Chapter 6
How Do Atoms Combine Together?

Very little of matter on the surface of this planet is in the form of elements. The vast majority is in the form of compounds. But why do elements combine together? How do they combine together? As you will see in this chapter, there are two ways in which atoms can combine.

6.1 Background

There is a simple experiment which helped chemists understand one of the ways in which atoms combined to form compounds. This is sometimes called the light-bulb experiment. An electrical circuit is constructed as shown in Figure 6.1.

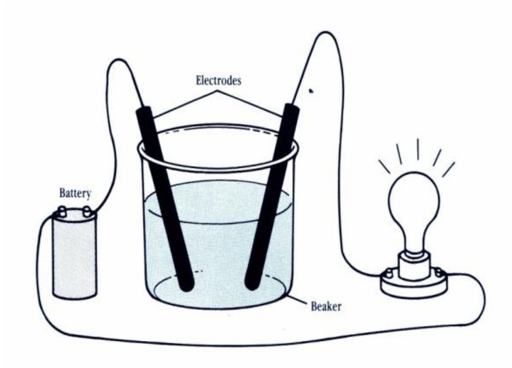


Figure 6.1 The equipment for the light-bulb experiment

If the beaker contains pure water, the bulb does not glow – the electrical circuit is not completed. If sugar is added to the water, the bulb still does not glow – there is still no electricity flowing around the circuit.
If salt is added to the water, the bulb glows brightly – an electric current is flowing through the solution.
Though both sugar and salt are compounds, in this experiment, they behave very differently. The way in which the atoms are held together in the two compounds must be fundamentally different.

It was the Swedish scientist, Svante Arrhenius who proposed the answer in 1884. Those compounds that, when dissolved in water, cause an electric current to flow, have themselves been split apart to form charged particles. It is these free-moving charged particles that convey the electric current through the solution from one electrode to the other. Such compounds are called *electrolytes*. Those compounds which, when dissolved in water, do not cause a current to flow, must still exist in their molecular form. They are called *non-electrolytes*.

6.2 Ions

This introduction to bonding will start with the compounds which are electrolytes. And for that, an explanation is needed as to why a compound would dissolve in water to give free-moving charged particles.

The answer to this question was discovered by American chemist Gilbert Lewis and the German chemist, Walter Kossel. They proposed that the driving force behind the formation of ions is the tendency to reach a lower energy state. The most stable electron energy state is to be found for the noble gases, that is the elements with full outer energy levels of electrons.

FORMATION OF AN ANION

For an element with many outer electrons, the easier way to acquire a noble-gas configuration is to gain one or more electrons. For example, fluorine with seven outer electrons only needs to gain one to acquire the same electron configuration as that of neon. By adding an electron, there is now one more electron than protons, that is it is no longer a neutral atom, instead it has become a negatively-charged ion, known as an *anion*.

The comparison between the neutral atom and the negative ion is illustrated below in Table 6.1.

Table 6.1 Comparison of atomic structure for the formation of the fluorine anion

Fluorine	**No. of protons**	**No. of electrons**	**Charge**
atom	9	9 (2 inner + 7 outer)	0
anion	9	10 (2 inner + 8 outer)	–1

FORMATION OF A CATION

For sodium, with one outer electron, it is easier to lose one electron and acquire the electron configuration of the preceding noble gas (neon) than to have to gain seven electrons to reach the electron configuration of argon. That is, by losing one electron, the inner layer of n = 2 now is the outermost layer. This positively-charged ion is known as a *cation*.

The comparison between the neutral atom and the positive ion is illustrated below in Table 6.2.

Table 6.2 Comparison of atomic structure for the formation of the sodium cation

Sodium	No. of protons	No. of electrons	Charge
atom	11	11 (10 inner + 1 outer)	0
cation	11	10 (10 inner + 0 outer)	+1

OVERVIEW

To summarize, atoms with one, two, or three electrons in their outer layer will lose those electrons to form positive ions and attain the electron configuration of the preceding noble gas. Those atoms with five, six, or seven electrons in their outer layer will gain electrons to form negative ions and attain the electron configuration of the following noble gas. These points are illustrated in Figure 6.2. As it is metals to the left of the Periodic Table (with few outer electrons) and nonmetals to the right of the Periodic Table (with closer to eight outer electrons), it is true to say that metal atoms lose electrons while nonmetal atoms gain electrons.

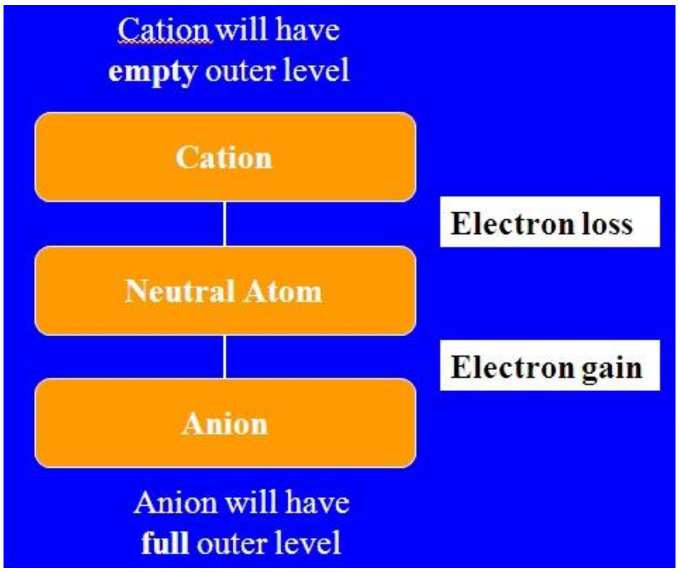
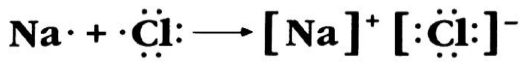


Figure 6.2 The relationship of ions to the original neutral atom.

6.3 The Formation of Ionic Compounds

The question then arises where the electrons come from or go to when positive and negative ions are formed. The answer is that the two processes have to occur simultaneously. One atom will lose electrons while another will gain those electrons. It is through the transfer of electrons that *ionic compounds* are formed.

For example, sodium chloride (common salt) is formed when sodium atoms lose an electron and chlorine atoms gain an electron. This process can be depicted using the electron-dot symbols as follows. An arrow is used to indicate the direction of the process (always from left to right).



These ions fit side-by-side in the solid phase and they are attracted to each other by the fact the neighbours possess opposite charges. The electrostatic attraction between ions of opposite charges is called an *ionic bond* and such a compound is called an *ionic compound*. The compound can be represented in a more compact form without the electron dots as: $[Na]^{+}[Cl]^{-}$. The way in which these ions pack together is referred to as a *crystal lattice* (Figure 6.3).

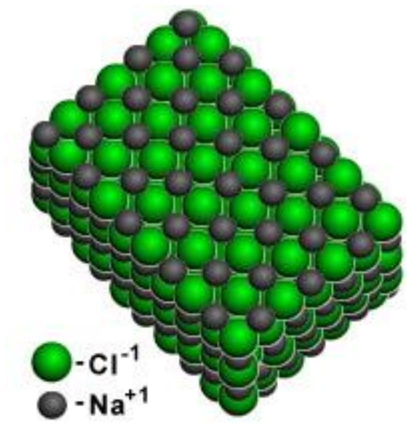


Figure 6.3 The arrangement of $[Na]^{+}$ and $[Cl]^{-}$ ions in a crystal lattice.

Figure 6.4 shows salt crystals. As the ratio of sodium and chlorine is one-to-one, a chemical formula can be written as NaCl. This simplest ratio is called the *formula unit*. A crystal of NaCl will contain an enormous number of these ions, but they will always be in the ratio of the formula unit, one $[Na^{+}]$ to one $[Cl^{-}]$.

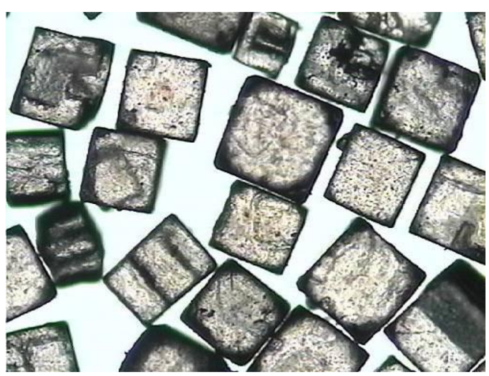
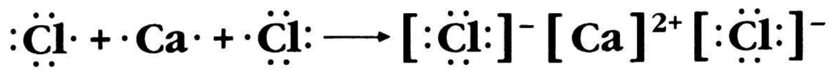


Figure 6.4 Crystals of salt

To balance the number of electrons given up by the metal with the number of electrons required by the nonmetal, it is sometimes necessary to have other than a one-to-one ratio. An example is that of calcium (with two outer electrons) and chlorine (with seven outer electrons). So each calcium atom needs to lose two electrons while each chlorine atom only needs to gain one. The solution is to use two chlorine atoms for each calcium atom, as can be seen below.



The compound can be represented in a more compact form without the electron dots as: $[Ca]^{2+}\cdot 2[Cl]^{-}$. The chemical formula of this compound will be $CaCl_2$, where the subscript two indicates the numbers of that ion in the formula unit (if there is no number, it is assumed to be one).

The combining of aluminum and oxygen illustrates the most complicated case. As aluminum is in Group 13, it has three electrons in its outer layer – it will need to lose all three to attain the same electron configuration as neon. Oxygen is in Group 16, it will need to gain two electrons to also gain the same electron configuration as neon. The only way all the electrons can be accounted for is if two aluminum atoms and three oxygen atoms combine together. The compound can be represented in a more compact form without the electron dots as: $2[Al]^{3+}\cdot 3[O]^{2-}$. The chemical formula of this compound will be Al_2O_3.

After becoming familiar with the concept of the formation of ions, it is not necessary to write out the electron-dot structures each time. The following two example illustrate this.

EXAMPLE 6.1

Determine the formula of the ionic compound formed between barium and sulfur.

Answer

Barium has two outer electrons (as it is in Group 2). To form an ion, these two outer electrons will be lost to give an empty outer electron layer. The cation formed will be: $\underline{Ba^{2+}}$.

Similarly, sulfur has six outer electrons. To form an ion, the atom will acquire two more electrons to give a full outer electron layer. The anion formed will be $\underline{S^{2-}}$.

As the charges are the same but of opposite sign, Ba^{2+} ions will combine with S^{2-} ions in a 1:1 ratio to form the compound $\underline{BaS}$.

EXAMPLE 6.2

Determine the formula of the ionic compound formed between lithium and nitrogen.

Answer

Lithium has one outer electron (as it is in Group 1). To form an ion, this outer electron will be lost to give an empty outer electron layer. The cation formed will be: $\underline{Li^{+}}$.

An atom of nitrogen has five outer electrons, three more are necessary to attain a full outer layer. The anion formed will be N^{3-}. To balance the charges, three times as many lithium cations will be needed to balance the nitrogen anions and the formula will be Li_3N.

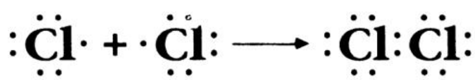
6.4 The Formation of Covalent Compounds

In the previous section, an explanation was proposed for the formation of compounds that are electrolytes. This concept stated that atoms could attain the more stable electron configuration of a noble gas by gaining or losing the electrons in the outer layer.

But this does not explain the formation of compounds that are non-electrolytes. These compounds consist of combinations of non-metals or a combination of a semimetal and a nonmetal. Thus each of the constituent elements has at least a half-full outer electron layer. To give-and-take so many electrons would not be feasible. Instead, these atoms attain the next noble-gas electron configuration by sharing electrons.

COVALENTLY-BONDED ELEMENTS

As a first example, chlorine atoms have seven electrons in their outer layer. It is known that the element exists in nature as pairs of atoms. We can account for this observation by assuming that for each pair of chlorine atoms, one electron is shared by each, so that each atom now has a total of eight electrons and the same electron configuration as argon.



This procedure is known as the formation of a *covalent bond*. The atoms are held together by the attraction of the positive nuclei of each atom towards the shared pair of electrons.

The chemical formula would be written as Cl_2. Unlike the ionic bond which is shared throughout the crystal lattice, a covalent bond is shared between pairs of atoms. To distinguish the two types of bonding, the Cl_2 unit is called a ***molecule***.

Hydrogen is another element in which the atoms exist in pairs. As the maximum number of electrons in the first layer is two, a bond can be formed by sharing the single electron of each hydrogen atom.

This sharing to give H_2 will result in each atom having two electrons – the same as the noble gas, helium.

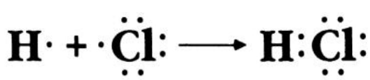
In Table 6.3, the gaseous elements are shown. For the noble gases (Group 18), as they have full outer layers of electrons, the elements are monatomic, that is, they exist as free atoms. For hydrogen, nitrogen, oxygen, fluorine, and chlorine, the elements obtain a full set of electrons in their outer layer by pairing up and sharing electrons to enable each atom to have a share in eight electrons. These elements, then, form diatomic molecules. In the case of oxygen, in addition to forming a diatomic molecule, under certain conditions can form a triatomic molecule, O_3, commonly called ozone.

Table 6.3 Formulas of the gaseous elements

		Group 15	Group 16	Group 17	Group 18
Period 1	H_2				He
Period 2		N_2	O_2 (O_3)	F_2	Ne
Period 3				Cl_2	Ar
Period 4					Kr
Period 5					Xe

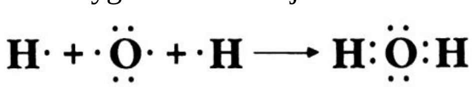
FORMATION OF COVALENT COMPOUNDS

Up to now, the examples have been of the same element joining together. Molecules are similarly formed when different nonmetal elements combine. The first example is the combination of hydrogen and chlorine. Hydrogen requires one electron to complete its electron layer while chlorine also needs one electron to complete its eight. Hydrogen will then have the same electron configuration as helium and chlorine that of argon.



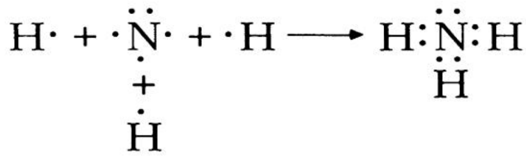
The resulting chemical formula will be HCl.

A second example will be that of hydrogen and oxygen. Hydrogen, as before, needs one more electron. Oxygen, with six electrons in its outer level, needs two more. So, in an analogous manner to the ionic bonding, each oxygen atom will join with two hydrogen atoms.

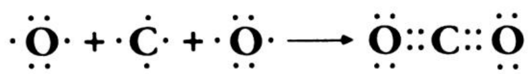


The resulting chemical formula will be H_2O.

A third example will be that of hydrogen and nitrogen. Hydrogen needs one more electron. Nitrogen, with five electrons in its outer level, needs three more. In this case, three hydrogen atoms will be needed to each share an electron with the nitrogen atom. The resulting chemical formula will be NH_3.

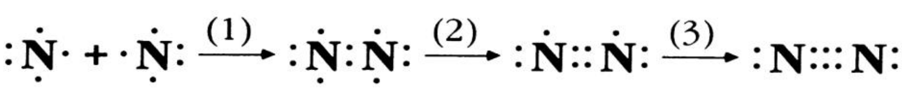


It is the role of chemistry to explain the world around us. So we know the formula of compounds and it is the application of the noble-gas configuration concept that helps explain why the compound can exist. For example, carbon and oxygen combine to form a gas of formula CO_2. Each carbon atom has four electrons in its outer level and needs a share in four more. Each oxygen atom has six electrons in its outer level and needs two more. The only way of accomplishing this task is if the atoms share two electrons each, instead of the one in the previous examples.



When four electrons are shared, chemists refer to this as a *double bond*, whereas in the earlier examples of sharing two electrons, they are called a *single bond*. The compound can be represented in a more compact form without the electron dots as: O=C=O. The resulting chemical formula will be CO_2.

Just as the element hydrogen is found as H_2 and chlorine as Cl_2, so the element nitrogen exists in nature as the molecule N_2. As each nitrogen atom has five electrons, each needs a share in three more to attain the noble-gas configuration of neon. In the diagram below, the process is shown stepwise. The sharing of six electrons is called a *triple bond*.



The compound can be represented in a more compact form without the electron dots as: N≡N. The resulting chemical formula will be N_2.

6.5 A Comparison of Ionic and Covalent Compounds

The type of bonding has a crucial influence on the properties of a compound. In Chapter 2, Section 2.4, the kinetic-molecular theory of matter was described. That Theory explained the melting of a solid as resulting from the vibrations of the particles overcoming the attractions between neighbouring particles.

For ionic compounds, the melting process requires the overcoming of the strong electrostatic attractions between the ions in the crystal lattice. Thus it is not surprising that ionic compounds are all solids at room temperature and require heating to very high temperatures to become liquid.

For covalent compounds, the electron-sharing between atoms is also very strong. However, this sharing is within each molecule. Between neighbouring molecules are weak *intermolecular attractions*. The nature of these weak attractions is beyond the scope of this beginning chemistry text. However, the ramifications of these weak attractions between molecules are of major importance, for it means that little energy is needed to overcome the attractions in the solid phase, and in the liquid phase. So many covalent (molecular) compounds are liquids or even gases at room temperature. Even those covalent compounds that are solid at room temperature, such as sugar (sucrose), melt on gentle heating. Two examples are compared in Table 6.4.

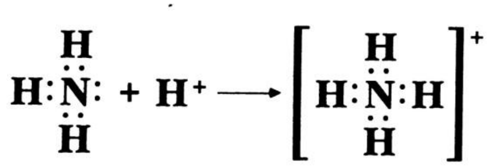
Table 6.4 A comparison of an ionic compound and a covalent (molecular) compound

Compound	NaCl	HCl
Type of bonding	ionic	covalent
Representation	$[Na]^+ [Cl]^-$	H−Cl
Phase at room temperature	solid	gas

6.6 Polyatomic Ions

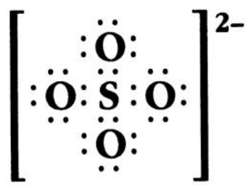
It is possible to have two or more atoms held together by covalent bonds and to have a charge as an ion. Such mixed species are called ***polyatomic ions***.

To illustrate how such a polyatomic ion theoretically forms, the NH_3 molecule, constructed in Section 6.4 will be used. Around the nitrogen atom, there are three pairs of electrons that are involved in forming covalent bonds with the hydrogen atoms. Each of these pairs of electrons is called a ***bonding pair***. In addition, there is one electron pair that is not involved in bonding. Those two electrons are known as a ***lone pair***. It is known that one of the polyatomic ions has the formula of $[NH_4]^+$. It can be imagined that a hydrogen cation – that is, a hydrogen atom which has lost its one electron – can attach itself to the lone pair. In this way, the nitrogen atom still has shares in eight electrons while the newly-acquired hydrogen atom shares both electrons formerly possessed by nitrogen alone.



The $[NH_4]^+$ ion is actually the only common polyatomic cation. All the others are polyatomic anions. The construction of the $[OH]^-$ polyatomic anion is the easiest to visualize. In Section 6.4, the combination of hydrogen atoms and oxygen atom to form H_2O was shown. An oxygen atom needs a share in two more electrons, while a hydrogen atom needs a share in one additional electron. If there is only one hydrogen atom available, then a second electron for the oxygen atom can be acquired from elsewhere – such as from an alkali metal. The oxygen atom will then complete the eight for its outer level, but the combination has a total of ten electrons compared with nine protons resulting in a net charge of −1. The combination of a hydrogen atom, an oxygen atom, and an ‘extra’ electron to give the $[OH]^{-}$ polyatomic anion is shown below.

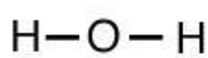
Many polyatomic anions contain four or five atoms. A common example is the $[SO_4]^{2-}$ ion. To imagine how this can be formed, the sulfur atom has six electrons, adding two more gives the $[S]^{2-}$ ion. Each of the four oxygen atoms can then add to a lone pair of the anion to give:



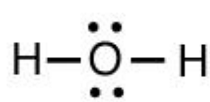
6.7 Structural Formulas

For covalently bonded molecules and polyatomic ions, lines instead of pairs of dots can be used to represent bonds. These representations are called ***structural formulas*** (or bond-line diagrams). As an example, would be represented as :

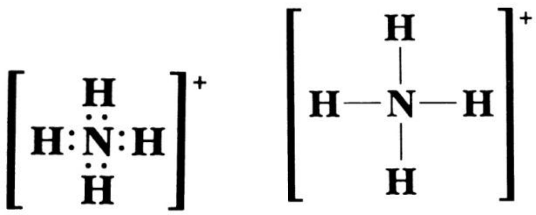
The lone pairs of electrons are usually not represented. The only exception are the lone pairs on a central atom. Usually, they are not shown, for example, water would normally be represented as a structural formula as:



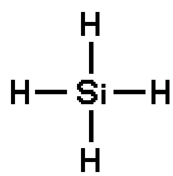
However, we will see in the next Section that the lone pairs on a central atom play a key role in determining the shape of the molecule. For these circumstances, those lone pairs on the central atom would be shown, as below.



Likewise, the $[NH_4]^{+}$ ion can be represented as both an electron-dot formula and as a structural formula. Even for the structural formula, the square brackets are necessary to indicate that it is a polyatomic ion, not a neutral molecule.



EXAMPLE 6.3
Determine the formula of the covalent compound formed between silicon and hydrogen. Then draw its structural formula.
Answer
An atom of silicon has four outer electrons. It will therefore need an additional four electrons to have a share in a total of eight electrons.
A hydrogen atom has one electron and it needs to share one other to complete its layer.
Thus four hydrogen atoms can provide one silicon atom with the required number of electrons.
The formula of the compound will be SiH_4.
The electron-dot structure is:

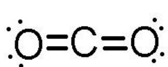


6.8 Shapes of Simple Molecules

Up to now, the electron-dot diagrams of covalently-bonded molecules have been constructed. The two-dimensional electron-dot diagrams drawn above have no relevance to the possible three-dimensional shape of the molecule. This Section will introduce the shapes of some simple molecules and discuss how the shapes can be explained.

CARBON DIOXIDE

The electron-dot diagram for carbon dioxide below shows the two double bonds around the central carbon atom. Conventionally, the diagram is drawn like this to make the molecule look symmetrical, nothing more.



Chemists have experimentally determined the shape of the carbon dioxide molecule and it is linear with a O=C=O bond angle of 180°. The shape of the molecule can be represented as a ***ball-and-stick model***. Originally, these were indeed constructed of wooden balls and metal rods. Now, computer-generated images are more commonly used. Figure 6.5 shows the representation for the carbon dioxide molecule. Chemists explain this shape in terms of the surrounding atoms preferring to be as far apart from each other as possible.

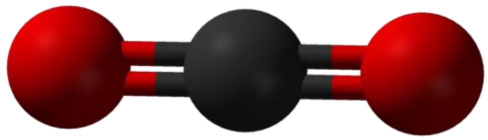


Figure 6.5 The ball-and-stick model for the carbon dioxide molecule

COLOUR CODING FOR MOLECULAR REPRESENTATIONS

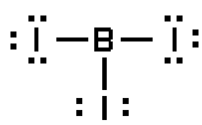
In Figure 6.5, the carbon atom is coloured grey-black and the oxygen atoms, red. There is a standard convention for the choice of colours in these molecular representations. The agreed standard colours for the most commonly used elements are listed in Table 6.5.

Table 6.5 The standard colour-code for atoms in molecular models

Element	**Colour**
Bromine	Dark red
Carbon	Black (or Grey)
Chlorine	Green
Fluorine	Light green
Hydrogen	White
Iodine	Violet
Nitrogen	Blue
Oxygen	Red
Phosphorus	Orange
Sulfur	Yellow

BORON TRIIODIDE

Here is the electron-dot structure of boron triiodide. Unusually, there are only three electron pairs around the central boron atom. For convenience, the electron-dot structure is usually drawn with the bonds at 90° angles.



Applying the same principle that the surrounding atoms would prefer to be as far apart as possible, the boron triiodide molecule would be predicted to be planar with the three I–B–I bond angle of 120°. This molecular shape is indeed what chemists find for boron triiodide.

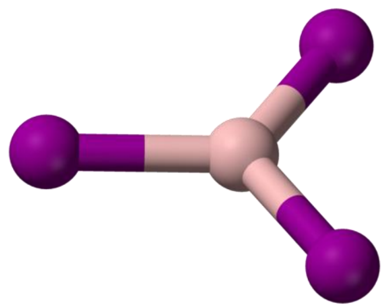
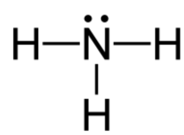


Figure 6.6 The ball-and-stick model for the boron triiodide molecule

AMMONIA

The next compound to be discussed is ammonia. In the previous Section, it was shown that the electron-dot structure is:



From a comparison with the boron triiodide molecule, one might think that ammonia, too, would be planar. But it is not. The actual molecular shape is shown in Figure 6.7, with the hydrogen atoms bent down below the plane. The H–N–H bond angle is significantly less than 120°.

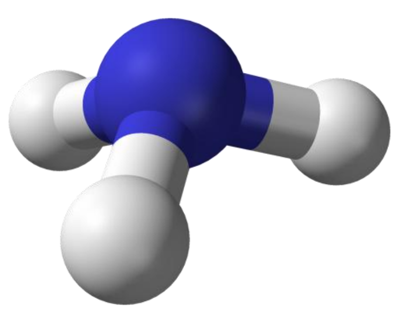


Figure 6.7 The ball-and-stick model for the ammonia molecule

How can we explain this shape? The clue is given by the electron-dot structure. In addition to the three covalent bonds to the hydrogen atoms, the nitrogen atom possesses a lone pair. This lone pair occupies space as if it were another atom. Lone pairs cannot be seen, it is the way in which the H–N bonds are pushed down that suggests the lone pair of electrons is there, 'doing the pushing.' We can represent the lone pair by an atom-shape with a pair of dots. So for ammonia, if we could see the lone pair, the molecule would look like Figure 6.8.

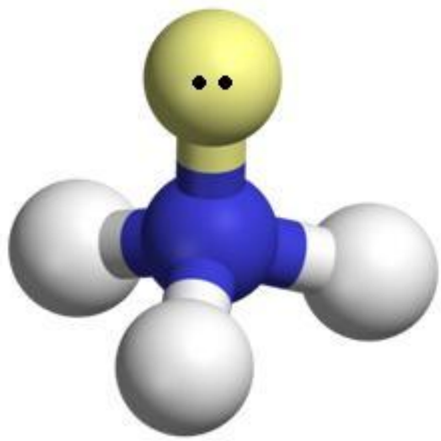
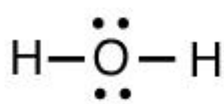


Figure 6.8 The ball-and-stick model for the ammonia molecule showing how the presence of the 'invisible' lone pair could result in the known shape of the molecule.

WATER

Finally, the water molecule will be considered. The electron-dot structure is shown below.



Simplistically, one might think that the molecule would be linear, like carbon dioxide. However, after considering the shape of the ammonia molecule, it would be apparent that the two lone pairs will play a vital role in determining the structure. This is indeed the case, with the water molecule being in the shape of a vee (Figure 6.9).

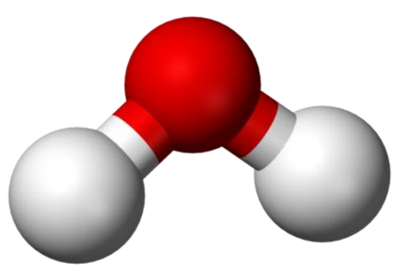


Figure 6.9 The ball-and-stick model for the water molecule

If the two lone pairs are added to the molecular shape, then it becomes obvious why the H–O–H bond should be 'squished down.'

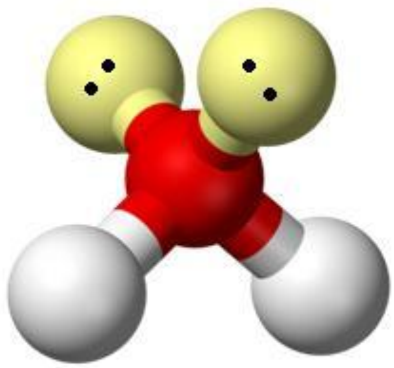
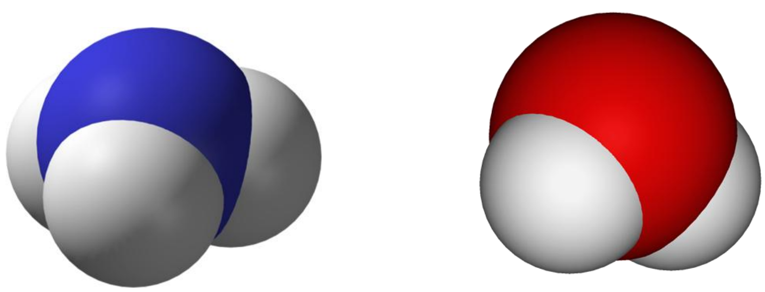


Figure 6.10 The ball-and-stick model for the water molecule showing how the presence of the two 'invisible' lone pairs could result in the known shape of the molecule.

SPACE-FILLING MODELS

Ball-and-stick models are very useful for representing the shape of a molecule. Nevertheless, atoms are not held together by 'molecular sticks.' In fact, it is known that in covalently bonded molecules, the electron pairs shared by atoms bring the atoms in close proximity. For this reason, space-filling models have been devised which are about as close to 'reality' as chemists can reach. The same colour-coding is used for these models as with the ball-and-stick models. As examples, the space-filling representations of ammonia and water are given below (Figure 6.11).



(a) (b)

Figure 6.11 The space-filling models of (a) ammonia and (b) water.

6.9 Where Next

Having been introduced to the two types of bonding in compounds, the next step is to see how compounds are named. Just as there are two categories of bonding, so there are two corresponding types of naming.