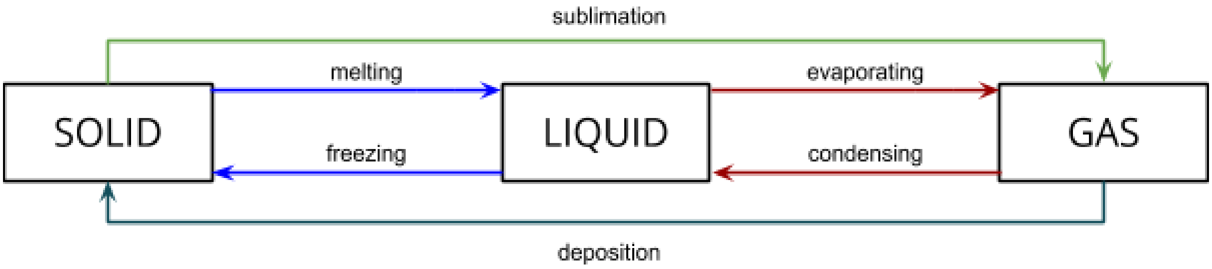
# 5th Form Chemistry

# 2018/19

Sodium acetate trihydrate

Notes taken in class; typed up by Timothy Langer

## Changes of State

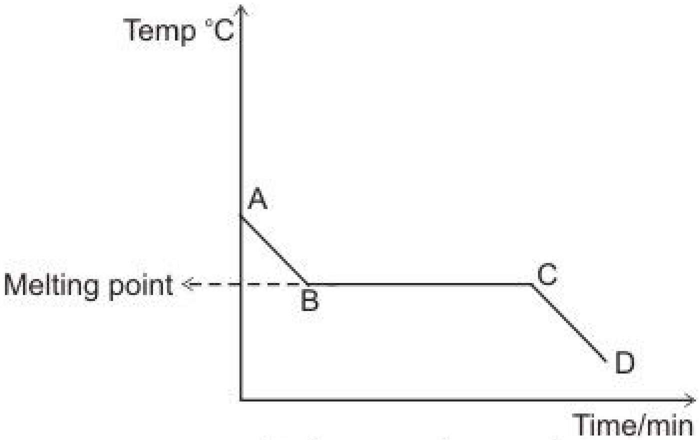



*What is the difference between boiling and evaporation?*

In boiling, one can see bubbles. A liquid only boils if its temperature is equal to or is greater than its boiling point. Bubbles form in the body of the liquid. Bubbles are seen as the gas is produced. In evaporation, only surface particles that have sufficient energy to break free from the neighbours become a gas.

## Cooling Curves

A cooling curve shows the temperature of a pure substance as it cools. These curves have a special shape.




Cooling curve of a pure substance

The flat bit happens at the melting or freezing point. As the liquid cools, the liquid particles lose energy more and more slowly. The means the temperature falls.

When the liquid freezes, new bonds are formed between the particles and the particles stick together in a solid.

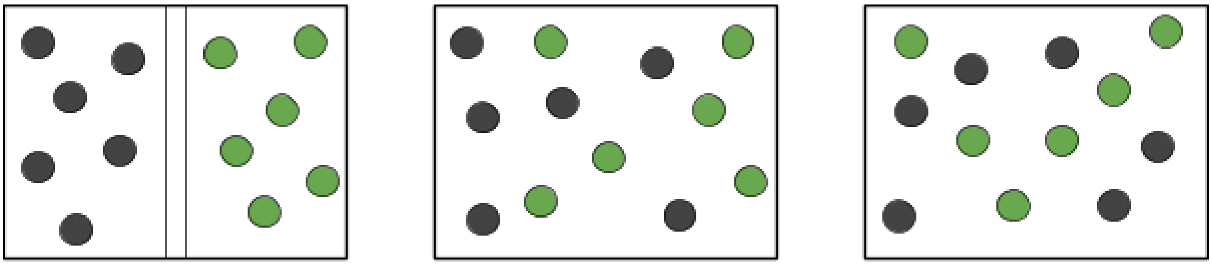
When these bonds are formed, energy is released. Because this energy is released, it stops the substance cooling.

Once the solid is formed, no more bonds are made and the cooling process continues.

## Diffusion

Diffusion is the movement of particles in a fluid from one region to another. It occurs because of two properties of fluids: the particles have spaces between them and are able to move about.

If we start off with two types of fluid which are separated, and then remove the barrier which is separating them then diffusion will occur. The particles from both regions will spread out and move in the spaces between other particles. The particles of each fluid will continue to mix until they are evenly distributed. Diffusion cannot happen in solids because their particles cannot move around.



1. The two types of particle are separated by a barrier.

2. The barrier is removed and the particles have begun to mix.

3. The particles are evenly distributed

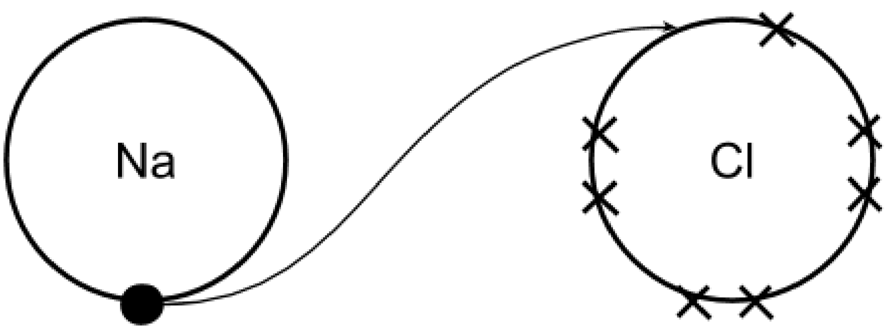
## Chemical Bonding

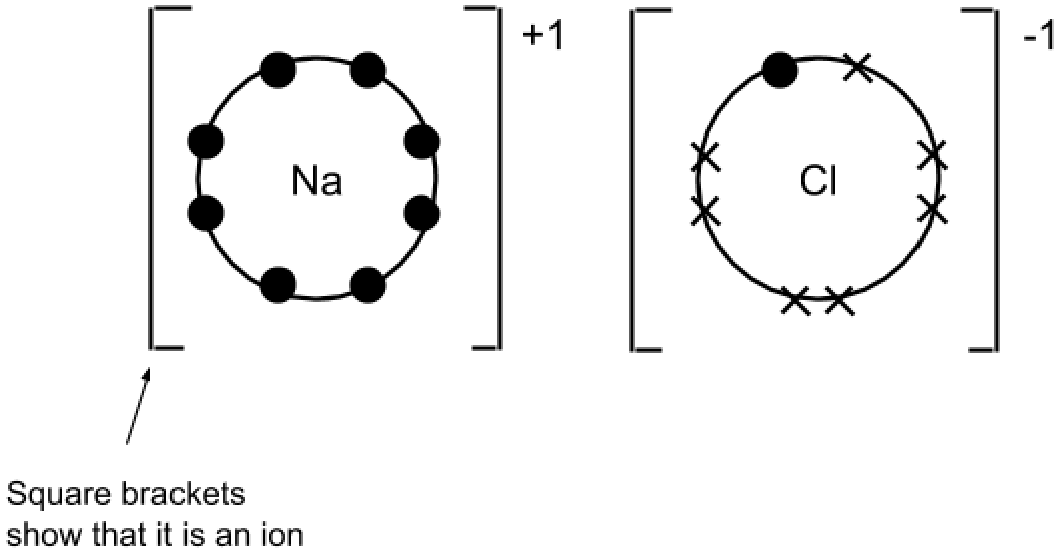
Whenever a chemical reaction takes place, chemical bonds are broken and made. Different elements form different types of bond. Whenever a metal reacts with a non-metal an **ionic bond** is formed.

All atoms react in order to get a full outer shell of electrons. Consider sodium reacting with chlorine, forming sodium chloride, for example:

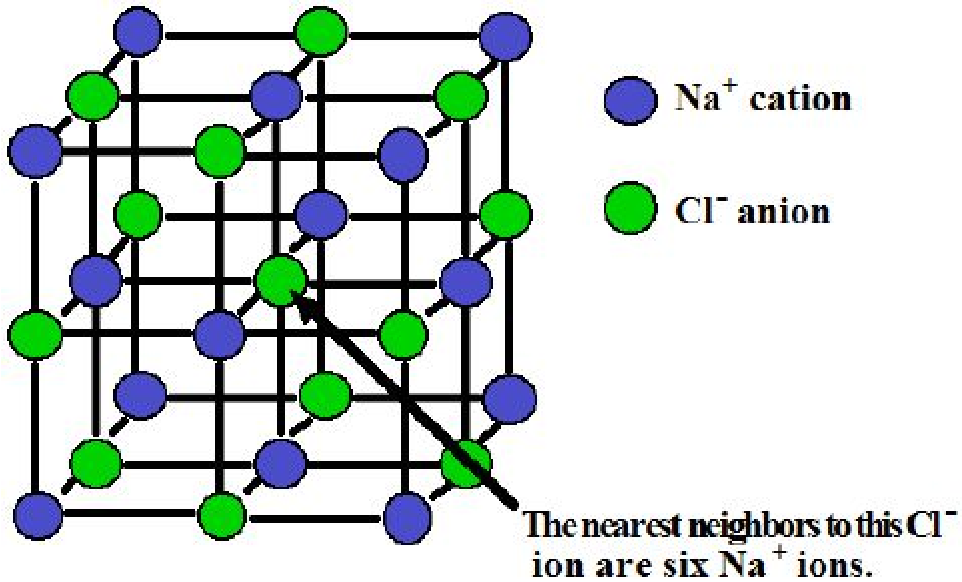
$^{11}$Na has 11e$^-$ & 11p$^+$
2, 8, 1 arrangement
1 in outer shell

$^{17}$Cl has 17e$^-$ & 17p$^+$
2, 8, 7 arrangement
7 in outer shell






## Ionic Structures




Metal atoms lose electrons to form positively charged ions or **cations**. (Remember that **cat**ions are **pussy**tive) Non-metal atoms gain electrons to form negatively charged ions or **anions**.

The oppositely charged ions that are formed are held in a regular three-dimensional lattice by electtrostatic attractions between the ions. The ions pack together in the most efficient way so there is little wasted space.

In NaCl, each $Na^+$ ion is surrounded by six $Cl^-$ ions, and each $Cl^-$ ion is surrounded by six $Na^+$ ions. This means that there are many strong electrostatic attractions between the ions within the giant lattice structure.

Examples of ionic compounds include:

- Sodium chloride
- Copper sulfate

The giant ionic lattices that are formed by ionic compounds account for their properties or characteristics:

1. **Melting and boiling points**

   Each ion is strongly bonded by electrostatic attraction to several others. These strong bonds are hard to break. A lot of **heat energy** is required and so the melting and boiling points are high.

2. **Crystalline structure**

   Salt has a hard crystalline structure. Crystals require a very regular arrangement with straight lines, meaning that the ions line up in straight lines, forming crystals with precise and straight edges and corners.

3. **Electrical conductivity**

   A flow of charged particles is creates an electrical current. If ions are able to move, they can conduct electricity. Ionic compounds such as NaCl contain ions. When they are solid, they cannot conduct electricity because the *ions* are unable to move. If the crystal is melted to become a liquid or dissolved in water to become a solution, the *ions* are free to move and they can carry an electrical charge.

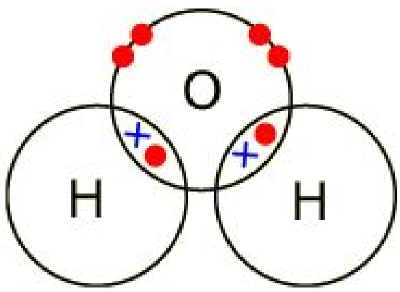
4. **Brittleness**

   Ionic crystals are **brittle**; they shatter when hit. This is because the hitting can displace a row of ions which are then forced next to ions of a similar charge. Ions with the same charge repel, so the crystal breaks apart.

## Covalent Bonding

This occurs between two or more non-metal atoms. The atoms all need to gain electrons, so they share. The sharing is *always* fair!

### How to draw a dot and cross diagram

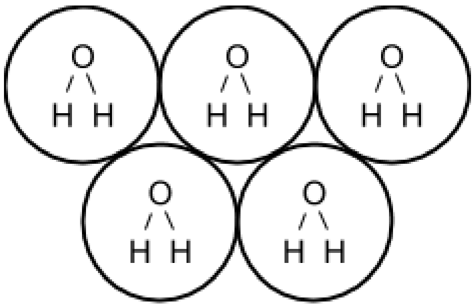



e.g. $H_2O$

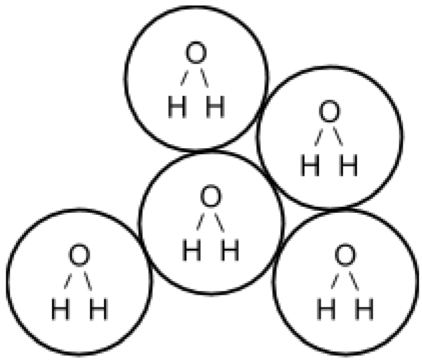
1. Draw atom circles first and label them.
2. Concentrate on the outside atoms and what they need to gain.
3. Fill in the shared areas fairly
4. Put in the rest of the outer shell electrons

## Simple Molecular Structures

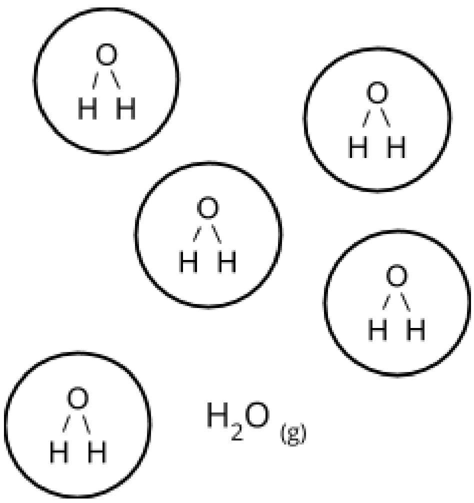
The vast majority of covalently bonded compounds exist as simple molecules. Water, $H_2O$, is an example. Consider the three states of water.



$H_2O_{(s)}$



$H_2O_{(l)}$




The water molecule $H_2O$ is the <u>same</u> in each state.

The covalent bonds **do not change**.

Simple molecular structures (e.g. $H_2O$, $CO_2$, $O_2$, $NH_3$, etc.) tend to have <u>low</u> melting and boiling points.

The covalent bonds are very strong, but the **intermolecular forces** that exist between different molecules are weak. These forces of attraction need only a small amount of heat energy to be broken, and so simple molecular substances are easily melted and boiled.

### Electrical conductivity of simple molecular structures

Simple molecular structures do not conduct electricity under **any** circumstances, as they have no ions, and all the electrons are bonded to the atoms and are not free to move. Water conducts electricity only when it contains impurities such as salt. Pure (distilled) water does **not** conduct.

## Giant Covalent Structures

The vast majority of non-metals bind together to form simple molecules. In a few cases (3 at iGCSE) they form giant covalent lattices.

### 1. Diamond

Diamond is an **allotrope** of carbon. An allotrope is a different physical form of the same element. Each carbon atom is covalently bonded to 4 others in a tetrahedral arrangement. These covalent bonds are very strong and its structure is hard and rigid.

### Properties of diamond

**a. Melting point / boiling point**

Diamond has a **very high** melting and boiling point. The whole structure is held together by many strong covalent bonds. These need a large amount of heat energy to break them.

*Fact: Diamond <u>sublimes</u> at roughly 3600°C*

**b. Electrical conductivity**

Diamond has no ions and no mobile electrons, so there is no way it can conduct.

**c. Crystalline structure**

The atoms are all lined up in a regular lattice. This makes diamond crystalline.

## 2. Graphite

Each carbon atom is covalently bonded to 3 others in the same layer. Loose electrons are found between the layers and hold the layers together as an intermolecular force.

### Properties of Graphite

**a. Melting point / boiling point**

Very high because you need to break many strong covalent bonds and this requires a lot of heat energy.

**b. Electrical conductivity**

Graphite *does* conduct electricity. The electrons between the layers are free to move, and therefore, because they are charged *and* mobile, they can conduct. Graphite is the **only** covalent conductor.

Graphite's layers are held very loosely in place. They slide off very easily, making graphite slippery to the touch. This means it can be used in "lead" pencils. Since graphite is a **lubricant** is used inside locks. It is also great for unsticking zippers.

## 3. Silicon dioxide

Silicon dioxide ($SiO_2$), also called silica, is a giant lattice in which each silicon atom is covalently bonded to four oxygen atoms and each oxygen atom is bonded to two silicon atoms.

*Note: Do not confuse with silicone. Silicone is used for implants.*

### Properties of silicon dioxide

a. **Melting point / boiling point**

Silicon dioxide is a giant structure. There are many strong covalent bonds between the atoms, so lots of heat energy is needed to break them. Therefore the melting and boiling points of silicon dioxide are **very high**.

b. **Electrical conductivity**

Silicon dioxide does not conduct electricity, as there are no ions and no free electrons, so nothing can carry a charge.

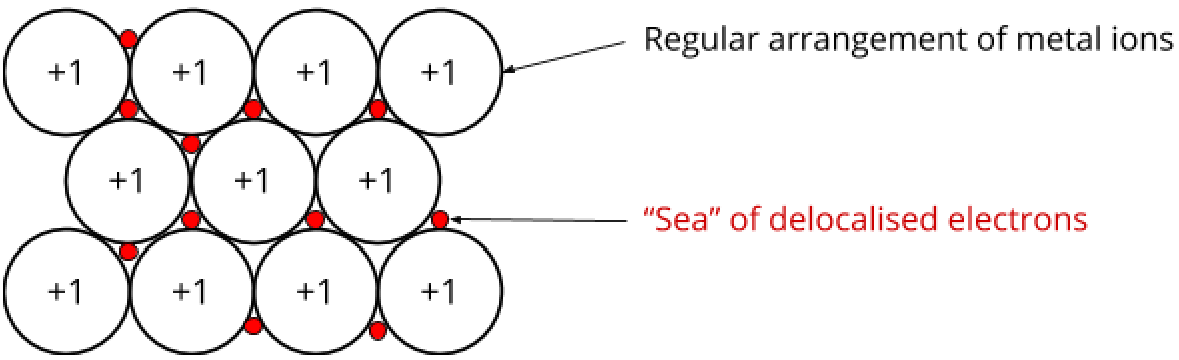
## Buckminsterfullerene

*Colloquially known as footballane due to its shape.*

Buckminsterfullerene is a simple molecular structure. It has a low melting and boiling point due to weak forces of attractions between the molecules. Each carbon atom is covalently bonded to 3 others. This is the same arrangement as in graphite. The 4th outer shell electron of each atom is delocalised and is free to move around the ball. However, these electrons cannot move from one ball to another, meaning that there is no electrical conductivity because there are no mobile charged particles. Buckminsterfullerene is used as miniature ball bearings (like in fidget spinners) in machinery and joints as it is suitable as a lubricant.

## Metallic Structures

All metals have a giant metallic structure. Metal atoms line up in a regular crystalline fashion. The outer shell electrons leave the atoms and move freely around the structure. These electrons are said to be **delocalised**. Consider sodium, for example:




The structure is held together by numerous electrostatic attractions between the metal ions and the sea of delocalised electrons. These attractions need a lot of heat energy to be overcome, therefore metals have high melting points. (at GCSE)

Metals are **good electrical conductors** as the electrons are free to move throughout the structure and thus they can carry a charge.

## Malleability

A malleable material can be hammered or bent into shape - it is **not brittle**. All pure metals are malleable and this is because the atoms are able to slide over each other, whilst maintaining the electrostatic attractions with the delocalised electrons. When given a shove rows of atoms move but they still hold together.