence electrons used for BPs and LPs, minus the number of electrons contributed to BPs by bonded atoms. Adjust this total to reflect any charge on an ion.
- Relate the number of valence electrons in the valence shell of the central atom to the group in which this element will be found in the periodic table.

**Act on Your Strategy**

(a) \(XF_3^+\) (trigonal pyramidal)

- The VSEPR notation is AX\(_3\)E.
- There are 3 BPs, and 1 LP. This requires eight valence electrons.
- Number of electrons contributed by central atom
  \[= (8 \text{e}^- - 3 \text{e}^- \text{(one e}^- \text{ from each F atom)}) + (1 \text{ e}^- \text{ for charge on ion})\]
  \[= 6 \text{ e}^-\]
- The central atom will be a group 16 element, such as O, S, Se, etc.

(b) \(XF_4^-\) (tetrahedral)

- The VSEPR notation is AX\(_4\).
- There are 4 BPs, which requires eight valence electrons.
- Number of electrons contributed by central atom
  \[= (8e^- - 4e^- \text{ (one } e^- \text{ from each } F \text{ atom)}) + (1e^- \text{ for charge on ion})\]
  \[= 5e^-\]
- The central atom will be a group 15 element, such as N, P, As, etc.

(c) \(XF_3\) (T-shaped)
- The VSEPR notation is \(AX_3E_2\).
- There are 3 BPs, and 2 LPs. This requires ten valence electrons.
  - Number of electrons contributed by central atom
    \[= 10e^- - 3e^- \text{ (one } e^- \text{ from each } F \text{ atom)}\]
    \[= 7e^-\]
- The central atom will be a group 17 element, such as Cl, Br, I, etc.

Check Your Solution
The VSEPR notation matches the total number of electrons in each example. These answers seem reasonable and correct.

Solutions for Practice Problems
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23. Problem
Use VSEPR theory to predict the shape of the following molecules. From the molecular shape and the polarity of the bonds, determine whether or not the molecule is polar.
(a) \(CH_3F\)
(b) \(CH_2O\)
(c) \(GaI_3\)

Solution
(a) Step 1
  Draw a preliminary Lewis structure for \(CH_3F\).
  \[
  \begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{C} \\
  \text{F} \\
  \end{array}
  \]
  Total number of valence electrons:
  \[= (1 \text{ C atom } \times 4e^-/C) + (1 \text{ F atom } \times 7e^-/F) + (3 \text{ H atoms } \times 1e^-/H)\]
  \[= 14e^-\]
  Total electrons needed for noble gas configuration:
  \[= (1 \text{ atom C } \times 8e^-/C) + (1 \text{ atom F } \times 8e^-/F) + (3 \text{ H atoms } \times 2e^-/H)\]
  \[= 22e^-\]
  Total electrons used in bonding:
  \[= 22e^- - 14e^-\]
  \[= 8e^-, \text{ or } 4 \text{ covalent bonds}\]
  Number of non-bonding electrons:
  \[= 14 \text{ valence } e^- - 8 \text{ bonding } e^-\]
  \[= 6e^-, \text{ or } 3 \text{ lone pairs}\]
  The Lewis structure for \(CH_3F\) will be
  \[
  \begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{C} \\
  \text{F} : \\
  \end{array}
  \]

Step 2
The Lewis structure shows four electron groups around the central carbon atom. All four are BPs.
Step 3 The geometric arrangement of the four electron groups is tetrahedral.

Step 4 The molecular shape for 4 BPs is tetrahedral.

Step 5 All of the bonds are polar. Their polarities do not cancel, however, since the C—F bond has a different polarity from the other three C—H bonds. The molecule will be polar.

(b) Step 1 From the sample problem shown in the textbook, the Lewis structure for this molecule is

\[
\begin{array}{c}
\text{O} \\
\text{H} \quad \text{C} \quad \text{H}
\end{array}
\]

Step 2 The Lewis structure has three electron groups (the double bond counts as one electron group). All are BPs.

Step 3 The geometric shape will be trigonal planar.

Step 4 The molecular shape is trigonal planar.

Step 5 Since the polarities of the three bonds are not equal, this molecule will be polar.

(c) Step 1 Draw a preliminary Lewis structure for GaI₃.

\[
\text{I} \quad \text{Ga} \quad \text{I}
\]

Total number of valence electrons:
\[= (1 \text{ Ga atom } \times 5 \text{ e}^-/\text{Ga}) + (3 \text{ I atoms } \times 7 \text{ e}^-/\text{I})\]
\[= 26 \text{ e}^-\]

Total electrons needed for noble gas configuration:
\[= 4 \text{ atoms } \times 8 \text{ e}^-/\text{atom}\]
\[= 32 \text{ e}^-\]

Total electrons used in bonding:
\[= 32 \text{ e}^- - 26 \text{ e}^-\]
\[= 6 \text{ e}^-, \text{ or 3 covalent bonds}\]

Number of non-bonding electrons:
\[= 26 \text{ valence e}^- - 6 \text{ bonding e}^-\]
\[= 20 \text{ e}^-, \text{ or 10 lone pairs}\]

The Lewis structure for GaI₃ will be

\[
\begin{array}{c}
\text{I} \\
\text{Ga} \\
\text{I}
\end{array}
\]

Step 2 The Lewis structure shows four electron groups around the central gallium atom. There are 3 BPs and 1 LP.

Step 3 The geometric arrangement of the four electron groups is tetrahedral.

Step 4 The molecular shape for 3 BPs and 1 LP is trigonal pyramidal.

Step 5 The polarities of the three Ga—I bonds are the same. They will not cancel, however, because of the molecular shape. This molecule is polar.

Check Your Solution
(a) This molecular shape corresponds to the VSEPR notation AX₃Y. It seems reasonable and correct.

(b) This molecular shape corresponds to the VSEPR notation AX₃Y. It seems reasonable and correct.

(c) This molecular shape corresponds to the VSEPR notation AX₃E. The answer seems reasonable and correct.
24. Problem
Freon-12, CCl₂F₂, was used as a coolant in refrigerators until it was suspected to be a cause of ozone depletion. Determine the molecular shape of CCl₂F₂ and discuss the possibility that the molecule will be a dipole.

Solution
Step 1
Draw a preliminary Lewis structure for CCl₂F₂.

\[
\text{Cl} \quad \text{C} \quad \text{F}
\]

Total number of valence electrons:
\[
= (1 \text{ C atom} \times 4 \text{ e}^-/\text{C}) + (2 \text{ F atoms} \times 7 \text{ e}^-/\text{F}) + (2 \text{ Cl atoms} \times 7 \text{ e}^-/\text{Cl})
= 32 \text{ e}^-
\]

Total electrons needed for noble gas configuration:
\[
= 5 \text{ atoms} \times 8 \text{ e}^-/\text{atom}
= 40 \text{ e}^-
\]

Total electrons used in bonding:
\[
= 40 \text{ e}^- - 32 \text{ e}^-
= 8 \text{ e}^- , \text{ or 4 covalent bonds}
\]

Number of non-bonding electrons:
\[
= 32 \text{ valence e}^- - 8 \text{ bonding e}^-
= 24 \text{ e}^- , \text{ or 12 lone pairs}
\]

The Lewis structure for CCl₂F₂ will be
\[
\text{F} \quad \text{Cl} \quad \text{C} \quad \text{F}
\]

Step 2
The Lewis structure shows four electron groups around the central carbon atom. All four are BPs.

Step 3
The geometric arrangement of the four electron groups is tetrahedral.

Step 4
The molecular shape for 4 BPs is tetrahedral.

Step 5
All of the bonds are polar. Their polarities do not cancel, however, since the C—F bond has a different polarity than the two C—Cl bonds. The molecule will be polar.

Check Your Solution
This molecular shape corresponds to the VSEPR notation AX₂Y₂. It seems reasonable and correct.

25. Problem
Which of the following is more polar? Justify your answer in each case.

(a) NF₃ or NCl₃
(b) ICl₄⁻ or TeCl₄

Solution
(a) Determine the molecular shape of each molecule, and compare the polarities of the bonds.

Both NF₃ and NCl₃ have the same central atom, and both F and Cl are in the same group of the periodic table. Therefore, these two molecules will have the same shape. (Both will have a trigonal pyramidal shape.) Since the N—F bonds are more polar than the N—Cl bonds, you can conclude that NF₃ is more polar.

(b) For the first molecule,
Step 1  Draw a preliminary Lewis structure for $\text{ICl}_4^-$.

\[
\text{Cl} \quad \text{I} \quad \text{Cl} \quad \text{Cl}
\]

Total number of valence electrons:
$= (1 \text{ I atom} \times 7 \text{ e}^-/\text{I}) + (4 \text{ Cl atoms} \times 7 \text{ e}^-/\text{Cl}) + (1 \text{ e}^- \text{ for charge on ion})$
$= 36 \text{ e}^-$

Total electrons needed for noble gas configuration:
$= 5 \text{ atoms} \times 8 \text{ e}^-/\text{atom}$
$= 40 \text{ e}^-$

Total electrons used in bonding:
$= 40 \text{ e}^- - 36 \text{ e}^-$
$= 4 \text{ e}^-$, or 2 covalent bonds

Since 8 e$^-$ are needed to bond the four fluorine atoms, there are not enough electrons. There must be an expanded octet around the iodine.

Number of non-bonding electrons:
$= 36 \text{ valence e}^- - 8 \text{ bonding e}^-$
$= 28 \text{ e}^-$, or 14 lone pairs

When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any lone pairs remaining, assign them to the central atom.

The Lewis structure for $\text{ICl}_4^-$ will be

Step 2  The Lewis structure shows six electron groups around the central iodine atom. There are 4 BPs and 2 LPs.

Step 3  The geometric arrangement of the six electron groups is octahedral.

Step 4  The molecular shape for 4 BPs and 2 LPs is square planar.

Step 5  A square planar molecular shape leads to a nonpolar molecule.

For the second molecule,

Step 1  Draw a preliminary Lewis structure for $\text{TeCl}_4$.

\[
\text{Cl} \quad \text{Te} \quad \text{Cl}
\]

Total number of valence electrons:
$= (1 \text{ Te atom} \times 6 \text{ e}^-/\text{Te}) + (4 \text{ Cl atoms} \times 7 \text{ e}^-/\text{Cl})$
$= 34 \text{ e}^-$

Total electrons needed for noble gas configuration:
$= 5 \text{ atoms} \times 8 \text{ e}^-/\text{atom}$
$= 40 \text{ e}^-$

Total electrons used in bonding:
$= 40 \text{ e}^- - 34 \text{ e}^-$
$= 6 \text{ e}^-$, or 3 covalent bonds
Since 8 e\(^-\) are needed to bond the four chlorine atoms, there are not enough electrons. There must be an expanded octet around the tellurium atom.

Number of non-bonding electrons:
= 34 valence e\(^-\) − 8 bonding e\(^-\)
= 26 e\(^-\), or 13 lone pairs

When there is an insufficient number of bonding electrons to give each atom an octet of valence electrons, assign lone pairs of electrons to all of the atoms that surround the central atom. If there are any lone pairs remaining, assign them to the central atom.

The Lewis structure for TeCl\(_4\) will be

\[
\begin{array}{c}
\text{Cl} \\
\text{Cl} \\
\text{Te} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]

**Step 2** The Lewis structure shows five electron groups around the central tellurium atom. There are 4 BPs and 1 LP.

**Step 3** The geometric arrangement of the five electron groups is trigonal bipyramidal.

**Step 4** The molecular shape for 4 BPs and 1 LP is seesaw.

**Step 5** All of the Te—Cl bonds are polar. Their polarities do not cancel out, however, because of the molecular shape. The molecule will be polar. Only TeCl\(_4\) is polar.

**Check Your Solution**
These results are consistent with the VSEPR shapes and are reasonable and correct.

### 26. Problem
A hypothetical molecule with formula XY\(_3\) is discovered, through experiment, to exist. It is polar. Which molecular shapes are possible for this molecule? Which shapes are impossible? Explain why in each case.

**Plan Your Strategy**
- Determine the VSEPR notations that can correspond to the molecular formula XY\(_3\).
- For each notation, determine the number of BPs and LPs that can give this molecular formula. From this information, determine the allowed molecular shapes that are possible.

**Act on Your Strategy**
- XY\(_3\) can have the VSEPR notations AX\(_3\)E and AX\(_3\)E\(_2\).
- The VSEPR notation AX\(_3\)E has 3 BPs and 1 LP. It is trigonal pyramidal. The VSEPR notation AX\(_3\)E\(_2\) has 3 BPs and 2 LPs. It is T-shaped. Since the molecule is polar, the VSEPR notation AX\(_3\) is not possible, since this would give a trigonal planar molecular shape that cannot be polar. Therefore, the molecular shape must be AX\(_3\)E, or trigonal pyramidal, or AX\(_3\)E\(_2\), or T-shaped.

**Check Your Solution**
These results are consistent with the VSEPR shapes and are reasonable and correct.