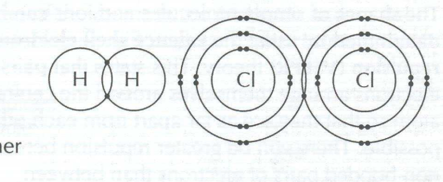
Covalent bonding

SINGLE COVALENT BONDS

Covalent bonding involves the sharing of one or more pairs of electrons so that each atom in the molecule achieves an inert gas configuration. The simplest covalent molecule is hydrogen. Each hydrogen atom has one electron in its outer shell. The two electrons are shared and attracted by both nuclei resulting in a directional bond between the two atoms to form a molecule. When one pair of electrons are shared the resulting bond is known as a single covalent bond. Another example of a diatomic molecule with a single covalent bond is chlorine, Cl_2.

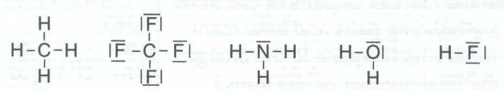


LEWIS STRUCTURES

In the Lewis structure (also known as electron dot structure) all the valence electrons are shown. There are various different methods of depicting the electrons. The simplest method involves using a line to represent one pair of electrons. It is also acceptable to represent single electrons by dots, crosses or a combination of the two. The four methods below are all correct ways of showing the Lewis structure of fluorine.

Sometimes just the shared pairs of electrons are shown, e.g. F–F. This gives information about the bonding in the molecule, but it is not the Lewis structure as it does not show all the valence electrons.

SINGLE COVALENT BONDS

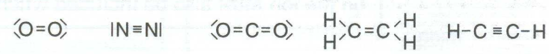


methane | tetrafluoromethane | ammonia | water | hydrogen fluoride

The carbon atom (electronic configuration 2.4) has four electrons in its outer shell and requires a share in four more electrons. It forms four single bonds with elements that only require a share in one more electron, such as hydrogen or chlorine. Nitrogen (2.5) forms three single bonds with hydrogen in ammonia leaving one non-bonded pair of electrons (also known as a lone pair). In water there are two non-bonded pairs and in hydrogen fluoride three non-bonded pairs.

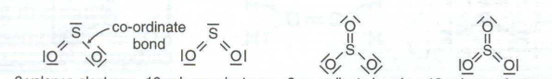
MULTIPLE COVALENT BONDS

In some compounds atoms can share more than one pair of electrons to achieve an inert gas configuration.



oxygen | nitrogen | carbon dioxide | ethene | ethyne

Note: electrons in the shared pair may originate from the same atom. This is known as a co-ordinate covalent bond, but in all other ways it is identical to a normal covalent bond. In some countries stress is laid upon trying to stick to the 'octet' rule so both sulfur dioxide and sulfur trioxide are shown as having a co-ordinate bond between sulfur and oxygen. In other countries sulfur 'expands its octet' to ten or twelve valence electrons and double bonds are shown between the sulfur and the oxygen. Both are acceptable in the International Baccalaureate.



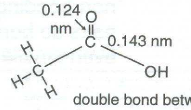
8 valence electrons around S | 10 valence electrons around S | 2 co-ordinate bonds, 8 valence electrons around S | 12 valence electrons around S

BOND LENGTH AND BOND STRENGTH

The strength of attraction that the two nuclei have for the shared electrons affects both the length and strength of the bond. Although there is considerable variation in the bond lengths and strengths of single bonds in different compounds, double bonds are generally much stronger and shorter than single bonds. The strongest covalent bonds are shown by triple bonds.

		Length / nm	Strength / kJ mol^{-1}
Single bonds	Cl–Cl	0.199	242
	C–C	0.154	348
Double bonds	C=C	0.134	612
	O=O	0.121	496
Triple bonds	C≡C	0.120	837
	N≡N	0.110	944

e.g. ethanoic acid:



double bond between C and O shorter and stronger than single bond

BOND POLARITY

In diatomic molecules containing the same element (e.g. H_2 or Cl_2) the electron pair will be shared equally, as both atoms exert an identical attraction. However, when the atoms are different the more electronegative atom exerts a greater attraction for the electron pair. One end of the molecule will thus be more electron rich than the other end, resulting in a polar bond. This relatively small difference in charge is represented by δ+ and δ–. The bigger the difference in electronegativities the more polar the bond.