

# 5th Form Chemistry

2018/19

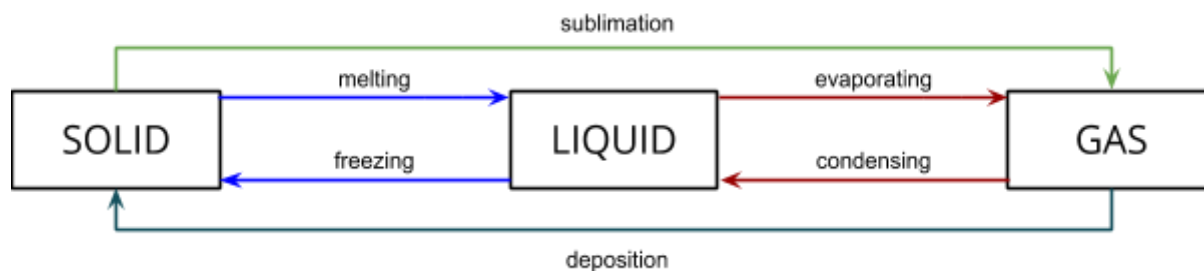


Synthetic bismuth metal crystal  
The surface is an iridescent very thin  
layer of oxidation.

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## Changes of State

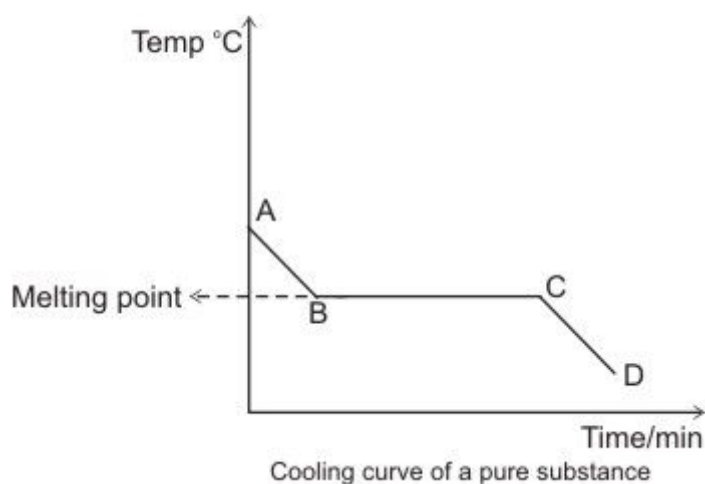


*What is the difference between boiling and evaporation?*

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.



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

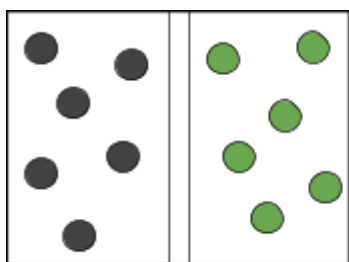
Once the solid is formed, no more bonds are made and the cooling process continues.

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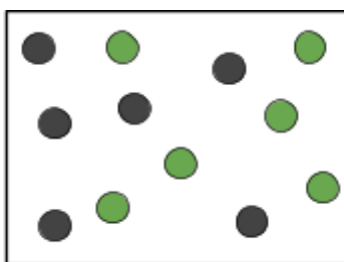
### Diffusion

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.

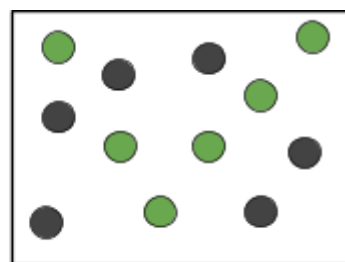
If we start off with two types of fluid which are separated, and then remove the barrier which is separating them then diffusion will occur. The particles from both regions will spread out and move in the spaces between other particles. The particles of each fluid will continue to mix until they are 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

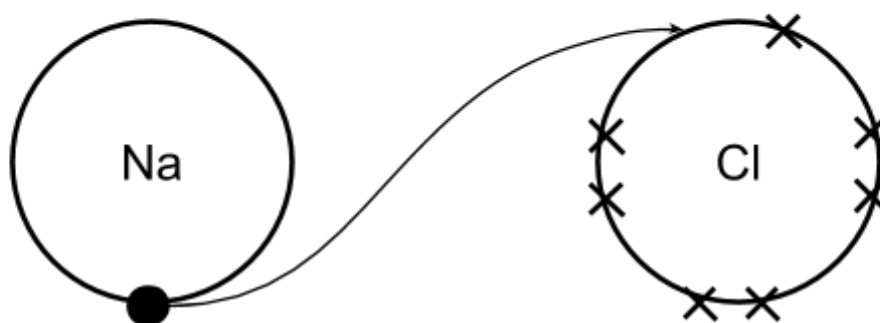
### Chemical Bonding

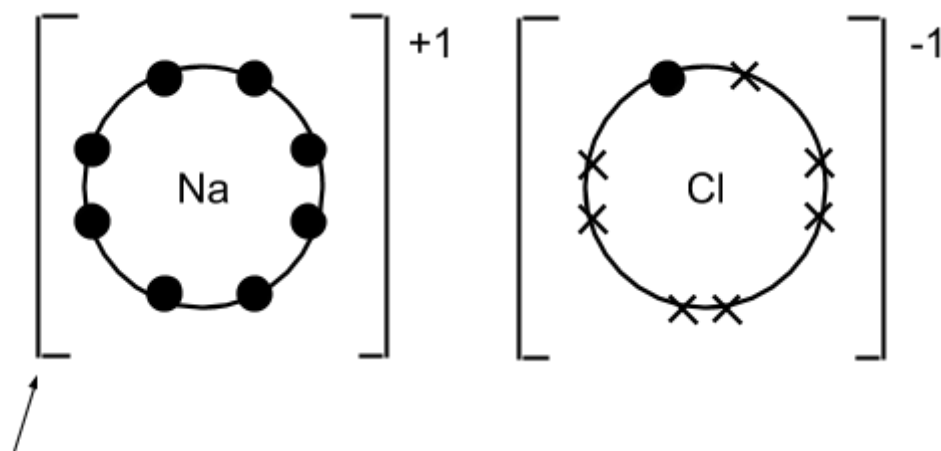
Whenever a chemical reaction takes place, chemical bonds are broken and made. Different elements form different types of bond. Whenever a metal reacts with a non-metal an **ionic bond** is formed.

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

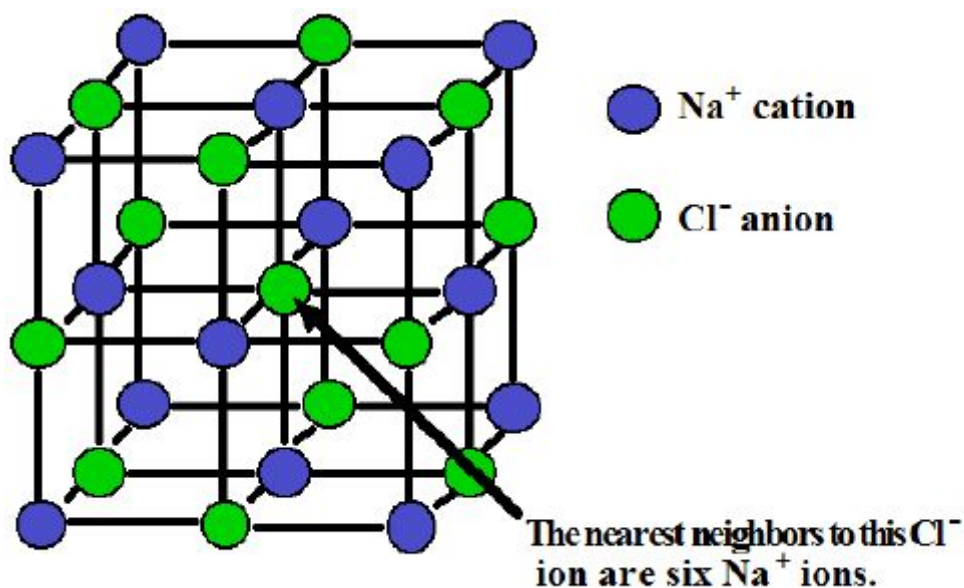
$^{11}\text{Na}$  has  $11\text{e}^-$  &  $11\text{p}^+$   
2, 8, 1 arrangement  
1 in outer shell

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





## Ionic Structures



Metal atoms lose electrons to form positively charged ions or **cations**. (Remember that **cations** 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 electrostatic attractions between the ions. The ions pack together in the most efficient way so there is little wasted space.

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

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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 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.

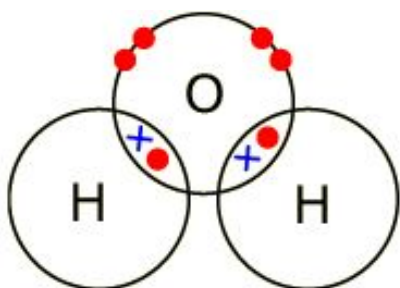
### 4. Brittleness

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

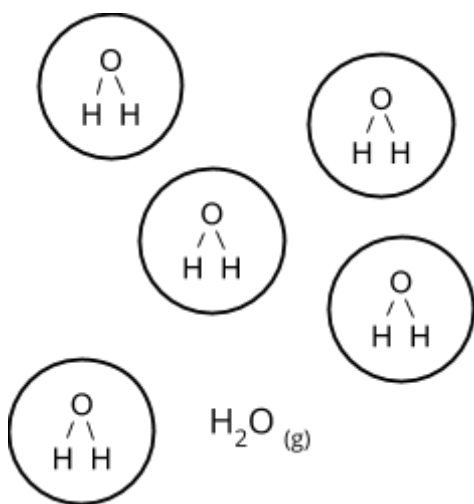
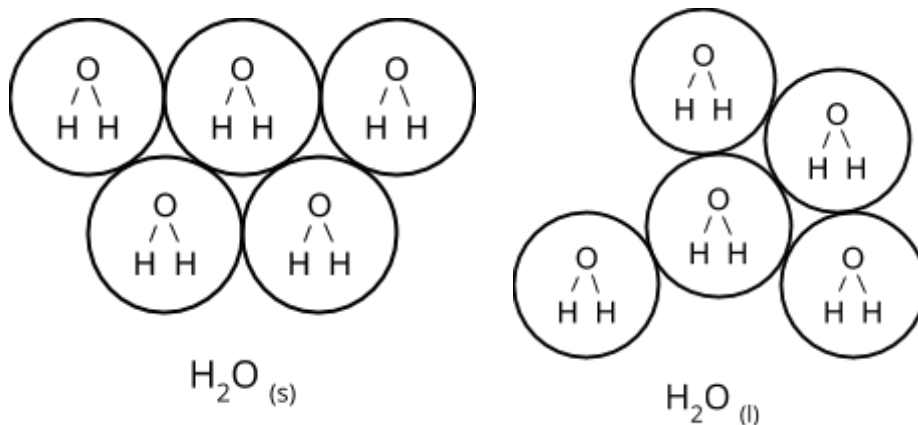


e.g. H<sub>2</sub>O

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,  $\text{H}_2\text{O}$ , is an example. Consider the three states of water.



The water molecule  $\text{H}_2\text{O}$  is the same in each state.

The covalent bonds **do not change**.

Simple molecular structures (e.g.  $\text{H}_2\text{O}$ ,  $\text{CO}_2$ ,  $\text{O}_2$ ,  $\text{NH}_3$ , etc.) tend to have low 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.

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### 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 sublimes 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 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 ( $\text{SiO}_2$ ), 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.*



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### 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.

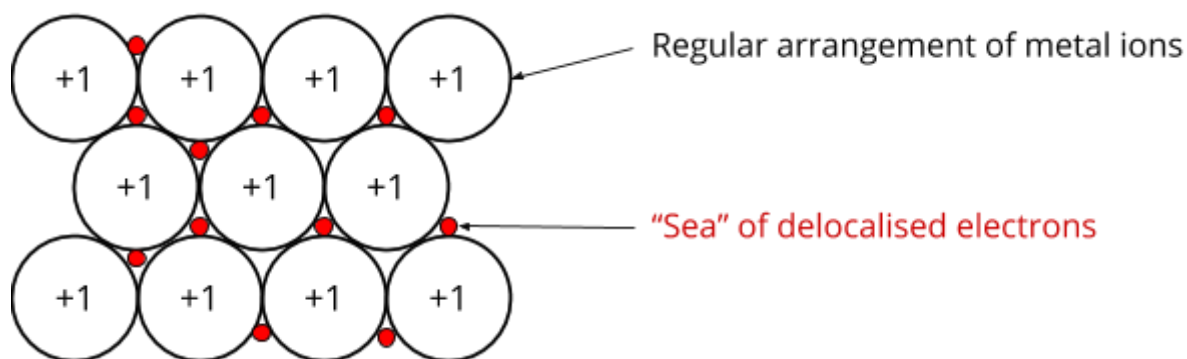
## Buckminsterfullerene

*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.

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### 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.

### The Periodic Table

The Periodic Table is a list of all known elements arranged in a very useful way. The elements occur in order of their **atomic number**. Rows going across the table are called **periods**. All elements in the same period have the same number of electron shells. Columns going down the table are called **groups**. Elements in the same group have the same number of electrons in the outer shell.

At GCSE we study 3 chemical groups:

- Group 0      Noble gases
- Group I      Alkali Metals
- Group VII    Halogens

#### Group 0 - Noble gases

This group is not very interesting. All members of the group have a full outer shell of electrons so they do not form any chemical bonds. This makes them exceedingly unreactive. However, this lack of reactivity means they are useful as such:

He	Helium	party balloons, airships, blimps
Ne	Neon	advertising signs, "neon lights"
Ar	Argon	incandescent light bulbs (stops filament from burning)
Kr	Krypton	used as a "flash" in photography (high-speed flash)
Xe	Xenon	used in "sunbed" tubes

#### Group I - Alkali Metals

This group of elements is *much* more interesting.

- Li
  - Na
  - K
  - Rb
  - Cs
  - Fr
- } available in school
- } available, but school doesn't buy them
- radioactive and only exists for fractions of a second

Group I elements react by losing an electron. As you go down the group, the atoms get bigger. The outer shell electrons gets further from the nucleus. The electron shells between the outer electrons and the nucleus act as "shields" and deflect some of the nuclear attraction. The outermost electron feels less attraction from the nucleus, and so it is lost more easily and reactions happen more readily.

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- **Reactivity increases** as you move down the group
- **Melting point decreases** as you move down the group
- **Hardness decreases** as you move down the group
- **Density increases** as you move down the group

### Group VII - The Halogens

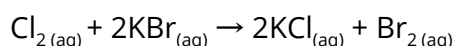
All the Halogens are non-metals, have 7 electrons in their outer shell and are all toxic (and will kill you if you let them)

Name	Formula	Colour	State (at RTP)
Fluorine	F <sub>2</sub>	Pale yellow	gas
Chlorine	Cl <sub>2</sub>	Green	gas
Bromine	Br <sub>2</sub>	Red / Brown	liquid
Iodine	I <sub>2</sub>	Grey / Black	solid
Astatine	At <sub>2</sub>	<i>Black*</i>	<i>solid*</i>

\* these are predictions, since Astatine is radioactive and decays too quickly for us to see

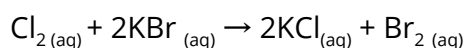
### Halogen Reactivity

Halogens react in a series of **displacement reactions**. In the following reaction, for example, the more reactive chlorine displaces the less reactive bromine from the compound:

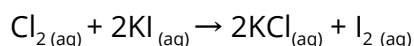


### Halogen Reactions [Practical]

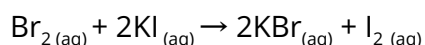
	KCl <sub>(aq)</sub>	KBr <sub>(aq)</sub>	KI <sub>(aq)</sub>
Cl <sub>2(aq)</sub>		Pale green → orange. Addition of cyclohexane turned it orange.	Pale green → brown. Addition of cyclohexane turned it pink.
Br <sub>2(aq)</sub>	No colour change		Orange turned brown. Addition of cyclohexane turned it pink.
I <sub>2(aq)</sub>	No colour change	No colour change	



colour in cyclohexane: **orange**



colour in cyclohexane: **pink**



colour in cyclohexane: **pink**

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**Conclusion:** Chlorine (Cl<sub>2</sub>) is more reactive > Bromine (Br<sub>2</sub>) is more reactive > Iodine (I<sub>2</sub>)

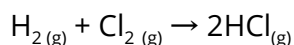
In each case, the more reactive halogen is able to displace the less reactive one from its compound.

### Explanation

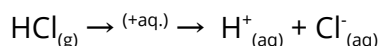
Halogens have **seven** electrons in their outer shell. When they react, they need to gain **one** electron. Fluorine is **2,7** so the incoming electron goes into the second shell. This is close to the nucleus and there is only **one** shield and so there is a stronger nuclear attraction. As you go down the group, the atoms get larger, so the outer shell gets further from the nucleus and the number of shields increases. The attraction between the incoming electron and the nucleus decreases. It becomes more difficult for the new electron to join the atom. Reactivity decreases.

### Hydrogen Halides

All halogens react with hydrogen gas to form hydrogen halides. For example:



All hydrogen halides are covalently bonded molecules and are gases at RTP. When any hydrogen halide dissolves in water (*only* water) it **dissociates** (splits into two) and turns into ions. For example:

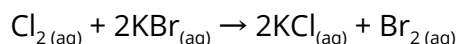


It is the H<sup>+</sup> ions which makes the solutions acidic, therefore HCl<sub>(aq)</sub> is hydrochloric *acid*.

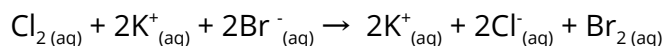
### Halogens, redox and ionic equations

[All high level stuff] Halogens react with halide ions in a series of displacement reactions. By looking at these in more detail, we can see that they are actually **redox** reactions, in which reduction and oxidation are both taking place. To look at the reactions more closely, we need to convert the balanced symbol equations into **ionic equations**. To do this, we follow a series of steps:

1. Write out the full balanced chemical equation with state symbols.



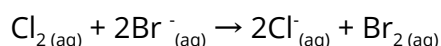
2. For all **compounds** that are *both aqueous and ionic*, split them into their ions.



3. Cancel out "spectator ions"

In this case, remove 2K<sup>+</sup><sub>(aq)</sub> from both sides as it does not change nor react.

4. Write out the ionic equation *with state symbols*.



How is this redox?

"Redox" is a reaction that involves simultaneous oxidation **and** reduction.

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What is oxidation?

Oxidation can be defined as

- When something reacts with and combines with oxygen
- When something loses hydrogen
- When something loses electrons

**For example:**  $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$

Bromide has lost electrons. It has been oxidised.

What is reduction?

Reduction can be defined as

- When something reacts with and gains hydrogen
- When something loses oxygen
- When something gains electrons

**For example:**  $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$

Chlorine has gained electrons. It has been reduced.

### Silver Halides

All aqueous halide ions react with silver nitrate solution to form **precipitates**. These precipitates are of characteristic colour, and so they can be used to identify the halide ions.

$\text{Cl}^-_{(\text{aq})} + \text{Ag}^+_{(\text{aq})} \rightarrow \text{AgCl}_{(\text{s})}$  a white precipitate (AgCl) is formed

$\text{Br}^-_{(\text{aq})} + \text{Ag}^+_{(\text{aq})} \rightarrow \text{AgBr}_{(\text{s})}$  a cream precipitate (AgBr) is formed

$\text{I}^-_{(\text{aq})} + \text{Ag}^+_{(\text{aq})} \rightarrow \text{AgI}_{(\text{s})}$  a yellow precipitate (AgI) is formed

### Organic Chemistry

Organic chemistry is the study of covalently bonded carbon compounds. The subject is divided into studying different types of organic compounds, which range from

Very simple e.g.  $\text{CH}_4$  (methane)

Very complicated e.g. haemoglobin

The simplest organic compounds are **hydrocarbons**. These are compounds that contain *only* hydrogen and carbon atoms. Many hydrocarbons are sources from **crude oil**.

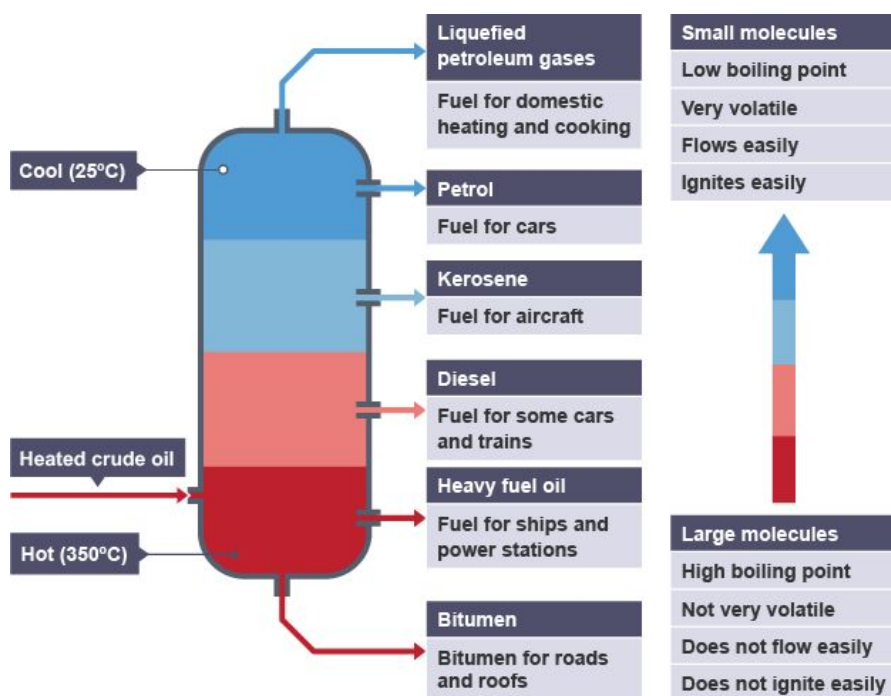
Crude oil is a fossil fuel. It has taken tens of millions of years to form. It is formed from dead plankton. These dead organisms lived in the prehistoric seas. When they died, they sank to the bottom and were covered in layers of sediment. Neither oxygen nor water was present: only the dead material and the sediment. (If there had been oxygen, it would have decomposed)

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At high pressure and temperature between 90°C and 120°C a series of very slow reactions turn the dead stuff into crude oil. Crude oil is extracted from rocks and processed at oil refineries.

### Refining crude oil

Crude oil is pumped into the refinery where it can undergo a variety of treatments. **Fractional distillation** occurs first. This occurs in a giant fractionating column which is between 80 and 100 metres tall.



Crude oil is heated until it boils and the vapours pass into a fractionating column, which is cooler at the top and hotter at the bottom. The vapours rise up the column. How far up the column a particular hydrocarbon moves depends on its boiling point. Smaller hydrocarbons have lower boiling points, while longer chain hydrocarbons have higher boiling points.

### Crude oil and its uses

Crude oil itself has no uses and it has to be separated into fractions before it can be used. These fractions are all mixtures, but each one contains a narrow range of sizes of hydrocarbons with similar boiling points.

#### Refinery gases

Refinery gases are a mixture of methane, ethane, propane and butane, which can be separated into individual gases if required. These gases are commonly used as liquefied petroleum gas (LPG) for domestic heating and cooking.

#### Gasoline (petrol)

As with all other fractions, petrol is a mixture of hydrocarbons with similar boiling points. It is used as fuel in cars.

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### Kerosene

Kerosene is used as a fuel for jet aircraft, as domestic heating oil and as “paraffin” for small heaters and lamps.

### Diesel

Used as fuel for buses, lorries, some cars and certain railway engines.

### Fuel oil

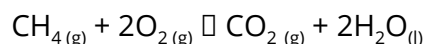
Fuel oil is used as a fuel for ships and for industrial heating.

### Bitumen

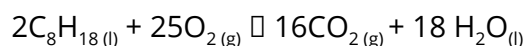
Bitumen is a thick sticky black material (also known as *residue*) which is used to make the top surface of roads.

## Hydrocarbons

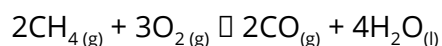
All hydrocarbons burn in air (oxygen) to form carbon dioxide and water and release heat in the process. The various fractions obtained from the column can therefore be used as **fuels**. A fuel is a substance which, when burned, releases heat energy. For example, the equation for burning methane is as follows:



Burning octane (present in gasoline/petrol):



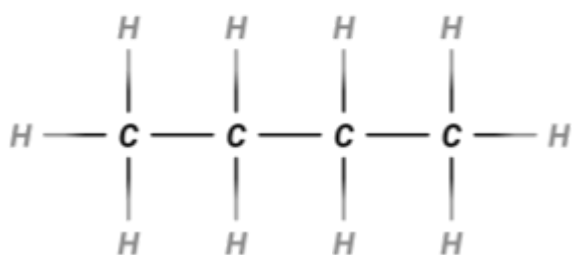
If there is not enough air (or oxygen), incomplete combustion occurs. This leads to the formation of carbon (soot) or carbon monoxide instead of carbon dioxide. For example, if methane burns in a badly maintained gas appliance, there may not be enough oxygen available to produce carbon dioxide, and so you get toxic carbon monoxide instead:



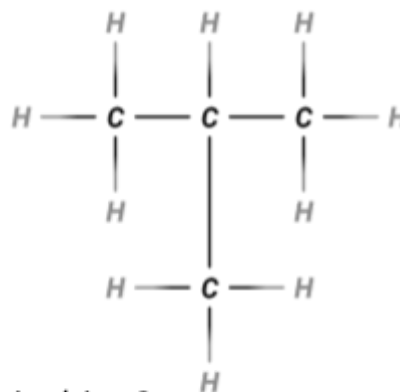
The formation of carbon monoxide from the incomplete combustion of hydrocarbons is very dangerous. Carbon monoxide is colourless and odourless and is very poisonous. Carbon monoxide is poisonous because it reduces the ability of the blood to carry oxygen around the body.

*Fact: Carbon monoxide has a higher affinity for haemoglobin than oxygen does. Once carbon monoxide has bonded to haemoglobin it stays there until the red blood cell dies. (Up to 3 months) This means that the supply of oxygen to the body's cells is reduced and cells can't respire. Cells die and, depending on how much carbon monoxide has been inhaled, the person can die as well.*

## Isomers



**Butane** is  $C_4H_{10}$

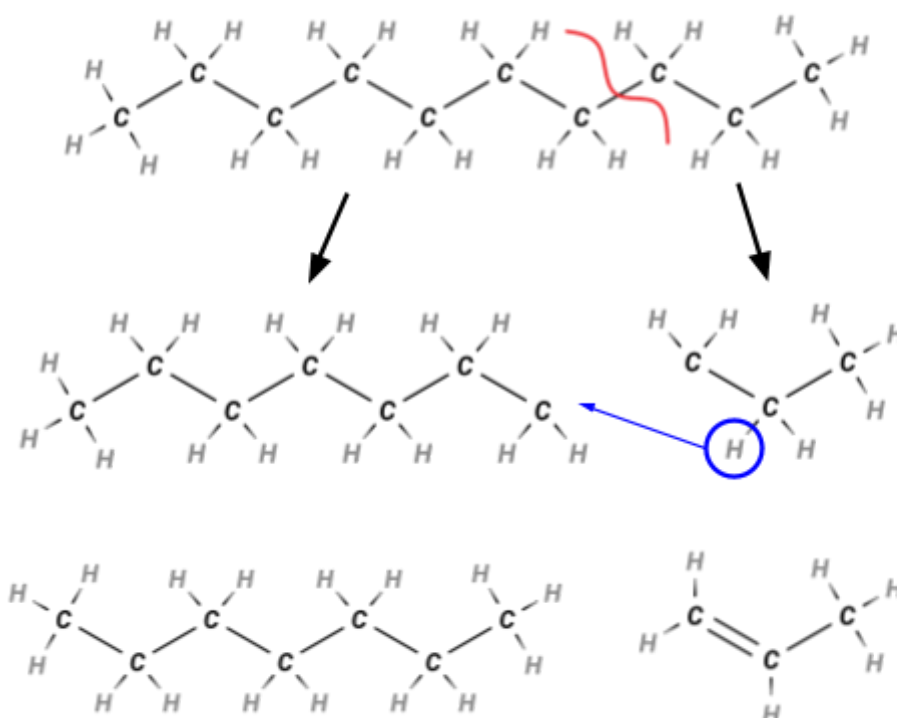


*So what is this...?*

This is an **isomer** of butane. It has the same molecular formula but its atoms are arranged in a different order. At GCSE, remember that the isomer of butane is called *2-methylpropane*.

## Cracking

Cracking is a chemical process in which a long chain alkane is broken down; it is a type of thermal decomposition. It is a way of producing valuable alkenes. Cracking is a totally random process. The alkane molecule can be cracked in any position and any numbers of times.



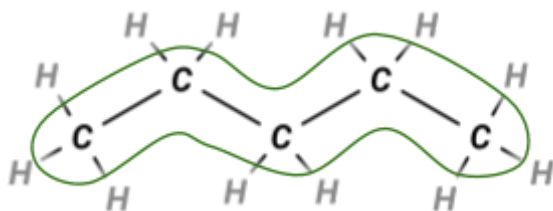
We can split the decane ( $C_{10}H_{22}$ ) molecule in any place.

However, once we have cracked it, we have two unhappy fragments. So, the H moves and C=C forms.

Now we have cracked decane to form heptane and propene.

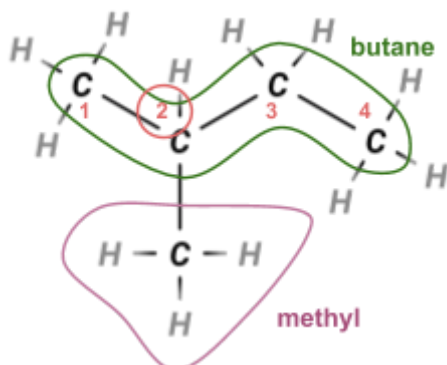


## Naming compounds

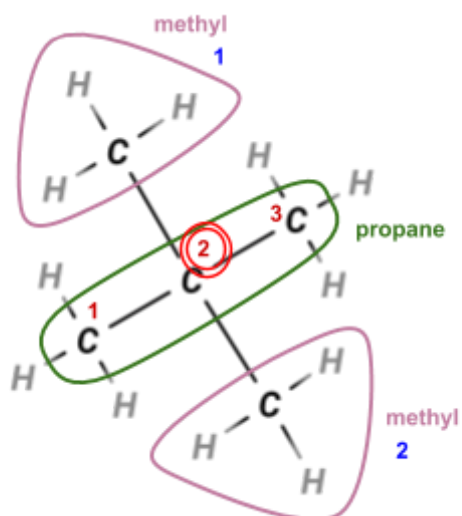


pentane

Longest chain  
of carbons is 5  
∴ **pent**



**2-methylbutane**



**2,2-dimethylpropane**

### To name a compound:

1. Decide what homologous series it is and do the ending. If it is an alkane, it will end in "ane". If it is an alkene, it will end in "ene".
2. Find the longest chain (it may go around corners) and put it at the end.
3. Identify the "dangly bits" and list them in alphabetical order.
4. Number the carbon atoms in the chain so that the dangly bits get the lowest numbers.
5. Complete the name using commas to separate numbers and using hyphens to separate numbers and letters.

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### Alkanes

Organic chemicals are arranged in families called **homologous series**. The alkanes are an example of a homologous series. The alkanes are described as being **saturated** (meaning they *only* contain single bonds) **hydrocarbons** (meaning they *only* contain hydrogen and carbon atoms).

When dealing with any organic compound, certain bonding rules are **always** obeyed.

C	4 bonds
H	1 bond
Halogen	1 bond
Oxygen	2 bonds

The alkanes, like all homologous series, share a **general formula**.

$C_nH_{2n+2}$  Where n is any whole number  $\geq 1$

n	Formula	Structure	Name
1	$CH_4$	$\begin{array}{c} H \\   \\ H - C - H \\   \\ H \end{array}$	methane
2	$C_2H_6$	$\begin{array}{ccccc} & H & & H & \\ &   & &   & \\ H & - C & - & C & - H \\ &   & &   & \\ & H & & H & \end{array}$	ethane
3	$C_3H_8$	$\begin{array}{ccccccc} & H & & H & & H & \\ &   & &   & &   & \\ H & - C & - & C & - & C & - H \\ &   & &   & &   & \\ & H & & H & & H & \end{array}$	propane

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4	$C_4H_{10}$	$  \begin{array}{ccccccc}  & H & & H & & H & & H \\  &   & &   & &   & &   \\  H & -C & - & C & - & C & - & C & -H \\  &   & &   & &   & &   \\  & H & & H & & H & & H  \end{array}  $	butane
5	$C_5H_{12}$	$  \begin{array}{ccccccccc}  & H & & H & & H & & H & & H \\  &   & &   & &   & &   & &   \\  H & -C & - & C & - & C & - & C & - & C & -H \\  &   & &   & &   & &   & &   \\  & H & & H & & H & & H & & H  \end{array}  $	pentane
6	$C_6H_{14}$	$  \begin{array}{ccccccccc}  & H & & H & & H & & H & & H & & H \\  &   & &   & &   & &   & &   & &   \\  H & -C & - & C & - & C & - & C & - & C & - & C & -H \\  &   & &   & &   & &   & &   & &   \\  & H & & H & & H & & H & & H & & H  \end{array}  $	hexane

Remember the first four alkanes!

- **M**onkeys **E**at **P**eanut **B**utter
- **M**y **E**lephant **P**refers **B**ananas

## Reactions of Alkanes

Alkanes are not very reactive when compared to other organic chemicals. They only do two types of reaction.

### Combustion

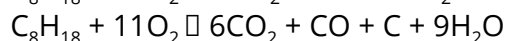
All alkanes burn very well and produce a lot of heat energy. The vast majority of them are used as fuels, and, provided the oxygen supply is sufficient, they undergo **complete combustion**.

1. **Propane  $C_3H_8$**   
 $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$
2. **Heptane  $C_7H_{16}$**   
 $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$
3. **Octane  $C_8H_{18}$**   
 $C_8H_{18} + 12\frac{1}{2}O_2 \rightarrow 8CO_2 + 9H_2O$
4. **Nonane  $C_9H_{20}$**   
 $C_9H_{20} + 14O_2 \rightarrow 9CO_2 + 10H_2O$

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Sometimes the oxygen supply is limited. If this happens, the alkane will still burn, but **incomplete combustion** takes place. Many different products can be formed depending on both the amount of oxygen and pure luck.

### Octane C<sub>8</sub>H<sub>18</sub>

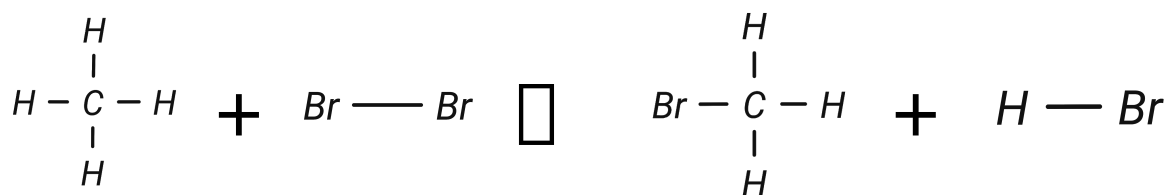


*Both are correct. Both can occur.*

Carbon monoxide has a higher affinity for haemoglobin than oxygen does. *For more information about carbon monoxide poisoning, see the section about hydrocarbons.*

## Substitution

When an alkane is mixed with a halogen (an element in Group VII) *in the presence of UV light* hydrogen atoms can be swapped for halogen atoms to form a produce called a **halogenoalkane**. There is only one such reaction on the syllabus, however:



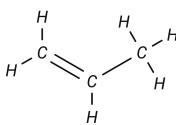
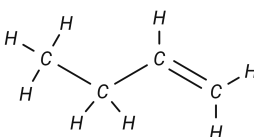
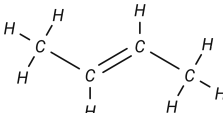
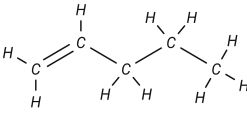
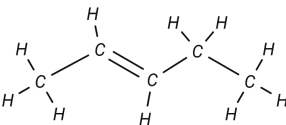
methane + bromine → bromomethane + hydrogen bromide

## Alkenes

The alkenes are another homologous series, and they too are hydrocarbons. Unlike alkanes, which are **saturated**, alkenes are **unsaturated**. A saturated compound contains *only* single bonds. An unsaturated compound contains (at least) one C=C, i.e. a double bond. The alkenes have a general formula C<sub>n</sub>H<sub>2n</sub>

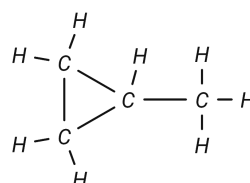
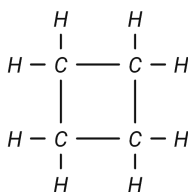
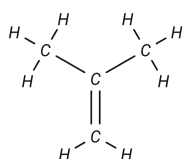
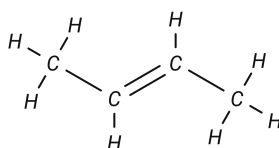
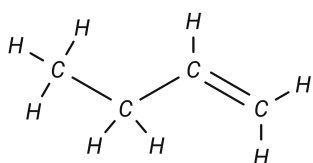
n	name	formula	structure
1	Does not exist, else there would not be a double bond.		
2	ethene	C <sub>2</sub> H <sub>4</sub>	$\begin{array}{ccc} \text{H} & & \text{H} \\ & \diagdown & / \\ & \text{C} = \text{C} & \\ & / & \diagdown \\ \text{H} & & \text{H} \end{array}$

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3	propene	$C_3H_6$	
4	but-1-ene	$C_4H_8$	
	but-2-ene		
5	pent-1-ene	$C_5H_{10}$	
	pent-2-ene		

One of the fun things about alkenes is that they have loads of isomers. Another homologous series, the cycloalkanes, also have a general formula  $C_nH_{2n}$

This means that  $C_4H_8$  has 5 isomers (7 at A-level)



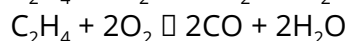
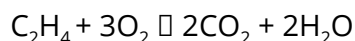
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### Reactions of Alkenes

#### Combustion

Like alkanes, alkenes can undergo complete or incomplete combustion.

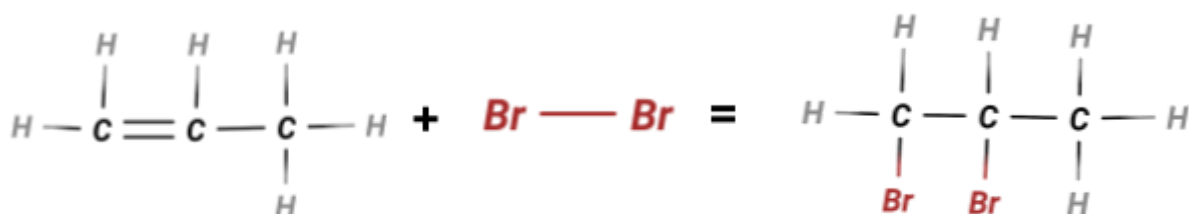
ethene  $C_2H_4$



However, unlike alkanes, alkenes are not used as fuels. They are too useful in other reactions and therefore they are too valuable to burn.

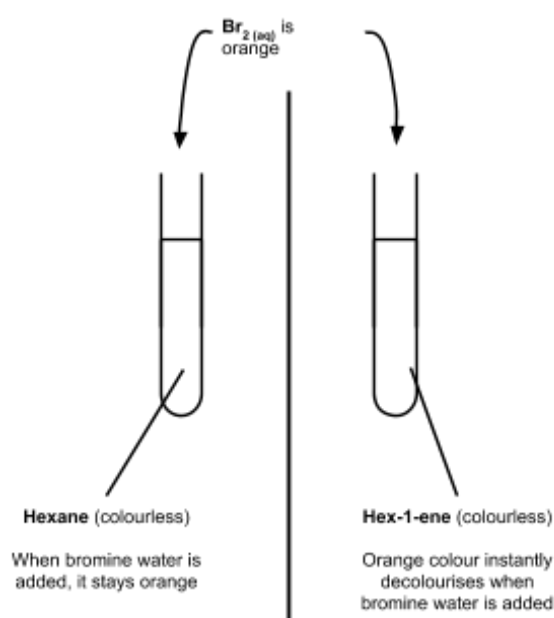
#### Addition reactions

Alkenes contain a  $C=C$ . This is made up of a regular single bond ( $C-C$ ), plus a second type of bond called a pi ( $\pi$ ) bond. The  $\pi$  bond is weaker than single bonds, and it breaks fairly readily. This means that other atoms can add onto the carbons.



Propene + bromine  $\rightarrow$  1,2-dibromopropane (colourless)

This reaction happens instantly. No UV light is needed. Bromine water is used as a test to distinguish between alkanes and alkenes.

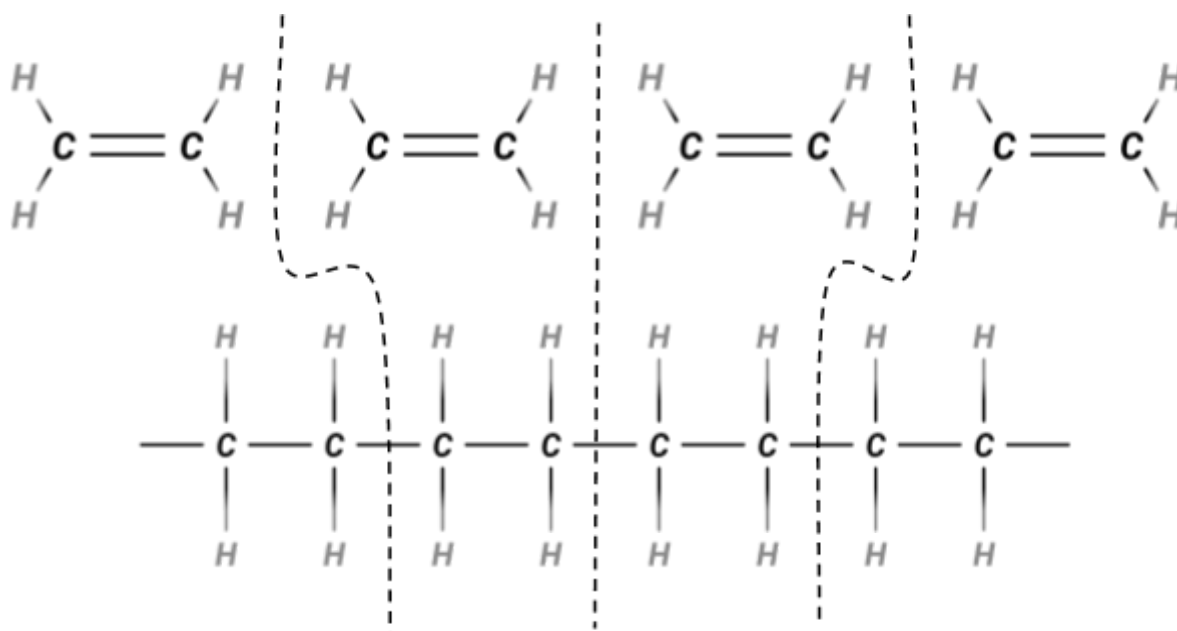


## Polymerisation

Alkenes are most widely used to make **polymers**. An alkene molecule is a **monomer**. When it reacts with many other monomers it makes a polymer. Most polymers are used as plastics. When any alkene polymerises, *only* the C=C reacts.

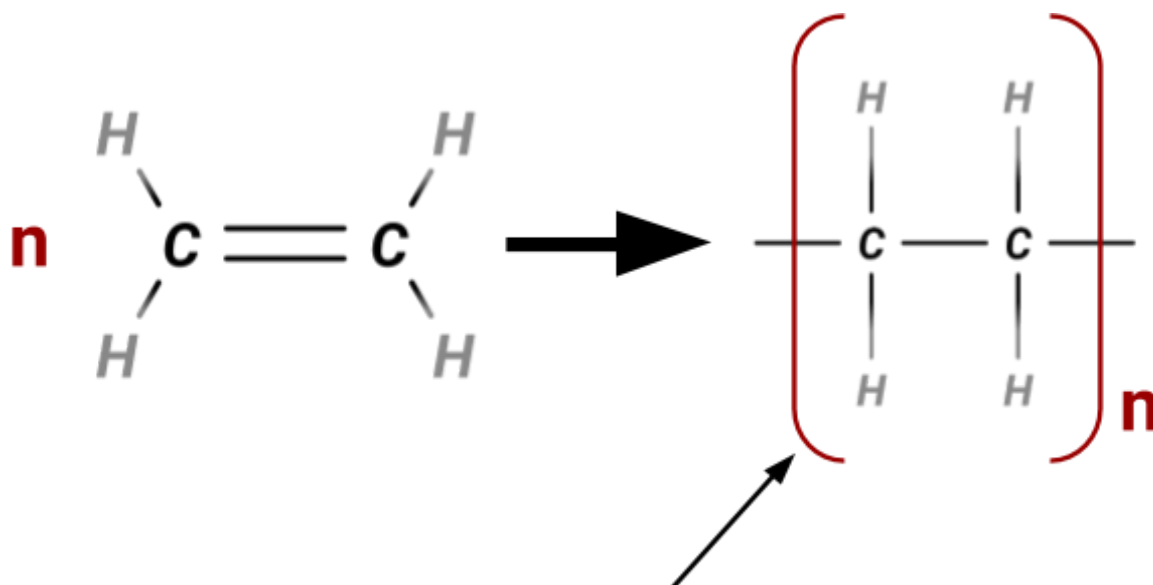


Consider ethene



Atom economy of 100%

To write this as a balanced symbol equation, we use 'n'



This is called a **repeat unit**.

## **5th Form Chemistry 2018/19**

### **The problem with polymers**

The big problem with polymers is that they are non-biodegradable. (Think plastic bags) They are all chemically unreactive. This makes plastic rubbish harmful to wildlife and ecosystems. (Turtles, swans, fish)

So, to dispose of them, we could burn the plastics, using them as fuel in power stations. However, burning plastics produces toxic fumes, including  $\text{CO}_2$ , CO and HCl.

So, we could alternatively recycle them. But this a very expensive process, especially due to the need for sorting and cleaning.

Or we could use chemistry to make life better, by developing biodegradable plastics, developing bacteria to “eat” plastics or by putting filters (such as NaOH) on chimneys to absorb toxic fumes when burning.