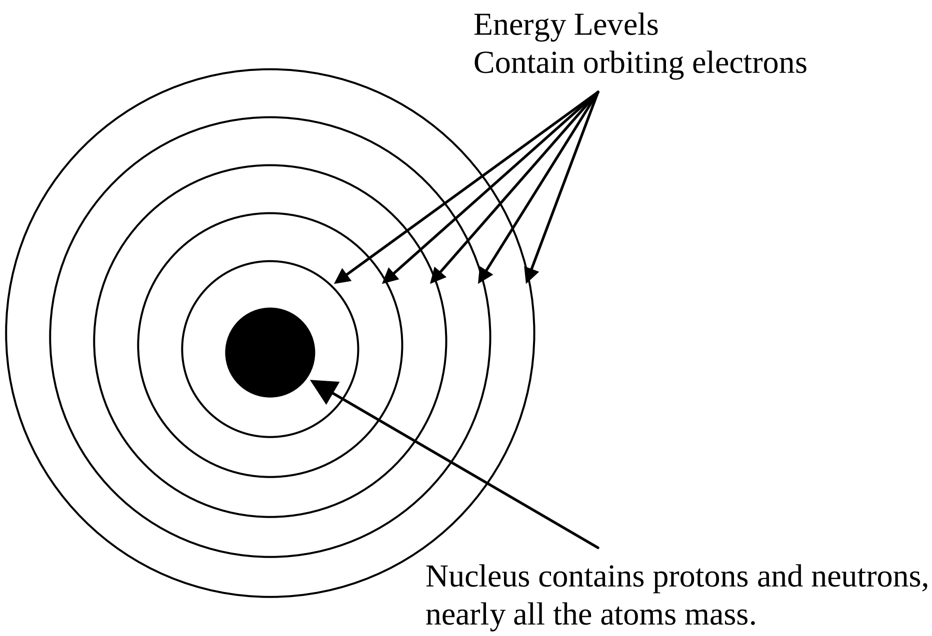
# Elemental Chemistry

## Atoms

Matter is anything that takes up space and has mass. All matter is made of atoms, in fact atoms are the smallest units of matter. However atoms are made of even smaller particles known as sub-atomic particles. The differences between atoms are explained by the differences in numbers of sub-atomic particles. A table showing sub-atomic particles, charge, mass and where they are found is shown below.

| Sub-atomic Particle | Charge | Relative Mass | Where Found |
|---|---|---|---|
| Proton | +1 | 1 | Nucleus |
| Neutron | 0 | 1 | Nucleus |
| Electron | -1 | 1/1840 | Energy Levels |

As a review the structure of an atom is shown below




### *Atomic Number and Mass Number*

Each different type of atom contains a different number of protons in the nucleus, this number is known as the atomic number. All atoms are electrically neutral, no overall charge, therefore the number of protons has to equal the number of electrons. Due to the greater weight of neutrons and protons and the fact they are found in the nucleus means the mass of an atom is concentrated in the nucleus. The number of protons and neutrons is called the mass number. The number of neutrons can be determined by subtracting the atomic number from the mass number.

Mass Number $\rightarrow$

$$^{23}_{11}\mathrm{Na}$$

Atomic Number $\rightarrow$

The nuclei of the same atoms must have the same number of protons, but can have differing numbers of neutrons. Atoms with the same number of protons but differing neutrons are called isotopes. Isotopes have the same atomic number but a different mass number.

| | $^{35}_{17}Cl$ | $^{37}_{17}Cl$ |
|---|---|---|
| Protons | 17 | 17 |
| Electrons | 17 | 17 |
| Neutrons | 18 | 20 |

Almost all elements have isotopes but their abundance varies greatly Cl-35 has a natural abundance of 75.77% whilst Cl-37 has abundance of 24.23%. For this reason mass numbers are often not whole numbers

## *Electronic Configuration*

The chemical properties of an atom depend on the number and arrangement of electrons in it's energy levels. The arrangement of electrons in an atom is known as the electronic configuration. We know that the electrons move around in energy levels or shells, there is a maximum of 7 shells labelled K – Q. The maximum number of electrons that can fit into each shell is determined by $2n^2$ where n equals the shell number.

| **Shell** | **Designation** | **Electron Capacity** |
|---|---|---|
| 1st | K | 2 |
| 2nd | L | 8 |
| 3rd | M | 18 |
| 4th | N | 32 |

Since the K shell has the lowest energy it is filled first and then fills out from there. The electrons in the outer shell are known as the valence electrons. These electrons are the ones involved in chemical reactions. Therefore similarities in the number of electrons in the outer shell between atoms will lead to similar chemical properties.

As a general rule the outermost shell will not hold more than 8 electrons, Therefore, apart from the first shell, the shells will fill to 8 electrons before moving on to the next shell. Take potassium for example, it has an atomic number of 19, therefore 19 electrons. 2 electrons can fit in the first shell, 8 in the second totalling 10. There are nine left to fill and even though all of them could fit in the 3rd shell, only 8 of them will fill in that shell leaving 1 to fill into the 4th shell. The electronic configuration of potassium is 2,8,8,1.

All atoms want a stable 8 electrons in the outer shell and in reactions will gain or lose electrons to get 8 in the outer shell.

# The Periodic Table

The elements in the periodic table are arranged in rows called periods and columns called groups. The group number corresponds to the number of electrons in the outer shell. And the period number corresponds to the number of electron shells or energy levels used.

Some groups have special names
Group I: alkali metals
Group II: alkaline earth metals
Group VII: halogens
Group VIII: noble (inert) gases

The block between groups II and III is known as the transition metals.

Metals are found on the bottom left of the table and non-metals on the top right of the table, the elements between these two in a staircase shape are known as the metalloids and have both metallic and non-metallic properties.

## *Trends in the Periodic Table*

Down a group

- Atomic number increases
- Atomic radius increases
- Metallic nature increases
- Electronegativity decreases

Across a period

- Atomic number increases
- Atomic radius decreases
- Metallic nature decreases
- Electronegativity increases

Electronegativity is the atoms ability to hold the valence electrons. It influences the rate at which a chemical reaction occurs. Elements with a low electronegativity lose electrons readily causing a fast reaction. Similarly those with a high electronegativity readily take in electrons and have a fast reaction rate. Those with an intermediate electronegativity will have a slower rate of reaction. It also affects the polarity of a substance, which will be discussed later

## *Electronic Configuration and the Periodic Table*

Due to the high level of organisation within the periodic table and the fact that they are arranged by similar properties, the electronic configuration is easily obtained, similarly we can easily identify the position on the periodic table given the electronic configuration. Take the examples below

| | | |
|---|---|---|
| Mg | 2,8,2 | Group 2 (2 electrons in outer shell) Period 3 (3 electron shells) |
| Ca | 2,8,8,2 | Group 2 (2 electrons in outer shell) Period 4 (4 electron shells) |

*Properties of Metals and Non-metals*

| Physical Properties | |
|---|---|
| **Metals** | **Non-metals** |
| Metals are high melting point solids (except mercury). | Non-metals are generally gases or liquids with a low melting point. |
| Metals are malleable and ductile | Non-metals are not malleable or ductile |
| Metals conduct heat and electricity well | Non-metals are poor conductors of heat and electricity (except graphite) |
| Metals are lustrous and can be polished | Non-metals have no lustre and cannot be polished. |
| Metals give off a note when struck | Non-metals do not give off a note when struck |
| Metals have high density | Non-metals have relatively low density with the exception of the solids. |
| Metals are not brittle | Non-metals are brittle and break easily |
| **Chemical Properties** | |
| **Metals** | **Non-metals** |
| Metals react by losing electrons to become positive ions | Non-metals react by gaining electrons to become negative ions |
| Metals are reducing agents | Non-metals can be oxidising or reducing agents |
| Metals can displace hydrogen from dilute acids | Non-metals cannot displace hydrogen from solution |
| Metals burn in oxygent to form basic oxides | Non-metals burn in oxygen to form either acidic or neutral oxides. |
| Metals react with halogens to form ionic halides | Non-metals react with halogens to form covalent halides. |

# Chemical Bonding

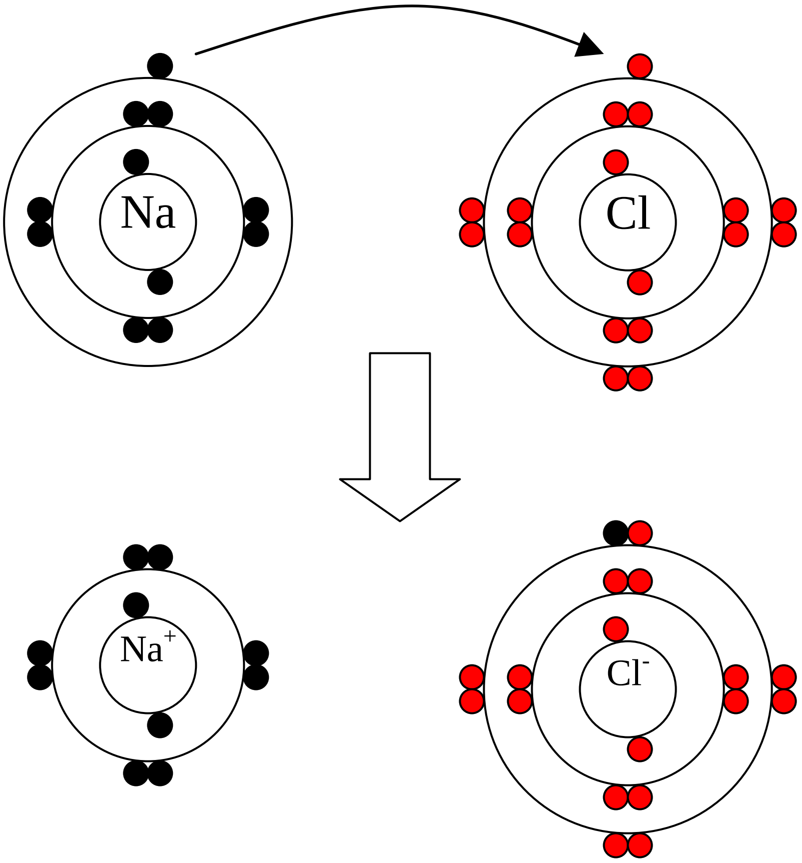
Chemical reactions are the result of moving electrons, the making and breaking of bonds. Atoms combine, or bond, to reach a more stable state. Of all the elements on the periodic table only the noble gases (group 8) rarely take place in chemical reactions (there are only a handful of noble gas compounds). This is due to the full 8 in the outer shell that all atoms try to achieve. All other atoms can achieve this by sharing electrons, or losing/gaining them.

## *Ionic Bonding*

Metals and non-metals generally react readily as the electrons lost by metals will be quickly taken up by the more electronegative non-metals. Lets take the example of sodium and chlorine.

Sodium has an electronic configuration of 2,8,1 and therefore has 1 in the outer shell. To achieve a stable 8 electrons in its outer shell it has two choices, it can lose 1 electron or gain 7. It requires less energy to lose 1 electron so that is what happens. Sodium has now become a sodium ion and since it has lost an electron there are now more protons than electrons, meaning it has become a positive ion.

On the other hand chlorine has an electronic configuration of 2,8,7 and therefore has 7 in the other shell. To achieve a stable 8 it requires less energy to gain 1 electron than lose 7 so it gains 1 in this case. Chlorine has now become a chloride ion and since it has gained an electron there is now more electrons than protons and it therefore has a negative charge.



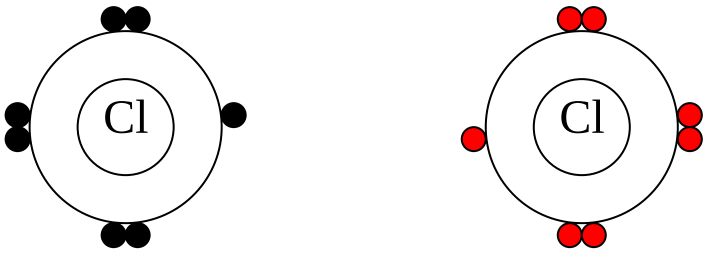

Once this transfer of electrons has taken place the electrostatic forces between the positive and negative ions hold the compound together. This electrostatic holding of atoms together is known as an ionic bond.

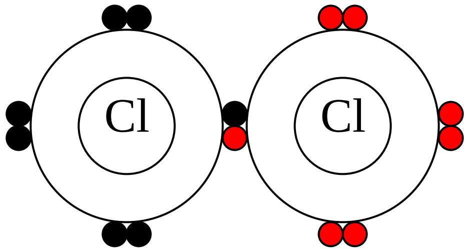
Ionic compounds will show the following properties.

- They usually have high melting and boiling points.  This is due to the large energy required to break the bonds between atoms.
- Because the ions are held in a lattice they are not mobile and therefore an ionic substance will not conduct electricity in a solid form.  However in a molten or aqueous (dissolved in water) form ions are free to move and will conduct electricity.
- Ionic solids are often soluble in water, this is a process called dissociation. They do not dissolve in non-polar substances such as ethanol and benzene.
- They are hard and not easily scratched.
- They are generally brittle and will break easily.
- Once ions form they show none of the properties of the parent atom.

## *Covalent Bonding*

Covalent compounds form between two or more non-metallic substances.  This means that all the atoms present in a covalent substance need to gain electrons.  They cannot transfer electrons, so they must be shared in order to make the stable 8. Covalent bonds exist as single, double or triple bonds.  In single bonds one electron is shared in double bonds two electrons are shared and in triple bonds three electrons are shared. Take the example using $Cl_2$ shown below (note only the outer shell electrons are used).




In the example shown above we can see that in the first line that each chlorine atom has 7 in its outer shell.  However when they come together in the bottom line the sharing allows for each to have 8

Properties of covalent substances include

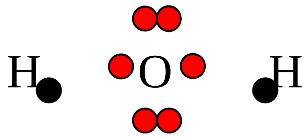
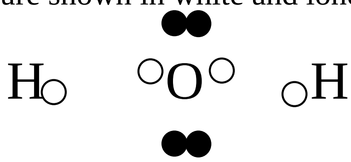
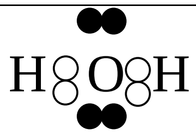
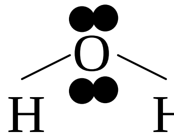
- They have low melting and boiling points.  This is due to the fact that they are discrete units, they don't exist in lattices and therefore not much energy is needed to break the weak secondary bonds.
- They do not conduct electricity in solid, liquid, gaseous or aqueous forms.
- They are not very soluble in water, but are more soluble in non-polar solvents
- Many are flammable.

# *Determining Bonding Diagrams for Covalent Substances*

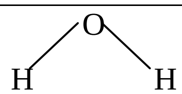
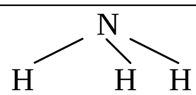
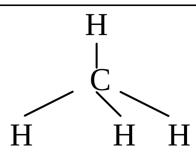
There are some simple rules to follow in drawing how covalent substances bond and the shape of the molecules. These rules are set out below.

1. Determine the correct formula of the molecule
2. Determine which is the central atom (the one with the most bonding pairs) and place the other atoms around it.
3. Determine the bonding pairs and lone pairs of electrons around the central atom.
4. Attach the atoms around the central atom to the bonding electron pairs so that each atom has 8 in its outer shell.
5. Count the number of electron pairs and arrange them so that they are as far apart as possible (remember it is a 3D shape).

Lets take the example of water

| | |
|---|---|
| Step 1 – Correct formula | The correct formula of water is $H_2O$ |
| Step 2 – Central Atom | O has the most bonding pairs and therefore must be in the centre   |
| Step 3 – Determine bonding and lone pairs | Bonding pairs are shown in white and lone pairs in black   |
| Step 4 – Attach all atoms so each has 8. |   |
| Step 5 Arrange atoms and lone pairs as far apart as possible |   |

The shapes of some common molecules are shown below

| **Compound** | **Number of bonding pairs around central atom** | **Number of lone pairs around central atom** | **Shape** | **Representative Shape** |
|---|---|---|---|---|
| HCl | 1 | 3 | Linear | H — Cl |
| $H_2O$ | 2 | 2 | V shaped |   |
| $NH_3$ | 3 | 1 | Trigonal pyramidal |   |
| $CH_4$ | 4 | 0 | Tetrahedral |   |

## *Polarity and Bonding*

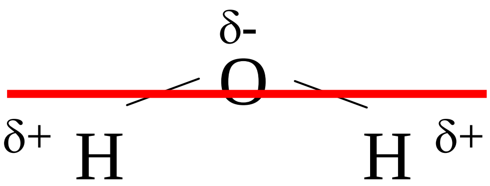
As stated earlier polarity can be affected by the electronegativity of the various elements. Electronegativity being an atoms ability to attract and hold electrons. Electronegativity can also affect the location of the electrons inside a molecule lets take two examples

$Cl_2$ which can be written as Cl — Cl is comprised of two chlorine atoms, obviously each of the atoms have the same electronegativity and therefore are pulling with equal strength on the electrons hence they stay in the middle, in this case we call it a <u>non-polar</u> covalent bond.

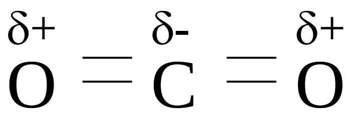
HCl which can be written as H — Cl is comprised of two atoms, chlorine and hydrogen. Due to the fact that they are different elements they have different electronegativity and therefore are not both pulling with equal strength on the electrons. In this case the chlorine, having the higher electronegativity, will pull harder on the electrons and therefore attract a slight negative charge to that end whilst the other end around the hydrogen will have a slight positive charge . In this case it is known as a <u>polar</u> covalent bond it is also known as a <u>bond dipole</u>. It can be represented as shown below.

$$\overset{\delta+}{H}—\overset{\delta-}{Cl}$$

The story is slightly different for entire molecules. A molecule comprised only of non-polar bonds will itself be non-polar. However if polar bonds are present it can either polar or non-polar depending on the distribution of charge. For a molecule with polar bonds to be non-polar it must be symmetrical and should not have a positive and a negative end.




**<u>Polar Molecule</u>**
Polar bonds and no symmetry along the red line




**<u>Non-Polar Molecule</u>**
Polar bonds and symmetry along all axis

List of common electronegativities is shown below

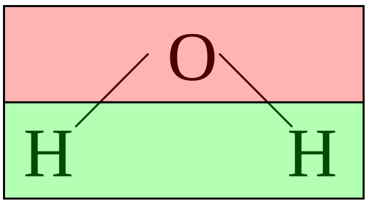
| | | | | | | |
|---|---|---|---|---|---|---|
| H<br>2.1 | | | | | | |
| Li<br>1.0 | Be<br>1.5 | B<br>2.0 | C<br>2.5 | N<br>3.0 | O<br>3.5 | F<br>4.0 |
| Na<br>0.9 | Mg<br>1.2 | Al<br>1.5 | Si<br>1.8 | P<br>2.1 | S<br>2.5 | Cl<br>3.0 |
| K<br>0.8 | Ca<br>1.0 | Ga<br>1.6 | Ge<br>1.8 | As<br>1.8 | Se<br>2.6 | Br<br>2.8 |
| Rb<br>0.8 | Sr<br>1.0 | In<br>1.7 | Sn<br>1.8 | Sb<br>1.8 | Te<br>2.1 | I<br>2.5 |
| Cs<br>0.7 | Ba<br>0.9 | Tl<br>1.8 | Pb<br>1.9 | Bi<br>1.9 | Po<br>2.0 | At<br>2.2 |
| Fr<br>0.7 | Ra<br>0.9 | | | | | |

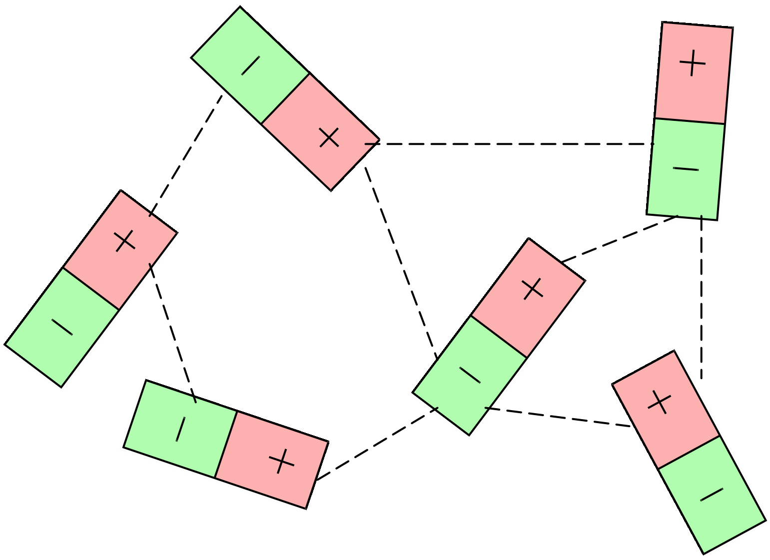
# Secondary Interactions

As stated earlier ionic substances are held in continuous lattices and are therefore normally solids at room temperature whereas covalent substances are often discrete units which we call molecules.  These molecular substances normally have a low melting/boiling point and are therefore generally gases at room temperature.  However when looking at substances like water and sucrose (found in sugar) we find that water is a liquid and sucrose is a solid, however they are both molecular substances, and therefore should have low boiling points.

From studies of matter in previous years we know that in gases the particles are furthest apart, in liquids the particles are quite close and in solids they are not only very close, but in many cases joined together.  With molecular substances this causes us to think about how separate molecules may be held together closely enough to form a liquid or a solid.

Secondary interactions, also called intermolecular bonds, act between molecules using forces of attraction to hold the molecules together.  The reason we call these interactions instead of bonds is unlike primary bonds, electrons are not transferred or shared.  Secondary interactions make use of a molecule's polarity in order to hold the molecules together.  Let us take the example of water below, we already know that water is polar and therefore has a slight positive charge on one end and a slight negative charge on the other end, we call this a **<u>dipole</u>** (two poles). Let us think of these polar water molecules as magnets

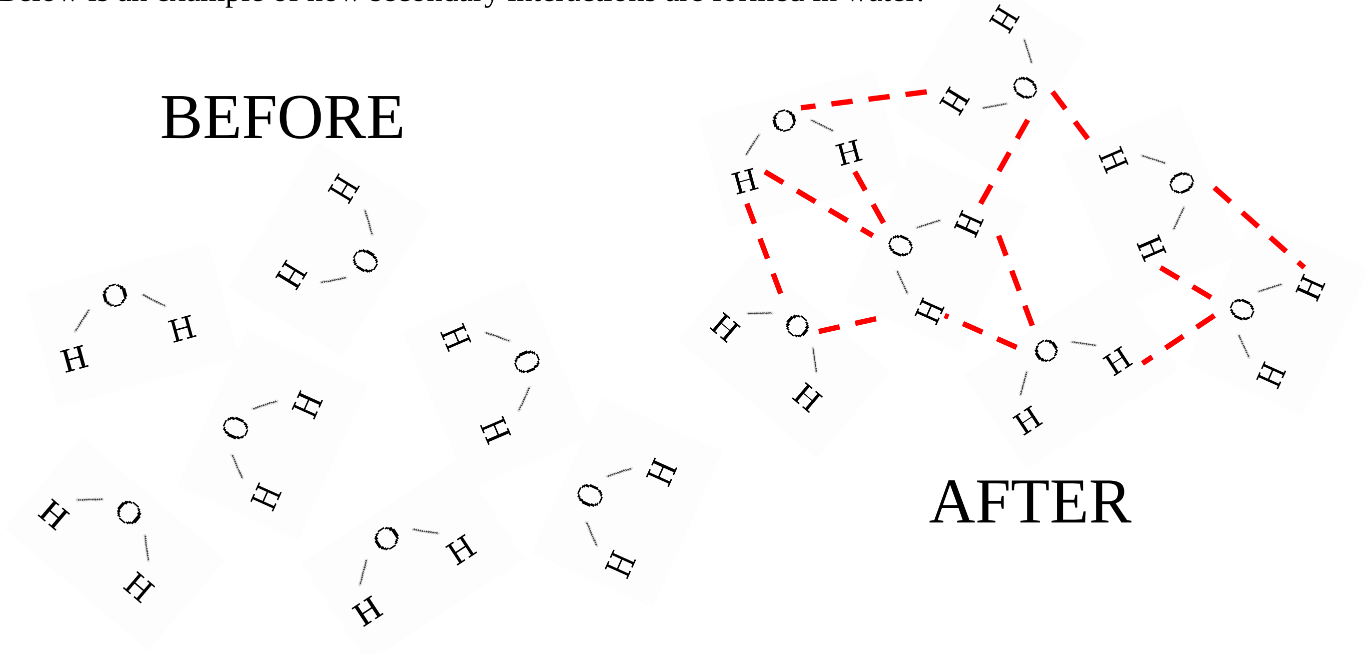



Thinking about the above example we can see both the primary covalent bonds and the secondary interactions.  In terms of the primary bonds, these are the bonds that hold the magnet itself together, the magnet would simply fall apart without these.  The primary bonds that hold the magnet together are very hard to break and that is why it is hard to break a magnet.  The dotted lines represent the forces of attraction between the ends of different magnets and therefore represent the secondary interactions.  If these magnets were allowed to come together they would stick, but they could easily be pulled apart.

This gives an indication of the difference in strength of primary bonds versus secondary interactions.  Primary bonds are strong and very hard to break whereas secondary interactions are easily broken with generally a little bit of heat.

Below is an example of how secondary interactions are formed in water.

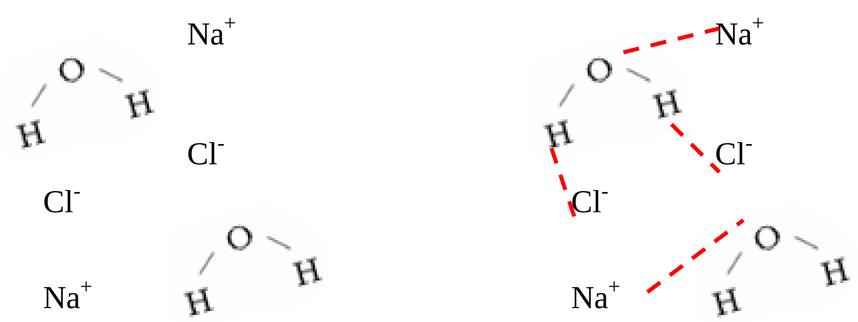



There are a few different types of secondary interactions that are outlined below

Dipole-Dipole Interactions - These occur when two dipolar molecules are attracted to each other the slight charge on each end causes electrostatic attraction between the opposing ends causing them to stick together. These interactions can only occur in polar molecules, they will not occur in non-polar ones as no dipole can be formed.

Hydrogen Bonds - Hydrogen bonds are the strongest of all secondary interactions and they are basically a special case of the dipole-dipole interactions. The function in the same way as dipole-dipole interactions but are stronger due to higher electronegativity differences. Hydrogen bonds will only form between oxygen, nitrogen or fluorine in one of the molecules and hydrogen attached to oxygen nitrogen or fluorine in the other molecule. The water above is an example of this.

Ion-Dipole Interactions - These occur when a polar molecule such as water is attracted to an ion. As they both have a charge this causes electrostatic attraction between the opposing charges. An example of this could be salt dissolving in water as shown below.




Dispersion Forces - Dispersion forces are the only secondary interactions that will occur between non-polar molecules. They occur due to an induced polarity in non-polar substances because the electrons are continuously moving. The larger the molecule, the larger the number of electrons and therefore the larger the dispersion forces holding them together.

# Chemical Formulae and Nomenclature

## *Classification of Substances*

In a very broad sense all substances can be described as either pure substances or mixtures.

In terms of your studies in chemistry mixtures are of very little uses as their composition varies. They are defined as two or more substances combined in any proportion. The characteristics of mixtures include:

- They are a physical blend of various pure substances (elements or compounds).
- Do not have a fixed and definite composition (cannot be defined by a formula).
- They have variable properties.
- Can be separated by physical changes.

If a substance cannot be separated by using a physical change then it is considered to be a pure substance. This can be further broken down into elements and compounds. Elements consist only of the one type of atom, they all have the same number of protons in the nucleus.

Compounds are substances in which two or more types of atoms have been chemically combined in a definite proportion. They can be split down further into molecular or covalent substances which are formed from only non-metallic atoms and ionic substances which are formed from metallic and non-metallic atoms.

## *Chemical Formulae*

Chemical formulae when written looks very similar for both ionic and covalent substances, and it is, however there is subtle differences in what it represents.

A molecular formula is used for covalent substances, it shows the number and type of atoms in one molecule of the substance.  For example for water the molecular formula is $H_2O$, it represents that each molecule of water has two hydrogen atoms and one oxygen atom.

A formula unit is used for ionic substances, because ionic substance exist as continuous lattices the formula shows the ration that the atoms combine in.  For example magnesium chloride, formula $MgCl_2$, has two chlorine atoms for every magnesium atom in the lattice.

Ions are positively or negatively charged particles that form when an atom gains or loses electrons.  Metal will lose electrons and form positive ions whilst non-metals will gain electrons and form negative ions. By knowing the number of electrons atoms will gain, lose or share we can figure out the formula of a chemical substance. Polyatomic ions are tightly bound groups of atoms that carry a charge and act like a single unit, for example $CO_3^{2-}$.

*<u>Determining Chemical Formulae</u>*

**Ionic Compounds** – Since ionic compounds involve the combination of metals and non-metals they have opposing charges metals positive and non-metals negative. To find the numbers of each type of atom (or ion) the total charge of the positive ions must equal the total charge of the negative ions.

Examples

**Sodium Oxide**: Comprised of $Na^+$ and $O^{2-}$
To balance the charges we need two Na for every O therefore the formula is $\mathbf{Na_2O}$

**Aluminium Sulfate**: Comprised of $Al^{3+}$ and $SO_4^{2-}$
To balance the charges we need two Al for every 3 $SO_4$ therefore the formula is $\mathbf{Al_2(SO_4)_3}$

**Covalent Compounds** – Since covalent compounds involve the combination of non-metals there is not going to be the opposing charges like ionic compounds. In this case the atoms will want to achieve a stable 8 in the outer shell by sharing electrons with other atoms. Looking at the charge on the ion identifies the number of electrons it will need to share. A 2- ion will need to share 2 electrons. It is these numbers that need to balance when determining the formula

Examples

**Methane**: Comprised of C and H
Since C shares 4 electrons and H shares 1 we need 4 H for every C therefore the formula is $\mathbf{CH_4}$

**Water**: Comprised of O and H
Since O shares 2 electrons and H shares 1 we need 2 H for every O therefore the formula is $\mathbf{H_2O}$

# Positive Ions

| Monovalent | +1 Charge | Divalent | +2 Charge | Trivalent | +3 Charge |
|---|---|---|---|---|---|
| Ammonium | $NH_4^+$ | Barium | $Ba^{2+}$ | Aluminium | $Al^{3+}$ |
| Caesium | $Cs^+$ | Cadmium | $Cd^{2+}$ | Chromium (III) | $Cr^{3+}$ |
| Copper (I) | $Cu^+$ | Calcium | $Ca^{2+}$ | Cobalt (III) | $Co^{3+}$ |
| Hydrogen | $H^+$ | Chromium (II) | $Cr^{2+}$ | Iron (III) | $Fe^{3+}$ |
| Lithium | $Li^+$ | Cobalt | $Co^{2+}$ | | |
| Mercury (I) | $Hg^+$ | Copper (II) | $Cu^{2+}$ | **Tetravalent** | **+4 Charge** |
| Potassium | $K^+$ | Iron (II) | $Fe^{2+}$ | | $Pb^{4+}$ |
| Rubidium | $Rb^+$ | Lead (II) | $Pb^{2+}$ | Lead (IV) | $Sn^{4+}$ |
| Sodium | $Na^+$ | Magnesium | $Mg^{2+}$ | Tin (IV) | $Ti^{4+}$ |
| Silver | $Ag^+$ | Manganese (II) | $Mn^{2+}$ | Titanium (IV) | |
| | | Mercury (II) | $Hg^{2+}$ | | |
| | | Nickel | $Ni^{2+}$ | | |
| | | Strontium | $Sr^{2+}$ | | |
| | | Tin (II) | $Sn^{2+}$ | | |
| | | Zinc | $Zn^{2+}$ | | |

# Negative Ions

| Monovalent | -1 Charge | Divalent | -2 Charge | Trivalent | -3 Charge |
|---|---|---|---|---|---|
| Bromide | $Br^-$ | Carbonate | $CO_3^{2-}$ | Borate | $BO_3^{3-}$ |
| Chloride | $Cl^-$ | Chromate | $CrO_4^{2-}$ | Nitride | $N^{3-}$ |
| Cyanide | $CN^-$ | Dichromate | $Cr_2O_7^{2-}$ | Phosphite | $PO_3^{3-}$ |
| Dihydrogen Phosphate | $H_2PO_4^-$ | Monohydrogen Phosphate | $HPO_4^{2-}$ | Phosphate | $PO_4^{3-}$ |
| Fluoride | $F^-$ | Oxalate | $C_2O_4^{2-}$ | | |
| Hydride | $H^-$ | Oxide | $O^{2-}$ | **Tetravalent** | **-4 Charge** |
| Hydrogen Carbonate | $HCO_3^-$ | Sulfide | $S^{2-}$ | | |
| Hydrogen Sulfate | $H_2SO_4^-$ | Sulfite | $SO_3^{2-}$ | Orthosilicate | $SiO_4^{4-}$ |
| Hydrogen Sulfite | $H_2SO_3^-$ | Sulfate | $SO_4^{2-}$ | | |
| Hydroxide | $OH^-$ | Thiosulfate | $S_2O_3^{2-}$ | | |
| Hypochlorite | $ClO^-$ | | | | |
| Iodide | $I^-$ | | | | |
| Nitrate | $NO_3^-$ | | | | |
| Nitrite | $NO_2^-$ | | | | |
| Perchlorate | $ClO_4^-$ | | | | |
| Permanganate | $MnO_4^-$ | | | | |
| Thiocyanate | $SCN^-$ | | | | |

# *Balancing Chemical Equations*

A chemical equation is a symbolic expression of a chemical reaction. They are used as a way of showing reactants, products and the ratios of each. A chemical reaction also must obey the law of conservation of mass which states that matter is neither created nor destroyed in a chemical reaction. This means that the number of each type of atom must be the same on each side of the reaction. Below are some rules regarding how to correctly balance equations.

1. Write a word equation which includes all reactants and products
2. Under each of the reactants and products in the word equation write the correct formula. If any formula are incorrect your equation will always be incorrect.
3. Alter the coefficients in front of each formula to balance the number of atoms of each element on both sides of the equation.
   - You must not alter either the subscripts or superscripts
   - You are only allowed to add numbers to the front of formulae eg **2** NaCl
4. Check that the number of atoms of each element is the same on both sides of the equation.
5. Write in the physical states for each species (aq) = aqueous, (g) = gas, (l) = liquid, (s) = solid