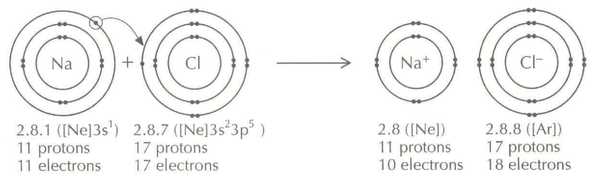
Ionic bonding

IONIC BOND

When atoms combine they do so by trying to achieve an inert gas configuration. Ionic compounds are formed when electrons are transferred from one atom to another to form ions with complete outer shells of electrons. In an ionic compound the positive and negative ions are attracted to each other by strong electrostatic forces, and build up into a strong lattice. Ionic compounds have high melting points as considerable energy is required to overcome these forces of attraction.

The classic example of an ionic compound is sodium chloride Na^+Cl^-, formed when sodium metal burns in chlorine. Chlorine is a covalent molecule, so each atom already has an inert gas configuration. However, the energy given out when the ionic lattice is formed is sufficient to break the bond in the chlorine molecule to give atoms of chlorine. Each sodium atom then transfers one electron to a chlorine atom to form the ions.



The charge carried by an ion depends on the number of electrons the atom needed to lose or gain to achieve a full outer shell.

Cations			Anions		
Group 1	**Group 2**	**Group 3**	**Group 5**	**Group 6**	**Group 7**
+1	+2	+3	−3	−2	−1
Li^+ Na^+ K^+	Mg^{2+} Ca^{2+}	Al^{3+}	N^{3-} P^{3-}	O^{2-} S^{2-}	F^- Cl^- Br^-

Thus in magnesium chloride two chlorine atoms each gain one electron from a magnesium atom to form $Mg^{2+}Cl^-{}_2$. In magnesium oxide two electrons are transferred from magnesium to oxygen to give $Mg^{2+}O^{2-}$. Transition metals can form more than one ion. For example, iron can form Fe^{2+} and Fe^{3+} and copper can form Cu^+ and Cu^{2+}.

FORMULAS OF IONIC COMPOUNDS

It is easy to obtain the correct formula as the overall charge of the compound must be zero.

lithium fluoride Li^+F^-
sodium oxide $Na^+{}_2O^{2-}$
potassium nitride $K^+{}_3N^{3-}$
magnesium chloride $Mg^{2+}Cl^-{}_2$
calcium sulfide $Ca^{2+}S^{2-}$
calcium phosphide $Ca^{2+}{}_3P^{3-}{}_2$
aluminium bromide $Al^{3+}Br^-{}_3$
iron(III) oxide $Fe^{3+}{}_2O^{2-}{}_3$
iron(II) oxide $Fe^{2+}O^{2-}$

Note: the formulas above have been written to show the charges carried by the ions. Unless asked specifically to do this it is common practice to omit the charges and simply write LiF, $MgCl_2$, etc.

IONS CONTAINING MORE THAN ONE ELEMENT

In ions formed from more than one element the charge is often spread (delocalized) over the whole ion. An example of a positive ion is the ammonium ion NH_4^+, in which all four N–H bonds are identical. Negative ions are sometimes known as acid radicals as they are formed when an acid loses one or more H^+ ions.

hydroxide OH^-
nitrate NO_3^- (from nitric acid, HNO_3)
sulfate SO_4^{2-}, hydrogen sulfate HSO_4^- } from sulfuric acid, H_2SO_4
carbonate CO_3^{2-}, hydrogen carbonate HCO_3^- } from carbonic acid, H_2CO_3
ethanoate CH_3COO^- (from ethanoic acid, CH_3COOH)

The formulas of the ionic compounds are obtained in exactly the same way. Note: brackets are used to show that the subscript covers all the elements in the ion.

sodium nitrate $Na^+NO_3^-$
ammonium sulphate $(NH_4^+)_2SO_4^{2-}$
calcium carbonate $Ca^{2+}CO_3^{2-}$
magnesium ethanoate $Mg^{2+}(CH_3COO^-)_2$
aluminium hydroxide $Al^{3+}(OH^-)_3$

IONIC OR COVALENT?

Ionic compounds are formed between metals on the left of the Periodic Table and non-metals on the right of the Periodic Table; that is, between elements in groups 1, 2, and 3 with a low electronegativity (electropositive elements) and elements with a high electronegativity in groups 5, 6, and 7. Generally the difference between the electronegativity values needs to be greater than about 1.8 for ionic bonding to occur.

	Al F	Al O	Al Cl	Al Br
Electronegativity	1.5 4.0	1.5 3.5	1.5 3.0	1.5 2.8
Difference in electronegativity	2.5	2.0	1.5	1.3
Formula	AlF_3	Al_2O_3	Al_2Cl_6	Al_2Br_6
Type of bonding	ionic	ionic	intermediate between ionic and covalent	covalent
M. pt / °C	1265	2050	Sublimes at 180	97