2.21 Ionic Bonding

100% ionic compounds do not exist but predominantly ionic compounds are formed when metals combine with non-metals.

Forming ions

Metal atoms lose electrons to form +ve ions. Non-metal atoms gain electrons to form -ve ions. Hydrogen is unusual in that it can form both an H+ and H- (hydride) ion.

The electronic configurations of the resulting ions usually correspond to a noble gas electronic configuration.

The diagram below shows that the main charge of some of the most common ions. The charges of the ions in group 1, 2, 3, and 5, 6, 7 can be related to their position in the periodic table, and this is how you should remember them.

The transition metals are more complicated in that these can form different ionic charges in different compounds. To specify which charge to use we put the charge in the name in brackets, e.g. Iron (III) chloride contains iron with a +3 charge.

All the non-metal ions end in –ide and should not be confused with -ates

e.g. Li3N is lithium nitride and contains the nitride ion N3- . It should not be confused with the nitrate ion NO3-.

Similarly don’t confuse phosphide ion P3- with the phosphate ion PO43-.
**Compound ions**

There are several compound ions you should learn the formulae and charge for. These ions have covalent bonds between the atoms in the compound ion. These ions will, however, bond ionically to metal ions.

**Compound ions to learn**
- Ammonium \((\text{NH}_4^+)\)
- Carbonate \((\text{CO}_3^{2-})\)
- Sulphate \((\text{SO}_4^{2-})\)
- Nitrate \((\text{NO}_3^-)\)
- Hydroxide \((\text{OH}^-)\)
- Phosphate \((\text{PO}_4^{3-})\)

**Summary of most important ions to know**

<table>
<thead>
<tr>
<th>+1</th>
<th>+2</th>
<th>+3</th>
<th>-3</th>
<th>-2</th>
<th>-1</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group 1</td>
<td>Group 2</td>
<td>Group 3</td>
<td>Group 5</td>
<td>Group 6</td>
<td>Group 7</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>Zinc</td>
<td>Iron (III)</td>
<td>Phosphate</td>
<td>Carbonate</td>
<td>Nitrate</td>
</tr>
<tr>
<td>Silver</td>
<td>Copper (II)</td>
<td></td>
<td>((\text{PO}_4^{3-}))</td>
<td>Sulphate</td>
<td>((\text{NO}_3^-))</td>
</tr>
<tr>
<td>Gold</td>
<td></td>
<td>Iron (II)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ammonium</td>
<td></td>
<td>Tin</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>((\text{NH}_4^+))</td>
<td></td>
<td>Lead</td>
<td></td>
<td></td>
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</tr>
</tbody>
</table>

**How to work out the formula from the ionic charge**

What is the formula of Lithium Sulphate?

1. Identify the ionic charges of the two ions
   - Lithium is in group 1 so has a +1 charge, \(\text{Li}^+\)
   - Sulphates have a -2 charge, \(\text{SO}_4^{2-}\)

2. Combine the ions together to get a neutral compound, i.e. combine so that the total +ve charge cancels out the total −ve charge

\[
\text{Li}^+ + \text{Li}^+ + \text{SO}_4^{2-} \quad \text{We need two lithium ions to cancel out the -2 charge on the sulphate}
\]

The formula is therefore \(\text{Li}_2\text{SO}_4\)

What is the formula of Calcium phosphate?

1. Identify the ionic charges of the two ions
   - Calcium is in group 2 so has a +2 charge, \(\text{Ca}^{2+}\)
   - Phosphates have a -3 charge, \(\text{PO}_4^{3-}\)

2. Combine the ions together to get a neutral compound, i.e. combine so that the total +ve charge cancels out the total −ve charge

We need to multiply up to get the same charge. Three calcium ions would produce +6 charge to cancel out the -6 charge on two phosphate ions

\[
\text{Ca}^{2+} + \text{Ca}^{2+} + \text{Ca}^{2+} + \text{PO}_4^{3-} + \text{PO}_4^{3-} \quad \text{We need 3 calcium ions to cancel out the -6 charge on 2 phosphate ions}
\]

The formula is therefore \(\text{Ca}_3(\text{PO}_4)_2\)

Only use brackets when there is more than one of the compound ion in the formula.

- e.g. Copper nitrate is \(\text{Cu(NO}_3)_2\)
- Calcium hydroxide is \(\text{Ca(OH)}_2\)
- Ammonium sulphate is \((\text{NH}_4)_2\text{SO}_4\)
Evidence for the existence of ions

How do we know ions exist? What is the evidence for it?

Migration of ions

There are a few simple experiments that can show the migration of coloured ions.

A drop of potassium Manganate solution or a small crystal of potassium mangante, which is purple, is placed on moist filter paper on a microscope slide and the ends of the slide are connected to a 24 V DC power supply. After ten minutes the purple colour of the \( \text{MnO}_4^- \) ion has migrated to the positive electrode.

You can see a video of this demo here: [https://www.youtube.com/watch?v=nQ3V6N6EzYM](https://www.youtube.com/watch?v=nQ3V6N6EzYM)

Physical properties of Ionic Compounds

The model of spherical ions strongly attracted to each other in a lattice is evidenced by the observed physical properties of ionic compounds.

- high melting points (there are strong attractive forces between the ions)
- non conductor of electricity when solid (ions are held together tightly and cannot move)
- conductor of electricity when in solution or molten. (Ions are free to move)
- brittle / easy to cleave apart

A little force will push the ions along and ions will be next to similar ions. There will be a force of repulsion between like ions, pushing the layers apart.
Structure of Ionic Substances

An Ionic bond is the electrostatic force of attraction between oppositely charged ions in a lattice.

Electron density maps obtained through X-ray diffraction also give evidence for the existence of ions and also the structure of ionic substances.

These maps show the likelihood of finding electrons in a region.

The contours are lines of equal electron density, with greater electron densities being on contours closer to the nucleus.

The maps show that for NaCl:

- The ions are arranged in a regular pattern.
- The chloride ions are larger than the sodium ions.

The ions in an ionic solid are arranged in a regular 3D pattern called a giant ionic lattice.

The sticks in this diagram are there to help show the arrangements of the ions. They do not represent the ionic bonds.

Ionic bonding is between ions and all their surrounding oppositely charged ions. Each sodium ion in this structure is surrounded and equally attracted by six chloride ions. The ionic bond is the attraction between all these ions.

Not all ionic substances have the same arrangement of ions in their lattices.
Ionic Radii

Positive ions are smaller compared to their atoms because it has one less shell of electrons and the ratio of protons to electrons has increased so there is greater net force on remaining electrons holding them more closely.

The negative ions formed from groups five to seven are larger than the corresponding atoms. The negative ion has more electrons than the corresponding atom but the same number of protons. So the pull of the nucleus is shared over more electrons and the attraction per electron is less, making the ion bigger.

Within a group the size of the ionic radii increases going down the group. This is because as one goes down the group the ions have more shells of electrons.

Strength of Ionic bonding in lattices

It is actually not easy to define what ‘one’ ionic bond is. We generally measure the strength of ionic bonding by the energy needed to form one mole of an ionic lattice from its gaseous ions (called the lattice enthalpy of formation) rather than trying to measure the strength of ‘one’ bond. You can find more on lattice enthalpy in chapter 3.15 Born Haber cycles

Ionic bonding is stronger and the melting points are higher when the ions are smaller and/or have higher charges. This is because the ions form stronger attractions when they have higher charges and can get closer to each other.

So MgO has a higher melting point than NaCl because the ions involved (Mg$^{2+}$ & O$^{2-}$) are smaller and have higher charges than those in NaCl (Na$^+$ & Cl$^-$)
Typical Physical properties of Ionic Compounds

• High melting points - There are strong electrostatic attractive forces between the oppositely charged ions in the lattice
• Non conductor of electricity when solid - The ions are held together tightly in the lattice and can not move so no charge is conducted
• Good conductor of electricity when in solution or molten - The ions are free to move when in solution and molten. Charge can be carried
• Brittle / easy to cleave apart

A little force will push the ions along and then ions will be next to ions of the same charge. There will be a force of repulsion between ions of the same charge, pushing the layers apart

- They are usually soluble in aqueous solvents.

Solubility of a solute in a solvent is a complicated balance of energy required to break bonds in the solute and solvent against energy given out making new bonds between the solute and solvent.

Ionic substances dissolving in water

When an ionic lattice dissolves in water it involves breaking up the bonds in the lattice and forming new bonds between the metal ions and water molecules.

The negative ions are attracted to the δ+ hydrogens on the polar water molecules and the positive ions are attracted to the δ- oxygen on the polar water molecules.

The higher the charge density the greater the hydration enthalpy (e.g. smaller ions or ions with larger charges) as the ions attract the water molecules more strongly.

Polar water molecules interacting with positive and negative ions of a salt.

• Ionic compounds that are insoluble in water have stronger attractions to each other than to water.

There is more detail on the energy changes in this process in chapter 3.17 enthalpies of Solution.