Assignment 9 Solutions

Chapter 8, #8.32, 36, 38, 42, 54, 56, 72, 100, 102, Chapter 10, #10.24, 40, 55, 63.

8.32. **Collect and Organize**

Of B^{3+}, I^-, Ca^{2+}, and Pb^{2+} we are to identify which have a complete valence-shell octet.

Analyze

Ion	Electron Configuration	Number of e^- in Valence Shell
B^{3+}	[He]	0
I^-	[Kr]$4d^{10}5s^25p^6$	8
Ca^{2+}	[Ar]	0
Pb^{2+}	[Xe]$4f^{14}5d^{10}6s^2$	2

Solve

The ions I^- and Ca^{2+} have complete valence-shell octets. B^{3+} does not have an octet but rather the duet of the He atom.

Think About It

Notice that cations are formed by loss of electrons to achieve a core noble gas configuration, which leaves no electrons in the valence shell, and anions are formed by gain of electrons to give 8 electrons in the valence shell.

8.36. **Collect and Organize**

For the diatomic species BN, HF, OH^-, and CN^-, we are to determine the total number of valence electrons.

Analyze

For each species we need to add the valence electrons for each atom. If the species is charged, we need to reduce or increase the number of electrons as necessary to form cations or anions, respectively.

Solve

(a) 3 valence e^- (B) + 5 valence e^- (N) = 8 valence e^-
(b) 1 valence e^- (H) + 7 valence e^- (F) = 8 valence e^-
(c) 6 valence e^- (O) + 1 valence e^- (H) + 1 e^- (negative charge) = 8 valence e^-
(d) 4 valence e^- (C) + 5 valence e^- (N) + 1 e^- (negative charge) = 10 valence e^-

8.38. **Collect and Organize**

We are to draw correct Lewis structures satisfying the octet rule for all atoms in the diatomic molecules and ions CO, O_2, ClO^-, and CN^-.

Analyze

To draw the Lewis structure, we first must determine the number of covalent bonds in the molecule by comparing the number of valence electrons in the molecule to the number of electrons the atoms need to complete the duet for H and the octet for all other atoms. After drawing the skeletal structure with bonds we distribute the remaining electrons in pairs to complete the octet for each atom. Finally we check our answer by counting the electrons in the Lewis structure drawn.

Solve

(a) For CO

[1] Have: $(1C \times 4\ e^-) + (1O \times 6\ e^-) = 10\ e^-$
Need: $(1C \times 8\ e^-) + (1O \times 8\ e^-) = 16\ e^-$

Difference: 16 e^- – 10 e^- = 6 e^- = 3 covalent bonds

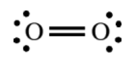
[2] Three covalent bonds use 6 e^-, leaving 4 electrons in 2 lone pairs to complete the octet for C and O:

[3] This Lewis structure has 10 e^-. Both carbon and oxygen have an octet, each with a triple bond and 1 lone pair.

(b) For O_2

[1] Have: (2O × 6 e^-) = 12 e^-
Need: (2O × 8 e^-) = 16 e^-
Difference: 16 e^- – 12 e^- = 4 e^- = 2 covalent bonds

[2] Two covalent bonds use 4 e^-, leaving 8 e^- for 4 lone pairs to complete the octets for the oxygen atoms:

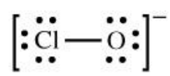


[3] This Lewis structure has 12 e^-. Both oxygen atoms are doubly bonded with 2 lone pairs.

(c) For ClO^-

[1] Have: (1Cl × 7 e^-) + (1O × 6 e^-) + 1 e^- = 14 e^-
Need: (1Cl × 8 e^-) + (1O × 8 e^-) = 16 e^-
Difference: 16 e^- – 14 e^- = 2 e^- = 1 covalent bond

[2] One covalent bond uses 2 e^-, leaving 12 e^- (6 lone pairs) to complete the octets on Cl and O.



[3] This Lewis structure has 14 e^- with each atom singly bonded with 3 lone pairs.

(d) For CN^-

[1] Have: (1C × 4 e^-) + (1N × 5 e^-) + 1 e^- = 10 e^-
Need: (1C × 8 e^-) + (1N × 8 e^-) = 16 e^-
Difference: 16 e^- – 10 e^- = 6 e^- = 3 covalent bonds

[2] Three covalent bonds use 6 e^-, leaving 4 e^- in 2 lone pairs to complete the octets on N and C:

[3] This Lewis structure has 10 e^-. Each atom is triple bonded and has 1 lone pair.

Think About It

When writing Lewis structures for ionic species, enclose the structure in brackets and indicate the charge on the ion as shown in this problem for ClO^- and CN^-.

8.42. **Collect and Organize**

Using the method described in the textbook, we are to draw Lewis structures for three greenhouse gases.

Analyze

To draw the Lewis structure, we first must determine the number of covalent bonds in the molecule by comparing the number of valence electrons in the molecule to the number of electrons the atoms need to complete the duet for H and the octet for all other atoms. We can then use one pair of electrons to form the bonds that link atoms together to form the skeletal structure of the molecule. We then distribute the remaining electrons in pairs to complete the octet for each atom. Finally we check our answer by counting the electrons in the Lewis structure drawn.

Solve

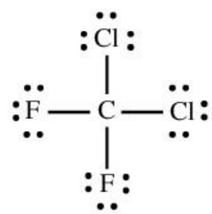
(a) CF_2Cl_2

[1] Have: $(1C \times 4\ e^-) + (2F \times 7\ e^-) + (2Cl \times 7\ e^-) = 32\ e^-$

Need: $(1C \times 8\ e^-) + (2F \times 8\ e^-) + (2Cl \times 8\ e^-) = 40\ e^-$

Difference: $40\ e^- - 32\ e^- = 8\ e^- = 4$ covalent bonds

[2] The central atom is usually listed first in the molecular formula, so carbon is central and bonds to two F and two Cl atoms. Four covalent bonds use 8 e^- leaving 24 e^- remaining to complete the octets for F and Cl. The octet for C is satisfied by forming 4 covalent bonds.



[3] This Lewis structure has 32 e^- with four bonding pairs and 12 lone pairs.

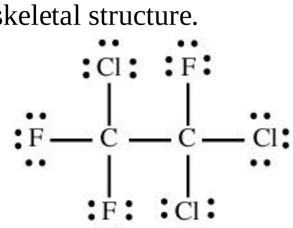
(b) Cl_2FCCF_2Cl

[1] Have: $(2C \times 4\ e^-) + (3Cl \times 7\ e^-) + (3F \times 7\ e^-) = 50\ e^-$

Need: $(2C \times 8\ e^-) + (3Cl \times 8\ e^-) + (3F \times 8\ e^-) = 64\ e^-$

Difference: $64\ e^- - 50\ e^- = 14\ e^- = 7$ covalent bonds

[2] We are given that there is a C–C bond in the molecule. From the formula given, we also see that one C is bound to 2 Cl atoms and 1 F atom. The other carbon is bound to 1 Cl atom and 2 F atoms. This requires 7 covalent bonds using 14 e^- leaving 36 e^- in 18 lone pairs to complete the octets on F and Cl. The carbon octets are satisfied through the 4 bonds it makes to form the skeletal structure.



[3] This Lewis structure has 50 e^- with 7 bonding pairs and 18 lone pairs.

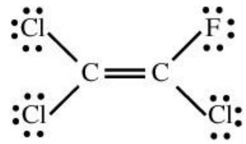
(c) C_2Cl_3F

[1] Have: $(2C \times 4\ e^-) + (3Cl \times 7\ e^-) + (1F \times 7\ e^-) = 36\ e^-$

Need: $(2C \times 8\ e^-) + (3Cl \times 8\ e^-) + (1F \times 8\ e^-) = 48\ e^-$

Difference: $48\ e^- - 36\ e^- = 12\ e^- = 6$ covalent bonds

[2] We are given that there is a C—C bond in the molecule. We also need 4 additional bonds to 3 Cl atoms and 1 F atom from the carbon atoms. This leaves one covalent bond left to be used for carbon—carbon double bonding to complete each C's octet. This leaves 24 e^- in 12 lone pairs to complete the octets on F and Cl.



[3] This Lewis structure has 36 e^- in 4 single bonds, 1 double bond, and 12 lone pairs.

Think About It
In determining the skeletal structure for these molecules, it is helpful to know that carbon commonly bonds to 4 atoms and can double or triple bond to itself.

8.54. **Collect and Organize**
We are asked to describe why electrons are not equally shared in the bonds between two different elements.
Solve
When atoms have different electronegativities, they do not equally pull on the electrons that contribute to the bond and the bond will be polar, with one atom being slightly negative (δ^-) and the other slightly positive (δ^+).
Think About It
Bonds between atoms of the same element will be nonpolar.

8.56. **Collect and Organize**
Of the bonds listed, we are to determine which is the least polar.
Analyze
The lower the difference in electronegativity between the bonded atoms, the less polar bond. We need the electronegativity values from Figure 8.5 in the textbook.
Solve

Bond	Electronegativity Difference
C—Se	2.5 – 2.4 = 0.1
C ═ O	3.5 – 2.5 = 1.0
Cl—Br	3.0 – 2.8 = 0.2
O ═ O	3.5 – 3.5 = 0
N—H	3.0 – 2.1 = 0.9
C—H	2.5 – 2.1 = 0.4

The O ═ O bond is nonpolar and, therefore, the least polar in the list.
Think About It
The most polar bond in the list is C ═ O because it has the greatest electronegativity difference between the atoms.

8.72. **Collect and Organize**
We are to draw Lewis structures, with resonance forms if appropriate, for urea [$H_2NC(O)NH_2$] and the nitrite ion (NO_2^-).
Analyze
We first draw one of the valid Lewis structures by the method described in the textbook, then redistribute the bonding pairs and lone pairs in the structure to draw resonance forms.

Solve

For urea, $H_2NC(O)NH_2$

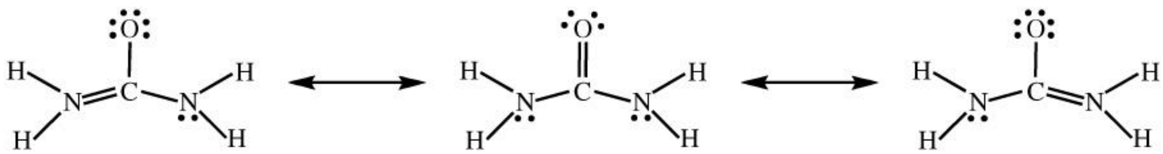
We are given in the problem that the carbon atom has a double bond to the oxygen atom in this compound. The hydrogen atoms will be terminal so as to not violate the duet rule.

Have: $(4H \times 1\ e^-) + (2N \times 5\ e^-) + (1C \times 4\ e^-) + (1O \times 6\ e^-) = 24\ e^-$

Need: $(4H \times 2\ e^-) + (2N \times 8\ e^-) + (1C \times 8\ e^-) + (1O \times 8\ e^-) = 40\ e^-$

Difference: $40\ e^- - 24\ e^- = 16\ e^- = 8$ covalent bonds

To connect the atoms we need 7 covalent bonds. The eighth bond forms the double bond between C and O. This would use 16 e^-, leaving 8 e^- in 4 lone pairs to complete the octets on O and N. This molecule has two additional resonance structures.



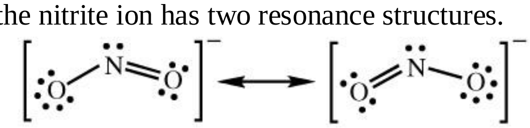
For nitrite ion, NO_2^-

Have: $(1N \times 5\ e^-) + (2O \times 6\ e^-) + 1\ e^- = 18\ e^-$

Need: $(1N \times 8\ e^-) + (2O \times 8\ e^-) = 24\ e^-$

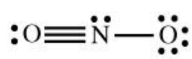
Difference: $24\ e^- - 18\ e^- = 6\ e^- = 3$ covalent bonds

To connect the atoms we need 2 covalent bonds. The third may be used to form a double bond between the N atom and one of the O atoms. This would use 6 e^-, leaving 12 e^- in 6 lone pairs to complete the octets on O and N. Because either of the oxygen atoms may be drawn as double bonded to the nitrogen, the nitrite ion has two resonance structures.



Think About It

We cannot draw a resonance form for NO_2^- in which an oxygen atom forms a triple bond to the nitrogen atom because the N atom would have more than an octet in the structure.



4 bonding pairs
<u>5 lone pairs</u>
18 e^-

8.100. **Collect and Organize**

For each molecule combining N with O we are to determine which are odd-electron molecules.

Analyze

To answer this we need only add up the valence electrons for each molecule.

Solve

(a) NO has $(1N \times 5\ e^-) + (1O \times 6\ e^-) = 11\ e^-$
(b) NO_2 has $(1N \times 5\ e^-) + (2O \times 6\ e^-) = 17\ e^-$
(c) NO_3 has $(1N \times 5\ e^-) + (3O \times 6\ e^-) = 23\ e^-$
(d) N_2O_4 has $(2N \times 5\ e^-) + (4O \times 6\ e^-) = 34\ e^-$
(e) N_2O_5 has $(2N \times 5\ e^-) + (5O \times 6\ e^-) = 40\ e^-$

The odd-electron molecules are NO, NO_2, and NO_3 (a–c).

Think About It

For these nitrogen–oxygen molecules, notice that the odd-electron species have an odd number of N atoms in their formulas.

8.102. **Collect and Organize**

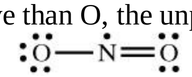
For the species named, we are to decide which atom in the molecule is most likely to have the unpaired electron.

Analyze

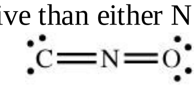
An unpaired electron in a molecule is a lone electron. This occurs when a molecule has an odd number of valence electrons. The atom that bears the odd electron is the least electronegative. The electronegativities of the atoms are N = 3.0, O = 3.5, C = 2.5, Cl = 3.0.

Solve

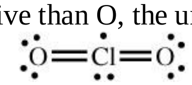
(a) For NO_2, since N is less electronegative than O, the unpaired electron is on N:



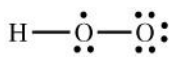
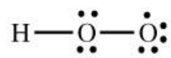
(b) For CNO, since C is less electronegative than either N or O, the unpaired electron is on C:



(c) For ClO_2, since Cl is less electronegative than O, the unpaired electron is on Cl:



(d) For HO_2, since H is less electronegative than O, we expect the unpaired electron to be on H. Hydrogen, however, must be bonded to one of the O atoms and obey the duet rule, so the unpaired electron must be on one of the O atoms:

 or 

10.24. **Collect and Organize**

Given two liquids, one polar and one nonpolar, we are to predict which has the higher boiling point.

Analyze

The greater the intermolecular forces, the higher the boiling point of the substance. Polar molecules attract each other through dipole–dipole interactions. Nonpolar molecules attract each other through dispersion forces.

Solve

Because dipole–dipole interactions are stronger than dispersion forces, the polar liquid has a higher boiling point than the nonpolar liquid.

Think About It

If a nonpolar molecule, however, is very large (high molar mass) the sum of all the weak dispersion forces may be greater in total strength than the dipole–dipole forces of a smaller molecule.

10.40. **Collect and Organize**

We are asked to name the types of intermolecular forces that must be overcome when CO_2 sublimes (changes from solid phase to gas phase), when $CHCl_3$ boils (liquid phase to gas phase), and when ice (H_2O) melts (solid phase to liquid phase).

Analyze

For each of these covalent substances, we need to determine whether the strongest intermolecular force between the molecules is dipole–dipole, hydrogen bonding (a special type of dipole–dipole interaction), or induced dipole–induced dipole (dispersion) forces. To determine this, we need to consider the geometry of the molecule (polar or nonpolar) and whether hydrogen is bonded to F, O, or N to give the possibility of hydrogen bonds.

Solve

(a) CO_2 is a linear molecule and as such is nonpolar. For sublimation of CO_2, only the weak dispersion forces are broken between the molecules.

(b) $CHCl_3$ is a polar, tetrahedral molecule. In order for this liquid to boil, dipole–dipole interactions between the molecules must be broken. Dispersion forces are also present, but they are weaker than the dipole–dipole interactions.

(c) H_2O is a polar, bent molecule with H bonded to the very electronegative atom O. In order to melt, hydrogen bonds (a special class of dipole–dipole interaction) between the water molecules must break. Dispersion forces are also present, but they are weaker than the hydrogen bonds.

Think About It

Breaking hydrogen bonds during melting of ice is only partial. Some hydrogen bonding exists in liquid water. In boiling water, however, the hydrogen bonds must be completely broken for water to go into the gas phase. Amazingly, though, the water molecules can still form water clusters (groups of water molecules) in the gas phase, which speaks to the strength of the hydrogen bond.

10.55. **Collect and Organize**

For each pair of compounds, we are to determine which is more soluble in H_2O.

Analyze

Water is a polar solvent capable of forming hydrogen bonds to dissolved substances with X—H bonds (X = F, O, N). In each pair of compounds, the more soluble is the more polar molecule or the one that forms hydrogen bonds. In considering whether a salt is soluble in water, we have to consider the relative strengths of the ionic bonds as well as the relative strengths of the ion–dipole interactions formed on dissolution.

Solve

(a) $CHCl_3$ is polar whereas CCl_4 is not. $CHCl_3$ is more soluble in water.

(b) CH_3OH is more polar because it has a smaller hydrocarbon chain compared to $C_6H_{11}OH$. CH_3OH is more soluble in water.

(c) NaF has a weaker ionic bond than MgO. NaF is more soluble in water.

(d) BaF_2 has a weaker ionic bond than CaF_2 because Ba^{2+} is larger than Ca^{2+}. BaF_2 is more soluble in water.

Think About It

Solubility is determined by many factors: polarity, ability to hydrogen bond, and strength of the intermolecular forces between molecules of the solute.

10.63. **Collect and Organize**

Of the ionic compounds listed [NaCl, KI, $Ca(OH)_2$, and CaO], we are to determine which would be most soluble in water.

Analyze

The weaker the ionic bond, the easier the bond breaks for the cation and anion to dissolve in water. Ionic bonds are weakest for large ions of low charge.

Solve

KI (b) has the largest ions of lowest (1+ and 1–) charge, so it is the most soluble in water because it has the weakest ion–ion bond.

Think About It

CaO with a 2+ cation and 2– anion would be expected to be the least soluble in water.