4.1 Types of Chemical Bonds

With a few rare exceptions, everything that you can see, taste, or touch is made up of atoms or ions of elements connected to other atoms or ions by chemical bonds. Chemical bonds hold the world as we know it together. What exactly is a chemical bond? In earlier science courses, you learned that a chemical bond is the electrical attraction that holds atoms or ions together in a molecular element or in an ionic or molecular compound (**Figure 1**). When bonds form, the atoms or ions that are joined function as a unit. Atoms combine to form molecular elements or compounds. Molecular elements are molecules that consist of atoms of the same element, such as nitrogen gas, $N_2(g)$. Molecular compounds comprise atoms of 2 or more different elements. Finally, an ionic compound is a pure substance composed of 2 or more ions combined in a fixed ratio.

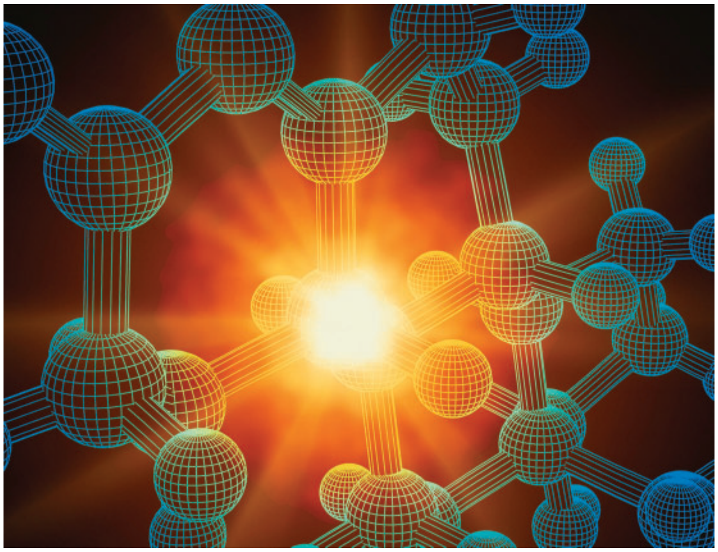


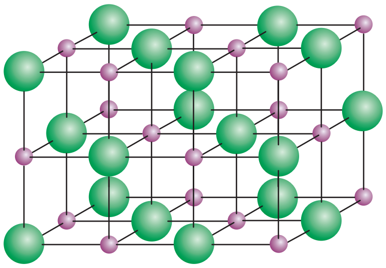
Figure 1 Artist's representation of chemical bonds that hold atoms together

Ionic Compounds and Ionic Bonding

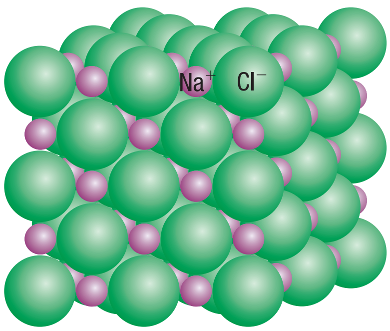
ionic bond the electrostatic attraction between oppositely charged ions

When solid sodium, Na(s), and chlorine gas, $Cl_2(g)$, react to form solid sodium chloride, NaCl(s), a transfer of electrons occurs from the sodium atoms to the chlorine atoms to form sodium ions, Na^+, and chloride ions, Cl^-. The electron transfer occurs because chlorine atoms have a much higher electronegativity than sodium atoms, and hence have a very strong attraction for the sodium atoms' single valence electron. The sodium and chloride ions now have opposite electric charges. As you know, oppositely charged objects attract each other. An **ionic bond** is a chemical bond between oppositely charged ions. Ionic compounds form when an atom that loses electrons relatively easily reacts with an atom that gains electrons relatively easily. Typically, an ionic compound results when a metal element reacts with a non-metal element to form positively charged and negatively charged ions that are held together by electrostatic attraction.

The electrostatic attraction between positively charged sodium ions and negatively charged chloride ions creates a crystal lattice of sodium chloride, NaCl(s) (**Figure 2**). The chemical formula for sodium chloride, NaCl, is an example of a formula unit: It represents the smallest quantity of sodium chloride that has this chemical formula. The formula unit of sodium chloride indicates that sodium and chlorine ions combine in a 1:1 ratio. Note that crystal formation only occurs when a large number of positive and negative ions come together. This configuration is more stable than if the ions were separated into individual formula units. The ions in a crystal lattice are arranged so that attractions between opposite charges are maximized and repulsions between like charges are minimized (Figure 2(b)).



(a)



(b)

Figure 2 The crystal lattice structure of sodium chloride. (a) Ball-and-stick model. Note that, while not shown here, 6 Cl^- ions surround each Na^+ ion. (b) Representation of the ions as space-filling spheres.

To illustrate the behaviour of electrons in ionic compounds, consider the ionic compound that forms from elemental calcium, Ca(s), and oxygen, O_2(g). You can predict the compound that will form by examining the valence electron configurations of both calcium and oxygen:

Ca: [Ar]$4s^2$

O: [He]$2s^22p^4$

Recall that the electron configurations of the noble gases are generally the most stable. Atoms and ions are more stable if they have the same number of electrons as the noble gas closest to them in the periodic table; that is, if they are **isoelectronic** with a noble gas. So, calcium would be more stable if it lost some electrons to become isoelectronic with argon ($1s^22s^22p^63s^23p^6$). Similarly, oxygen would be more stable if it gained some electrons to become isoelectronic with neon ($1s^22s^22p^6$). If calcium transfers 2 of its electrons to oxygen, both atoms can achieve a noble gas configuration for the electrons in their valence shells. The new electron configuration for oxygen now becomes the same as that of neon, because its valence orbital now has 8 electrons. By losing 2 electrons, calcium achieves the electron configuration of argon. This configuration results in calcium also having 8 electrons in its valence orbital.

isoelectronic having the same number of electrons per atom, ion, or molecule

$$\underbrace{\mathrm{Ca} + \mathrm{O}}_{2e^-} \longrightarrow \mathrm{Ca}^{2+} + \mathrm{O}^{2-}$$

To predict the formula of the ionic compound, remember that chemical compounds are always electrically neutral. The number of positive charges balances the number of negative charges. In this case, both Ca^{2+} and O^{2-} ions have equal but opposite charges. Therefore, the formula unit for calcium oxide is CaO. WEB LINK

Molecular Elements and Compounds, and Covalent Bonding

When a metal and a non-metal react to form oppositely charged ions, an ionic bond develops. What kind of bond develops between 2 atoms that are similar, or even identical? Consider what happens when two hydrogen atoms move close together.

Remember that a hydrogen atom has 1 proton and 1 electron. The attraction of each electron by the proton of the other hydrogen atom creates a force that pulls both atoms toward each other. When the hydrogen atoms are a certain optimum distance from each other, the attractive proton–electron force balances the repulsive proton–proton and electron–electron forces. The resulting hydrogen molecule is more stable and therefore has lower energy than the individual atoms that comprise it (**Figure 3**).

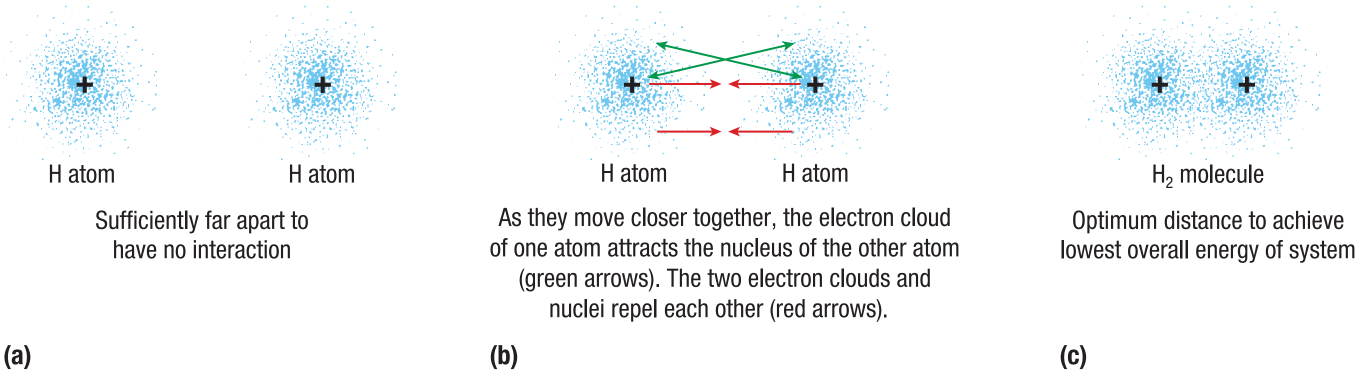


Figure 3 The interaction of two hydrogen atoms as they move closer together. (a) When the atoms are far apart, they have no interaction. (b) The atoms begin to interact as they move closer to each other. (c) The atoms interact to form the H_2 molecule at the distance that minimizes their energy.

covalent bond a chemical bond in which atoms share the bonding electrons

bonding electron pair an electron pair that is involved in bonding, found in the space between 2 atoms

The type of bond that forms when atoms share electrons is called a **covalent bond**. These shared electrons are called **bonding electron pairs**. Covalent bonds typically form between atoms of non-metal elements.

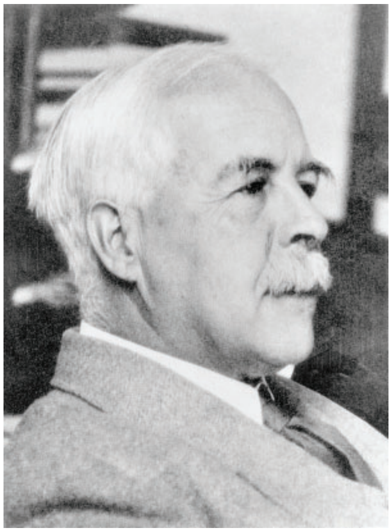


Figure 4 G.N. Lewis (1875–1946)

Lewis Theory of Bonding

In 1916, Gilbert Lewis (**Figure 4**) provided the first straightforward view of chemical bonding. He gathered information about chemical formulas, valence electrons, and models of the atom in an attempt to understand chemical bonds. From the data he collected, Lewis made predictions about the nature of bonding in what is now known as the Lewis theory of bonding. Ideas from this theory include the following:

- Atoms and ions are stable if they have a full valence shell of electrons.
- Electrons are most stable when they are paired.
- Atoms form chemical bonds to achieve a full valence shell of electrons.
- A full valence shell of electrons may be achieved by an exchange of electrons between metal and non-metal atoms.
- A full valence shell of electrons may be achieved by the sharing of electrons between non-metal atoms.
- The sharing of electrons results in a covalent bond.

Lewis Structures

Lewis structure a diagram that represents the arrangement of covalent electrons and bonds in a molecule or polyatomic ion

Based on his theory, Lewis drew structures to represent pure substances. A **Lewis structure** shows the arrangement of electrons and bonds in a molecule or polyatomic ion. After making observations of thousands of pure substances, chemists have identified rules for drawing Lewis structures. As you have already learned, the most important requirement for the formation of a stable substance is that the participating atoms are isoelectronic with noble gases. Hydrogen, in the first period of the periodic table, follows the **duet rule**: a hydrogen atom forms stable configurations when it shares 2 electrons, represented by dots, to obtain a full valence shell (**Figure 5**).

duet rule the observation that the complete outer shell of valence electrons when hydrogen and period 2 metals are involved in bonding

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

Figure 5 Two hydrogen atoms combine to form the H_2 molecule.

In this manner, hydrogen atoms become isoelectronic with helium atoms (**Figure 6**). A helium atom has 2 valence electrons so its valance shell is already full ($1s^2$). Consequently, helium does not form bonds.

$$\mathrm{He{:}}$$

Figure 6 Lewis symbol of a helium atom. A Lewis symbol shows the valence electrons in an atom of each element. Lewis symbols (of atoms) are used to draw Lewis structures (of pure substances).

octet rule the observation that many atoms tend to form the most stable substances when they are surrounded by 8 electrons in their valence shells

Atoms of non-metals in period 2 of the periodic table (carbon through fluorine) form stable configurations when they obey the octet rule. The **octet rule** states that many atoms form the most stable substances when they are isoelectronic with neon. Neon, a noble gas, does not form bonds because each neon atom already has an octet of valence electrons (**Figure 7**).

Figure 7 Lewis symbol of a neon atom

In atoms of period 2 non-metals, the valence shell consists of the single 2*s* and three 2*p* orbitals. Each orbital can hold at most 2 electrons, so 8 electrons fill the valence shell. For example, the F_2 molecule has a total of 8 valence electrons surrounding each atom (**Figure 8**). Each fluorine atom is now isoelectronic with neon.

Figure 8 An atom of fluorine has 7 electrons in its valence shell. If it shares one electron with another fluorine atom in a fluorine molecule, each atom will have a total of 8 valence electrons, and become isoelectronic with neon.

Even though each fluorine atom shares 1 of its electrons with the other fluorine atom in F_2, it also has 3 pairs of electrons that are not involved in bonding. Pairs of electrons that do not participate in chemical bonds are called **lone electron pairs**.

lone electron pair a pair of valence electrons that is localized to a given atom but not involved in bonding

Follow the steps below to draw Lewis structures.

Steps for Drawing Lewis Structures

1. Identify the central atom, which is usually the element with the highest bonding capacity. Write the symbol for the central atom, then arrange the symbols of the atoms for the rest of the elements in the pure substance around it.
2. Add up the number of valence electrons available in an atom of each of the elements. This number represents the total number of electrons (dots) you will draw in your Lewis structure. If the structure is a polyatomic ion, add 1 electron for each unit of negative charge, or subtract 1 for each unit of positive charge.
3. Place 1 pair of electrons between each adjacent pair of atoms. Every 2 of these dots represent a bonding electron pair that forms a single covalent bond.
4. Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms (not the central atom). Follow the duet rule for hydrogen atoms and the octet rule for all other atoms.
5. Determine how many electrons are still available by subtracting the number of electrons you have used so far from the total number of valence electrons.
6. Place the remaining electrons on the central atom in pairs.
7. If the central atom(s) does not have a full octet, move lone pairs from the surrounding atoms into a bonding position between those atoms and the central atom until all octets are complete.
8. If the surrounding atoms have complete octets and there are electrons remaining, add these electrons as lone pairs onto the central atom(s). Check the finished structure. All atoms (except hydrogen) should have a complete octet, counting shared and lone pairs to complete the Lewis structure. If you are representing a polyatomic ion, place square brackets around the entire structure and write the charge outside the brackets.

You can write the structural formula of a molecule or polyatomic ion based on a Lewis structure. Remove the dots representing the lone electron pairs, and replace the dots representing bonding electron pairs with solid lines to represent covalent bonds. Use double or triple lines for double or triple bonds.

Tutorial 1 Drawing Lewis Structures

In this Tutorial, you will practise drawing Lewis structures for various substances containing elements from the first two periods of the periodic table. You will also write structural formulas for substances based on their Lewis structures.

Sample Problem 1: Drawing a Lewis Structure for a Molecule

Methanal, or formaldehyde, is a gas with a distinct odour. Formaldehyde is used in the production of cosmetics, pharmaceuticals, and textiles (**Figure 9**). The chemical formula of this molecule is H_2CO. Draw the Lewis structure and then write the structural formula for this molecule.

Figure 9 Formaldehyde may be used to make fibres wrinkle-resistant, such as this permanent-press acetate cloth.

Solution

Step 1. Identify the central atom, which is the element with the highest bonding capacity. Write the symbol for the central atom, then arrange the symbols of the atoms for the rest of elements in the substance around it.

The central atom for methanal is carbon.

$$\begin{array}{lll} H & & \\ & C & O \\ H & & \end{array}$$

Step 2. Add up the number of valence electrons available in an atom of each of the elements. This number represents the total number of electrons (dots) you will draw in your Lewis structure. Since methanal is not a polyatomic ion, you do not need to make changes to this number to compensate for charge.

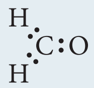
1 C atoms: $1(4e^-) = 4e^-$

2 H atoms: $2(1e^-) = 2e^-$

1 O atoms: $1(6e^-) = 6e^-$

Total: $12e^-$

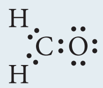
Step 3. Place 1 pair of electrons between each adjacent pair of atoms to represent the bonding electron pairs. Record the total number of electrons available and the number of electrons used so far.



Start: $12e^-$

Used: $6e^-$

Step 4. Place pairs of the remaining valence electrons as lone pairs on the surrounding atoms (not the central atom). Follow the duet rule for any hydrogen atoms and the octet rule for all other atoms.



Step 5. Determine how many electrons are still available by subtracting the number of electrons you have used so far from the total number of valence electrons.

Start: $12e^-$

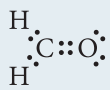
Used: $12e^-$

Remaining: $12e^- - 12e^- = 0e^-$

Step 6. Place the remaining electrons on the central atom(s) in pairs.

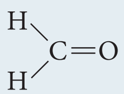
There are no remaining electrons in this case.

Step 7. If the central atom(s) does not have a full octet, move lone pairs from the surrounding atoms into a bonding position between those atoms and the central atom until all octets are complete.



Step 8. If the surrounding atoms have complete octets and there are electrons remaining, add these electrons as lone pairs onto the central atom(s). Check the finished structure. All atoms (except hydrogen) should have a complete octet, counting shared and lone pairs. This is not a polyatomic ion, so no square brackets are needed. Your Lewis structure is therefore complete.

Step 9. To draw the structural formula, remove the dots representing the lone pairs. Replace the dots representing bonding electron pairs with solid lines to represent covalent bonds. Use double or triple lines for double or triple bonds.



Methanal is a neutral molecule. You therefore do not need to represent a charge, so your structural formula for methanal is complete.

Sample Problem 2: Drawing a Lewis Structure for a Polyatomic Ion

Sodium nitrate, $NaNO_3$, is a versatile compound used in many products, such as fireworks, pottery enamel, and even as a preservative in smoked meats (**Figure 10**). It contains the polyatomic ion called nitrate, NO_3^-. Draw the Lewis structure and the structural formula for nitrate.

Solution

Step 1. Identify the central atom and write its symbol. Then, arrange the symbols of the atoms for the rest of elements around it.

Nitrogen is the central atom.

O N O
O

Step 2. Add up the number of valence electrons available in each of the atoms. This number represents the total number of electrons (dots) you will draw in your Lewis structure. If the structure is a polyatomic ion, add 1 electron for each unit of negative charge, or subtract 1 for each unit of positive charge.

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

3 O atoms: $3(6e^-) = 18e^-$

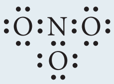
charge $= 1e^-$

Total: $24e^-$

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

Start: $24e^-$
Used: $6e^-$

Step 4. Place pairs of the remaining valence electrons, or lone pairs, on the surrounding atoms, following the duet rule or octet rule appropriately.



Step 5. Determine how many electrons are still available.

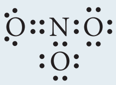
Start: $24e^-$

Used: $24e^-$

$24e^- - 24e^- = 0e^-$ left

Step 6. There are no remaining electrons to place on the central atom.

Step 7. The central N atom has an incomplete octet. Move a lone pair on one oxygen atom to between nitrogen and oxygen.



Step 8. There are no electrons remaining. Since you are representing a polyatomic ion, place brackets around the entire structure and write the charge outside the brackets. Your Lewis structure is now complete.

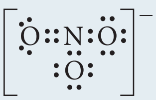
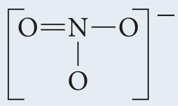


Figure 10 Sodium nitrate is used in the curing of smoked meat. This compound slows the growth of bacteria, extending the shelf life of the product. Sodium nitrate also occurs naturally in some vegetables. For example, 1 kg of celery contains about 1 g of sodium nitrate.

Step 9. Remove the dots representing lone electron pairs, and replace the dots representing bonding electron pairs with solid lines.



Your structural formula for the nitrate ion is now complete.

Practice

1. Draw the Lewis structure and the structural formula for each of the following molecules: T/I C
   (a) NBr_3
   (b) CF_4
   (c) N_2
   (d) C_2H_4
2. Draw the Lewis structure and the structural formula for each of the following polyatomic ions: T/I C
   (a) PO_3^{3-}
   (b) CN^-
   (c) NO_2^-
   (d) ClO^-

Figure 11 Boron trifluoride gas is used to improve the quality of aluminum casts, such as those used to make automobile parts.

Exceptions to the Octet Rule

When drawing Lewis structures for molecules, hydrogen atoms always obey the duet rule, and carbon, nitrogen, oxygen, and fluorine atoms (and other halogens) obey the octet rule in all circumstances. However, there are exceptions to the octet rule. The first exception applies to molecules with central atoms that are surrounded by fewer than 8 electrons. These central atoms are said to have underfilled octets. The second exception applies to molecules with central atoms that are surrounded by more than 8 electrons. These central atoms are said to have overfilled octets.

Boron forms compounds in which the boron atom has fewer than 8 electrons around it. The boron atom therefore has an underfilled octet. Boron trifluoride, BF_3, is a gas at normal temperatures and pressures (**Figure 11**). It reacts very energetically with molecules such as water and ammonia, which have lone pairs of electrons that do not participate in a chemical bond. The boron atom has 3 valence electrons, and each fluorine atom has 7 valence electrons. Therefore, boron trifluoride has a total of 24 valence electrons (**Figure 12**).

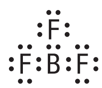


Figure 12 The Lewis structure for boron trifluoride

In this structure, each fluorine atom has 3 lone pairs of electrons, which total to 18 electrons. The remaining 6 electrons surround the boron atom, causing it to be electron deficient. The reactivity of boron trifluoride with electron-rich molecules such as water and ammonia, NH_3, occurs because the boron atom is electron deficient.

The molecule that forms from the reaction of boron trifluoride with ammonia has the chemical formula H_3NBF_3. **Figure 13** shows the reactants and products of this reaction using simplified Lewis structures. In a **simplified Lewis structure**, dots representing bonding electron pairs are replaced by solid lines and any dots representing lone electron pairs are kept. In the H_3NBF_3 molecule, the boron atom obtains an octet of electrons.

simplified Lewis structure a Lewis structure in which bonding electron pairs are represented by solid lines and lone electron pairs by dots

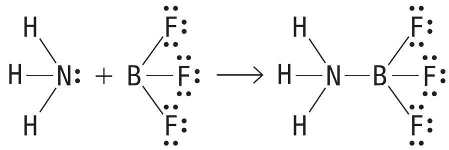
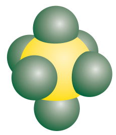


Figure 13 Simplified Lewis structures showing the formation of H_3NBF_3

Some atoms can exceed the octet rule (have an overfilled octet). These atoms use nearby vacant *d* orbitals to exceed their octet. Therefore, you do not observe this behaviour for elements in periods 1 or 2 of the periodic table. An example of an atom exceeding the octet rule is the sulfur atom in sulfur hexafluoride, SF_6. A **space-filling model** shows the relative sizes of the atoms and their relative orientations in three-dimensional space. **Figure 14** shows a space-filling model of sulfur hexafluoride. Sulfur hexafluoride is a very stable molecule. The sum of the valence electrons in this unusual molecule is as follows:

space-filling model a model of a molecule showing the relative sizes of the atoms and their relative orientations



SF_6

Figure 14 A space-filling model of a molecule of SF_6

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

6 F atoms: $6(7e^-) = 42e^-$

Total valence electrons: $6e^- + 42e^- = 48e^-$

Figure 15 shows the arrangement of electrons around the sulfur atom in sulfur hexafluoride.

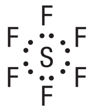


Figure 15 The sulfur atom in a sulfur hexafluoride molecule has an overfilled octet.

Notice in Figure 15 that 12 electrons form the S–F bonds. The total number of valence electrons is 48, so there are 36 electrons unaccounted for. Since fluorine atoms always follow the octet rule, you can complete the octet for each of the 6 fluorine atoms to give the Lewis structure shown in **Figure 16**.

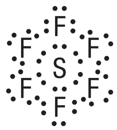
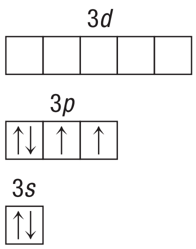


Figure 16 Lewis structure of a sulfur hexafluoride molecule

This structure uses all 48 valence electrons for sulfur hexafluoride, but the sulfur atom has 12 electrons around it, which exceeds the octet rule. How is this possible?

To answer this question, consider the types of valence orbitals that are available for second- and third-period elements. The second-period elements have 2*s* and 2*p* valence orbitals. The third-period elements have 3*s*, 3*p*, and 3*d* orbitals. As you examine the electron distribution of the atoms in the third period, notice that as you move across the periodic table from sodium to argon, the 3*s* and 3*p* orbitals fill with electrons but the 3*d* orbitals remain empty. The valence orbital diagram for a sulfur atom is



The sulfur atom in a sulfur hexafluoride molecule can have 12 electrons in its valence shell by using the 3*s*, 3*p*, and 3*d* orbitals.

Tutorial 2 Drawing Simplified Lewis Structures of Molecules That Are Exceptions to the Octet Rule

In this Tutorial, you will draw simplified Lewis structures for molecules that contain atoms that have underfilled or overfilled octets.

Sample Problem 1: Molecules with Underfilled Octets

Beryllium forms compounds in which the beryllium atom has fewer than 8 electrons around it. Draw the simplified Lewis structure for a molecule of beryllium chloride, $BeCl_2$.

Solution

Step 1. Identify the central atom, and write its symbol. Then, arrange the symbols of the atoms for the rest of the elements in the molecule around it.

Cl Be Cl

Step 2. Count all the valence electrons for all the atoms in the molecule. The total represents the total number of electrons (dots) you will draw in your Lewis structure.

1 Be atom: $1(2e^-) = 2e^-$

2 Cl atoms: $2(7e^-) = 14e^-$

Total: $16e^-$

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

Cl:Be:Cl

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

:Cl:Be:Cl:

Step 5. Determine the number of electrons still available.

Start: $16e^-$

Used: $16e^-$

$16e^- - 16e^- = 0\ e^-$ left

Step 6. Place the remaining electrons on the central atom(s) in pairs.
There are no remaining electrons.

Step 7. The central Be atom has an incomplete octet. You could move 1 lone pair of electrons from each chlorine atom to complete the octet for the beryllium atom. However, since you know that the beryllium atom will be underfilled when it forms a compound, the Lewis structure of beryllium chloride is best represented as the following:

:Cl:Be:Cl:

Step 8. There are no electrons remaining, so the Lewis structure for beryllium chloride is complete.

Step 9. Replace the dots between atoms with lines to represent bonds to simplify the Lewis structure.

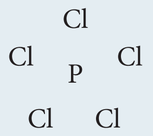
:Cl—Be—Cl:

Sample Problem 2: Molecules with Overfilled Octets

Phosphorus forms molecular compounds in which the phosphorus atom shares more than 8 electrons. Draw the simplified Lewis structure for a molecule of phosphorus pentachloride, PCl_5.

Solution

Step 1. Identify the central atom and write its symbol. Then, arrange the symbols of the atoms for the rest of the elements in the molecule around it.



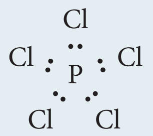
Step 2. Count all the valence electrons for all the atoms in the molecule. The total represents the total number of electrons (dots) to draw in your Lewis structure.

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

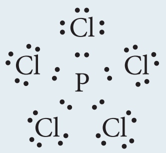
5 Cl atoms: $5(7e^-) = 35e^-$

Total: $40e^-$

Step 3. Place 1 pair of electrons between each adjacent pair of atoms, forming single covalent bonds. Draw 1 pair of electrons between each chlorine atom and phosphorus atom. You will have a total of 10 electrons around the phosphorus atom. Since you know that the phosphorus atom is overfilled, you can accept this arrangement.



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



Step 5. Determine the number of electrons still available.

Start: $40e^-$

Used: $40e^-$

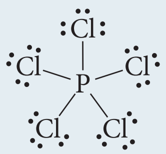
$40e^- - 40e^- = 0e^-$ left

Step 6. Place the remaining electrons on the central atom(s) in pairs.
There are no remaining electrons.

Step 7. Each chlorine atom has a full octet and the phosphorus atom is overfilled, so you do not need to move any electrons.

Step 8. There are no electrons remaining, so the Lewis structure for phosphorus pentachloride is complete.

Step 9. Replace the dots between atoms with lines to represent bonds to simplify the Lewis structure.



Practice

1. Draw the simplified Lewis structure for each of the following molecules. T/I C
   (a) ICl_5
   (b) $RnCl_2$
2. Draw the simplified Lewis structure for each of the following: T/I C
   (a) ClF_3
   (b) NO^+
   (c) ICl_4^-

Coordinate Covalent Bonding

coordinate covalent bond a covalent bond in which the electrons involved in bonding are from one atom

When 2 hydrogen atoms form a covalent bond, each hydrogen atom donates 1 electron to form the bond. However, sometimes the covalent bond that forms between 2 atoms involves 2 electrons donated by a single atom. This type of bond is called a **coordinate covalent bond**. One example of this type of bond occurs in the ammonium ion, NH_4^+. In this ion, the nitrogen atom contributes 5 electrons, and the 3 hydrogen atoms each contribute 1 electron to the structure. A single covalent bond forms between the nitrogen atom and each of the 3 hydrogen atoms to form a molecule of ammonia, NH_3. Thus, 2 valence electrons of the nitrogen atom do not participate in bond formation. When a hydrogen ion, H^+, interacts with a molecule of ammonia, it forms a covalent bond with this remaining electron pair (**Figure 17**). Both of the electrons that make up this covalent bond come from the nitrogen atom.

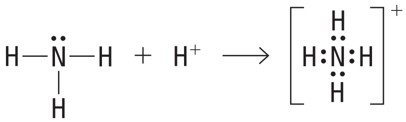


Figure 17 Lewis structures of the reaction between ammonia and a hydrogen ion to form the ammonium ion

Aluminum chloride can exist in two forms, given by the formulas $AlCl_3$ and Al_2Cl_6. The molecule Al_2Cl_6 has a coordinate covalent bond. A molecule of Al_2Cl_6 forms from 2 molecules of $AlCl_3$. A Lewis structure of $AlCl_3$ reveals that the aluminum atom does not have an octet of valence electrons. Both sets of the electrons involved in the coordinate covalent bonds come from chlorine atoms (**Figure 18**).

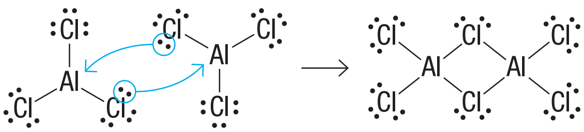


Figure 18 Simplified Lewis structures of the reaction between two molecules of $AlCl_3$ to form Al_2Cl_6

4.1 Review

Summary

- Chemical bonds hold atoms together to form molecules and ionic compounds.
- Chemical bonds form when a group of atoms has a lower total energy.
- Noble gas configurations of atoms are generally the most stable.
- The two main types of chemical bonds are ionic and covalent bonds.
- Lewis structures can be used to represent molecules and polyatomic ions.
- Most Lewis structures contain atoms that obey the octet rule or the duet rule (for the hydrogen atom).
- There are exceptions to the octet rule in which the central atom has an underfilled or an overfilled octet.
- A coordinate covalent bond is a covalent bond in which the electrons involved in bonding are from one atom.

Questions

1. Carbon tetrafluoride gas, $CF_4(g)$, is used as a refrigerant. It is very stable when released into the atmosphere and is a known potent greenhouse gas. Calcium fluoride, $CaF_2(s)$, is the main source of the world's supply of fluorine. It is used in the production of specialized lenses and window materials. K/U
   (a) Describe the bonding in a molecule of carbon tetrafluoride.
   (b) Compare and contrast the bonding of carbon tetrafluoride and calcium fluoride.
2. Some plant fertilizers contain the following compounds: ammonium sulfate, $(NH_4)_2SO_4$; calcium phosphate, $Ca_3(PO_4)_2$; potassium oxide, K_2O; diphosphorus pentoxide, P_2O_5; and potassium chloride, KCl. K/U
   (a) Which compounds contain ionic bonds?
   (b) Which compounds contain covalent bonds?
   (c) Do any compounds contain both ionic and covalent bonds? Explain how this is possible.
3. Aluminum metal is produced by heating aluminum oxide, $Al_2O_3(s)$, and other substances to almost 1000 °C until the mixture melts. Molten aluminum oxide is an excellent conductor. An electric current passed through liquid aluminum oxide can provide the electrons needed to convert the aluminum ions to neutral metal atoms. Explain the following properties of aluminum oxide in terms of ionic bonding: K/U
   (a) its high electrical conductivity when molten
   (b) its high melting point
4. Draw Lewis structures for each of the following molecules or polyatomic ions. State whether they obey the octet rule and if any coordinate covalent bonds form. In each case, the first atom listed is the central atom. K/U C
   (a) $POCl_3$
   (b) SO_4^{2-}
   (c) PO_4^{3-}
   (d) ClO_4^{-}
5. An exception to the octet rule is a central atom with fewer than 8 valence electrons. Beryllium hydride, BeH_2, and borane, BH_3, are both examples. Draw the Lewis structures of these 2 molecules. K/U C
6. Common exceptions to the octet rule are compounds and polyatomic ions with central atoms having more than 8 electrons around them. Phosphorus pentafluoride, PF_5; sulfur tetrafluoride, SF_4; xenon tetrafluoride, XeF_4; and tri-iodide ion, I_3^{-}, are all examples of exceptions to the octet rule. K/U C
   (a) Draw the Lewis structures of these substances.
   (b) For which elements in these substances can the atoms have more than 8 electrons around them?
   (c) How can the atoms of the elements you identified in Part (b) be surrounded by more than 8 electrons?
7. A classmate does not understand the concept of coordinate covalent bonding. Make up and write an analogy to help him or her understand. T/I C