Chapter 9

Basic Concepts of Chemical Bonding

9.1 Outline

- Lewis symbols and valence electrons
- Ionic bonding - electrostatic attractions between ions of opposite charge
- Covalent bonding - sharing of one or more electron pairs between atoms
- Bond polarity and electronegativity
- Drawing Lewis structures
- Resonances structures, exceptions to the octet rule, and strengths of covalent bonds

9.2 Chemical Bonds

Chemical bond – a strong attractive fource that exists between atoms in a molecule.
The three types of chemical bonds are as follows:

ionic bond a bond between oppositely charged ions. The ions are formed from atoms by transfer of one or more electrons.

covalent bond a bond formed between two or more atoms by sharing of electrons.

metallic bond Bonding, usually in solid metals, in which bonding electrons are relatively free to move throughout the three-dimensional structure.

9.3 Lewis Symbols

- The valence electrons, those that reside in the outermost shell of an atom, are responsible for chemical bonding.
- Lewis symbol (electron dot symbol) The chemical symbol for an element, with a dot for each valence electron.

- Dots are placed on the four sides of the chemical symbol, where each side can accommodate up to two electrons.

Table 9.1: Lewis Symbols

Electron	**Electron Configuration**	**Lewis Symbol**	**Electron**	**Electron Configuration**	**Lewis Symbol**
Li	$[He]2s^1$	Li·	Na	$[Ne]3s^1$	Na·
Be	$[He]2s^2$	·Be·	Mg	$[Ne]3s^2$	·Mg·
B	$[He]2s^2 2p^1$	·Ḃ·	Al	$[Ne]3s^2 3p^1$	·Ȧl·
C	$[He]2s^2 2p^2$	·Ċ̣·	Si	$[Ne]3s^2 3p^2$	·Ṩi·
N	$[He]2s^2 2p^3$	·Ṇ̇:	P	$[Ne]3s^2 3p^3$	·Ṗ̣:
O	$[He]2s^2 2p^4$	:Ọ̇:	S	$[Ne]3s^2 3p^4$	:Ṩ:
F	$[He]2s^2 2p^5$	·F̤̈:	Cl	$[Ne]3s^2 3p^5$	·C̤̈l:
Ne	$[He]2s^2 2p^6$	:N̤̈e:	Ar	$[Ne]3s^2 3p^6$	:Ä̤r:

9.4 Ionic Bonding

- The combination of sodium metal and chlorine gas results in a violent reaction. The product of this very exothermic reaction is sodium chloride ($NaCl$).

$$Na\,(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s) \qquad \Delta H_f = -410.9\,kJ$$

- Sodium chloride is comprised of Na^+ and Cl^- ions (see Figure 9.1).
- Recall that metals (e.g., Na) have a tendency to lose electrons to form cations, whereas nonmetals (e.g., Cl) gain electrons to become anions.

Na· + ·C̤̈l: ⟶ Na^+ + [:C̤̈l:]⁻

9.5 Structure of Sodium Chloride (NaCl)

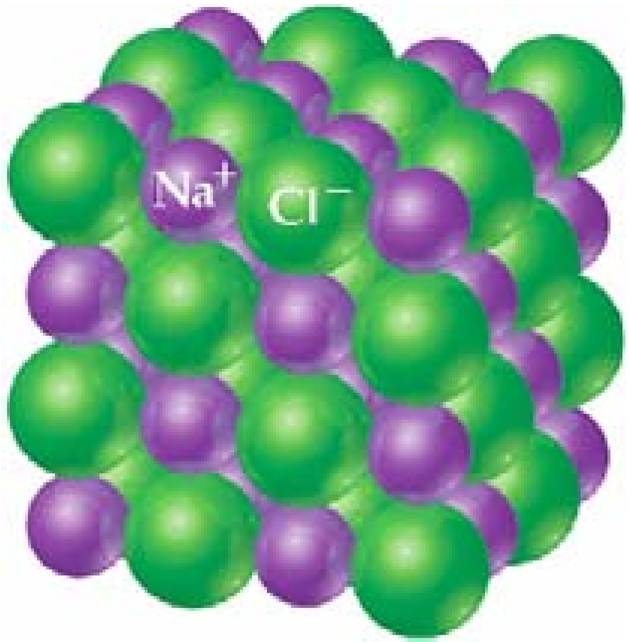


Figure 9.1: Structure of Sodium Chloride (NaCl)

In this three-dimensional array of ions, each Na^+ ion is surrounded by siz Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions.

9.6 Formation of NaCl

- What drives the reaction between Na (s) and Cl_2 (g) to form NaCl?
- Na has a relatively low first ionization energy (i.e., it is perfectly happy to give up its valence electron). Ask yourself, what is the electron configuration of the resulting Na^+ cation?
- Cl has a strong tendency to gain an electron, which is manifested in its large negative electron affinity ($E_a = -349\frac{\text{kJ}}{\text{mol}}$). Ask yourself, what is the electron configuration of the resulting Cl^- cation?
- In addition, as the Na^+ and Cl^- ions are drawn together to form NaCl, a substantial amount of energy is released, known as the lattice energy.

9.7 Lattice Energy

- Lattice energy – The energy required to completely separate a mole of a solid ionic compound into its gaseous ions.

- Lattice energies are positive as energy is required to overcome attractive forces between oppositely charged ions. For example,

$$\mathrm{NaCl(s)} \longrightarrow \mathrm{Na^{+}(g) + Cl{-}(g)} \qquad \Delta H_{\mathrm{lattice}} = +788\,\frac{\mathrm{kJ}}{\mathrm{mol}}$$

- The energy associated with electrostatic interactions is governed by Coulomb's Law:

$$E_{el} = \frac{\kappa Q_1 Q_2}{d} \tag{9.1}$$

- Lattice energy increases with the charge on the ions.
- It also increases with decreasing size of ions.
- See the worked example entitled Magnitudes of Lattice Energies.

9.8 Magnitudes of Lattice Energies

Which substance would you expect to have the greatest lattice energy, MgF_2, CaF_2, or ZrO_2?

$$\mathrm{MgF_2(s)} \longrightarrow \mathrm{Mg^{2+}(g) + 2\,F^{-}(g)}$$

Because the product of the charge, Q_1Q_2, appears in the numerator of the equation above, the lattice energy will increase dramatically when the charges of the ions increase. Thus,

$$\begin{array}{rlll} \mathrm{MgF_2} & Q_1 = {}^{+}2 & Q_2 & = -1 \\ \mathrm{CaF_2} & Q_1 = {}^{+}2 & Q_2 & = -1 \\ \mathrm{ZrO_2} & Q_1 = {}^{+}4 & Q_2 & = -2 \end{array}$$

$$\mathrm{CaF_2 < MgF_2 < ZrO_2}$$

Table 9.2: Lattice Energies for Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
LiF	1030	$MgCl_2$	2326
LiCl	834	$SrCl_2$	2127
LiI	730		
NaF	910	MgO	3795
NaCl	788	CaO	3414
NaBr	732	SrO	3217
NaI	682		
KF	808	ScN	7547
KFCl	701		
KBr	671		
CsCl	657		
CsI	600		

9.9 Covalent Bonding

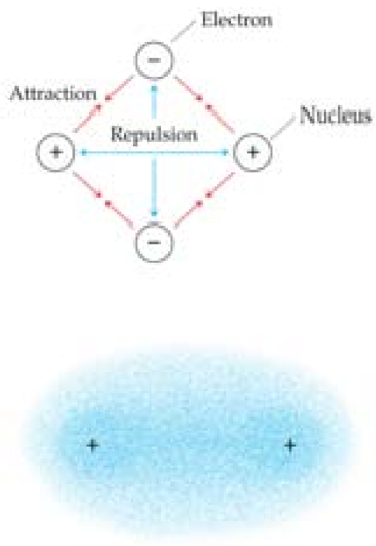


Figure 9.2: Covalent Bonding

- In covalent bonds, atoms share electrons.
- There are several electrostatic interactions in these bonds:
  - Attractions between electrons and positive nuclei.
  - Repulsions between electrons
  - Repulsions between nuclei
  - Attractive forces must outweigh the repulsive ones

9.10 Lewis Structures

- Consider two Hydrogen atoms coming together to form a covalently bonded H_2 molecule:

$$\mathrm{H}\cdot + \cdot\mathrm{H} \longrightarrow \mathrm{H}{:}\mathrm{H}$$

- The H_2 molecule on the right, with its two electrons, exhibits the noble-gas configuration of Helium. The shared pair of electrons can be represented by a single bond, as shown below¿

$$\mathrm{H}{:}\mathrm{H} = \mathrm{H} \text{——} \mathrm{H}$$

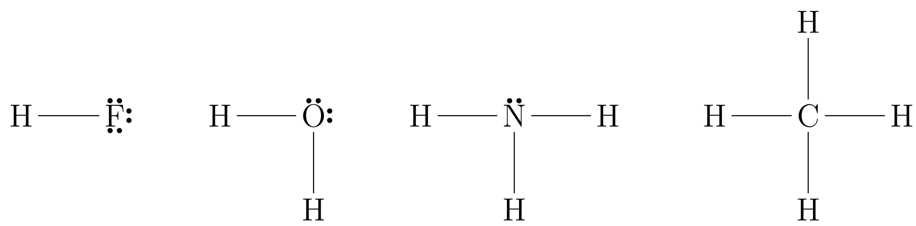
- Consider two chlorine atoms coming together to form a covalently bonded Cl_2 molecule:

$$:\underset{..}{\ddot{\mathrm{Cl}}}\cdot + \cdot\underset{..}{\ddot{\mathrm{Cl}}}: \longrightarrow :\underset{..}{\ddot{\mathrm{Cl}}}{:}\underset{..}{\ddot{\mathrm{Cl}}}:$$

- Each chlorine atom on the right now has a *complete octet* of electrons by sharing the bonding electron pair. It achieves the noble gas configuration of argon (Ar). Again, the shared pair of electrons can be represented by a single bond, as shown below.

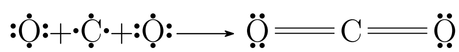
9.11 Typical Bonding Motifs

Typical bonding motifs above: hydrogen and halogens typically form one bond, oxygen two bonds, nitrogen three bonds, and carbon four bonds.

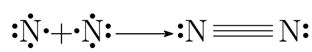


9.12 Multiple Bonds

- Double bonds in CO_2:



- Triple bond in N_2:



9.13 Bond Polarity and Electronegativity

- Molecules such as H_2, N_2, Cl_2, etc are said to be **nonpolar**.
- A nonpolar covalent bond is one in which the electrons are shared equally between two atoms.
- On the other hand, a polar covalent bond is one in which one of the atoms exerts a greater attraction for the bonding electrons than the other.
- In other words, there exists a bond between atoms of different electronegativities (e.g., O-H, C-H, N-H, etc.).

9.14 Electronegativity

- Electronegativity – the ability of atoms in a molecule to attract electrons to themselves.
- On the periodic table, electronegativity increases as you go. . .
  - from left to right across a row
  - from bottom to the top of a column.

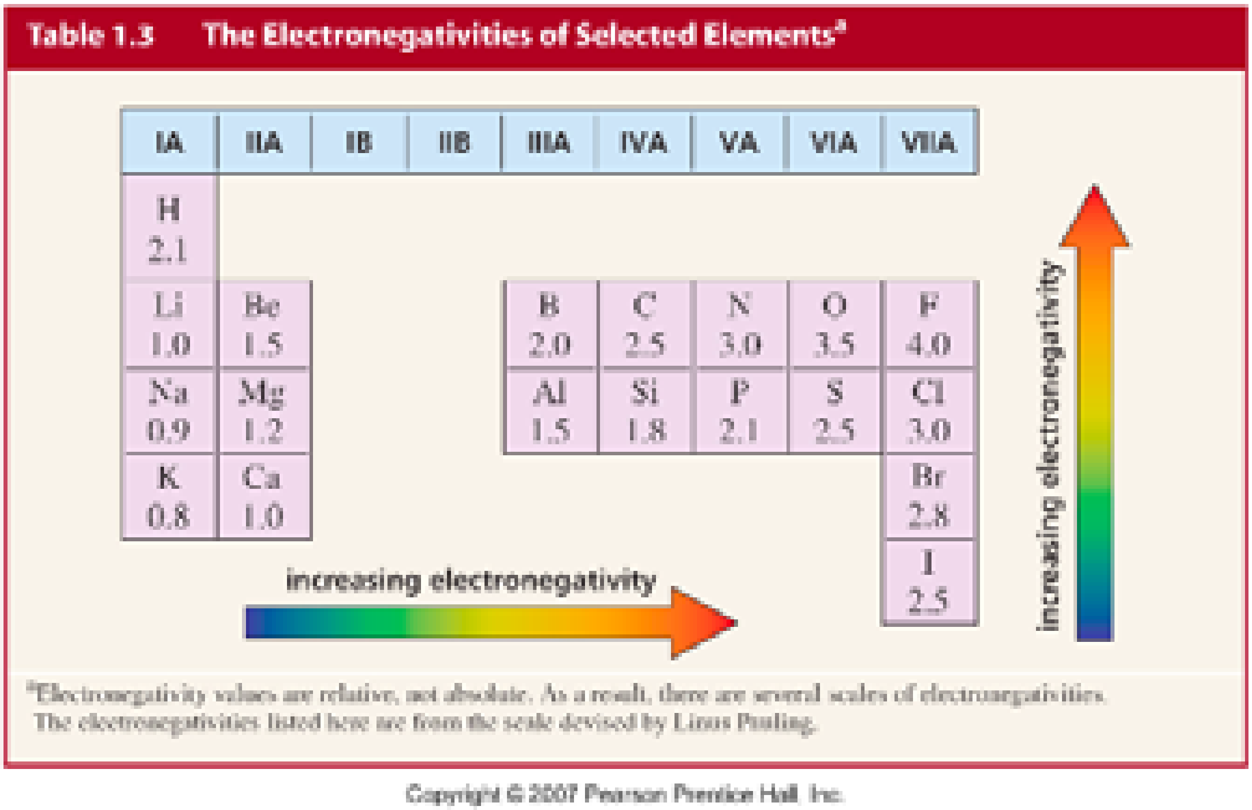


Table 1.3 The Electronegativities of Selected Elements[a]

IA	IIA	IB	IIB	IIIA	IVA	VA	VIA	VIIA
H 2.1								
Li 1.0	Be 1.5			B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2			Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0							Br 2.8
								I 2.5

[a]Electronegativity values are relative, not absolute. As a result, there are several scales of electronegativities. The electronegativities listed here are from the scale devised by Linus Pauling.

Figure 9.3: The Electronegativities of Selected Elements

9.15 Polar Covalent Bonds

- Though atoms often form compounds by sharing electrons, the electrons are not always shared equally.
- Fluorine pulls harder on the electrons it shared with Hydrogen than Hydrogen does.
- Therefore, the Fluorine end of the molecule has more electron density than the hydrogen end.

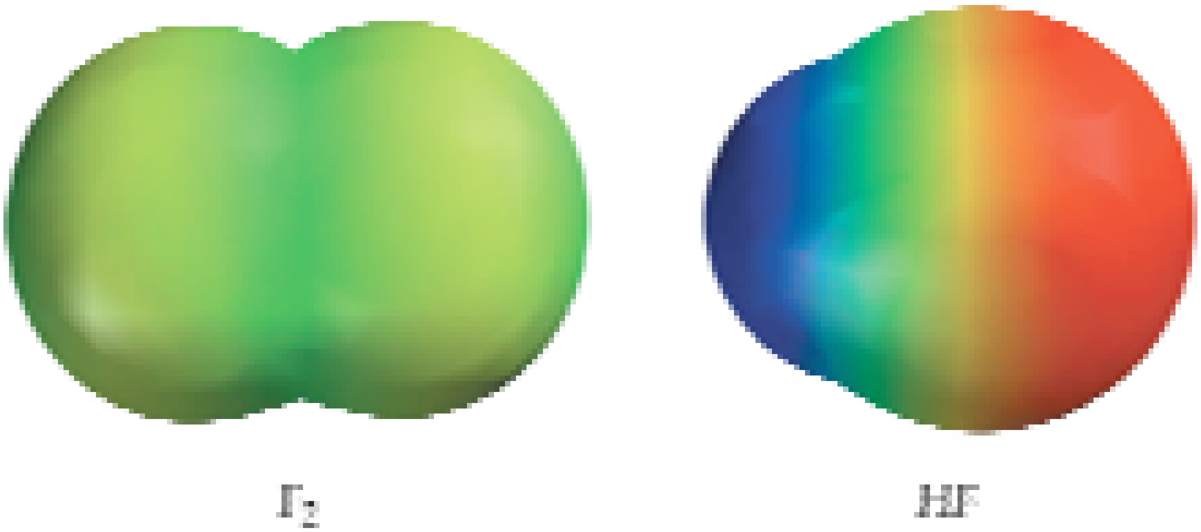


Figure 9.4: Polar Covalent Bonds

Table 9.3: Electronegativity and Bond Polarity

Compound	F_2	HF	LiF
Electronegativity	4.0 - 4.0 = 0	4.0 - 2.1 = 1.9	4.0 - 1.0 = 3.0
Type of bond	Nonpolar covalent	Polar covalent	Ionic

- When two atoms share electrons unequally, a bond dipole results.
- The dipole moment, μ, produced by two equal but opposite charges separated by a distance r is calculated as:

$$\mu = Qr \tag{9.2}$$

- The dipole is measured in debye (D) units.

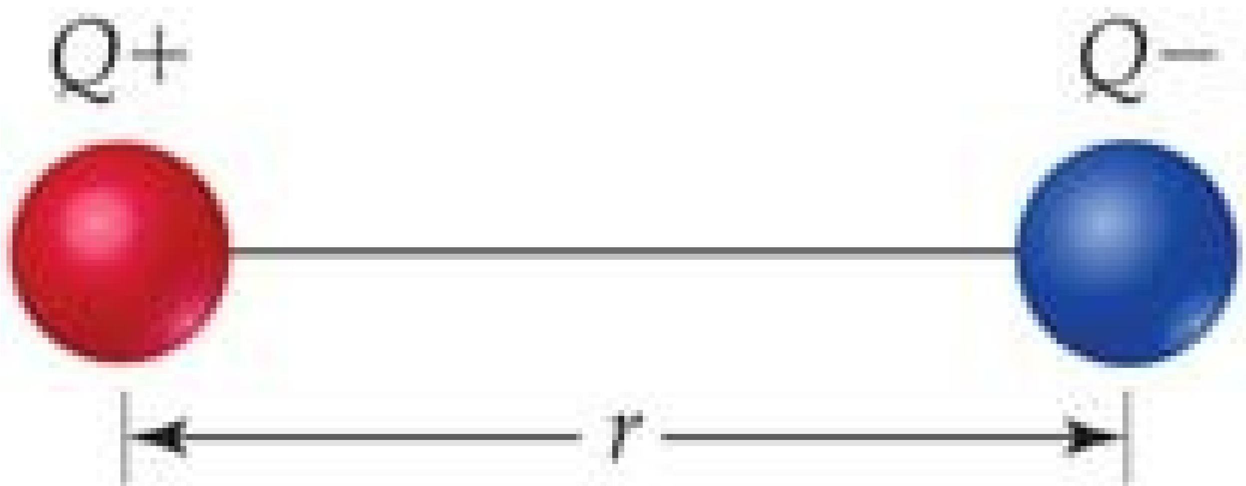


Figure 9.5: Dipole Moment μ

Table 9.4: Polar Covalent Bonds

Compound	**Bond Length (Å)**	**Electronegativity**	**Dipole Moment (D)**
HF	0.92	1.9	1.82
HCl	1.27	0.9	1.08
HBr	1.41	0.7	0.82
HI	1.61	0.4	0.44

The greater the difference in electronegativity, the more polar the bond is.

9.16 Writing Lewis Structures

- Lewis structures are representations of molecules showing all electrons, bonding and nonbonding.

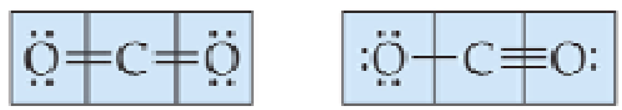
1. Find the sum of valence electrons of all atoms in the polyatomic ion or molecule.
   - If it is an anion, add one electron for each negative charge.
   - If it is a cation, subtract one electron for each positive charge.
   - Total number of valence electrons for PCl_3 is:____ $5 + (7*3) = 26$
2. The central atom is the least electronegative element that isn't Hydrogen. Connect the other atoms to it by single bonds.
3. Fill the octets of the outer atoms.
   - How many electrons have you accounted for in the above structure? 24
   - How many do you have left? 2

- Fill in the octet of the central atom.
- If you run out of electrons before the central atom has an octet: form multiple bonds until it does

$$H - C - \ddot{\underset{\cdot\cdot}{N}}: \longrightarrow H - C \equiv N:$$

- Then assign formal charges.
- For each atom, count the electrons in lone pairs and half the electrons it shares with other atoms.
- Subtract that from the number of valence electrons for that atom: the difference is its formal charge (see (9.3)).

$$\text{formal charge} = \#\text{ of valence electrons} - \#\text{ of lone-pair electrons} - \frac{1}{2}\#\text{ of bonding electrons} \tag{9.3}$$



Valence electrons:	6	4	6	6	4	6
−(Electrons assigned to atom):	6	4	6	7	4	5
Formal charge:	0	0	0	−1	0	+1

Figure 9.6: Formal Charges for CO_2

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- The best Lewis structure...
  - ... is the one with the fewest charges.
  - ... puts a negative charge on the most electronegative atom.

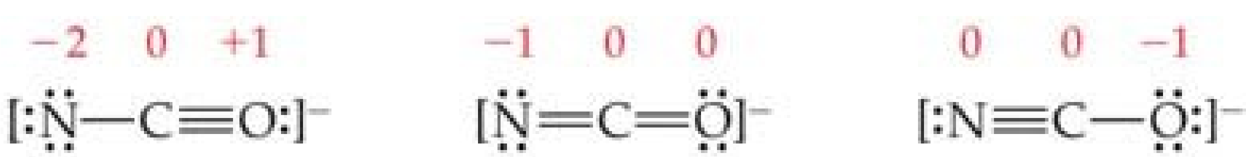
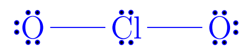


Figure 9.7: Choosing the best formal charge for CON.

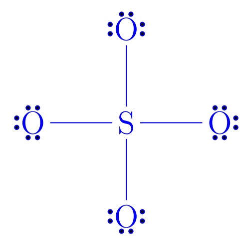
9.17 Lewis Structures for Polyatomic Ions

Draw the Lewis structures for:

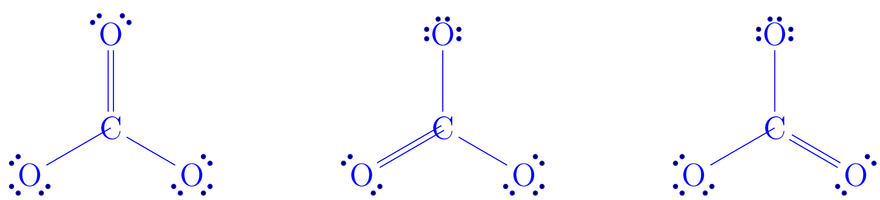
(a) ClO_2^- $7 + 2(6) + 1 = 20$ valence electrons



(b) SO_4^{2-} $6 + 4(6) + 2 = 32$ valence electrons

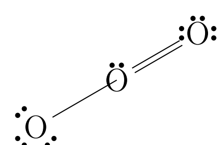


(c) CO_3^{2-} $4 + 3(6) + 2 = 24$ valence electrons



9.18 Resonance

- One possible Lewis structure of the ozone (O_3) molecule is depicted below.



- But this is not consistent with the observed structure of ozone, whereby...
  - both O−O bonds are the same length.
- One Lewis structure cannot accurately describe a molecule like ozone.
- We use multiple structures (i.e. resonance structures) to describe the molecule.
- Placement of atoms remain the same, but the placement of electrons are different.
- A double headed arrow is used to indicate that the real molecule is described by an average of the resonance structures.

9.18.1 An Analogy to Color Mixing

Just as green is a *blend* of blue and yellow... the real structure of ozone is a *blend* of the two resonance structures.

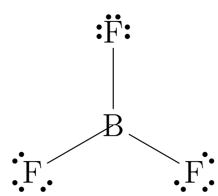
9.19 Exceptions to the Octet Rule

- The three types of systems that don't follow the octet rule are as follows:
  - Ions or molecules with an odd number of electrons
  - Ions or molecules with less than an octet
  - Ions or molecules with more than eight valence electrons (an expanded octet)

9.19.1 Odd Number of Electrons

Ions and molecules with an odd number of electrons (e.g., ClO_2, NO, NO_2, and O_2).

9.19.2 Fewer than Eight Electrons



- Consider BF_3
  - Notice how the central boron does not have a complete octet?
  - What about resonance?
  - See worked example below:

Draw the Lewis structure for boron trifluoride, BF_3, and explain why it does not obey the octet rule.
These resonance structures are less important than the first because they put positive charge on the most electronegative fluorine atoms!

9.19.3 More than Eight Electrons

- The only way PCl_5 can exist is if phosphorus has 10 valence electrons around it.
- It is allowed to expand the octet of atoms on or below the 3rd row/period.
- Ask yourself, does NCl_5 exist? No.
- Most likely the *d* orbitals in these atoms participate in bonding.

9.20 Covalent Bond Strength

$$:\underset{\cdot\cdot}{\ddot{Cl}}-\underset{\cdot\cdot}{\ddot{Cl}}:(g) \longrightarrow 2:\underset{\cdot\cdot}{\ddot{Cl}}\cdot(g)$$

- The strength of a bond is measured by determining how much energy is required to break the bond.
- This is the bond enthalpy.
- The bond enthalpy for a Cl-Cl bond is measured to be 242 kJ/mol.