SCH4U-C

Introduction to Chemical Bonding

Introduction

Iceland is an island nation in the north Atlantic that was once very wealthy. In 2009, however, Iceland was on the verge of bankruptcy as its banks had taken too many risks and had become insolvent. Fortunately, Iceland has an abundant natural resource that could again make it one of the wealthiest countries per capita in the world. This resource isn't gold or oil, but a limitless supply of geothermal energy, as Iceland is situated on top of a volcano! This means that energy-intensive industries, such as the production of hydrogen, can be done more cheaply in Iceland than in most other countries. Entrepreneurs in Iceland hope to use geothermal energy to convert sea water into hydrogen:

$$2H_2O_{(l)} + \text{energy} \rightarrow 2H_{2(g)} + O_{2(g)}$$

Hydrogen is an energy-rich molecule that has the potential to replace fossil fuels as the fuel of choice for the twenty-first century. Like any fuel, hydrogen has potential chemical energy stored within its structure. When hydrogen burns, it reacts with oxygen to release this energy and water vapour:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)} + \text{energy}$$

The great advantage of hydrogen as a fuel is that it's non-polluting—its only waste product is water. Unlike non-renewable fuels like oil, the supply of hydrogen on earth is almost limitless—wherever there's water and energy, you can make hydrogen. You'll learn more about hydrogen in future lessons.

Planning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Bond Formation	½
Lewis Structures or Electron Dot Diagrams	1½
Hybrid Orbitals	1
The Bonding Continuum	½
Key Questions	1

What You Will Learn

After completing this lesson, you will be able to

- draw orbital and electron dot representations of bonding
- use energy-level diagrams to predict the number and type of bonds that elements form
- use the concept of electronegativity to predict the type of bond that two elements form

Bond Formation

To better understand how energy is stored within compounds like hydrogen, you need to take a closer look at the nature of chemical bonds and how they form. Hydrogen exists as a diatomic molecule with the chemical formula H_2. Hydrogen is so reactive that it is rarely observed in nature. Instead, hydrogen can be made, or extracted, from more complex molecules in the laboratory. Regardless of how you make it, hydrogen always forms H_2 molecules—never H_3 or just H. To understand why, begin by looking at the hydrogen atom. Neutral hydrogen atoms consist of a proton in the nucleus surrounded by an "electron shell" which, in this case, has one electron in the 1*s* orbital. The electrostatic attraction between the proton and the electron keep the electron from flying out of the atom's orbital.

Now consider the potential energy changes that occur when two hydrogen atoms approach or collide with each other (Figure 2.1). Recall from earlier science courses that potential energy is the energy stored within a system by virtue of the stresses in that system. For example, a system consisting of a stretched elastic band has more potential energy than one consisting of a limp elastic band. Similarly, a system consisting of two magnets being held apart by a force has more potential energy than one consisting of two magnets that have been allowed to join together. As soon as you release the force holding the magnets apart, they fly together.

Take a look at Figure 2.1, in which two hydrogen atoms collide and either become a molecule or do not. Initially, there is no attraction between the atoms on the right-hand side of Figure 2.1 because they are too far apart. As a result, this two-atom system has zero potential energy at point (a), while having an even lower "negative potential energy" as they move closer.

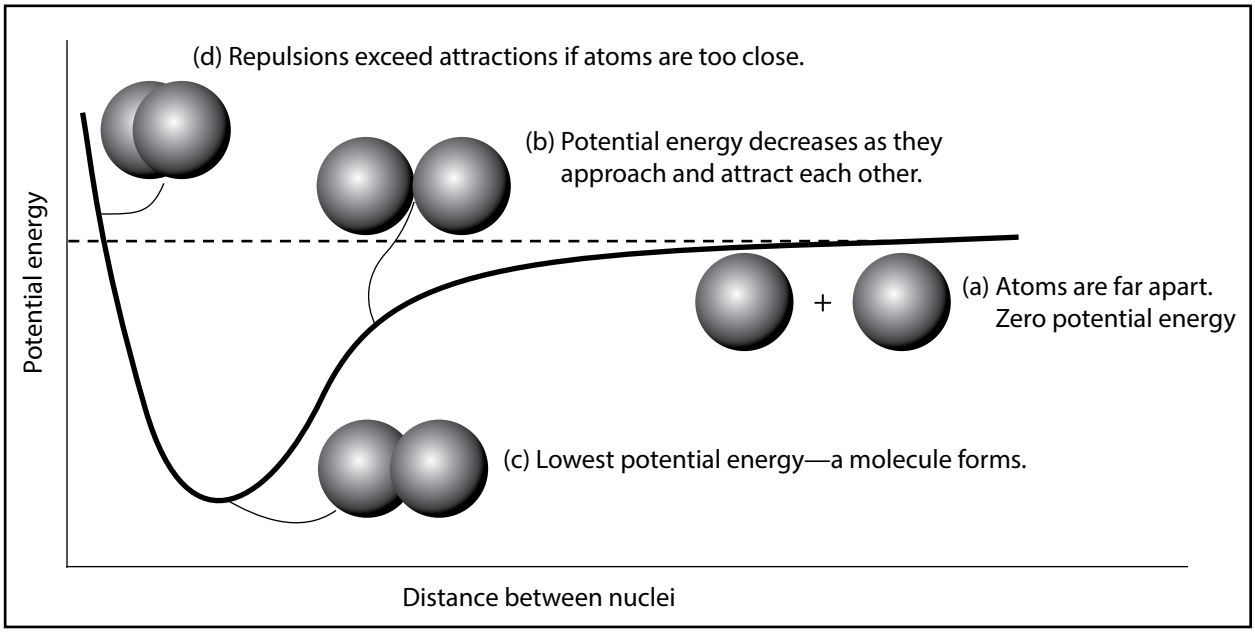

Figure 2.1: Potential energy changes that occur during the collision of hydrogen atoms

As the two hydrogen atoms approach each other at point (b), each atom's proton begins to exert an attraction for the electron of the other atom. This causes the atoms to move even closer together. As they do, the potential energy of the system decreases. If the nuclei become too close, as they do at point (d), they quickly bounce away. The repulsions between the two orbitals and between the two protons will cause the atoms to go their separate ways. If the distance is "just right," the attractive and repulsive forces equal each other, so the nuclei reach a stable point. This stable distance is the average bond length for the hydrogen molecule, and its potential energy level is indicated at the lowest point on the curve—point (c)—which is defined as negative (less than zero) energy potential. Because H_2 is so much more stable than H, unpaired hydrogen atoms are practically non-existent at standard temperature and pressure.

After an H_2 molecule forms, the pair of electrons in the molecule remain simultaneously attracted to both nuclei. Both atoms now have access to two electrons, which is the number required to fill its orbital and achieve stability. The result is that the atoms "share" these electrons. A chemical bond resulting from sharing electrons is a strong type of bond called a covalent bond. The difference in energy between the initial and the minimum potential energy state is the energy released from the system when a hydrogen molecule forms from its atoms. In general, the formation of any chemical bond results in a decrease of potential energy. If the atoms are forced together, closer than the average bond length, then the electron–electron and proton–proton repulsions exceed the attractive forces and potential energy increases.

Bonding Representations

When a covalent bond forms, the orbitals, which contain the bonding electrons, overlap. In the case of the hydrogen molecule, this involves the overlap of the 1*s* orbitals of each atom. Once orbital overlap occurs, the shared electron pair can travel throughout the entire molecule (Figure 2.2).

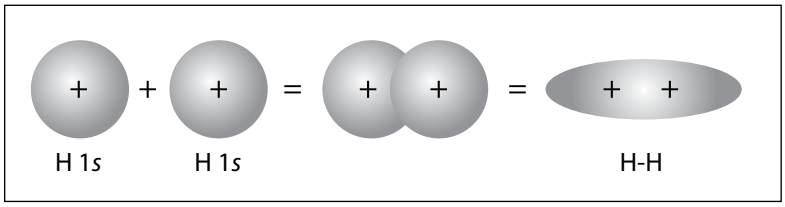

Figure 2.2: The formation of a covalent bond involves the overlap of orbitals. In this case, 1*s* orbitals overlap as the hydrogen molecule forms.

Showing the overlapping orbitals in a small molecule like hydrogen is easy. However, showing the overlap of orbitals of larger atoms can be very difficult. Luckily, chemists have devised simpler methods of representing bonding.

Orbital Overlap Diagram

In an orbital overlap diagram, overlapping orbitals are shown connected to each other with a rectangle (Figure 2.3). Note that only the outermost or valence electrons of oxygen are given in this diagram. The inner electrons are not involved in bonding.

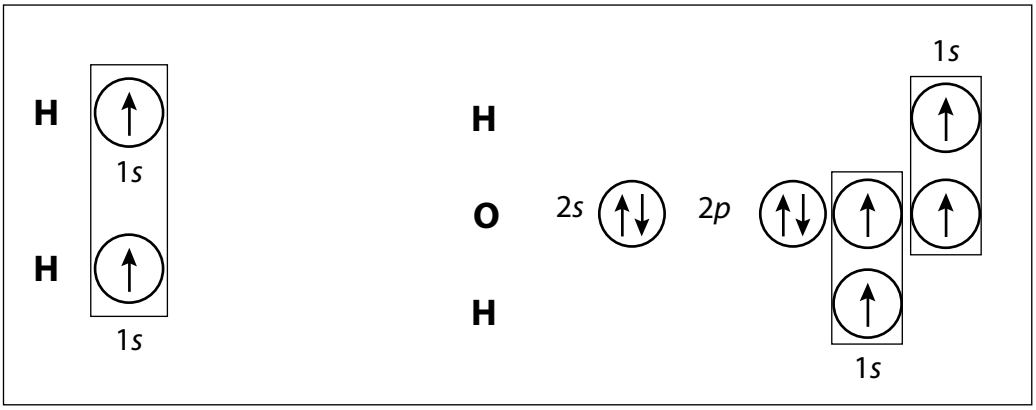

Figure 2.3: Orbital diagrams of molecules of hydrogen (left) and water (right). Each arrow represents an electron, each circle represents an orbital, and overlapping orbitals are connected with a rectangle.

The Octet Rule

In Lesson 1, you learned that atoms gain or lose electrons to achieve the stability of a noble gas electronic configuration. The discussion in Lesson 1 centred around the formation of ions. Atoms such as hydrogen and oxygen can also achieve stability by sharing their valence electrons. For example, the orbital diagram of the overlap of the hydrogen molecule (Figure 2.3) shows that both hydrogen atoms have access to two electrons—the same number of electrons as helium. Similarly, oxygen in the water molecule now has two electrons in 2*s* and six electrons in 2*p*. It achieves stability by overlapping its *p* orbitals with the 1*s* orbital of hydrogen. In doing so, the oxygen atom fills its outer orbitals with eight electrons—just like neon's filled outer orbitals.

An atom is most stable when it can obtain a noble gas electronic configuration in its outer orbitals. The outer orbitals are collectively called the valence orbitals and the outer electrons are called valence electrons. Typically, the valence orbitals in an atom are the outermost *s* and *p* orbitals. These orbitals get involved in the bonding process with other atoms.

With the exception of hydrogen, most non-metals will gain or share one or more electrons in order to have eight valence electrons. Most metals will lose or share one or more electrons in order to have eight valence electrons. Chemists refer to this trend—this tendency to lose or gain electrons to obtain eight in the valence orbit—as the octet rule. This rule also states that in the case of hydrogen and helium, they need two electrons in the valence orbital. The hydrogen atom is one example of an atom that does not achieve the stable octet of electrons but only has two electrons when stable, like helium.

Support Question

Be sure to try the Support Questions on your own before looking at the suggested answers provided.

13. Draw orbital overlap diagrams to show the bonding in hydrogen chloride, HCl, and oxygen, O_2. Suggested answer

Lewis Structures or Electron Dot Diagrams

Lewis structures or electron dot diagrams are one of the simplest methods of representing bonding. Gilbert Lewis first introduced Lewis structures in 1916, prior to the development of quantum mechanics. Lewis structures are based on the following fundamental ideas:

1. The noble gases (group 18) are stable because they have filled outer or valence orbitals. For example, with the exception of helium, all noble gases have eight electrons in their valence, *s* and *p*, orbitals.
2. Atoms gain, lose, or share electrons to achieve an electron arrangement like that of the nearest noble gas. In doing so, atoms "complete their octet." For example, magnesium reacts to lose its two valence electrons. In doing so, it forms a stable Mg^{2+} ion.
3. Metal atoms tend to lose electrons to achieve stability, while non-metals gain electrons.
4. Non-metals can also share electrons to achieve stability.

Electron Dot Diagrams of Elements

The electron dot diagrams for the elements of the second period are shown in Figure 2.4.

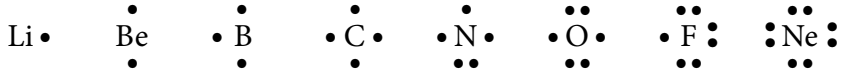

Figure 2.4: Electron dot diagrams of the second period (row) elements

The dots in these diagrams represent only the valence electrons of the atom. For example, since oxygen has atomic number 8, it has eight electrons. The electron configuration for oxygen would be [$1s^2$, $2s^2$, $2p^4$]. The valence orbitals are the *s* and *p* orbitals. We see here that oxygen has six electrons in those valence orbitals. It is these six electrons that are shown on the dot diagram for the element oxygen. That's why the electron dot diagram for oxygen has six electrons, instead of eight. Since sulfur is in the same group as oxygen, it also has the same electron dot diagram. With the exception of helium, the electron dot diagrams in Figure 2.4 are the same

for all of the elements in the periodic table that are found in the groups (columns) beneath each of these elements. For example, chlorine, bromine, and iodine are in the same group (group 17) and therefore have the same electron dot diagram as fluorine in Figure 2.4. The difference between paired and single electrons in these diagrams is significant. Atoms share their single electrons with other atoms to form covalent bonds. Paired electrons, which are called lone pairs, rarely participate in bonding.

Electron Dot Diagrams of Simple Ions

Figure 2.4 can also be used to predict the electron dot diagram for the ions that these elements form. The electrons in these diagrams are the valence electrons in these atoms. In Lesson 1, you learned that atoms gain, lose, or share electrons to achieve an electron arrangement like that of the nearest noble gas. Consequently, lithium loses its single valence electron to form an Li^+ ion. Therefore, the electron dot diagram of Li^+ consists of only the chemical symbol (Figure 2.5). Square brackets are usually used in the electron dot diagrams of ions. When lithium reacts with chlorine, lithium's valence electron is transferred, filling chlorine's valence orbit. Hence, the electron dot diagram of the chloride ion contains eight electrons, rather than seven (Figure 2.5).

Figure 2.5: Electron dot diagrams of lithium and chloride ions

Lewis Structures of Molecules and Complex Ions

Lewis structures and structural formulas are perhaps the simplest methods for predicting and representing the bonding in a molecule or complex ion. Being able to draw these diagrams will be useful for determining the shape and polarity of molecules in Lesson 3. For example, Figure 2.6 shows the Lewis structure and structural formula for water. Note that the covalent bond between the atoms is represented by a solid line in the structural formula. Lone pairs of electrons are generally not drawn in the structural formula. Both diagrams show how the atoms in the water molecule are bonded. However, neither representation shows the actual shape of the water molecule. In Lesson 3, you'll learn how to determine shape from the electron dot diagram and what implications shape has on the properties of water.

Figure 2.6: Electron dot diagram and structural formula for water

The following steps outline a systematic way of determining the electron dot diagram of a molecule or complex ion. This is commonly called creating a Lewis structure or a Lewis dot diagram.

The Lewis structure will include all valence electrons found in the structure, including the lone pairs that are not involved in bonding. Just showing a structural formula (drawing) need not include lone pairs.

Example

Draw the Lewis structure and structural formula for the nitrate ion, NO_3^-.

Solution

Step 1: Count all of the valence electrons in the molecule or ion. In the case of an ion, add or subtract electrons to account for the ionic charge.

Nitrogen	$5e^-$	(N has 5 valence electrons)
Oxygen	$18e^-$	(6 valence electrons/atom × 3 atoms = $18e^-$)
Ionic charge	$1e^-$	(Ionic charge of nitrate)
=	**$24e^-$**	**(Total valence electrons)**

Step 2: Arrange the peripheral atoms symmetrically around the central atom. Use a pair of electrons to form a bond that links these atoms to the central atom. In this case, since there is only one nitrogen atom and three oxygen atoms, it is likely that nitrogen is the central atom in Figure 2.7.

$24e^-$
$-6e^-$
$18e^-$ remain

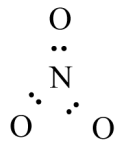

Figure 2.7

Step 3: Add pairs of electrons to complete the octet of the peripheral atoms.

$18e^-$
$-18e^-$
$0e^-$ remain

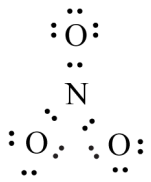

Figure 2.8

Step 4: Place any unassigned electrons on the central atom. For this example, all electrons have already been used. Note that the central nitrogen atom has only six electrons. Therefore, its octet is incomplete.

Step 5: If the octet of the central atom is incomplete, move a lone pair of electrons from a peripheral atom to a new position between the central and peripheral atom. In this case, a lone pair from any of the oxygen atoms is moved to complete nitrogen's octet, which forms a double bond.

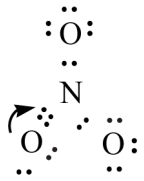

Figure 2.9 The Lewis structure showing all electrons

Step 6: Draw the structural formula. For ions, surround the ion with square brackets and place the ionic charge outside the bracket.

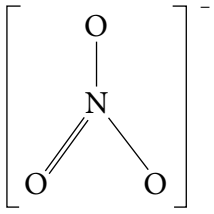

Figure 2.10 The structural formula showing bonds and charge

Support Questions

14. Draw the electron dot diagrams for:

a) P^{3-}

b) Ca^{2+}

Suggested answer

15. Draw the Lewis structure and structural formula for:

a) NH_3

b) CH_3Cl

c) CO_2

d) ClO_3^-

e) PO_4^{3-}

Suggested answer

Hybrid Orbitals

So far, you have learned that atomic orbitals overlap to form covalent bonds. This allows two atoms to simultaneously attract an electron, resulting in a covalent bond. This bonding theory is very useful to predict the number and type of bonds that many elements form. But it is not without its limitations, one of which you are now going to examine in greater detail. In Lesson 1, you learned that carbon has two single electrons in its 2*p* orbitals of the energy-level diagram (Figure 2.11).

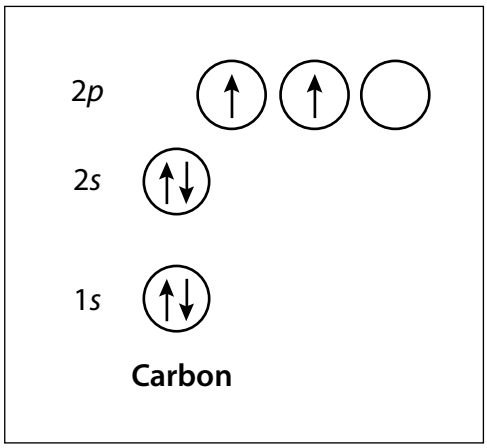

Figure 2.11: Energy-level diagram of carbon in what is called its ground state

As a result, carbon should form two bonds by overlapping its 2*p* orbitals with the orbitals of another atom(s). Therefore, carbon should bond with two hydrogen atoms. In fact, the molecule CH_2 does exist, but is very unstable, mainly due to the empty 2*p* orbital. Methane, CH_4, which is a major component of natural gas, is a much more stable and abundant compound. Analysis of the structure of methane reveals that its central carbon atom is surrounded by four hydrogen atoms. Furthermore, experimental evidence also suggests that the four covalent bonds from the central carbon atom are identical in both strength and length. According to the bonding theory discussed earlier, this can only be possible if the four orbitals that carbon uses for bonding are identical. In order to explain how carbon forms four identical bonds, chemists have proposed a modification to the bonding theory called orbital hybridization.

Orbital Hybridization

The basic idea behind orbital hybridization is that in order to produce four identical bonds, carbon must have four identical orbitals. In order to achieve this goal, one of the 2*s* electrons must first move (or be "promoted") into the empty 2*p* orbital (Figure 2.12). This process is feasible since the energy difference between these two orbitals is slight. Presumably, the electron could absorb this energy from the environment.

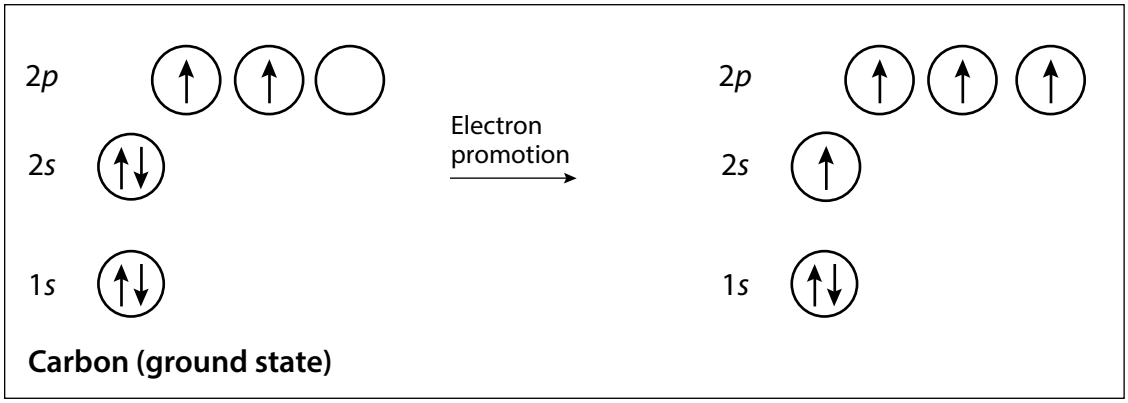

Figure 2.12: The first step in hybridization is the promotion of a 2*s* electron into the empty 2*p* orbital.

As a result of this promotion, carbon can now form four bonds. However, three of the bonds use *p* orbitals, while the fourth involves an *s* orbital. Since *s* and *p* orbitals have different shapes, you would expect the bonds that they form to have different characteristics, such as a difference in strength and length. Consequently, the bonding theory has to be "tweaked" one more time. To produce identical orbitals, the 2*s* and 2*p* orbitals of carbon blend together to form four new identical orbitals called sp^3 hybrid orbitals (Figure 2.13). The name of these orbitals comes from the orbitals that were used to make them. For example, since one *s* orbital combined with three *p* orbitals, the result is four new orbitals called s^1p^3 orbitals. However, in chemistry it is common to omit the number "1" in chemical formulas, as the 1 is assumed to be there.

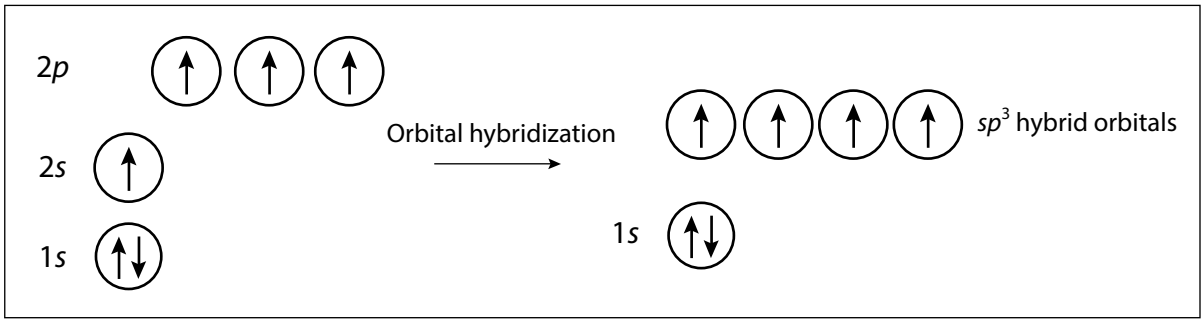

Figure 2.13: Orbital hybridization in carbon involves 2*s* and 2*p* orbitals "blending" or hybridizing to form four new hybrid orbitals.

Since the new orbitals are a mixture of *s* and *p*, they are a blend or "hybrid" of both. This is analogous to blending red and white paint together to produce a blend or "hybrid" of the two colours—pink. Figure 2.14 shows the resultant shape and direction of the four sp^3 hybrid orbitals.

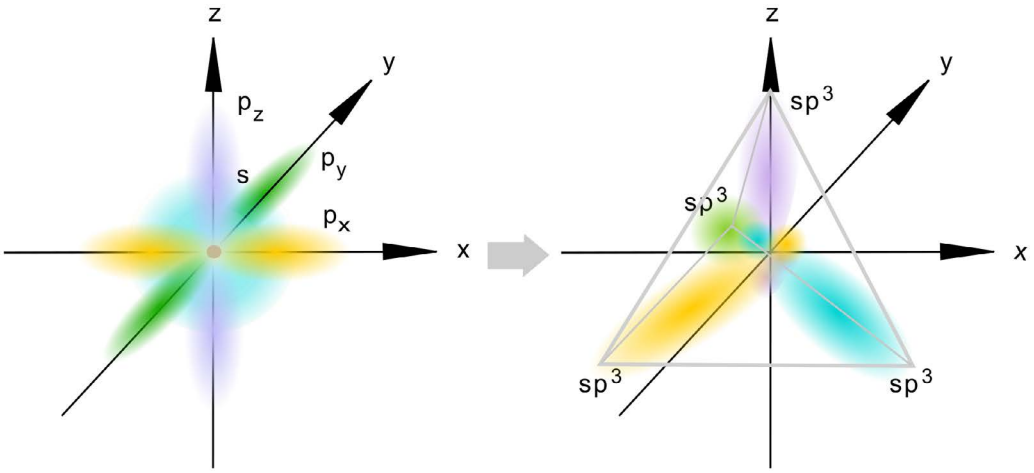

Figure 2.14: The formation of four identical sp^3 hybrid orbitals. These orbitals form the shape of a pyramid (also called a tetrahedron). The shape of these orbitals determines the shape of many of the molecules that carbon forms.

Source: http://commons.wikimedia.org/wiki/File:Hybrydyzacja_sp3.svg

It is crucial to note that orbital hybridization is a theoretical alteration to bonding theory. The alteration (hybridization) exists so that the theory agrees with what is observed in nature. It allows us to account for the fact that the carbon atom can bond with four hydrogen atoms and form methane. Hybridization is not a natural process that can be observed and measured. Describing the real behaviour of electrons is beyond the scope of this course as it requires very advanced mathematics.

Hybridization in Boron

Orbital hybridization is necessary to explain the bonding of a number of elements. Boron, for example, should only form one bond, according to its energy-level diagram. This is because it has only one electron (in the 2*p* orbital) that it can share. Based on this diagram, you would expect boron to become stable by forming one bond with hydrogen. In fact, the compound BH does exist, but it is extremely reactive, again due to the presence of the empty, 2*p*, valence orbitals. However, BH_3 also exists and is less reactive than BH. Experimental evidence shows that the three bonds in BH_3 are identical. To account for the observed evidence, the *s* and *p* orbitals in boron must first be hybridized. First, a 2*s* electron is promoted to an empty 2*p* orbital (Figure 2.15). Note that one of the *p* orbitals remains empty.

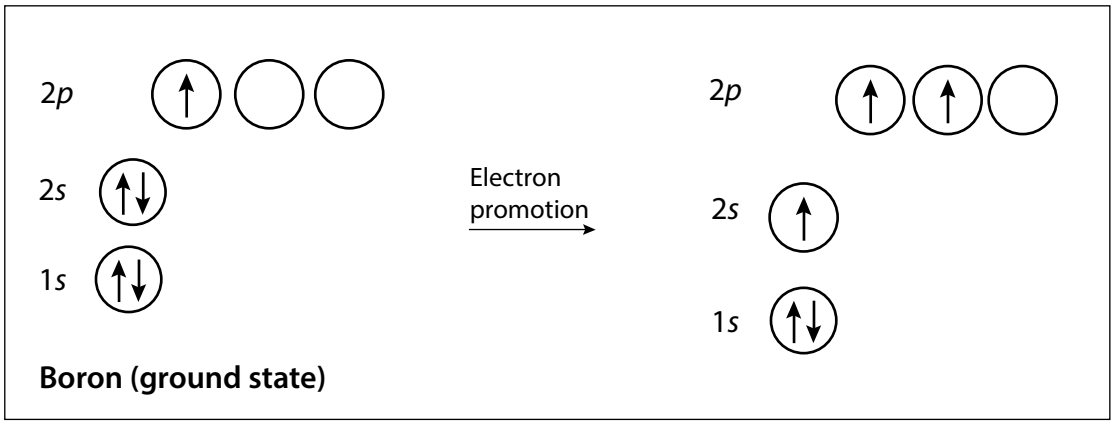

Figure 2.15: First step in the hybridization of boron's orbitals. One 2*s* electron is promoted to an empty 2*p* orbital. Notice that one of the *p* orbitals remains empty.

Then, the 2*s* orbital and two 2*p* orbitals hybridize to form three identical hybrid orbitals called sp^2 hybrid orbitals (Figure 2.16).

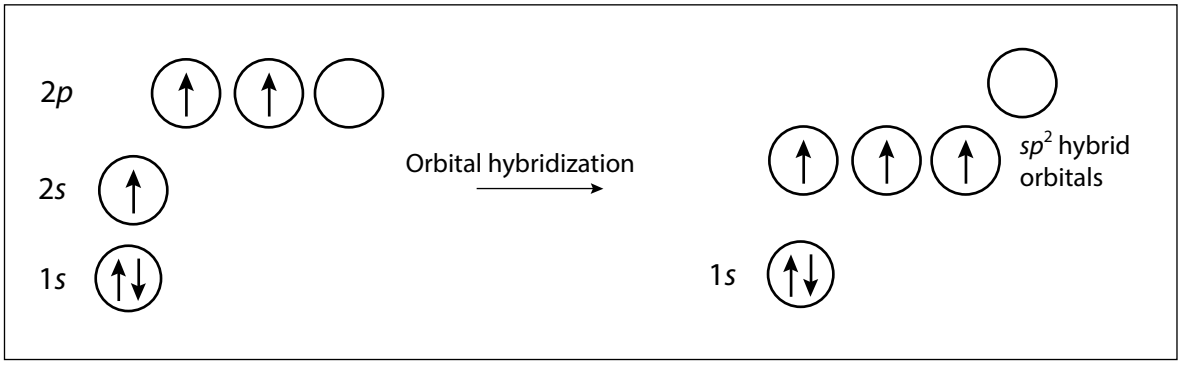

Figure 2.16: The formation of sp^2 hybrid orbitals in boron

Unlike methane, CH_4, boron trihydride, BH_3, is quite reactive. This is caused by the empty *p* orbital. As you saw earlier, many molecules and ions have paired electrons, often called lone pairs. BH_3 readily reacts with another entity that has a lone pair of electrons that can be inserted into the empty 2*p* orbital. For example, BH_3 readily reacts with the hydride ion, H^-:

$$BH_3 + H^- \rightarrow BH_4^-$$

The two electrons of the hydride ion fill the 2*p* orbital of boron, forming a fourth covalent bond.

Support Questions

16. Explain why the concept of orbital hybridization is necessary to explain the bonding in carbon.

Suggested answer

17. **a)** Use an energy-level diagram to show why the element beryllium, in its ground state, should not form covalent compounds with hydrogen.

b) Using energy-level diagrams and the concept of orbital hybridization, explain why beryllium dihydride, BeH_2, exists.

c) Refer to your diagrams in (b) to explain why BeH_2 is very reactive.

Suggested answer

The Bonding Continuum

Earlier in this lesson, a covalent bond was described as the simultaneous attraction of two nuclei for a pair of electrons. In a covalent bond, each atom contributes one of the electrons to the shared pair. Due to these attractions, the two nuclei are in a continuous tug-of-war for the shared electron pair. If the bonding atoms are identical, as in the case of the hydrogen molecule, H_2, then you can assume that each atom attracts the electrons equally. This is perfectly fair sharing of the electron pair, so each atom has the use of the shared pair equally.

Polar and Non-polar Covalent Bonds

What if the bonding atoms are different (rather than a hydrogen atom bonding with another hydrogen atom)? When one of the atoms has a stronger attraction for electrons than the other, the electron sharing that occurs is unequal or lopsided. The shared electron pair is more likely to be found near the atom with the stronger attraction. This is what occurs in the hydrogen chloride molecule, HCl. The electron dot diagram for this compound predicts that hydrogen and chlorine are linked by a single covalent bond (Figure 2.17).

Figure 2.17: Electron dot diagram and structural formula of hydrogen chloride, HCl, with an x to represent hydrogen's electron forming a single covalent bond

Experimental evidence shows that chlorine has a stronger attraction for electrons than hydrogen. Therefore, the sharing of the bonding electron pair is unequal, in favour of chlorine.

Since the bonding electrons are found closer to chlorine more often than they are found near hydrogen, the chlorine end of the molecule is more negative than the hydrogen end. The result is a partial charge separation, with the chlorine end of the molecule being slightly negative, and the hydrogen end being slightly positive. We say that the shared pair of negative electrons has shifted over toward the Cl side of the bond. Covalent bonds in which electrons are shared unequally are called polar covalent bonds (Figure 2.18). Bonds in which electron sharing is equal are called non-polar covalent bonds. As you'll see in Lessons 3 and 4, the polarity of molecules has a profound impact on physical properties.

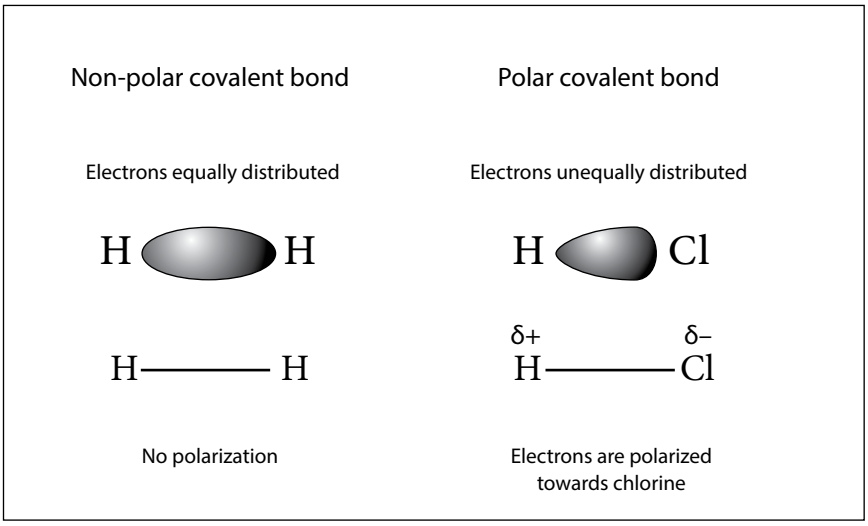

Figure 2.18: A comparison of polar and non-polar bonds. Polar bonds result from the unequal sharing of electrons. The Greek letter delta, δ, is used to indicate the two poles of the polar covalent bond. The negative end of the bond is identified with delta negative, or δ–, while the positive end is identified with delta positive, or δ+. The term "poles" is analogous to a bar magnet that has a negative pole and a positive pole.

Electronegativity

The measure of an atom's ability to attract a shared pair of electrons in a bond is called electronegativity. For example, since chlorine has a stronger attraction for the bonding electrons in hydrogen chloride, chlorine is said to be more electronegative than hydrogen. A table of electronegativity values is a useful tool for determining the type of bonding between two atoms. Figure 2.19 shows the trend of electronegativity values in the periodic table. Note that the electronegativity of an element increases as you move to the right across a period or row on the periodic table, and decreases as you descend within a group (a column or family). Fluorine is the most electronegative element in the periodic table.

The difference in electronegativity between two elements in a bond is an indicator of how polar the bond is. The covalent bond between two fluorine atoms is non-polar because the electronegativity difference between the elements is zero. This means that the bonding pair of electrons in the fluorine molecule is shared equally. A bond between oxygen and fluorine is slightly polar because of the slight difference in the electronegativity values of these elements. The bond between hydrogen and fluorine is fairly polar because the electronegativity difference between these elements is great.

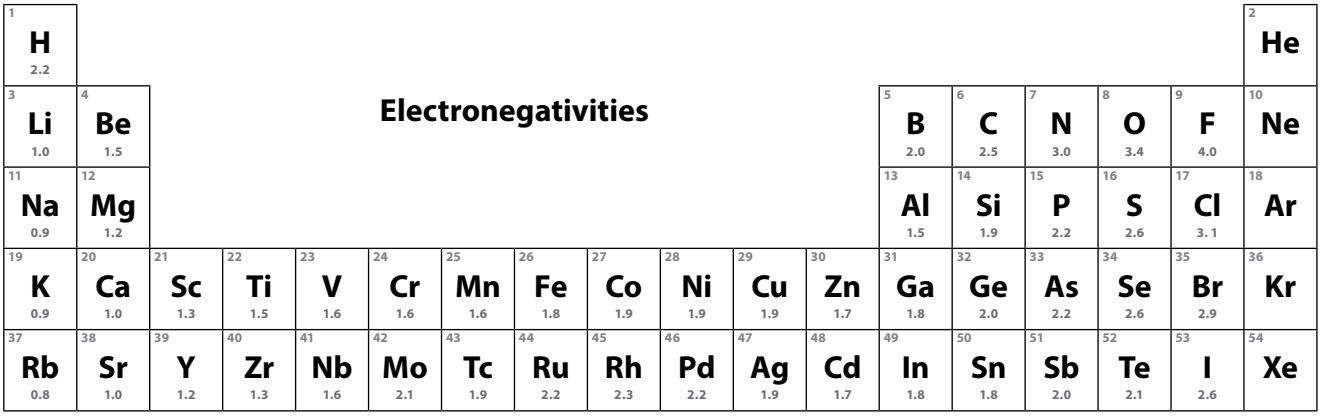

Figure 2.19: Electronegativity values for most of the elements of the periodic table

Predicting the Type of Bond between Elements

As the electronegativity difference between two elements increases, the polarity of the covalent bond between the elements also increases. In the extreme, the bond becomes so polar that no electron sharing occurs at all. In these cases, the electron is completely transferred from the element with a low electronegativity to the element with a much greater electronegativity. As a result, the atom that lost the electron becomes a positively charged ion or cation. The atom that gained the electron becomes a negatively charged ion or anion. These ions are bonded to each other as a result of the positive-negative (or electrostatic) attraction between them. Bonds between oppositely charged ions in a compound are called ionic bonds.

For example, the electronegativities of sodium and fluorine are 0.9 and 4.0. The difference in electronegativity between these elements is 4.0 – 0.9 = 3.1. Because the difference is so large, no electron sharing occurs. Instead, the single sodium electron in the 3*s* orbital transfers to fluorine to fill its 2*p* orbital. Therefore, sodium fluoride is an ionic compound. The difference in electronegativity between hydrogen and bromine is 2.9 – 2.2 = 0.7. Therefore, hydrogen bromide is a molecular compound that contains a polar covalent bond. The bond in hydrogen bromide is less ionic and more covalent in character than the bond in sodium fluoride. The electronegativity difference between elements in hydrogen, H_2, and phosphorus trihydride, PH_3, is 2.2 – 2.2 = 0. Therefore, both substances are molecular substances containing non-polar covalent bonds. Hydrogen is a molecular, diatomic, element, while phosphorus trihydride is a molecular compound. The following graphic summarizes the relationship between electronegativity difference and bond type (Figure 2.20).

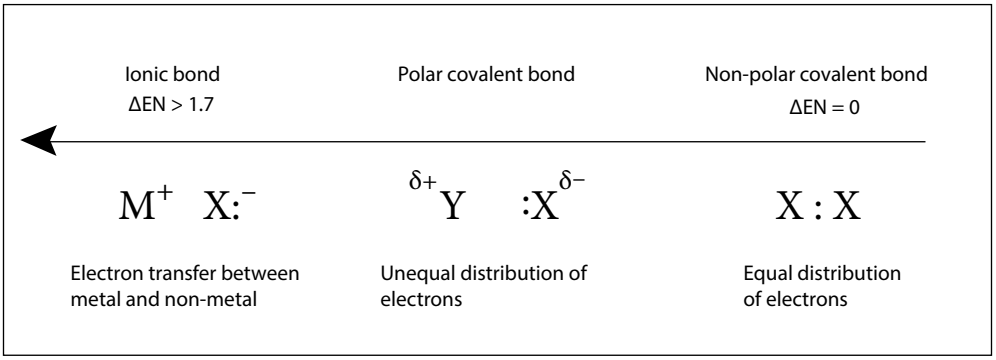

Figure 2.20: The bonding continuum. The "rule of thumb" in chemistry is that if the electronegativity difference between two atoms is greater than 1.7, the electrons will not be shared. Ions will form because one or more electrons will have been completely taken by one of the atoms. This typically occurs when metals meet non-metals.

Support Questions

18. **a)** Consider the bonding between the following pairs of elements. Using electronegativity difference, rank the bonds in order of increasing ionic character.

K and Br Na and Cl S and F H and I

b) Which bond is most covalent in character? Why?

c) Predict the bonding that will occur in each of the following compounds as being either non-polar covalent, polar covalent, or ionic. Classify each compound as being either an ionic or a molecular compound. Justify your prediction by calculating differences in electronegativity.

i) CO

ii) CaF_2

iii) K_2O

iv) PH_3

v) HBr

Suggested answer

Key Questions

Now work on your Key Questions in the online submission tool. You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

4. Draw an orbital overlap diagram to represent the bonding in ammonia, NH_3. **(5 marks)**

5. Draw the Lewis structure and structural formula for:

a) HCN

b) SO_3^{2-}

(8 marks total)

6. Experimental evidence suggests that the nitrogen atom in ammonia, NH_3, has four identical orbitals in the shape of a pyramid or tetrahedron.

a) Draw an energy-level diagram to show the formation of these hybrid orbitals. (**Hint:** No electron promotion is required.) **(4 marks)**

b) Name the type of hybrid orbitals found in NH_3. Of the four hybrid orbitals on the N atom, how many will take part in bonding? Explain. **(2 marks)**

c) Draw for yourself the energy-level diagram showing the hybrid orbitals formed in the C atom when it bonds. Now look at those hybrid orbitals and those of the N atom, and describe how the bonding with a N atom will differ with the bonding that occurs with a C atom, even though both atoms have four hybrid orbitals oriented in a tetrahedral shape. **(2 marks)**

7. **a)** Calculate the differences in electronegativity between the elements in each of these compounds. **(3 marks)**

i) $CaBr_2$

ii) Na_3N

iii) CH_4

b) Predict whether each of these compounds would be an ionic or a molecular compound and justify each prediction. **(3 marks)**

c) Rank the bonds in these compounds in order of decreasing ionic character. Where do we always find compounds containing metals, in this ranking order? **(2 marks)**

Now go on to Lesson 3. Send your answers to the Key Questions to the ILC when you have completed Unit 1 (Lessons 1 to 4).