Atomic combinations: Electronegativity and ionic bonding*

Free High School Science Texts Project

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1 Electronegativity

Electronegativity is a measure of how strongly an atom pulls a shared electron pair towards it. The table below shows the electronegativities (obtained from www.thecatalyst.org/electabl.html) of a number of elements:

Element	**Electronegativity**
Hydrogen (H)	2.1
Sodium (Na)	0.9
Magnesium (Mg)	1.2
Calcium (Ca)	1.0
Chlorine (Cl)	3.0
Bromine (Br)	2.8

Table 1: Table of electronegativities for selected elements

Definition 1: Electronegativity
Electronegativity is a chemical property which describes the power of an atom to attract electrons towards itself.

The greater the electronegativity of an element, the stronger its attractive pull on electrons. For example, in a molecule of hydrogen bromide (HBr), the electronegativity of bromine (2.8) is higher than that of hydrogen (2.1), and so the shared electrons will spend more of their time closer to the bromine atom. Bromine will have a slightly negative charge, and hydrogen will have a slightly positive charge. In a molecule like hydrogen (H_2) where the electronegativities of the atoms in the molecule are the same, both atoms have a neutral charge.

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NOTE: The concept of electronegativity was introduced by *Linus Pauling* in 1932, and this became very useful in predicting the nature of bonds between atoms in molecules. In 1939, he published a book called 'The Nature of the Chemical Bond', which became one of the most influential chemistry books ever published. For this work, Pauling was awarded the Nobel Prize in Chemistry in 1954. He also received the Nobel Peace Prize in 1962 for his campaign against above-ground nuclear testing.

1.1 Non-polar and polar covalent bonds

Electronegativity can be used to explain the difference between two types of covalent bonds. **Non-polar covalent bonds** occur between two identical non-metal atoms, e.g. H_2, Cl_2 and O_2. Because the two atoms have the same electronegativity, the electron pair in the covalent bond is shared equally between them. However, if two different non-metal atoms bond then the shared electron pair will be pulled more strongly by the atom with the highest electronegativity. As a result, a **polar covalent bond** is formed where one atom will have a slightly negative charge and the other a slightly positive charge. This is represented using the symbols δ^+ (slightly positive) and δ^- (slightly negative). So, in a molecule such as hydrogen chloride (HCl), hydrogen is H^{δ^+} and chlorine is Cl^{δ^-}.

1.2 Polar molecules

Some molecules with polar covalent bonds are **polar molecules**, e.g. H_2O. But not *all* molecules with polar covalent bonds are polar. An example is CO_2. Although CO_2 has two polar covalent bonds (between $C^{\delta+}$ atom and the two $O^{\delta-}$ atoms), the molecule itself is not polar. The reason is that CO_2 is a linear molecule and is therefore symmetrical. So there is no difference in charge between the two ends of the molecule. The *polarity* of molecules affects properties such as *solubility*, *melting points* and *boiling points*.

Definition 2: Polar and non-polar molecules
A **polar molecule** is one that has one end with a slightly positive charge, and one end with a slightly negative charge. A **non-polar molecule** is one where the charge is equally spread across the molecule.

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<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=polarmolecules-100512064938-phpapp01&stripped_title=polar-molecules&userName=kwarne>

Figure 1

1.2.1 Electronegativity

1. In a molecule of hydrogen chloride (HCl),
   a. What is the electronegativity of hydrogen
   b. What is the electronegativity of chlorine?
   c. Which atom will have a slightly positive charge and which will have a slightly negative charge in the molecule?
   d. Is the bond a non-polar or polar covalent bond?
   e. Is the molecule polar or non-polar?
2. Complete the table below:

Molecule	**Difference in electronegativity between atoms**	**Non-polar/polar covalent bond**	**Polar/non-polar molecule**
H_2O			
HBr			
F_2			
CH_4			

Table 2

2 Ionic Bonding

2.1 The nature of the ionic bond

You will remember that when atoms bond, electrons are either *shared* or they are *transferred* between the atoms that are bonding. In covalent bonding, electrons are shared between the atoms. There is another type of bonding, where electrons are *transferred* from one atom to another. This is called **ionic bonding**.

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 from one atom to the other.

Definition 3: Ionic bond
An ionic bond is a type of chemical bond based on the electrostatic forces between two oppositely-charged ions. 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.

Example 1:

In the case of NaCl, the difference in electronegativity 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 a Na^+ ion. Chlorine gains an electron and forms a Cl^- ion. The attractive force between the positive and negative ion holds the molecule together.

The balanced equation for the reaction is:

$$Na + Cl \rightarrow NaCl \tag{1}$$

This can be represented using Lewis notation:

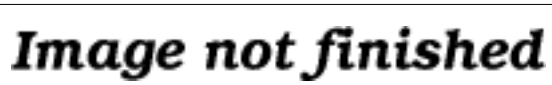

Figure 2: Ionic bonding in sodium chloride

Example 2:

Another example of ionic bonding takes place between magnesium (Mg) and oxygen (O) to form magnesium oxide (MgO). Magnesium has two valence electrons and an electronegativity of 1.2, while oxygen has six valence electrons and an electronegativity of 3.5. 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 molecule together.

The balanced equation for the reaction is:

$$2Mg + O_2 \rightarrow 2MgO \tag{2}$$

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 molecules of magnesium oxide (MgO).

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Figure 3: Ionic bonding in magnesium oxide

TIP: Notice that the number of electrons that is either lost or gained by an atom during ionic bonding, is the same as the **valency** of that element

2.1.1 Ionic compounds

1. Explain the difference between a *covalent* and an *ionic* bond.
2. Magnesium and chlorine react to form magnesium chloride.
   a. What is the difference in electronegativity between these two elements?
   b. Give the chemical formula for:
      - a magnesium ion
      - a choride ion
      - the ionic compound that is produced during this reaction
   c. Write a balanced chemical equation for the reaction that takes place.
3. Draw Lewis diagrams to represent the following ionic compounds:
   a. sodium iodide (NaI)
   b. calcium bromide ($CaBr_2$)
   c. potassium chloride (KCl)

2.2 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 a crystal lattice is shown in Figure 4.

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Figure 4: The crystal lattice arrangement in an ionic compound (e.g. NaCl)

2.3 Properties of Ionic Compounds

Ionic compounds have a number of properties:

- Ions are arranged in a lattice structure
- Ionic solids are crystalline at room temperature
- The ionic bond is a strong electrical attraction. This means that ionic compounds are often hard and have high melting and boiling points
- Ionic compounds are brittle, and bonds are broken along planes when the compound is stressed
- Solid crystals don't conduct electricity, but ionic solutions do