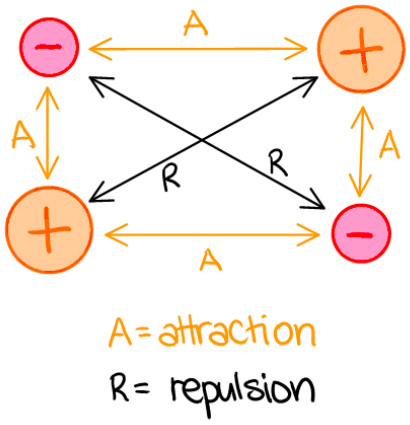
4.2 - Covalent Bonding

4.2.1 - Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively-charged nuclei

Outer shell electrons interact and rearrange themselves into a **more stable arrangement** that has lower chemical energy.

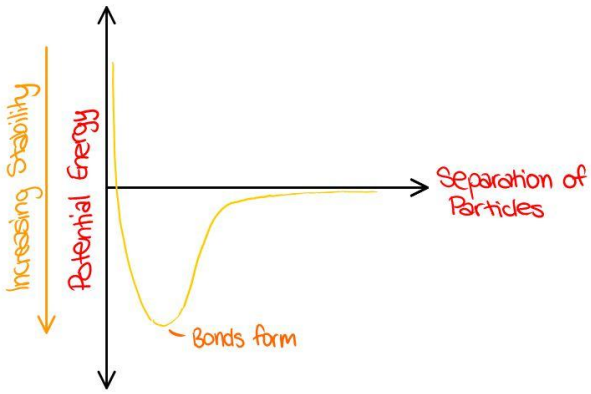


The **positively-charged nucleus** of an atom is attracted to the **negatively charged electrons**. When two atoms come together to form a covalent bond, the nuclei will be attracted to the electron pairs of the other atom.

However, there is **repulsion** between all the electrons, as they have the same charge. The same is true for the nuclei, which also repel each other.

To maintain the covalent bond, a balance must be reached between attraction and repulsion.

A **molecule** can be defined as a discrete group of non-metal atoms covalently bonded to one another. Molecules contain atoms in a set ratio.



4.2.2 - Describe how the covalent bond is formed as a result of electron sharing

There is significant overlap in the radii of the atoms when a molecule is formed.

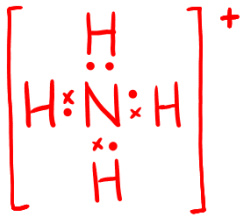
As the two atoms approach each other, electrostatic attractions and repulsions occur between the nuclei and electrons.

Covalent bonds involve the sharing of the electrons. The electrons form **bonding pairs**. When only one pair occupies the space between the electrons, this is a single covalent bond. Any other pairs of valence electrons are called **non-bonding pairs**, or lone pairs. These will help to determine the shape of the molecule, which in turn affects its properties.

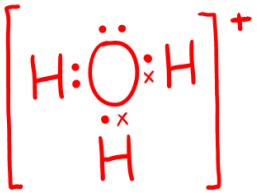
When both the electrons in a bonding pair come from the same atom, they form a **dative covalent bond**, such as in CO, NH_4^+ and H_3O^+

CO

NH_4^+



H_3O^+



This sharing of electrons allows each of the atoms to fill their outer shell.

4.2.3 - Deduce the Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom

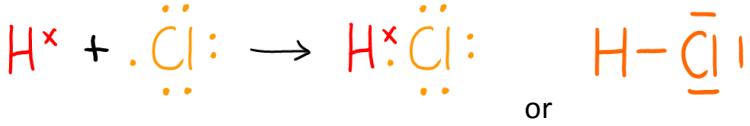
This can be done using dots, crosses or lines

Electron shell diagrams, also called **Lewis or electron dot structures**, can be constructed for covalently bonded molecules. In these, all the valence electrons are drawn, as they form part of the bonding, including the non-bonding electrons. They are used to show how a full outer shell is obtained.

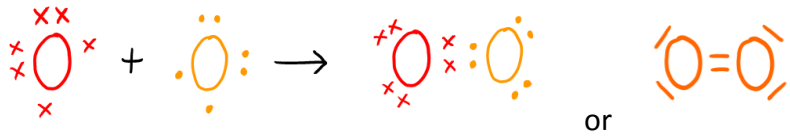
In Lewis structures, the valence electrons are represented by **dots or crosses**. Pairs of dots or crosses represent non-bonding pairs, while a dot and a cross represent a bond.

For diatomic molecules, there are only two atoms bonded to fill the outer shell. In chlorine, which has seven valence electrons, has a single bond

In HCl, only a single bond is required. In both cases, the other six electrons do not take part, as they are non-bonding valence electrons.



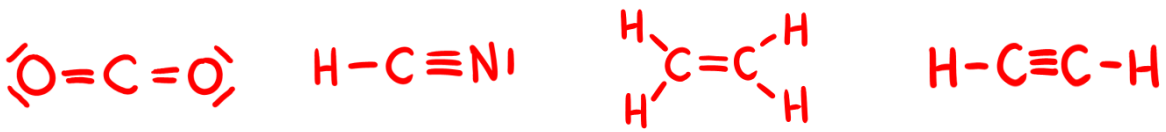
Group six elements have six electrons so they must be sharing another two electrons to have their full outer shell. For the diatomic molecule 0_2, a **double bond** is formed to fill its shell.



Elements in group five need a triple covalent bond to get the eight electrons. This leaves only a single pair of non- bonding electrons.

Other examples include:

CO_2 **HCN** **C_2H_4** **C_2H_2**



Lewis structure can be drawn for any molecules, which become useful in determining their shape, according to this procedure:

1. Determine how many valence electrons there are in each atom in the molecule
2. Find how many electrons are required to fill the valence shell - the number of bonds the atom will form
3. Draw electron dot diagrams, pairing up all the electrons except the number that will be used to form bonds.
    - The atom with the most bonds will be the central atom
4. Arrange the outer atoms around the central atom so that their single dots are near the central atom
    - Pair up single electrons between the central atom and outer atoms to form covalent bonds
5. Each atom in the structure should now have a total of 8 valence electrons, except hydrogen, which has 2.

4.2.4 - State and explain the relationship between the number of bonds, bond length and bond strength

Number of Bonds	Bond Length	Bond Strength
Single bond	Long	Strong
Double bond	Shorter	Stronger
Triple bond	Shortest	Strongest

Bond length decreases as there are more electron pairs involved, causing greater attractive force between the two nuclei.

Bond strength increases because more energy is required to break them.

For example, looking at the bonds between carbon atoms:

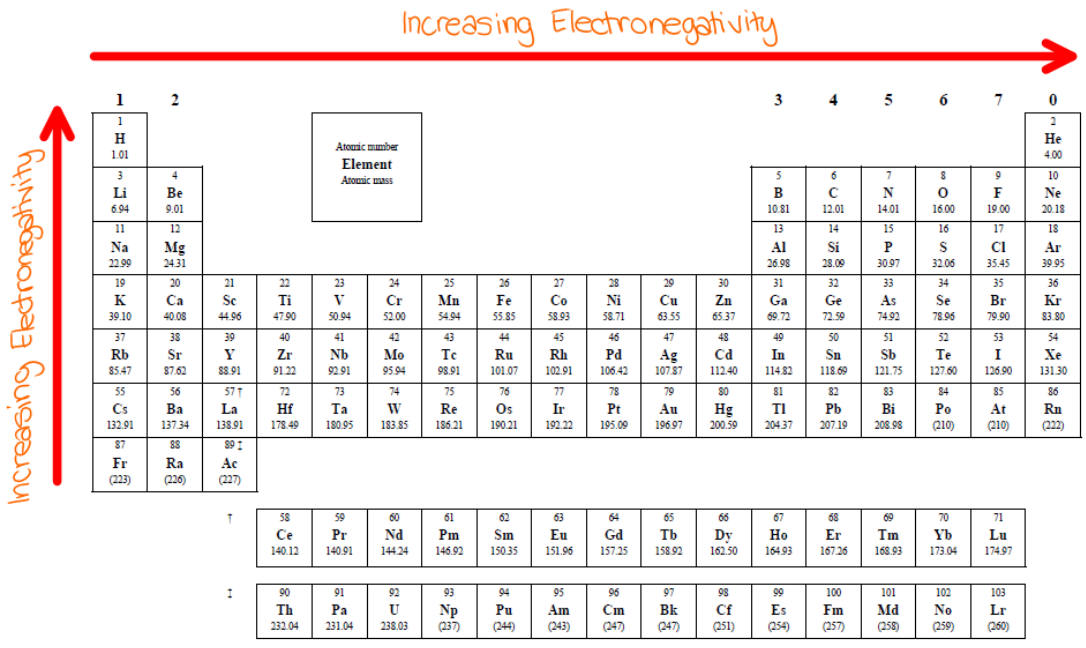
Number of Bonds	**Bond Length (pm)**	**Bond Strength** Dissociation Enthalpy (kJ mol^{-1})
Single bond	154	364
Double bond	134	602
Triple bond	120	835

In carboxylic acid:

Bond	**Bond Length (pm)**	**Bond Strength** Dissociation Enthalpy (kJ mol^{-1})
C=O	120	799
C-O	143	358

4.2.5 - Predict whether a compound of two elements would be covalent from the position of the elements on the periodic table or from their electronegativity values

When two or more different elements are bonded, the sharing of the electrons is not exactly equal. This is because their electronegativity values are different. Electronegativity is a measure of the ability of an atom to attract electron in a bond.

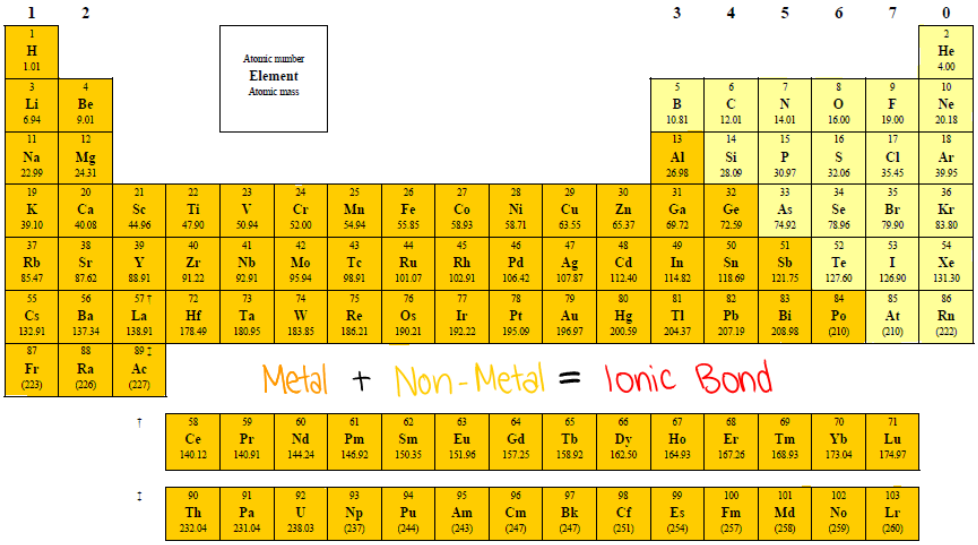


Electronegativity is the highest at the top of the period table, as these have their valence electrons closer to the nuclei. As we move down, the other hand, the valence shell becomes further away. Moving across, the nuclear charge increases, and so does the attraction between the nucleus and valence electrons.

So, the electronegativity **increases as we move up and across** the periodic table. However, the noble gases have an undefined electronegativity, as they already have a full shell.

- Non-metals have higher electronegativity than metals
    - This leads to ionic bonding.

If two non-metals bond, it will be covalent because the electronegativity values are closer together.

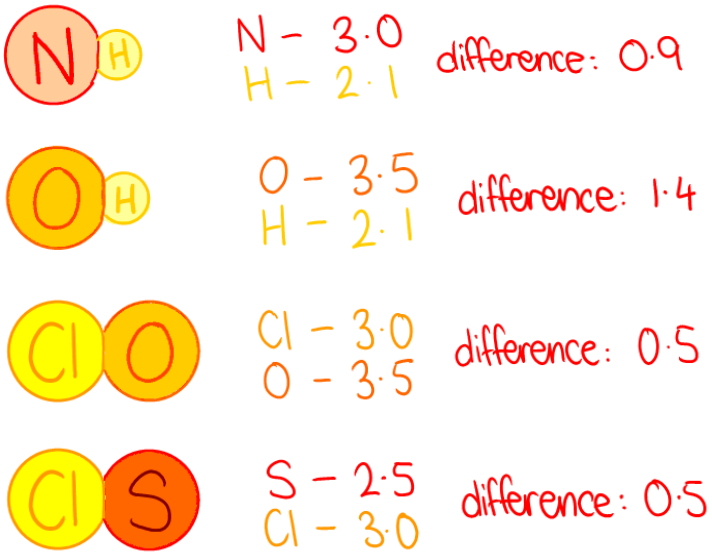


4.2.6 - Predict the relative polarity of bonds from electronegativity values

Every element has a different electronegativity. The polarity of bonds is determined by the difference in electronegativity of the constituent atoms.

Polar molecules have a slight charge on each end.

Difference in Electronegativity	Bond Type
0.0 - 0.4	Non-Polar Covalent
0.5 - 2.0	Polar Covalent
> 2.0	Ionic



4.2.7 - Predict the shape and bond angles for species with four, three and two negative charge centres on the central atom using the valence shell electron pair repulsion theory (VSEPR)

The **structural formula** is the most useful representation of molecules. The actual shape of the molecule is shown, with the electron pair drawn as simple lines, though the non-bonding pairs can also be shown as two dots. The shape has an important role in the chemical and physical properties of the molecules.

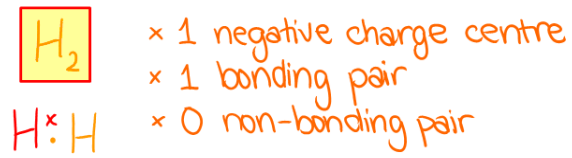
Valence Shell Electron Pair Repulsion theory (VSEPR) is based on the fact that each pair of electrons will be repelled from the others, causing them to move as far away from them as possible in the three-dimensional space. The electrostatic repulsion of pairs determines the geometry of the atoms in the molecule. The space between the electron pairs usually goes in this order

non-bonding/non-bonding > non-bonding/bonding > bonding/bonding

The shape is determined by the number of bonding and non-bonding pairs of electrons on the central atom. Although non-bonding pairs affect the shape, they are not considered when naming the shape of the molecule.

A **negative charge centre** or region refers to pairs of electron on the central atom, including both bonding and non-bonding pairs. Double or triple bonds are still counted as one charge centre.

Diatomic (linear)

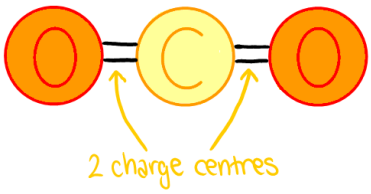


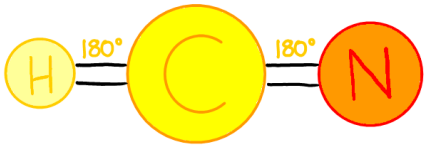
Other examples:

HCl **O_2** **N_2**

Linear

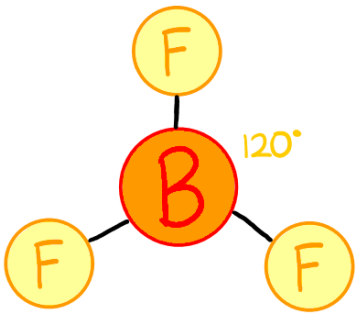
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
2	4	0	180^o

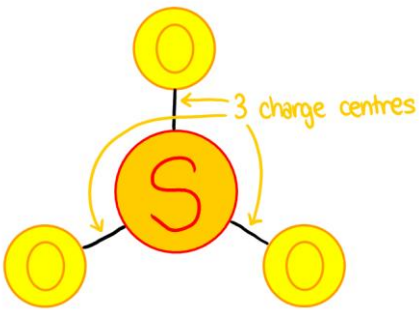




Trigonal Planar

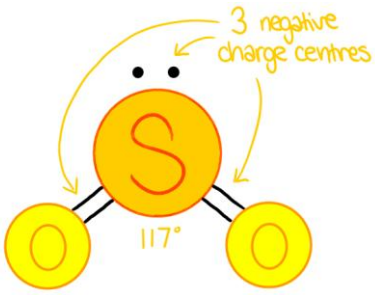
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
3	3	0	120^o





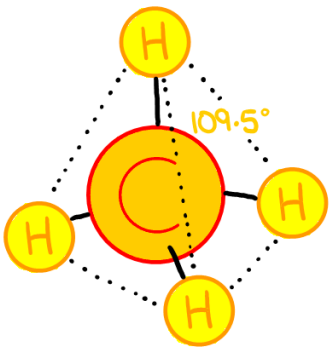
V-Shaped

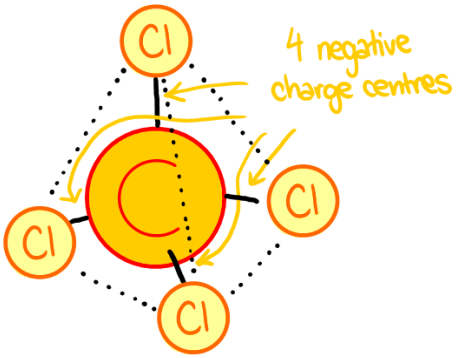
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
3	4	1	117^{o}



Tetrahedral

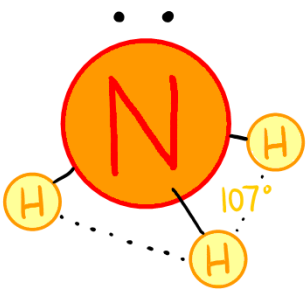
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	4	0	109.5^{o}

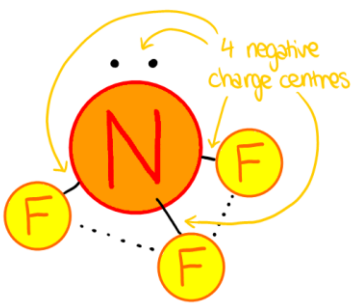




Trigonal Pyramidal

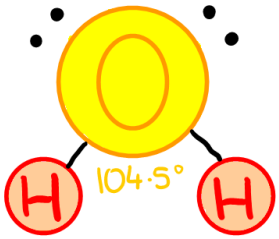
Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	3	1	107^{o}

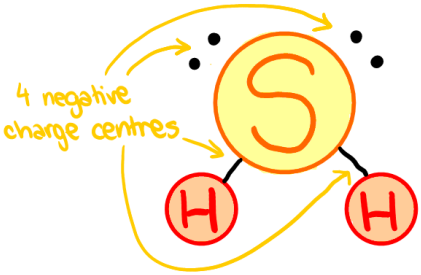




Bent Linear

Negative Charge Centres	Bonding Pairs	Non-Bonding Pairs	Bond Angle
4	2	2	104.5^{o}





4.2.8 - Predict whether or not a molecule is polar from its molecular shape and bond polarities

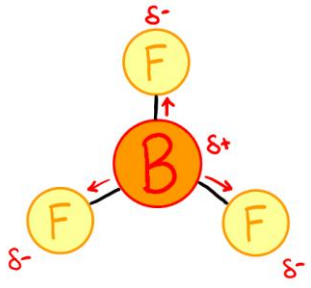
For a molecule to be non-polar, the bonded atoms must be within 0.4 of each other in electronegativity. The electrons would be evenly shared between them. Molecules with **permanent dipoles** that cancel each other out are also non- polar.

If the electrons are shared unevenly, the bond is polar, or a permanent dipole. The symbol delta δ is used to indicate the slight charge that result, especially **δ-** and **δ+**

The electrons are more likely to be found around the atom with the negative dipole.

In **diatomic molecules**, one part will be more negative, as one atom attracts the electrons towards it more.

When there is more than one polar covalent bond, the shape must be considered, as dipoles can cancel each other out. Any molecule that is perfectly symmetrical with be non-polar overall.



4.2.9 - Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and C_{60} fullerene)

The different forms of carbon are known as **allotropes**, and they have vastly different appearance and physical properties. The way they are covalently bonded makes them so different.

As well as the three below, carbon may bond to form

- Carbon nanotubes
- Lonsdaleites
- Chaoite (existence disputed)
- Fullerite

And many others, whose existence has not yet been confirmed

Diamond

In diamond, **all the valence electrons of each carbon atoms are bonded** in a giant covalent network lattice structure. As a result, it cannot conduct electricity. The bonding forces are all very strong, making it very hard. It also has a high sublimation point of 3550^{o}C, at which point it turns straight into a gas.

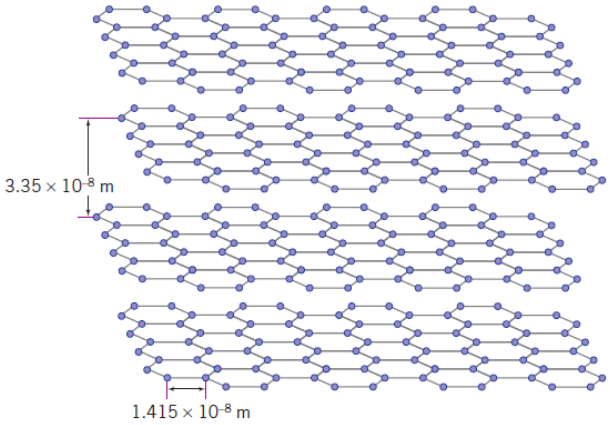
The lattice is very regular, so diamond can be cut in very specific direction known as **cleavage panes**, the lines along which the atoms in the lattice align perfectly. Diamond cutters must identify these panes to optimise the beauty of the cut diamond.

Diamonds are the hardest known naturally occurring substance. Small or flawed stones, known as industrial diamonds, are used on the tips of cutting equipment like drills.

The uses of diamonds include:

- Jewellery
- Cutting tools
- Drills

Graphite



In the structure of graphite, each carbon is **bonded to three others**, forming layers of hexagons. The free electrons explain why it is such a good conductor of electricity. The layers have weak forces between them, making it easy for them to slide past each other. Graphite has a high melting point of 3730°C, so it can still **lubricate** at the high temperatures inside engines and other moving mechanical structures. Its uses include:

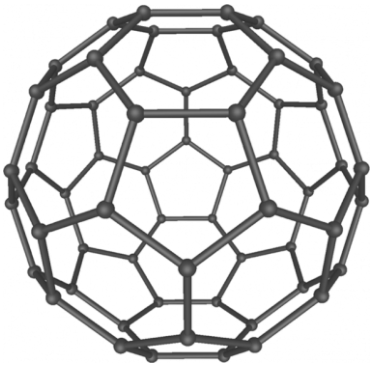
- Good conductor of electricity
- Lubricant
- Additive to rubber and plastic to make them more flexible
- Moderator in nuclear reactors
- High-strength composite material that is light and flexible
    - Tennis racquets
    - Fishing rods
- Pencils
- Electrodes

Fullerenes (Buckyballs)

Buckyballs have a roughly spherical group of covalently bonded carbons arranged in hexagons and pentagons, similar to a soccer ball.

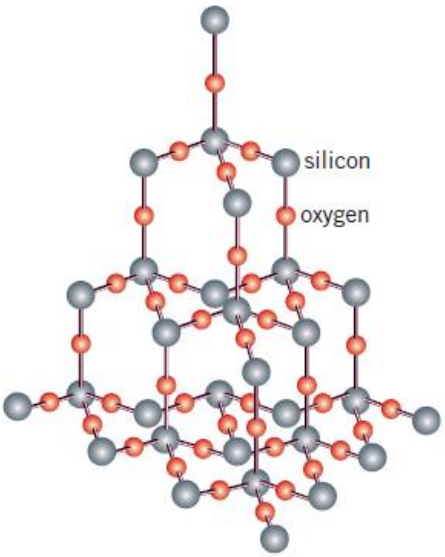
Each carbon atom is bonded to three others, leaving **one free electron per carbon atom**. This enables fullerenes to conduct electricity. They have many uses including

- Superconductivity
- Broad-spectrum lasers
- Catalyst
- Medical uses
- Electromagnetic devices



4.2.10 - Describe the structure of and bonding in silicon and silicon dioxide

Silicon is a member of **group 4** on the periodic table, enabling it to make **four covalent bonds** with other silicon atoms. As silicon is larger than carbon, its bond length is longer. Less energy is required to break a Si-Si bond, so it is more reactive than diamond.



Silicon dioxide (or silica) is a major constituent of sand and is **used to make glass**. It also has a network lattice structure that is made up of alternating silicon and oxygen atoms. It will occur in sandstone, silica sand or quartzite. Its crystal form is quartz. Silica is one of the most abundant oxide materials in the Earth's crust.

Each oxygen atom is covalently bonded to two silicon atoms, whilst each silicon atom is bonded to four oxygen atoms.