

Covalent bonding

Searchable typed version of the AS Chemistry (9701) class slides on covalent bonding (syllabus section 3.2). The slides were saved as pictures, so their text could not be searched or copied; this version carries the same text, with the diagrams described. The original slides, with the dot-and-cross diagrams and orbital pictures, are on the same page of megalecture.com.

Syllabus: 3 Chemical bonding

This topic introduces the different ways by which chemical bonding occurs and the effect this can have on physical properties.

3.2 Covalent bonding and co-ordinate (dative covalent) bonding, including shapes of simple molecules. Candidates should be able to:

1. describe, including the use of 'dot-and-cross' diagrams: (i) covalent bonding, in molecules such as hydrogen, oxygen, chlorine, hydrogen chloride, carbon dioxide, methane, ethene; (ii) co-ordinate (dative covalent) bonding, such as in the formation of the ammonium ion and in the Al_2Cl_6 molecule;
2. describe covalent bonding in terms of orbital overlap, giving σ and π bonds, including the concept of hybridisation to form sp , sp^2 and sp^3 orbitals (see also Section 14.3);
3. explain the shapes of, and bond angles in, molecules by using the qualitative model of electron-pair repulsion (including lone pairs), using as simple examples: BF_3 (trigonal), CO_2 (linear), CH_4 (tetrahedral), NH_3 (pyramidal), H_2O (non-linear), SF_6 (octahedral), PF_5 (trigonal bipyramidal);
4. predict the shapes of, and bond angles in, molecules and ions analogous to those specified in 3.2(b) (see also Section 14.3).

(Outcomes c and d are greyed out on the slide: they are covered in the shapes-of-molecules notes.)

Slide 1: Chemical bonding

When two or more atoms form a chemical compound, the atoms are held together in a characteristic arrangement by attractive forces.

The chemical bond is the force of attraction between any two atoms in a compound. The attraction is the force that overcomes the repulsion of the positively charged nuclei of the two atoms.

Interactions involving valence electrons are responsible for the chemical bond. We shall focus our attention on these electrons and the electron arrangement of atoms both before and after bond formation.

Slide 2: Covalent bonding

When electrons are shared rather than transferred, the shared electron pair is referred to as a covalent bond.

Covalent bonds tend to form between atoms with similar tendencies to gain or lose electrons. The most obvious examples are the diatomic molecules H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 and I_2 .



(Diagrams: $F + F \rightarrow F-F$ with one shared pair; $O + O \rightarrow O=O$ with two shared pairs; $N + N \rightarrow N \equiv N$ with three shared pairs; $H + F \rightarrow H-F$.)

Slide 3: Covalent bonding

(Diagram: two overlapping atoms; the shared electrons sit between the two protons, attracted to both.)

A covalent bond is the electrostatic force of attraction between the positively charged nuclei of both atoms and their shared pair(s) of electrons.

(Diagram: the attraction between the nuclei and the electrons in the bond is greatest for a triple bond, then a double bond, then a single bond.)

Slide 4: 'Dot-and-cross' diagrams

Dot-and-cross diagrams and displayed formulae are shown for: chlorine $Cl-Cl$, water $H-O-H$, ammonia $H-N(H)-H$, oxygen $O=O$, carbon dioxide $O=C=O$, ethene $H_2C=CH_2$, nitrogen $N \equiv N$, and ethyne $H-C \equiv C-H$.

Slide 5: Covalent bonding

The bond is the force of attraction between the pair of electrons and the two positive nuclei of the atoms.

This increased negative charge in the centre holds the two positively charged nuclei together, thus forming the bond.

Depending on the number of electrons involved, the bonding is classified as single, double and triple: $H-H$, $H-Cl$, $O=O$, $N \equiv N$.

Slide 6: Covalent compounds

Displayed formulae of methane (C with four H), ammonia (N with three H and one lone pair) and water (O with two H and two lone pairs).

Slide 7: Skill check 1

Draw dot-and-cross diagrams of PCl_3 , N_2 , CS_2 , C_2H_4 , NH_3 , H_2O_2 , N_2H_4 , SO_3 .

Slide 8: Examples of molecules

(Diagrams of covalently bonded molecules: a puckered ring of eight S atoms (S_8); a tetrahedron of four P atoms (P_4); the giant covalent lattice of silicon, each Si atom bonded to four others.)

Slide 9: Skill check 2

When barium metal burns in oxygen, the ionic compound barium peroxide, BaO_2 , is formed. Which dot-and-cross diagram represents the electronic structure of the peroxide anion in BaO_2 ? (Four diagrams A to D of two oxygen atoms with electrons marked from the first oxygen, the second oxygen

and the barium atom.)

Slide 10: Skill check 3

Dicarbon monoxide, C_2O , is found in dust clouds in space. Analysis of it shows that the sequence of atoms in this molecule is $C-C-O$. All bonds are double bonds and there are no unpaired electrons. How many lone pairs of electrons are present in a molecule of C_2O ? A 1 B 2 C 3 D 4

Slide 11: Covalent bonding

Atoms share electrons to get the nearest noble gas electronic configuration.

- Some don't achieve an "octet" as they haven't got enough electrons, e.g. Al in $AlCl_3$.
- Others share only some; if they share all they will exceed their "octet", e.g. NH_3 and H_2O .
- Atoms of elements in the 3rd period onwards can exceed their "octet" if they wish, as they are not restricted to eight electrons in their "outer shell", e.g. PCl_5 , SO_2 , SO_3 and SF_6 .

Only in period 2 are the elements restricted to form an octet. However, in period 3, more than 8 electrons can be taken in the outermost shell due to the s, p and d orbitals, which take up 2, 6 and 10 electrons respectively.

Slides 12 and 13: Defying the octet

(Dot-and-cross diagrams: boron trifluoride, $3 F + B \rightarrow BF_3$, boron with only six electrons around it, trigonal shape; sulfur hexafluoride, $6 F + S \rightarrow SF_6$, sulfur with twelve electrons around it, octahedral shape; PF_3 and PF_5 ; ClF_5 .)

Slide 14: Molecular orbitals

Covalent bonding is brought about by the atomic orbitals of each atom overlapping with each other, coalescing and forming a molecular orbital.

Only individually filled orbitals take part in such bonding, so that the resultant molecular orbital has only two electrons. In the molecular orbital the two electrons circulate around and are attracted to two nuclei of the two bonded atoms.

(Diagram: atomic orbital + atomic orbital $\rightarrow$ molecular orbital; an orbital containing 1 electron plus another containing 1 electron gives 2 electrons paired up in a molecular orbital.)

Slide 15: Orbital overlap

Covalent bonds are formed when orbitals, each containing one electron, overlap. This forms a region in space where an electron pair can be found; new molecular orbitals are formed.

(Diagram: an orbital containing 1 electron and another orbital containing 1 electron overlap to provide a region in space which can contain a pair of electrons: the molecular orbital.)

The space that these shared electrons move within is called a molecular orbital. A molecular orbital is made when two atomic orbitals overlap.

Slides 16 and 17: Sigma bonds

A sigma bond (σ bond) is formed when orbitals overlap head-on in a covalent bond.

In a sigma bond, the electron density is concentrated in the orbital overlap volume between the two nuclei.

A sigma bond can be rotated without breaking the bond.

The region where the two orbitals overlap is the molecular orbital which contains electrons of the sigma bond.

Sigma bonds can form by overlap of two s-orbitals, an s-orbital and a p-orbital, or two p-orbitals. (Diagrams: $1s + 1s \rightarrow \text{H-H}$; $1s + 3p \rightarrow \text{H-Cl}$; $3p + 3p \rightarrow \text{Cl-Cl}$.)

Slides 18 to 21: Pi bonds

A pi bond (π bond) is formed when orbitals overlap sideways in a covalent bond. Pi bonds are found in molecules with double and triple bonds.

(Diagram: two 2p orbitals on neighbouring O atoms overlap sideways, above and below the line joining the nuclei.)

A single π bond is drawn as two electron clouds, one arising from each lobe of the p orbitals. The two clouds of electrons in a π bond represent one bond consisting of a total of two electrons. (Diagrams: 2p + 2p orbitals on two C atoms interact and overlap in two regions, giving the π bond; the end view shows a single p orbital.)

In a pi bond, the electron density is concentrated in the orbital overlap volumes above and below the line joining the two nuclei.

A pi bond cannot be rotated without breaking the bond. (Diagram: rotation is possible around the single C–C bond of ethane; no rotation is possible about the C=C bond of ethene.)

Slide 22: Skill check 4

What is always involved in a carbon–carbon π bond? A a shared pair of electrons B a sideways overlap of p orbitals C delocalised electrons

Slide 23: Double bonds

A double bond consists of one sigma bond and one pi bond, for example the C=C double bond in ethene and the O=O double bond in oxygen.

A typical triple bond, for example in nitrogen, consists of one sigma bond and two pi bonds in two mutually perpendicular planes.

Pi bonds are weaker than sigma bonds.

Slides 24 to 26: Multiple bonds: ethene; ethane and ethene

An example of a pi bond is the C=C bond in ethene. This bond consists of two parts. One part consists of two p orbitals of carbon overlapping in a sigma bond. In the other part of the bond, two p orbitals

overlap in a pi bond.

(Diagrams: the σ framework of ethene, C–C and four C–H σ bonds; the side-on overlap of the remaining p orbitals above and below the σ bond to give the π bond. Ethane: all σ bonds, H–C–H angles 109.5° . Ethene: p orbitals on each carbon overlap to form the π bond, H–C–H angles 120° .)

Slides 27 and 28: Triple bond: ethyne and nitrogen

(Diagrams: in ethyne, H–C $\equiv$ C–H, each carbon's p_y and p_z orbitals overlap sideways to give π_y and π_z , two π bonds at right angles around the C–C σ bond. In nitrogen, N $\equiv$ N, the same $p_y + p_z$ overlap gives $\pi_y + \pi_z$, with a lone pair on each N.)

Slide 29: Two doubles: carbon dioxide

(Diagram: in O=C=O the carbon's p_y and p_z orbitals overlap with a p orbital on each oxygen, giving one π_z bond to one oxygen and one π_y bond to the other.)

Slide 30: Skill check 5

Which statements about covalent bonds are correct? A A triple bond consists of one π bond and two σ bonds. B The electron density in a σ bond is highest along the axis between the two bonded atoms. C A π bond restricts rotation about the σ bond axis.

Slides 31 to 34: Dative covalent bonding

A dative covalent bond (or co-ordinate covalent bond) is a covalent bond between two atoms in which the shared electrons are contributed by only one of the atoms.

e.g. boron trifluoride and ammonia, NH₃BF₃. Boron has an incomplete shell in BF₃ and can accept/share a pair of electrons donated by ammonia.

(Diagrams: H₃N: + BF₃ → H₃N→BF₃, the arrow showing the dative bond from N to B; H₃N: + H⁺ → [NH₄]⁺; H₂O: + H⁺ → [H₃O]⁺; carbon monoxide, C $\equiv$ O with one of the three shared pairs donated by oxygen.)

Slides 35 and 36: Al₂Cl₆

(Diagrams: two AlCl₃ molecules join to form Al₂Cl₆: a lone pair on a chlorine of each AlCl₃ forms a dative bond to the aluminium of the other, giving a bridged structure with two chlorine atoms shared between the two aluminium atoms; shown both as a dot-and-cross diagram and with arrows for the two dative bonds.)

Slide 37: Dative covalent bonding

A dative covalent bond differs from a covalent bond only in its formation.

Both electrons of the shared pair are provided by one species (donor) and it shares the electrons with the acceptor.

Donor species will have lone pairs in their outer shells. Acceptor species will be short of their "octet" or

maximum.

Chemists call it a dative covalent bond because the word dative means 'giving' and one atom gives both the electrons to make the bond. Once formed, there is no difference between a dative bond and any other covalent bond.

Slides 38 and 39: Bond energy

The strength of a covalent bond is measured by the bond energy.

The bond energy is the amount of energy required to break a covalent bond, per mole of bonds. The greater the bond energy, the stronger the bond.

Very large bond energies can make molecules unreactive. The nitrogen molecule (N_2) is unreactive because of the very large $\text{N}\equiv\text{N}$ bond energy of 944 kJ mol^{-1} .

Bond forming releases energy, and is therefore exothermic. Bond breaking requires energy, and is therefore endothermic. For the same two atoms, triple bonds are stronger than double bonds, which in turn are stronger than single bonds.

Bond	Bond energy (kJ mol^{-1})
C–C	350
C=C	610
C≡C	840

Slide 40: Bond length

The covalent bond length is the distance between the nuclei of the two atoms linked by one or more covalent bonds.

For the same two atoms, triple bonds are shorter than double bonds, which in turn are shorter than single bonds.

(Diagram: (a) two nuclei close together with the shared electrons between them, greater attraction; (b) the nuclei further apart, less attraction.)

Checked by Mega Lecture

This typed version carries the text of all 40 slides. Diagrams (dot-and-cross diagrams, orbital pictures, molecule models) are described in words and are not reproduced; use the original slides for those. One unit corrected: " 944 KJ mol^{-1} " is written kJ mol^{-1} . Spotted a mistake? Tell us on WhatsApp: +92 336 780 1123.