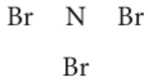
## Section 4.1: Types of Chemical Bonds

### Tutorial 1 Practice, page 200

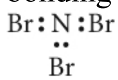
**1. (a)** Lewis structure for $NBr_3$:

**Step 1.** The central atom for nitrogen tribromide is bromine.



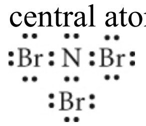
**Step 2.** Add up the number of valence electrons available.
1 N atom: $1(5e^-) = 5e^-$
3 Br atoms: $3(7e^-) = 21e^-$
Total: $26e^-$

**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms to represent each bonding electron pair.



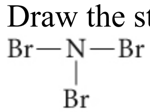
Start: $26e^-$
Used: $6e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding and central atoms.



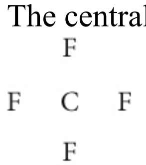
**Step 5.** Determine how many electrons are still available.
Start: $26e^-$
Used: $26e^-$
Remaining: $26e^- – 26e^- = 0e^-$

**Step 6.** Draw the structural formula.



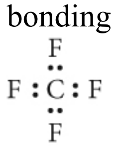
**(b)** Lewis structure and structural formula for $CF_4$:

**Step 1.** The central atom in the tetrafluoromethane molecule is carbon.



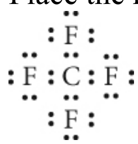
**Step 2.** Add up the number of valence electrons available.
1 C atom: $1(4e^-) = 4e^-$
4 F atoms: $4(7e^-) = 28e^-$
Total: $32e^-$

**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms to represent the bonding electron pairs.



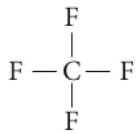
Start: $32e^-$
Used: $8e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding atoms.



**Step 5.** Determine how many electrons are still available.
Start: $32e^-$
Used: $32e^-$
Remaining: $32e^- - 32e^- = 0e^-$

**Step 6.** Draw the structural formula.



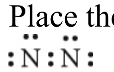
**(c)** Lewis structure and structural formula for $N_2$:

**Step 1.** There are two N atoms in the nitrogen gas molecule.
N N

**Step 2.** Add up the number of valence electrons available.
2 N atoms: $2(5e^-) = 10e^-$

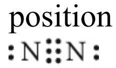
**Step 3.** Place 1 pair of electrons between the two N atoms.
N:N
Start: 10 $e^-$
Used 2 $e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the two atoms.



**Step 5.** Determine how many electrons are still available.
Start: 10 $e^-$
Used 10 $e^-$
Remaining: $10e^- - 10e^- = 0e^-$

**Step 6.** The two N atoms do not have a full octet. Move lone pairs into a bonding position between those atoms until all octets are complete.



**Step 7.** Draw the structural formula.
$N \equiv N$

**(d)** Lewis structure and structural formula for $C_2H_4$:

**Step 1.** The central atoms in the ethylene molecule are two carbons.

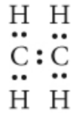
H H

C C

H H

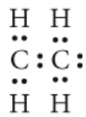
**Step 2.** Add up the number of valence electrons available.
2 C atoms: $2(4e^-) = 8e^-$
4 H atoms: $4(1e^-) = 4e^-$
Total: $12e^-$

**Step 3.** Place 1 pair of electrons to represent the bonding electron pairs. Follow the duet rule for the hydrogen atoms and the octet rule for the carbon atoms.



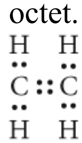
Start: $12e^-$
Used: $10e^-$

**Step 4.** Place the remaining two valence electrons as a lone pair on one of the C atoms.

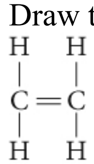


**Step 5.** Determine how many electrons are still available.
Start: $12e^-$
Used: $12e^-$
Remaining: $12e^- - 12e^- = 0e^-$

**Step 6.** One of the two C atoms does not have a full octet. Move the lone pair from the other C atom into a bonding position between the two C atoms to complete the octet.

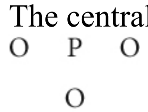


**Step 7.** Draw the structural formula.



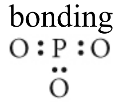
**2. (a)** Lewis structure for $PO_3^{3-}$:

**Step 1.** The central atom in the phosphite ion is phosphorus.



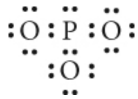
**Step 2.** Add up the number of valence electrons available. Add 1 electron for each unit of negative charge.
1 P atom: $1(5e^-) = 5e^-$
3 O atoms: $3(6e^-) = 18e^-$
charge: 3e–
Total: $26e^-$

**Step 3.** Place 1 pair of electrons between each adjacent pairs of atoms to represent the bonding electron pairs.



Start: $26e^-$
Used: $6e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding and central atoms.



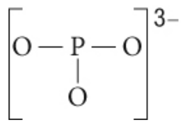
**Step 5.** Determine how many electrons are still available.

Start: $26e^-$

Used: $26e^-$

Remaining: $26e^- - 26e^- = 0e^-$

**Step 6.** Draw the structural formula. Since $PO_3^{3-}$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



**(b)** Lewis structure for $CN^-$:

**Step 1.** The central atom in the cyanide ion is carbon.

C N

**Step 2.** Add up the number of valence electrons available. Add 1 electron for each unit of negative charge.

1 C atom: $1(4e^-) = 4e^-$

1 N atom: $1(5e^-) = 5e^-$

charge: 1e–

Total: $10e^-$

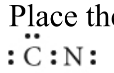
**Step 3.** Place 1 pair of electrons between the two atoms.

C : N

Start: 10 $e^-$

Used 2 $e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the two atoms.



**Step 5.** Determine how many electrons are still available.

Start: 10 $e^-$

Used 10 $e^-$

Remaining: $10e^- - 10e^- = 0e^-$

**Step 6.** The N atom does not have a full octet. Move lone pairs from the C atom into a bonding position between the two atoms to complete the octet.

**Step 7.** Draw the structural formula. Since $CN^-$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.

$[C \equiv N]^-$

**(c)** Lewis structure for $NO_2^-$:

**Step 1.** The central atom in the nitrite ion is nitrogen.

O N O

**Step 2.** Add up the number of valence electrons available. Add 1 electron for each unit of negative charge.

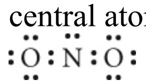
1 N atom: $1(5e^-) = 5e^-$
2 O atoms: $2(6e^-) = 12e^-$
charge: 1e–
Total: $18e^-$

**Step 3.** Place 1 pair of electrons between each pair of adjacent atoms to represent the bonding electron pairs.

O : N : O

Start: $18e^-$
Used: $4e^-$

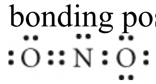
**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding and central atoms.



**Step 5.** Determine how many electrons are still available.

Start: 18 $e^-$
Used 18 $e^-$
Remaining: $18e^- - 18e^- = 0e^-$

**Step 6.** The N atom does not have a full octet. Move a lone pair on one O atom into a bonding position between the two atoms so the N atom has a complete octet.



**Step 7.** Draw the structural formula. Since $NO_2^-$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.

$\left[\text{O}=\text{N}-\text{O}\right]^-$

**(d)** Lewis structure for $ClO^-$:

**Step 1.** There is one Cl atom and one O atom in the hypochlorite ion.

Cl O

**Step 2.** Add up the number of valence electrons available. Add 1 electron for each unit of negative charge.

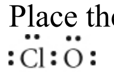
1 Cl atom: $1(7e^-) = 7e^-$
1 O atom: $1(6e^-) = 6e^-$
charge: 1e–
Total: $14e^-$

**Step 3.** Place 1 pair of electrons between the two atoms.

Cl : O

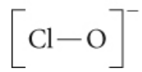
Start: $14e^-$
Used: $2e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the two atoms.



**Step 5.** Determine how many electrons are still available.
Start: $14e^{-}$
Used: $14e^{-}$
Remaining: $14e^{-} - 14e^{-} = 0e^{-}$

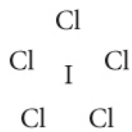
**Step 6.** Draw the structural formula. Since $ClO^{-}$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



### Tutorial 2 Practice, page 204

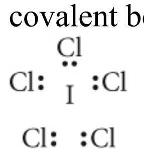
**1. (a)** Simplified Lewis structure for iodine pentachloride, $ICl_5$:

**Step 1.** Write the symbol of the central atom, and write the symbols for the other atoms around it.

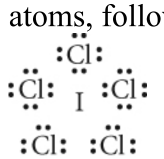


**Step 2.** Count all the valence electrons for all the atoms in the molecule.
1 I atom: $1(7e^{-}) = 7e^{-}$
5 Cl atoms: $5(7e^{-}) = 35e^{-}$
Total: $42e^{-}$

**Step 3.** Place one pair of electrons between each adjacent pair of atoms, forming single covalent bonds.

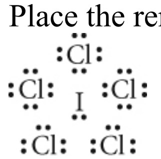


**Step 4.** Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.

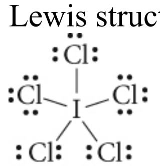


**Step 5.** Determine the number of electrons still available.
Start: $42e^{-}$
Used: $40e^{-}$
$42e^{-} - 40e^{-} = 2e^{-}$ left

**Step 6.** Place the remaining pair of electrons on the central atom.



**Step 7.** Replace the dots between atoms with lines representing bonds to simplify the Lewis structure.



**(b)** Simplified Lewis structure for radon dichloride, $RnCl_2$:

**Step 1.** Write the symbol of the central atom, and write the symbols for the other atoms around it.

Cl Rn Cl

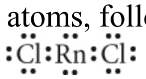
**Step 2.** Count all the valence electrons for all the atoms in the molecule.

1 Rn atom: $1(8e^-) = 8e^-$
2 Cl atoms: $2(7e^-) = 24e^-$
Total: $32e^-$

**Step 3.** Place one pair of electrons between each adjacent pair of atoms, forming single covalent bonds.

Cl:Rn:Cl

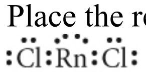
**Step 4.** Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.



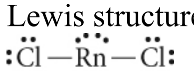
**Step 5.** Determine the number of electrons still available.

Start: $32e^-$
Used: $30e^-$
$32e^- - 30e^- = 2e^-$ left

**Step 6.** Place the remaining pair of electrons on the central atom.



**Step 7.** Replace the dots between atoms with lines representing bonds to simplify the Lewis structure.



**2. (a)** Simplified Lewis structure for chlorine trifluoride, $ClF_3$:

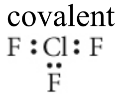
**Step 1.** Write the symbol of the central atom, and write the symbols for the other atoms around it.

F Cl F
F

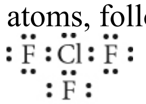
**Step 2.** Count all the valence electrons for all the atoms in the molecule.

1 Cl atom: $1(7e^-) = 7e^-$
3 F atoms: $3(7e^-) = 21e^-$
Total: $28e^-$

**Step 3.** Place one pair of electrons between each adjacent pair of atoms, forming single covalent bonds.



**Step 4.** Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.

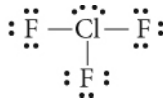


**Step 5.** Determine the number of electrons still available.

Start: $28e^-$
Used: $26e^-$
$28e^- - 26e^- = 2e^-$ left

**Step 6.** Place the remaining pair of electrons on the central atom.

**Step 7.** Replace the dots between atoms with lines representing bonds to simplify the Lewis structure.



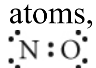
**(b)** Simplified Lewis structure for the nitrosonium ion, $NO^+$:

**Step 1.** There is one N atom and one O atom in the polyatomic ion.
N O

**Step 2.** Count the valence electrons for the two atoms in the molecule. Subtract 1 electron for each unit of positive charge.
1 N atom: $1(5e^-) = 5e^-$
1 O atom: $1\ (6e^-) = 6e^-$
positive charge: $-1e^-$
Total: $10e^-$

**Step 3.** Place one pair of electrons between the two atoms, forming a single covalent bond.
N : O

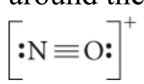
**Step 4.** Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.



**Step 5.** Determine the number of electrons still available.
Start: $10e^-$
Used: $10e^-$
$10e^- - 10e^- = 0e^-$ left

**Step 6.** The atoms do not have full octets. Move a lone pair from each atom into a bonding position between the two atoms so both atoms have a complete octet.
:N:::O:

**Step 7.** Draw the structural formula. Since $NO^+$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



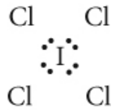
**(c)** Simplified Lewis structure for the iodine tetrachloride ion, $ICl_4^-$:

**Step 1.** Write the symbol of the central atom, and write the symbols for the other atoms around it.

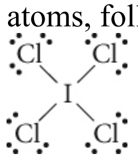
Cl Cl
I
Cl Cl

**Step 2.** Count all the valence electrons for all the atoms in the molecule. Add 1 electron for each unit of negative charge.
1 I atom: $1(7e^-) = 7e^-$
4 Cl atoms: $4(7e^-) = 28e^-$
negative charge: $1e^-$
Total: $36e^-$

**Step 3.** Place one pair of electrons between each adjacent pair of atoms, forming single covalent bonds.

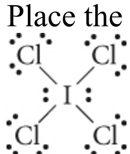


**Step 4.** Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.

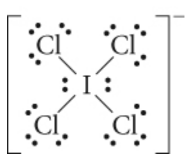


**Step 5.** Determine the number of electrons still available.
Start: $36e^-$
Used: $32e^-$
$36e^- - 32e^- = 4e^-$ left

**Step 6.** Place the remaining electrons on the central atom in pairs.



**Step 7.** Replace the dots between atoms with lines representing bonds to simplify the Lewis structure. Since $ICl_4^-$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



## Section 4.1 Questions, page 205

**1. (a)** The bonds between the carbon and fluoride atoms in carbon tetrafluoride are covalent.

**(b)** Both calcium fluoride and carbon tetrafluoride involve the element fluorine. However, carbon tetrafluoride involves fluorine atoms covalently bonded to a central carbon atom while calcium fluoride contains fluoride anions ionically bonded to a calcium cation.

**2.** Ammonium sulfate, $(NH_4)_2SO_4$: $\Delta EN_{N–H} = 3.0 – 2.2 = 0.8$, so the N–H bonds are covalent. $\Delta EN_{O–S} = 3.4 – 2.6 = 1.2$, so the S–O bonds are covalent. Sulfate, $SO_4^{2-}$, is a polyatomic ion. So, there is covalent bonding between sulfur and oxygen in the sulfate ion and between nitrogen and hydrogen in the ammonium ion as well as ionic bonding between ammonium and sulfate.

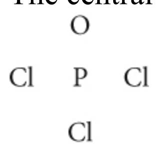
Calcium phosphate, $Ca_3(PO_4)_2$: phosphate, $PO_4^{3-}$, is a polyatomic ion, so $Ca_3(PO_4)_2$ is an ionic compound. $\Delta EN_{O–P} = 3.4 – 2.2 = 1.2 \leq 1.7$, so the P–O bonds are covalent. There is covalent bonding between phosphorous and oxygen as well as ionic bonding between calcium and phosphate.

$K_2O$: $\Delta EN_{O–K} = 3.4 – 0.8 = 2.2 \geq 1.7$; therefore $K_2O$ is an ionic compound.

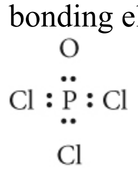
$P_2O_5$: $\Delta EN_{O–P} = 1.2 \leq 1.7$; therefore the P–O bonds are covalent.

KCl: $\Delta EN_{Cl–K} = 3.2 – 0.8 = 2.0 \geq 1.7$; therefore KCl is an ionic compound.

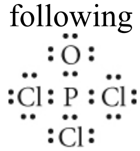
**(a)** $(NH_4)_2SO_4$, $Ca_3(PO_4)_2$, $K_2O$, and KCl contain ionic bonds
**(b)** $(NH_4)_2SO_4$, $Ca_3(PO_4)_2$, and $P_2O_5$ contain covalent bonds.
**(c)** $(NH_4)_2SO_4$, and $Ca_3(PO_4)_2$ contain both ionic and covalent bonds.
**3. (a)** Aluminum oxide has high electrical conductivity when molten because when an ionic compound melts, the ions are free to move, and thus to carry electric charges. (In an ionic solid, the ions are held together too tightly for them to move, so they cannot conduct an electric current.) If electric current is applied to the molten compound, the positively charged ions move toward the cathode while the negatively charged ions move toward the anode.
**(b)** Aluminum oxide has a high melting point because the large charge difference between the two ions in the compound ($Al^{3+}$ and $O^{2-}$) results in a particularly strong ionic bond. Consequently, a great deal of thermal energy is required to break this bond and cause aluminum oxide to melt.
**4. (a)** Lewis structure for phosphoryl trichloride, $POCl_3$:
**Step 1.** The central atom is phosphorus.



**Step 2.** Add up the number of valence electrons available.
1 P atom: $1(5e^-) = 5e^-$
1 O atom: $1\ (6e^-) = 6e^-$
3 Cl atoms: $3(7e^-) = 21e^-$
Total: $32e^-$
**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms to represent the bonding electron pairs.

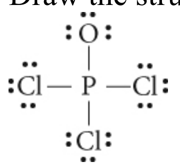


Start: $32e^-$
Used: $8e^-$
**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.

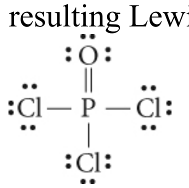


**Step 5.** Determine how many electrons are still available.
Start: $32e^-$
Used: $32e^-$
Remaining: $32e^- - 32e^- = 0$

**Step 6.** Draw the structural formula.



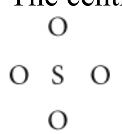
This molecule obeys the octet rule. However, experimental evidence shows that phosphorus tends to form five bonds rather than four. It is thought this occurs as a result of the overlap of an orbital containing an oxygen electron pair with an empty d orbital in phosphorus, resulting in a coordinate covalent bond. The resulting Lewis structure consistent with this theory is:



The prediction of this Lewis structure is beyond the scope of this course.

**(b)** Lewis structure for the sulfate ion, $SO_4^{2-}$:

**Step 1.** The central atom is sulfur.



**Step 2.** Add up the number of valence electrons available.

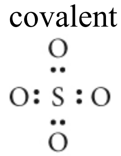
1 S atom: $1(6e^-) = 6e^-$

4 O atoms: $4(6e^-) = 24e^-$

negative charge: $2e^-$

Total: $32e^-$

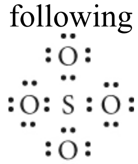
**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms, forming single covalent bonds.



Start: $32e^-$

Used: $8e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.



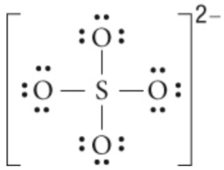
**Step 5.** Determine how many electrons are still available.

Start: $32e^-$

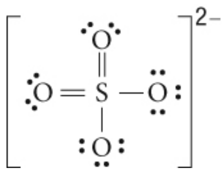
Used: $32e^-$

Remaining: $32e^- - 32e^- = 0$

**Step 6.** Draw the structural formula. Since $SO_4^{2-}$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



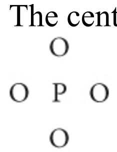
This ion obeys the octet rule. However, experimental evidence shows that sulfur tends to form six bonds rather than four. Further evidence shows that two of the sulfur-oxygen bonds are actually double bonds. The Lewis structure that agrees with this evidence is:



The prediction of this Lewis structure is beyond the scope of this course.

**(c)** Lewis structure for the phosphate ion, $PO_4^{3-}$:

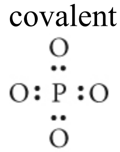
**Step 1.** The central atom is phosphorus.



**Step 2.** Add up the number of valence electrons available.

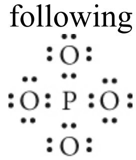
1 P atom: $1(5e^-) = 5e^-$
4 O atoms: $4(6e^-) = 24e^-$
negative charge: $3e^-$
Total: $32e^-$

**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms, forming single covalent bonds.



Start: $32e^-$
Used: $8e^-$

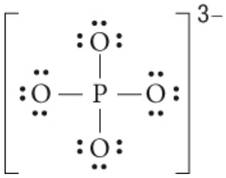
**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.



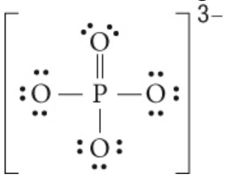
**Step 5.** Determine how many electrons are still available.

Start: $32e^-$
Used: $32e^-$
Remaining: $32e^- - 32e^- = 0$

**Step 6.** Draw the structural formula. Since $PO_4^{3-}$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.



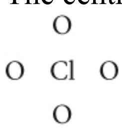
This ion obeys the octet rule. However, experimental evidence shows that phosphorus tends to form five bonds rather than four. Further evidence shows that one of the phosphorus-oxygen bonds is actually a double bond. The Lewis structure that agrees with this evidence is:



The prediction of this Lewis structure is beyond the scope of this course.

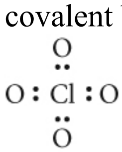
**(d)** Lewis structure for the perchlorate ion, $ClO_4^-$:

**Step 1.** The central atom is chlorine.



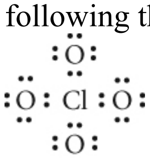
**Step 2.** Add up the number of valence electrons available.
1 Cl atom: $1(7e^-) = 7e^-$
4 O atoms: $4(6e^-) = 24e^-$
negative charge: $1e^-$
Total: $32e^-$

**Step 3.** Place 1 pair of electrons between each adjacent pair of atoms, forming single covalent bonds.



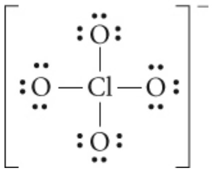
Start: $32e^-$
Used: $8e^-$

**Step 4.** Place the remaining valence electrons as lone pairs on the surrounding atoms, following the octet rule.

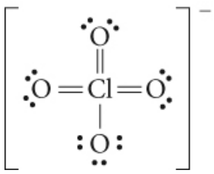


**Step 5.** Determine how many electrons are still available.
Start: $32e^-$
Used: $32e^-$
Remaining: $32e^- - 32e^- = 0$

**Step 6.** Draw the structural formula. Since $ClO_4^-$ is a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets.

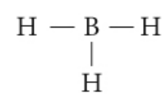


This ion obeys the octet rule.
However, experimental evidence shows that the perchlorate ion contains three chlorine-oxygen double bonds. The Lewis structure that agrees with this evidence is:



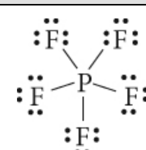
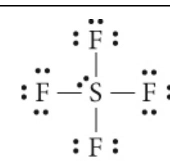
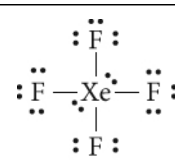
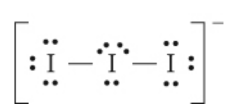
The prediction of this Lewis structure is beyond the scope of this course.

**5. Some Common Exceptions to the Octet Rule**

| | Substance | Valence electrons | Lewis structure |
|---|---|---|---|
| **(a)** | beryllium hydride, $BeH_2$ | 1 Be atom: $1(2e^-) = 2e^-$<br>2 H atoms: $2(1e^-) = 2e^-$<br>Total: $4e^-$ | H — Be — H |
| **(b)** | borane, $BH_3$ | 1 B atom: $1(3e^-) = 3e^-$<br>3 H atoms: $3(1e^-) = 3e^-$<br>Total: $6e^-$ |  |

**6. (a)**

**Some Common Exceptions to the Octet Rule**

| Substance | Valence electrons | Lewis structure |
|---|---|---|
| phosphorus pentafluoride, $PF_5$ | 1 P atom: $1(5e^-) = 5e^-$<br>5 Cl atoms: $5(7e^-) = 35e^-$<br>Total: $40e^-$ |  |
| sulfur tetrafluoride, $SF_4$ | 1 S atom: $1(6e^-) = 6e^-$<br>4 F atoms: $4(7e^-) = 28e^-$<br>Total: $34e^-$ |  |
| xenon tetrafluoride, $XeF_4$ | 1 Xe atom: $1(8e^-) = 8e^-$<br>4 F atoms: $4(7e^-) = 28e^-$<br>Total: $36e^-$ |  |
| tri-iodide ion, $I_3^-$ | 3 I atoms: $3(7e^-) = 21e^-$<br>negative charge: $1e^-$<br>Total: $22e^-$ |  |

**(b)** The following atoms can have more than 8 electrons around them: P in $PF_5$; S in $SF_4$; Xe in $XeF_4$; and I in $I^{3-}$.
**(c)** These elements all have vacant *d* orbitals that they can use to exceed their octet.
**7.** Answers will vary. Sample answer: Coordinate covalent bonding is like you and a friend each needing to bring an identical item, and if your friend does not bring his or hers, then you must supply both. This is like a coordinate covalent bond—one atom is electron deficient so the other atom supplies both electrons for the bonding to occur.