CHAPTER 6

Chemical Bonding

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August 26, 2021

1 INTRODUCTION

When you look at everything around you and what it is made of, you will realise that atoms seldom exist on their own. More often, the things around us are made up of different atoms that have been joined together. This is called **chemical bonding**. Chemical bonding is one of the most important processes in chemistry because it allows all sorts of different molecules and combinations of atoms to form, which then make up the objects in the complex world around us.

1.1 What happens when atoms bond?

A **chemical bond** is formed when atoms are held together by attractive forces. This attraction occurs when electrons are shared between atoms, or when electrons are exchanged between the atoms that are involved in the bond. The sharing or exchange of electrons takes place so that the outer energy levels of the atoms involved are filled, making the atoms are more stable. If an electron is **shared**, it means that it will spend its time moving in the electron orbitals around both atoms. If an electron is **exchanged** it means that it is transferred from one atom to another. In other words one atom gains an electron while the other *loses* an electron.

DEFINITION : Chemical bond

A chemical bond is the physical process that causes atoms and molecules to be attracted to each other and held together in more stable chemical compounds.

The type of bond that is formed depends on the elements that are involved. In this chapter, we will be looking at three types of chemical bonding: **covalent, ionic** and **metallic bonding**.

You need to remember that it is the valence electrons (those in the outermost level) that are involved in bonding and that atoms will try to fill their outer energy levels so that they are more stable. The noble gases have completely full outer energy levels, so are very stable and do not react easily with other atoms.

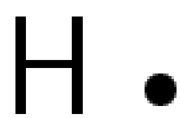
2 LEWIS STRUCTURES

Lewis notation uses dots and crosses to represent the **valence electrons** on different atoms. The chemical symbol of the element is used to represent the nucleus and the inner electrons of the atom. To determine which are the valence electrons we look at the last *energy level* in the atom's electronic structure (Chapter 4). For example, chlorine's electronic structure can be written as: $1\text{s}^2 2\text{s}^2 2\text{p}^6 3\text{s}^2 3\text{p}^5$ or $[\text{Ne}] 3\text{s}^2 3\text{p}^5$. The last energy level is the third one and it contains 7 electrons. These are the valence electrons.

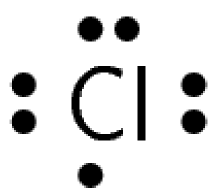
> **TIP**
>
> *If we write the condensed electron configuration, then we can easily see the valence electrons.*

For example:

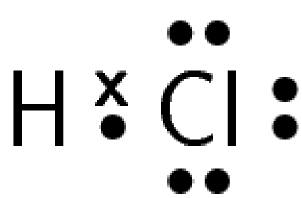
A hydrogen atom (one valence electron) would be represented like this:



A chlorine atom (seven valence electrons) would look like this:



A molecule of hydrogen chloride would be shown like this:



The dot and cross in between the two atoms, represent the pair of electrons that are shared in the covalent bond.

Table 1 below gives some further examples of Lewis diagrams.

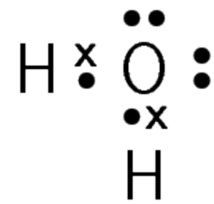
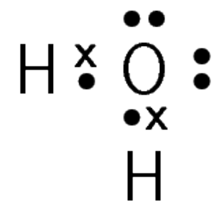
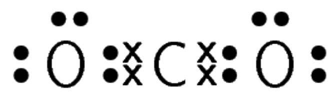
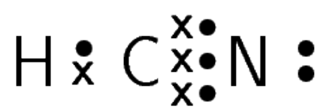
Iodine	I_2	
Water	H_2O	
Carbon dioxide	CO_2	
Hydrogen cyanide	HCN	

Table 1: Lewis diagrams for some simple molecules

For carbon dioxide, you can see how we represent a double bond in Lewis notation. As there are two bonds between each oxygen atom and the carbon atom, two pairs of valence electrons link them. Similarly, hydrogen cyanide shows how to represent a triple bond.

3 COVALENT BONDING

3.1 The nature of the covalent bond

Covalent bonding occurs between the atoms of **non-metals**. The outermost orbitals of the atoms overlap so that unpaired electrons in each of the bonding atoms can be shared. By overlapping orbitals, the outer energy shells of all the bonding atoms are filled. The shared electrons move in the orbitals around both atoms. As they move, there is an attraction between these negatively charged electrons and the positively charged nuclei. This attractive force holds the atoms together in a covalent bond.

DEFINITION : Covalent bond

Covalent bonding is a form of chemical bonding where pairs of electrons are shared between atoms.

You will have noticed in Table 1 that the number of electrons that are involved in bonding varies between atoms.

TIP

There is a relationship between the valency of an element and its position on the periodic table. For the elements in groups 1 and 2, the valency is the group number. For the elements in groups 13–18, the valency is the group number minus 10. For the transition metals, the valency can vary. In these cases we indicate the valency by a roman numeral after the element name, e.g. iron (III) chloride.

We can say the following:

- A **single covalent** bond is formed when two electrons are shared between the same two atoms, one electron from each atom.
- A **double covalent** bond is formed when four electrons are shared between the same two atoms, two electrons from each atom.
- A **triple covalent** bond is formed when six electrons are shared between the same two atoms, three electrons from each atom.

You should also have noticed that compounds can have a mixture of single, double and triple bonds and that an atom can have several bonds. In other words, an atom does not need to share all its valence electrons with one other atom, but can share its valence electrons with several different atoms.
We say that the **valency** of the atoms is different.

DEFINITION : Valency

The number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.

Below are a few examples. Remember that it is only the valence electrons that are involved in bonding and so when diagrams are drawn to show what is happening during bonding, it is only these electrons that are shown.

WORKED EXAMPLE 1: COVALENT BONDING

QUESTION

How do hydrogen and chlorine atoms bond covalently in a molecule of hydrogen chloride?

SOLUTION

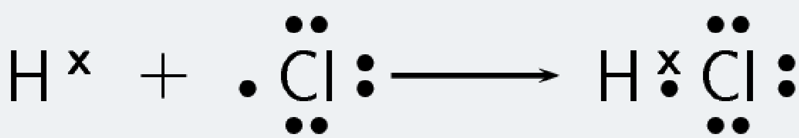
Step 1: Determine the electron configuration of each of the bonding atoms

A chlorine atom has 17 electrons and an electron configuration of $[\text{Ne}]3\text{s}^2 3\text{p}^5$. A hydrogen atom has only one electron and an electron configuration of 1s^1.

Step 2: Determine how many of the electrons are paired or unpaired

Chlorine has seven valence electrons. One of these electrons is unpaired. Hydrogen has one valence electron and it is unpaired.

Step 3: Work out how the electrons are shared The hydrogen atom needs one more electron to complete its outermost energy level. The chlorine atom also needs one more electron to complete its outermost energy level. Therefore one pair of electrons must be shared between the two atoms. A single covalent bond will be formed.



WORKED EXAMPLE 2: COVALENT BONDING INVOLVING MULTIPLE BONDS

QUESTION

How do nitrogen and hydrogen atoms bond to form a molecule of ammonia (NH_3)?

SOLUTION

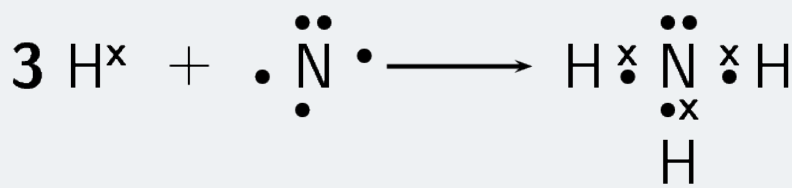
Step 1: Give the electron configuration

A nitrogen atom has seven electrons, and an electron configuration of $[He]2s^2 2p^3$. A hydrogen atom has only one electron, and an electron configuration of $1s^1$.

Step 2: Give the number of valence electrons

Nitrogen has five valence electrons. Three of these electrons are unpaired. Hydrogen has one valence electron and it is unpaired.

Step 3: Work out how the electrons are shared Each hydrogen atom needs one more electron to complete its valence energy shell. The nitrogen atom needs three more electrons to complete its valence energy shell. Therefore three pairs of electrons must be shared between the four atoms involved. Three single covalent bonds will be formed.



WORKED EXAMPLE 3: COVALENT BONDING INVOLVING A DOUBLE BOND

QUESTION

How do oxygen atoms bond covalently to form an oxygen molecule?

SOLUTION

Step 1: Determine the electron configuration of the bonding atoms.

Each oxygen atom has eight electrons, and their electron configuration is $[\text{He}]2\text{s}^2 2\text{p}^4$

Step 2: Determine the number of valence electrons for each atom and how many of these electrons are paired and unpaired.

Each oxygen atom has six valence electrons. Each atom has two unpaired electrons.

Step 3: Work out how the electrons are shared Each oxygen atom needs two more electrons to complete its valence energy shell. Therefore two pairs of electrons must be shared between the two oxygen atoms so that both outermost energy levels are full. A double bond is formed.

3.2 Properties of covalent compounds

Covalent compounds have several properties that distinguish them from ionic compounds and metals. These properties are:

1. The melting and boiling points of covalent compounds are generally lower than those of ionic compounds.
2. Covalent compounds are generally more flexible than ionic compounds. The molecules in covalent compounds are able to move around to some extent and can sometimes slide over each other (as is the case with graphite, which is why the lead in your pencil feels slightly slippery). In ionic compounds, all the ions are tightly held in place.
3. Covalent compounds generally are not very soluble in water, for example plastics are covalent compounds and many plastics are water resistant.
4. Covalent compounds generally do not conduct electricity when dissolved in water, for example iodine dissolved in pure water does not conduct electricity.

4 IONIC BONDING

4.1 The nature of the ionic bond

When electrons are transferred from one atom to another it is called **ionic bonding**.

Electronegativity is a property of an atom, describing how strongly it attracts or holds onto electrons. Ionic bonding takes place when the difference in electronegativity between the two atoms is more than $1,7$. This usually happens when a metal atom bonds with a non-metal atom. When the difference in electronegativity is large, one atom will attract the shared electron pair much more strongly than the other, causing electrons to be transferred to the atom with higher electronegativity. When ionic bonds form, a metal donates one or more electrons, due to having a low electronegativity, to form a positive ion or cation. The non-metal atom has a high electronegativity, and therefore readily gains electrons to form a negative ion or anion. The two ions are then attracted to each other by electrostatic forces.

DEFINITION : Ionic bond

An ionic bond is a type of chemical bond where one or more electrons are transferred from one atom to another.

Example 1:

In the case of NaCl, the difference in electronegativity between Na $(0,93)$ and Cl $(3,16)$ is $2,1$. Sodium has only one valence electron, while chlorine has seven. Because the electronegativity of chlorine is higher than the electronegativity of sodium, chlorine will attract the valence electron of the sodium atom very strongly. This electron from sodium is transferred to chlorine. Sodium loses an electron and forms an Na^{+} ion.

$$\mathrm{Na}^{\bullet} \longrightarrow \mathrm{Na}^{+} + \text{electron}$$

Chlorine gains an electron and forms a Cl^{-} ion.

> **NOTE**
>
> *Chlorine is a diatomic molecule and so for it to take part in ionic bonding, it must first break up into two atoms of chlorine. Sodium is part of a metallic lattice and the individual atoms must first break away from the lattice.*

The electron is therefore transferred from sodium to chlorine:

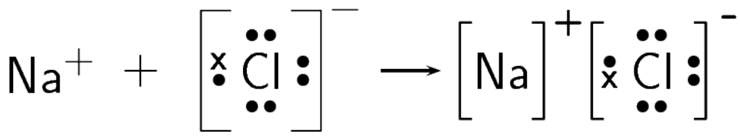


Figure 1: Ionic bonding in sodium chloride

The balanced equation for the reaction is:

$$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$$

Example 2:
Another example of ionic bonding takes place between magnesium (Mg) and oxygen (O_2) to form magnesium oxide (MgO). Magnesium has two valence electrons and an electronegativity of $1,31$, while oxygen has six valence electrons and an electronegativity of $3,44$. Since oxygen has a higher electronegativity, it attracts the two valence electrons from the magnesium atom and these electrons are transferred from the magnesium atom to the oxygen atom. Magnesium loses two electrons to form Mg^{2+}, and oxygen gains two electrons to form O^{2-}. The attractive force between the oppositely charged ions is what holds the compound together.

The balanced equation for the reaction is:

$$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$$

Because oxygen is a diatomic molecule, two magnesium atoms will be needed to combine with one oxygen molecule (which has two oxygen atoms) to produce two units of magnesium oxide (MgO).

4.2 The crystal lattice structure of ionic compounds

Ionic substances are actually a combination of lots of ions bonded together into a giant molecule. The arrangement of ions in a regular, geometric structure is called a **crystal lattice**. So in fact (NaCl) does not contain one (Na) and one (Cl) ion, but rather a lot of these two ions arranged in a crystal lattice where the ratio of Na to Cl ions is $1 : 1$.

The structure of the crystal lattice is shown below.

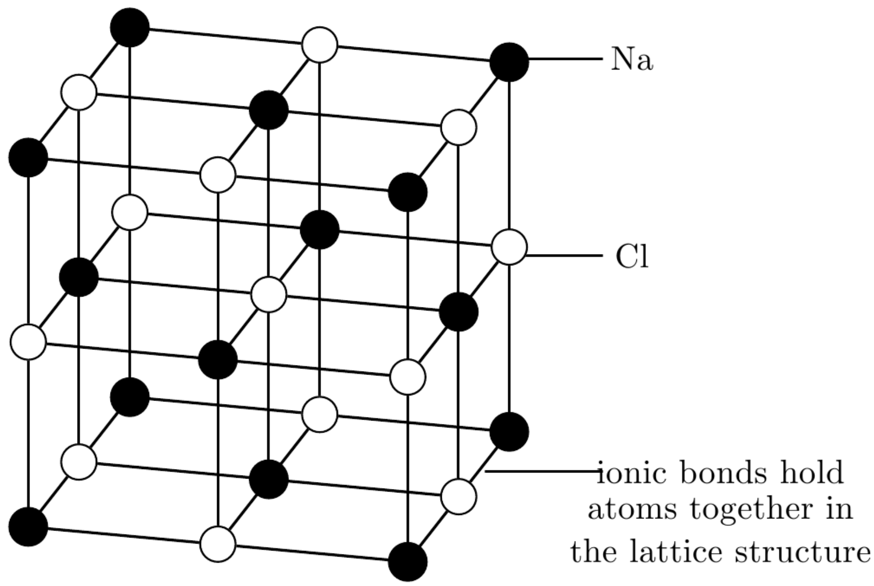


Figure 2: The crystal lattice arrangement in NaCl

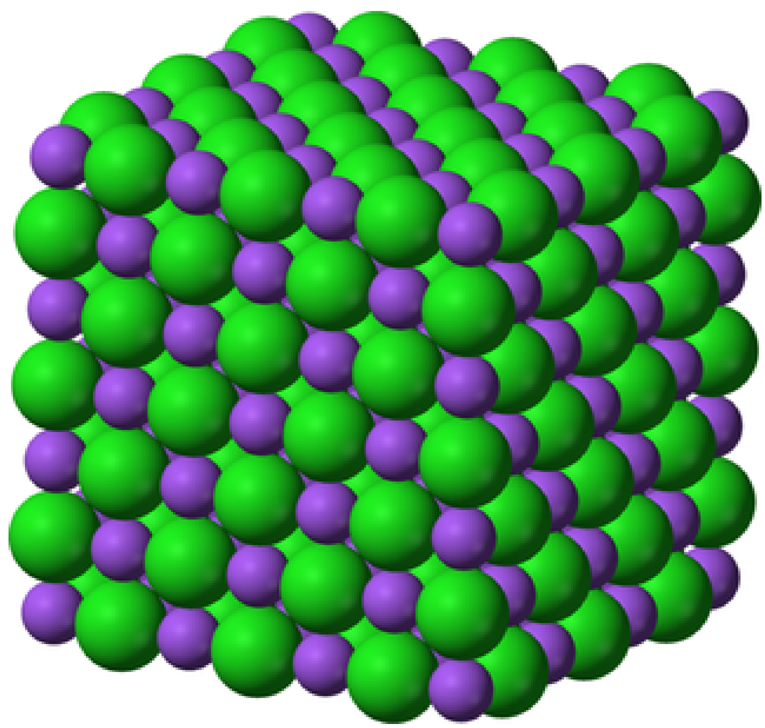


Figure 3: A space filling model of the sodium chloride lattice

4.3 Properties of ionic compounds

Ionic compounds have a number of properties:

1. Ions are arranged in a lattice structure
2. Ionic solids are crystalline at room temperature
3. The ionic bond is a strong electrostatic attraction. This means that ionic compounds are often hard and have high melting and boiling points
4. Ionic compounds are brittle and bonds are broken along planes when the compound is put under pressure (stressed)
5. Solid crystals do not conduct electricity, but ionic solutions do

5 METALLIC BONDING

5.1 The nature of the metallic bond

The structure of a metallic bond is quite different from covalent and ionic bonds. In a metallic bond, the valence electrons are delocalised, meaning that an atom's electrons do not stay around that one nucleus. In a metallic bond, the positive atomic nuclei (sometimes called the "atomic kernels") are surrounded by a sea of delocalised electrons which are attracted to the nuclei (see Figure 4 below).

DEFINITION : Metallic bond

Metallic bonding is the electrostatic attraction between the positively charged atomic nuclei of metal atoms and the delocalised electrons in the metal.

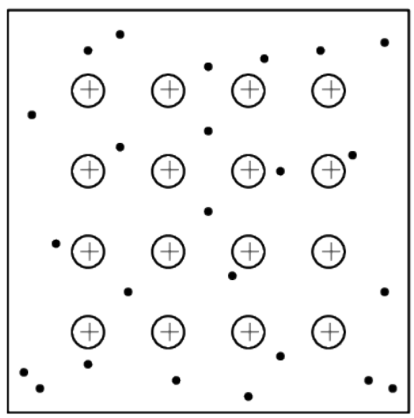


Figure 4: Positive atomic nuclei (+) surrounded by delocalised electrons (•)

Figure 5: Ball and stick model of copper

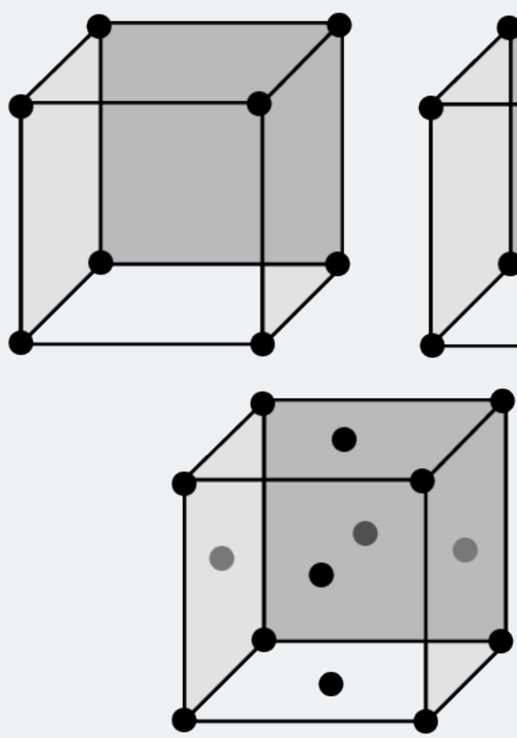
5.2 Properties of metals

1. Metals are *shiny*.
2. Metals *conduct electricity* because electrons are free to move.
3. Metals conduct heat because the positive nuclei are packed closely together and can easily transfer the heat.
4. Metals have a *high melting point* because the bonds are strong and a high density because of the tight packing of the nuclei.

ACTIVITY : Building models

Using coloured balls (or jellytots) and sticks (or toothpicks) build models of each type of bonding. Think about how to represent each kind of bonding. For example, covalent bonding could be represented by simply connecting the balls with sticks to represent the molecules, while for ionic bonding you may wish to construct part of the crystal lattice.

Do some research on types of crystal lattices (although the section on ionic bonding only showed the crystal lattice for sodium chloride, many other types of lattices exist) and try to build some of these. Share your findings with your class and compare notes to see what types of crystal lattices they found. How would you show metallic bonding?



6 WRITING FORMULAE

In Chapter 2 you learnt about the writing of chemical formulae. Table 2 shows some of the common anions and cations that you should know.

Name of compound ion	**formula**	**Name of compound ion**	**formula**
Acetate (ethanoate)	CH_3COO^-	Manganate	MnO_4^{2-}
Ammonium	NH_4^+	Nitrate	NO_3^-
Carbonate	CO_3^{2-}	Nitrite	NO_2^-
Chlorate	ClO_3^-	Oxalate	C_2O4^{2-}
Chromate	CrO_4^-	Oxide	O_2^-
Cyanide	CN^-	Permanganate	MnO_4^-
Dihydrogen phosphate	$H_2PO_4^-$	Peroxide	O_2^{2-}
Hydrogen carbonate	HCO_3^-	Phosphate	PO_4^{3-}
Hydrogen phosphate	HPO_4^{3-}	Phosphide	P^{3-}
Hydrogen sulphate	HSO_4^-	Sulphate	SO_4^{2-}
Hydrogen sulphite	HSO_3^-	Sulphide	S^{2-}
Hydroxide	OH^-	Sulphite	SO_3^{2-}
Hypochlorite	ClO^-	Thiosulphate	$S_2O_3^{2-}$

Table 2: Table showing common compound ions and their formulae

6.1 Chemical compounds: names and masses

In Chapter 4 you learnt about atomic masses. In this chapter we have learnt that atoms can combine to form compounds. Molecules are formed when atoms combine through covalent bonding, for example ammonia is a molecule made up of three hydrogen atoms and one nitrogen atom. The **relative molecular mass** (M) of ammonia (NH_3) is:

$$\begin{aligned} M &= \text{relative atomic mass of one nitrogen} + \text{relative atomic mass of three hydrogens} \\ &= 14,0 + 3(1,01) \\ &= 17,03 \end{aligned}$$

One molecule of NH_3 will have a mass of $17,03$ units. When sodium reacts with chlorine to form sodium chloride, we do not get a molecule of sodium chloride, but rather a sodium chloride crystal lattice. Remember that in ionic bonding molecules are not formed. We can also calculate the mass of one unit of such a crystal. We call this a **formula unit** and the mass is called the **formula mass**. The formula mass for sodium chloride is:

$$\begin{aligned} M &= \text{ relative atomic mass of one sodium atom } + \text{ relative atomic mass of one chlorine atom} \\ &= 23,0 + 35,45 \\ &= 58,45 \end{aligned}$$

The formula mass for NaCl is $58,45$ units.

7 CHAPTER SUMMARY

- A **chemical bond** is the physical process that causes atoms and molecules to be attracted to each other and held together in more stable chemical compounds.
- Atoms are more **reactive**, and therefore more likely to bond, when their outer electron orbitals are not full. Atoms are less reactive when these outer orbitals contain the maximum number of electrons. This explains why the noble gases do not react.
- **Lewis** notation is one way of representing molecular structure. In Lewis notation, dots and crosses are used to represent the valence electrons around the central atom.
- When atoms bond, electrons are either shared or exchanged.
- **Covalent bonding** occurs between the atoms of non-metals and involves a sharing of electrons so that the orbitals of the outermost energy levels in the atoms are filled.
- A **double** or **triple bond** occurs if there are two or three electron pairs that are shared between the same two atoms.
- The **valency** is the number of electrons in the outer shell of an atom which are able to be used to form bonds with other atoms.
- Covalent compounds have lower melting and boiling points than ionic compounds. Covalent compounds are also generally flexible, are generally not soluble in water and do not conduct electricity.
- An **ionic** bond occurs between atoms where there is a large difference in electronegativity. An exchange of electrons takes place and the atoms are held together by the electrostatic force of attraction between the resulting oppositely-charged ions.
- Ionic solids are arranged in a **crystal lattice** structure.
- Ionic compounds have high melting and boiling points, are brittle in nature, have a lattice structure and are able to conduct electricity when in solution.
- A **metallic bond** is the electrostatic attraction between the positively charged nuclei of metal atoms and the delocalised electrons in the metal.
- Metals are able to conduct heat and electricity, they have a metallic lustre (shine), they are both malleable (flexible) and ductile (stretchable) and they have a high melting point and density.
- We can work out the relative molecular mass for covalent compounds and the formula mass for ionic compounds and metals.

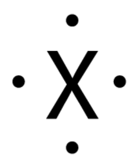
8 EXERCISES

8.1 Exercise 1

1. Study the following lewis diagrams and select a possible element for X:

   1.1

   1.2



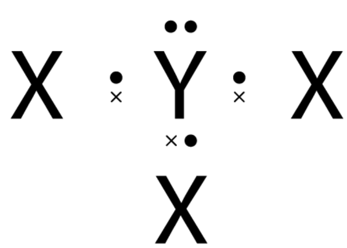
2. For Bromine gas(Br_2), draw the correct lewis diagram.

3. Which of the following molecules (CO_2,Br_2,HCN, HCl and N_2) contains a double bond?

4. Consider the chemical reaction between nitrogen and hydrogen that forms NH_3. Answer the following questions:

   4.1 What is the number of electrons in the outer energy level of nitrogen?

   4.2 The lewis structure of the product that is formed looks as follows:



   Is this true or false?

   4.3 What is the name of the product that is formed?

8.2 Exercise 2

1. For each of the following elements give the group number, the number of valence electrons and the number of electrons needed to fill the outer shell.

   1.1 He

      1.1.1 Give the group number:

      1.1.2 Give the number of valence electrons:

      1.1.3 Give the number of electrons needed to fill the outer shell:

   1.2 Li

      1.2.1 Give the group number:

      1.2.2 Give the number of valence electrons:

      1.2.3 Give the number of electrons needed to fill the outer shell:

   1.3 B

      1.3.1 Give the group number:

      1.3.2 Give the number of valence electrons:

      1.3.3 Give the number of electrons needed to fill the outer shell:

   1.4 F

      1.4.1 Give the group number:

      1.4.2 Give the number of valence electrons:

      1.4.3 Give the number of electrons needed to fill the outer shell:

   1.5 C

      1.5.1 Give the group number:

      1.5.2 Give the number of valence electrons:

      1.5.3 Give the number of electrons needed to fill the outer shell:

2. Draw the correct lewis diagrams for the following compounds:

   2.1 H_2S

   2.2 HCN

8.3 Exercise 3

1. Define ionic bond:

2. Magnesium and chlorine reacts to form magnesium chloride.

   2.1 What is the difference in electronegativity between these two elements?

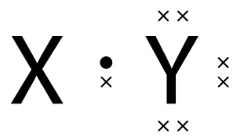
2.2 Give the formula for magnesium ion.

2.3 Give the formula for chloride ion.

2.4 Give the ionic compound that is produced during the reaction.

3. Is the following lewis diagrams correct for the given ionic compounds?

3.1 Sodium iodide (NaI)



3.2 Calcium bromide ($CaBr_2$)

8.4 Exercise 4

1. Study the following elements and compounds and answer the questions (H_2O, NaCl, Cu).

1.1 Give an example of covalent bonding.

1.2 Give an example of ionic bonding.

1.3 Give an example of metallic bonding.

2. What type of atoms are involved in the following bonds?

2.1 Covalent bond

2.2 Ionic bond

2.3 Metallic bond

3. Classify the following as true or false:

3.1 The melting point of compounds with covalent bonds are usually low.

3.2 The melting point of compounds with metallic bonds are usually high.

3.3 An ionic compound is able to conduct electricity.

4. Identify the type of bond (covalent, ionic or metallic) in each of the following compounds:

4.1 H_2SO_4

4.2 NaI

4.3 Zn

8.5 Exercise 5

1. Write the chemical formulae for each of the following compounds and calculate the relative molecular mass:

   1.1 Hydrogen cyanide

   1.2 Carbon dioxide

   1.3 Sodium carbonate

2. What is the formula and the name of the compound that forms between:

   2.1 NH_4^+ and S^{2-}

   2.2 Mg^{2+} and P^{3-}

   2.3 Al^{3+} and $Cr_2O_7^{2-}$

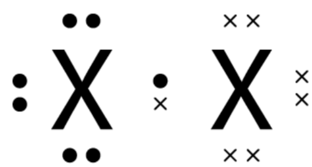
9 ANSWERS FOR EXERCISES

9.1 Exercise 1

1.1 Beryllium

1.2 Carbon

2. Bromine lewis diagram



3. CO_2

4.1 5

4.2 True

4.3 Ammonia

9.2 Exercise 2

1.1 He

1.1.1 4

1.1.2 4

1.1.3 4

1.2 Li

1.2.1 1

1.2.2 1

1.2.3 7

1.3 B

1.3.1 3

1.3.2 3

1.3.3 5

1.4 F

   1.4.1 7

   1.4.2 7

   1.4.3 1

1.5 C

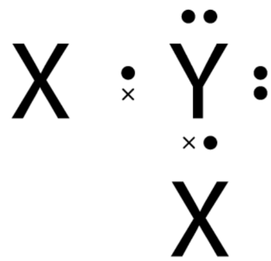
   1.5.1 7

   1.5.2 7

   1.5.3 1

2. Lewis diagram of :

   2.1 H_2S



   2.2 HCN

9.3 Exercise 3

1. It is a type of chemical bond where one or more electrons are transferred from one atom to another.

2.1 1.7

2.2 Mg^{2+}

2.3 Cl^{-}

2.4 $MgCl_2$

3.1 True

3.2 True

9.4 Exercise 4

1.1 H_2O

1.2 NaCl

1.3 Cu

2.1 Non-metal

2.2 Metals non-metals

2.3 Metals

3.1 True

3.2 True

3.3 True

4.1 Ionic

4.2 Ionic

4.3 Metallic

9.5 Exercise 5

1.1 HCN; 27

1.2 CO_2; 44

1.3 Na_2CO_3 ;106

2.1 $(NH_4)_2S$;Ammonium sulfide

2.2 Mg_3P_2;Magnesium phosphide

2.3 $Al_2(Cr_2O_7)_3$;Aluminum dichromate