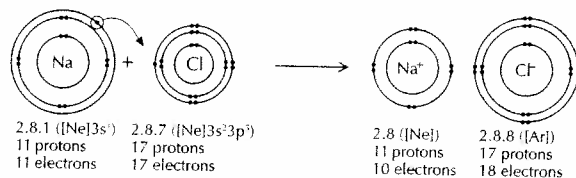# Ionic bonding

IONIC BOND

When atoms combine they do so by trying to achieve an inert gas configuration. Ionic compounds are formed when electrons are transferred from one atom to another to form ions with complete outer shells of electrons. In an ionic compound the positive and negative ions are attracted to each other by strong electrostatic forces, and build up into a strong lattice. Ionic compounds have high melting points as considerable energy is required to overcome these forces of attraction.

The classic example of an ionic compound is sodium chloride Na^+Cl^-, formed when sodium metal burns in chlorine. Chlorine is a covalent molecule, so each atom already has an inert gas configuration. However, the energy given out when the ionic lattice is formed is sufficient to break the bond in the chlorine molecule to give atoms of chlorine. Each sodium atom then transfers one electron to a chlorine atom to form the ions.

2.8.1 ([Ne]3s¹)   2.8.7 ([Ne]3s²3p⁵)   2.8 ([Ne])   2.8.8 ([Ar])
11 protons   17 protons   11 protons   17 protons
11 electrons   17 electrons   10 electrons   18 electrons

The charge carried by an ion depends on the number of electrons the atom needed to lose or gain to achieve a full outer shell.

Cations			Anions		
Group 1	**Group 2**	**Group 3**	**Group 5**	**Group 6**	**Group 7**
+1	+2	+3	–3	–2	–1
Li^+ Na^+ K^+	Mg^{2+} Ca^{2+}	Al^{3+}	N^{3-} P^{3-}	O^{2-} S^{2-}	F^- Cl^- Br^-

Thus in magnesium chloride two chlorine atoms each gain one electron from a magnesium atom to form $Mg^{2+}Cl^-_2$. In magnesium oxide two electrons are transferred from magnesium to oxygen to give $Mg^{2+}O^{2-}$. Transition metals can form more than one ion. For example, iron can form Fe^{2+} and Fe^{3+} and copper can form Cu^+ and Cu^{2+}.

FORMULAS OF IONIC COMPOUNDS

It is easy to obtain the correct formula as the overall charge of the compound must be zero.

lithium fluoride Li^+F^- magnesium chloride $Mg^{2+}Cl^-_2$ aluminium bromide $Al^{3+}Br^-_3$
sodium oxide $Na^+_2O^{2-}$ calcium sulfide $Ca^{2+}S^{2-}$ iron(III) oxide $Fe^{3+}_2O^{2-}_3$
potassium nitride $K^+_3N^{3-}$ calcium phosphide $Ca^{2+}_3P^{3-}_2$ iron(II) oxide $Fe^{2+}O^{2-}$

Note: the formulas above have been written to show the charges carried by the ions. Unless asked specifically to do this it is common practice to omit the charges and simply write LiF, $MgCl_2$, etc.

IONS CONTAINING MORE THAN ONE ELEMENT

In ions formed from more than one element the charge is often spread (delocalized) over the whole ion. An example of a positive ion is the ammonium ion NH_4^+, in which all four N–H bonds are identical. Negative ions are sometimes known as acid radicals as they are formed when an acid loses one or more H^+ ions.

hydroxide OH^-
nitrate NO_3^- (from nitric acid, HNO_3)
sulfate SO_4^{2-}
hydrogensulfate HSO_4^- { from sulfuric acid, H_2SO_4 }

carbonate CO_3^{2-}
hydrogencarbonate HCO_3^- { from carbonic acid, H_2CO_3 }
ethanoate CH_3COO^- (from ethanoic acid, CH_3COOH)

The formulas of the ionic compounds are obtained in exactly the same way. Note: brackets are used to show that the subscript covers all the elements in the ion.

sodium nitrate $Na^+NO_3^-$ calcium carbonate $Ca^{2+}CO_3^{2-}$
ammonium sulphate $(NH_4^+)_2SO_4^{2-}$ magnesium ethanoate $Mg^{2+}(CH_3COO^-)_2$ aluminium hydroxide $Al^{3+}(OH^-)_3$

IONIC OR COVALENT?

Ionic compounds are formed between metals on the left of the Periodic Table and non-metals on the right of the Periodic Table; that is, between elements in groups 1, 2, and 3 with a low electronegativity (electropositive elements) and elements with a high electronegativity in groups 5, 6, and 7. Generally the difference between the electronegativity values needs to be greater than about 1.8 for ionic bonding to occur.

	Al	F	Al	O	Al	Cl	Al	Br
Electronegativity	1.5	4.0	1.5	3.5	1.5	3.0	1.5	2.8
Difference in electronegativity		2.5		2.0		1.5		1.3
Formula	AlF_3		Al_2O_3		Al_2Cl_6		Al_2Br_6	
Type of bonding	ionic		ionic		intermediate between ionic and covalent		covalent	
M. pt / °C	1265		2050		Sublimes at 180		97	

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 electrostatically by both positive nuclei resulting in a directional bond between the two atoms to form a molecule. When one pair of electrons is 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.
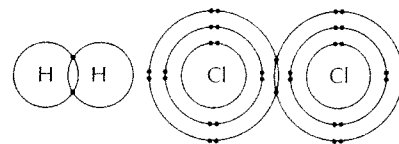
LEWIS STRUCTURES

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.
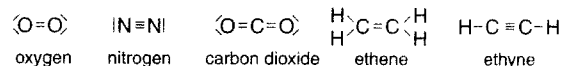
SINGLE COVALENT BONDS

H	$\mathsf{\bar{F}}$			
H–C–H	F–C–F	H–N–H	H–O	H–F
H	F	H	H	
methane	tetrafluoro-methane	ammonia	water	hydrogen fluoride
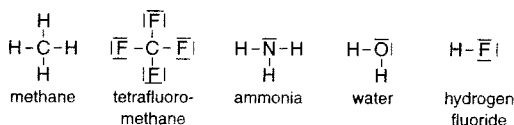
The carbon atom (electronic configuration 2.4) has four electrons in its outer shell and requires a share in four more electrons. It forms four single bonds with elements that only require a share in one more electron, such as hydrogen or chlorine. Nitrogen (2.5) forms three single bonds with hydrogen 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.
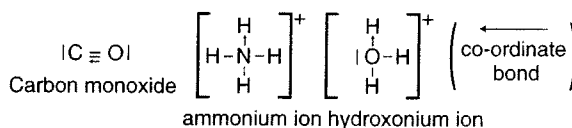
BOND LENGTH AND BOND STRENGTH

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

MULTIPLE COVALENT BONDS

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

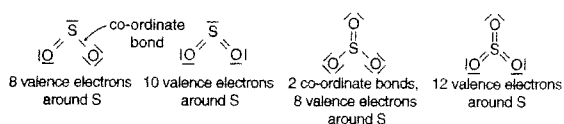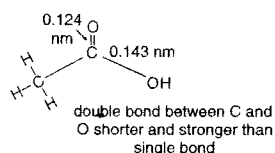| O=O | N≡N | O=C=O | H₂C=CH₂ | H–C≡C–H |
| oxygen | nitrogen | carbon dioxide | ethene | ethyne |

CO-ORDINATE (DATIVE) BONDS

The electrons in the shared pair may originate from the same atom. This is known as a coordinate covalent bond.

C≡O
Carbon monoxide

ammonium ion   hydroxonium ion

Sulfur dioxide and sulfur trioxide are both sometimes shown as having a coordinate bond between sulfur and oxygen or they are shown as having double bonds between the sulfur and the oxygen. Both are acceptable.

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 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.
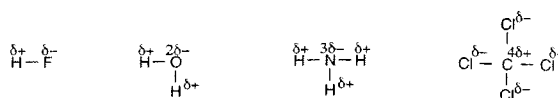
Shapes of simple molecules and ions

VSEPR THEORY

The shapes of simple molecules and ions can be determined by using the **valence shell electron pair repulsion (VSEPR)** theory. This states that pairs of electrons arrange themselves around the central atom so that they are as far apart from each other as possible. There will be greater repulsion between non-bonded pairs of electrons than between bonded pairs. Since all the electrons in a multiple bond must lie in the same direction, double and triple bonds count as one pair of electrons. Strictly speaking the theory refers to negative charge centres, but for most molecules this equates to pairs of electrons.

This results in five basic shapes depending on the number of pairs.

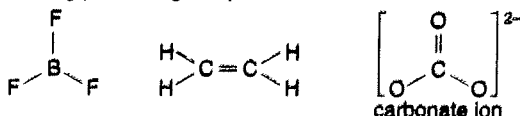| No. of charge centres	Shape	Name of shape	Bond angle(s)
2		linear	180°
3		trigonal planar	120°
4		tetrahedral	109.5°
5		trigonal bipyramidal	90°, 120°, 180°
6		octahedral	90°, 180°

WORKING OUT THE ACTUAL SHAPE

To work out the actual shape of a molecule calculate the number of pairs of electrons around the *central* atom, then work out how many are bonding pairs and how many are non-bonding pairs. (For ions the number of electrons which equate to the charge on the ion must also be included when calculating the total number of electrons.)

2 NEGATIVE CHARGE CENTRES

Cl—Be—Cl   O=C=O   H—C≡C—H   H—C≡N

double bond counts as one pair

triple bond counts as one pair
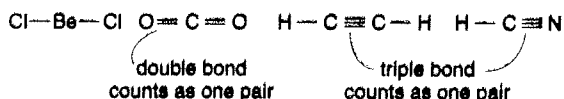
3 NEGATIVE CHARGE CENTRES

3 bonding pairs – trigonal planar

carbonate ion
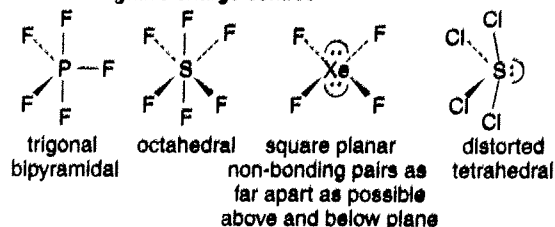
2 bonding pairs, 1 non-bonded pair – bent or V-shaped

nitrite ion

5 AND 6 NEGATIVE CHARGE CENTRES
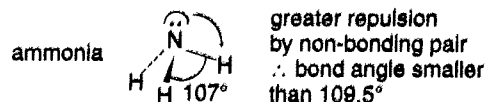
5 and 6 negative charge centres

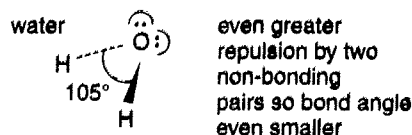trigonal bipyramidal   octahedral   square planar non-bonding pairs as far apart as possible above and below plane   distorted tetrahedral

4 NEGATIVE CHARGE CENTRES

4 bonding pairs – tetrahedral

ammonium ion   tetrafluoroborate ion
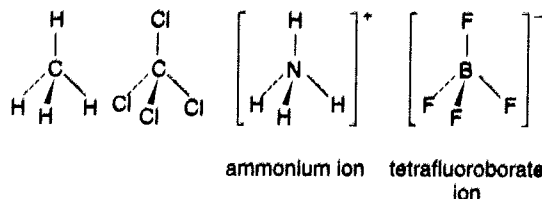
3 bonding pairs, 1 non-bonding pair – trigonal pyramid

ammonia

greater repulsion by non-bonding pair
∴ bond angle smaller than 109.5°

H H 107°

2 bonding pairs, 2 non-bonding pairs – bent or V-shaped

water

105°

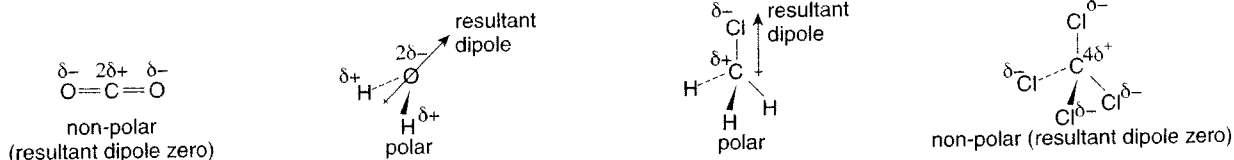even greater repulsion by two non-bonding pairs so bond angle even smaller

Intermolecular forces and allotropes of carbon

MOLECULAR POLARITY
Whether a molecule is polar, or not, depends both on the relative electronegativities of the atoms in the molecule and on its shape. If the individual bonds are polar then it does not necessarily follow that the molecule will be polar as the resultant dipole may cancel out all the individual dipoles.
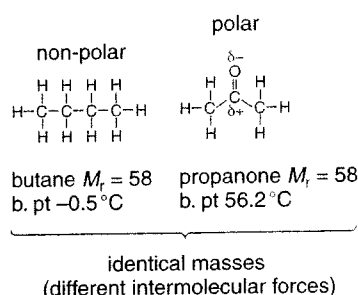
$$\delta- \quad 2\delta+ \quad \delta-$$
$$O = C = O$$
non-polar
(resultant dipole zero)

resultant dipole

$\delta+$ H $\quad$ $2\delta-$ $\emptyset$ $\quad$ $H^{\delta+}$
polar

resultant dipole

$\delta-$ Cl $\quad$ $\delta+$ C $\quad$ H $\quad$ H $\quad$ H
polar

$\delta-$ Cl $\quad$ $4\delta+$ C $\quad$ $\delta-Cl$ $\quad$ $Cl^{\delta-}$ $\quad$ $Cl^{\delta-}$
non-polar (resultant dipole zero)

Van der Waals' forces
Even in non-polar molecules the electrons can at any one moment be unevenly spread. This produces temporary instantaneous dipoles. An instantaneous dipole can induce another dipole in a neighbouring particle resulting in a weak attraction between the two particles. Van der Waals' forces increase with increasing mass.

increasing van der Waals' forces →

	F_2	Cl_2	Br_2	I_2
M_r	38.0	70.9	160	254
b. pt / °C	−188	−34.0	58.0	183

increasing van der Waals' forces →

	CH_4	C_2H_6	C_3H_8	C_4H_{10}
M_r	16.0	30.0	44.0	58.0
b. pt / °C	−162	−88.6	−42.2	−0.5

Dipole:dipole forces
Polar molecules are attracted to each other by electrostatic forces. Although still relatively weak the attraction is stronger than van der Waals' forces.

non-polar

polar

butane $M_r = 58$
b. pt −0.5 °C

propanone $M_r = 58$
b. pt 56.2 °C

identical masses
(different intermolecular forces)

INTERMOLECULAR FORCES
The covalent bonds between the atoms *within* a molecule are very strong. The forces of attraction *between* the molecules are much weaker. These intermolecular forces depend on the polarity of the molecules.

Hydrogen bonding
Hydrogen bonding occurs when hydrogen is bonded directly to a small highly electronegative element, such as fluorine, oxygen, or nitrogen. As the electron pair is drawn away from the hydrogen atom by the electronegative element, all that remains is the proton in the nucleus as there are no inner electrons. The proton attracts a non-bonding pair of electrons from the F, N, or O resulting in a much stronger dipole:dipole attraction. Water has a much higher boiling point than the other group 6 hydrides as the hydrogen bonding between water molecules is much stronger than the dipole:dipole bonding in the remaining hydrides. A similar trend is seen in the hydrides of group 5 and group 7. Hydrogen bonds between the molecules in ice result in a very open structure. When ice melts the molecules can move closer to each other so that water has its maximum density at 4 °C.

boiling temperature / K

the ice lattice
≈ = hydrogen bond
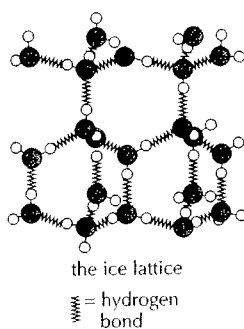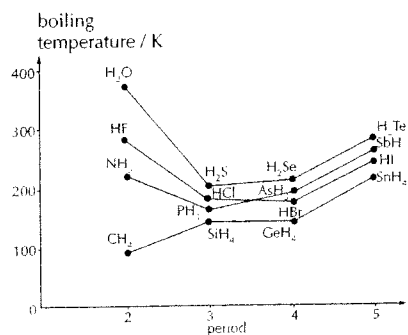
ALLOTROPES OF CARBON
Allotropes occur when an element can exists in different crystalline forms. In diamond each carbon atom is covalently bonded to four other carbon atoms to form a giant covalent structure. All the bonds are equally strong and there is no plane of weakness in the molecule so diamond is exceptionally hard and because all the electrons are localized it does not conduct electricity. Both silicon and silicon dioxide, $SiO2$, form similar giant tetrahedral structures.

diamond
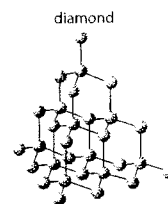
In graphite each carbon atom has very strong bonds to three other carbon atoms to give layers of hexagonal rings. There are only very weak bonds between the layers. The layers can slide over each other so graphite is an excellent lubricant and because the electrons are delocalized between the layers it is a good conductor of electricity.

graphite
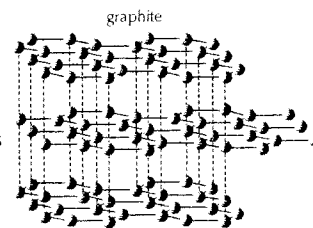
A third allotrope of carbon is buckminsterfullerene. Sixty carbon atoms are arranged in hexagons and pentagons to give a geodesic spherical structure similar to a football. Following the initial discovery of buckminsterfullerene many other similar carbon molecules have been isolated. This has led to a new branch of science called nanotechnology.
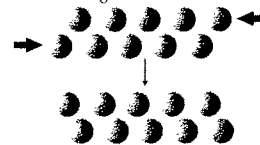
Metallic bonding and physical properties related to bonding type

METALLIC BONDING

The valence electrons in metals become detached from the individual atoms so that metals consist of a close packed lattice of positive ions in a sea of delocalized electrons. A metallic bond is the attraction that two neighbouring positive ions have for the delocalized electrons between them. Metals are malleable, that is, they can be bent and reshaped under pressure. They are also ductile, which means they can be drawn out into a wire.

delocalized electrons

nucleus and inner shells
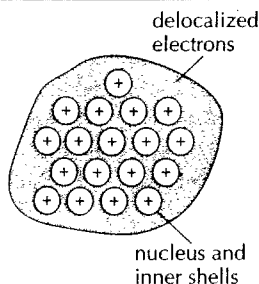
Metals are malleable and ductile because the close-packed layers of positive ions can slide over each other without breaking more bonds than are made.

Impurities added to the metal disturb the lattice and so make the metal less malleable and ductile. This is why alloys are harder than the pure metals they are made from.

TYPE OF BONDING AND PHYSICAL PROPERTIES

Melting and boiling points

When a liquid turns into a gas the attractive forces between the particles are completely broken so boiling point is a good indication of the strength of intermolecular forces. When solids melt the crystal structure is broken down, but there are still some attractive forces between the particles. Melting points are affected by impurities. These weaken the structure and result in lower melting points.

Covalent macromolecular structures have extremely high melting and boiling points. Metals and ionic compounds also tend to have relatively high boiling points due to ionic attractions. Hydrogen bonds are in the order of $\frac{1}{10}$ th the strength of a covalent bond whereas van der Waals' forces are in the order of less than $\frac{1}{100}$ of a covalent bond. The weaker the attractive forces the more volatile the substance.

diamond lattice

NaCl
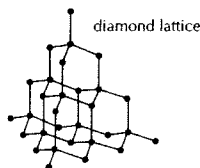
Diamond (melting point over 4000 °C)
All bonds in the macromolecular structure covalent

Sodium chloride (melting point 801 °C)
Ions held strongly in ionic lattice
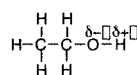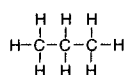
Compound	propane	ethanal	ethanol
M_r	44	44	46
M. pt / °C	−42.2	20.8	78.5
Polarity	non-polar	polar	polar
Bonding type	van der Waals'	dipole:dipole	hydrogen bonding

Solubility

'Like tends to dissolve like'. Polar substances tend to dissolve in polar solvents, such as water, whereas non-polar substances tend to dissolve in non-polar solvents, such as heptane or tetrachloromethane. Organic molecules often contain a polar head and a non-polar carbon chain tail. As the non-polar carbon chain length increases in an homologous series the molecules become less soluble in water. Ethanol itself is a good solvent for other substances as it contains both polar and non-polar ends.

CH_3OH
C_2H_5OH decreasing
C_3H_7OH solubility in water
C_4H_9OH
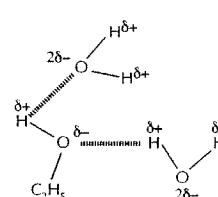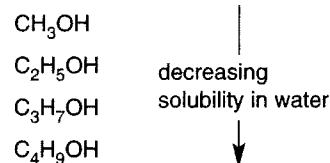
Ethanol is completely miscible with water as it can hydrogen-bond to water molecules.

Conductivity

For conductivity to occur the substance must possess electrons or ions that are free to move. Metals (and graphite) contain delocalized electrons and are excellent conductors. Molten ionic salts also conduct electricity, but are chemically decomposed in the process. Where all the electrons are held in fixed positions, such as diamond or in simple molecules, no electrical conductivity occurs.

When a potential gradient is applied to the metal, the delocalized electrons can move towards the positive end of the gradient carrying charge.
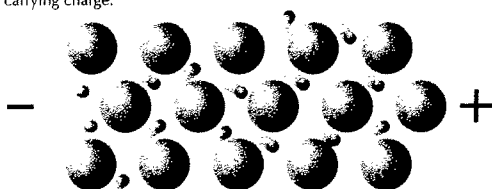
When an ionic compound melts, the ions are free to move to oppositely charged electrodes. Note: in molten ionic compounds it is the ions that carry the charge, not free electrons.

NaCl   heat
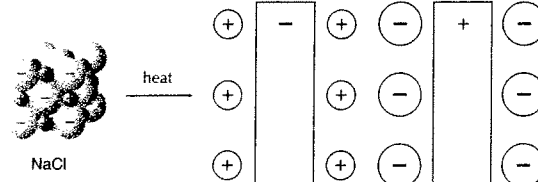
ⓗⓛ Molecular orbitals and hybridization (1)

COMBINATION OF ATOMIC ORBITALS TO FORM MOLECULAR ORBITALS

Although the Lewis representation is a useful model to represent covalent bonds it does make the false assumption that all the valence electrons are the same. A more advanced model of bonding considers the combination of atomic orbitals to form molecular orbitals.
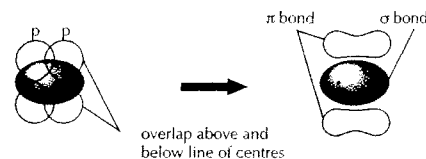
σ bonds

A σ (sigma) bond is formed when two atomic orbitals on different atoms overlap along a line drawn through the two nuclei. This occurs when two s orbitals overlap, an s orbital overlaps with a p orbital, or when two p orbitals overlap 'head on'.
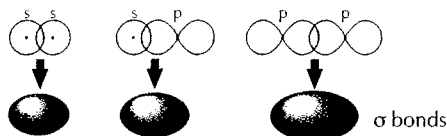
π bonds

A π (pi) bond is formed when two p orbitals overlap 'sideways on'. The overlap now occurs above and below the line drawn through the two nuclei. A π bond is made up of two regions of electron density.
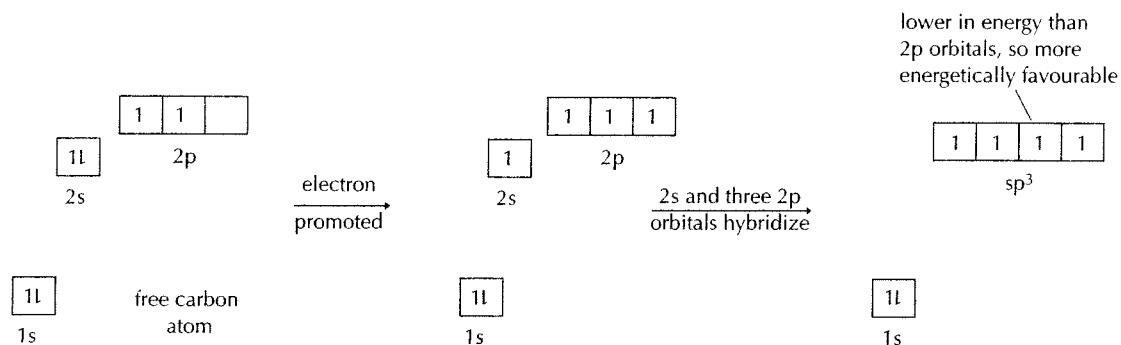

σ bonds

overlap above and below line of centres

HYBRIDIZATION (1)

sp³ hybridization

Methane provides a good example of sp³ hybridization. Methane contains four equal C–H bonds pointing towards the corners of a tetrahedron with bond angles of 109.5°. A free carbon atom has the configuration $1s^2 2s^2 2p^2$. It cannot retain this configuration in methane. Not only are there only two unpaired electrons, but the p orbitals are at 90° to each other and will not give bond angles of 109.5° when they overlap with the s orbitals on the hydrogen atoms.

When the carbon bonds in methane one of its 2s electrons is promoted to a 2p orbital and then the 2s and three 2p orbitals *hybridize* to form four new hybrid orbitals. These four new orbitals arrange themselves to be as mutually repulsive as possible, i.e. tetrahedrally. Four equal σ bonds can then be formed with the hydrogen atoms.

lower in energy than 2p orbitals, so more energetically favourable

sp³ orbitals of C atom overlap with s orbital on H atoms to form four equal σ bonds.

methane
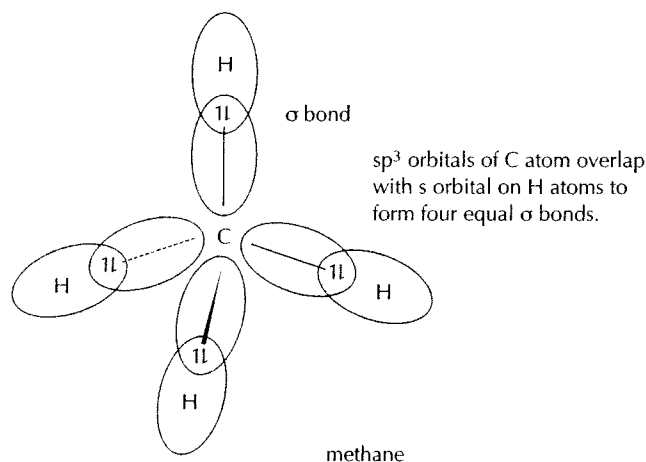
HYBRIDIZATION (2)
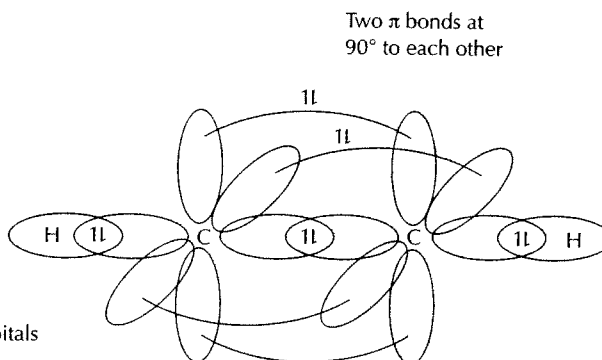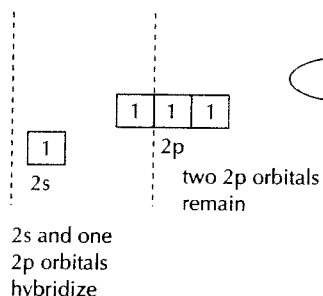
sp² hybridization

sp² hybridization occurs in ethene. After a 2s electron on the carbon atom is promoted the 2s orbital hybridizes with two of the 2p orbitals to form three new planar hybrid orbitals with a bond angle of 120° between them. These can form σ bonds with the hydrogen atoms and also a σ bond between the two carbon atoms. Each carbon atom now has one electron remaining in a 2p orbital. These can overlap to form a π bond. Ethene is thus a planar molecule with a region of electron density above and below the plane.
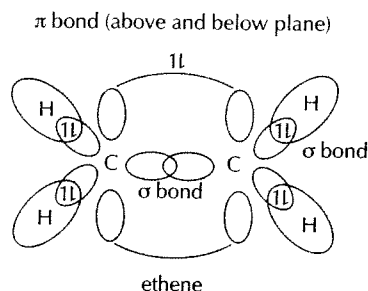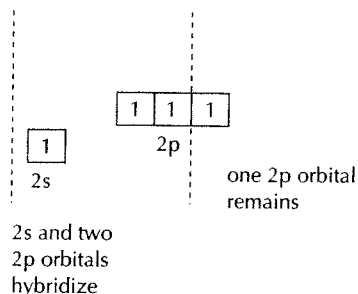
sp hybridization

sp hybridization occurs when the 2s orbital hybridizes with just one of the 2p orbitals to form two new linear sp hybrid orbitals with an angle of 180° between them. The remaining two p orbitals on each carbon atom then overlap to form two π bonds. An example is ethyne.

RELATIONSHIP BETWEEN TYPE OF HYBRIDIZATION, LEWIS STRUCTURE, AND MOLECULAR SHAPES

Molecular shapes can be arrived at either by using the VSEPR theory or by knowing the type of hybridization. Hybridization can take place between any s and p orbital in the same energy level and is not just restricted to carbon compounds. If the shape and bond angles are known from using Lewis structures then the type of hybridization can be deduced. Similarly if the type of hybridization is known the shape and bond angles can be deduced.

| Hybridization	Regular bond angle	Examples
sp³	109.5°	
sp²	120°	
sp	180°	

Delocalization of electrons

RESONANCE STRUCTURES

When writing the Lewis structures for some molecules it is possible to write more than one correct structure. For example, ozone can be written:


These two structures are known as resonance hybrids. They are extreme forms of the true structure, which lies somewhere between the two. Evidence that this is true comes from bond lengths, as the bond lengths between the oxygen atoms in ozone are both the same and are intermediate between an O=O double bond and an O–O single bond. Resonance structures are usually shown with a double headed arrow between them. Other common compounds which can be written using resonance structures are shown here.
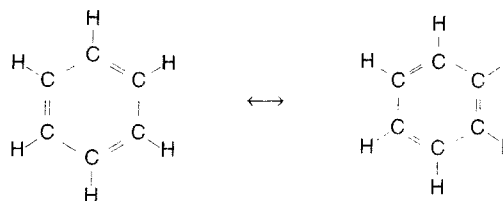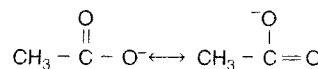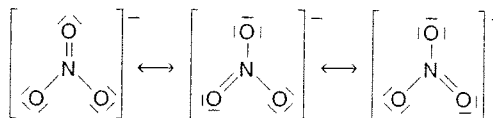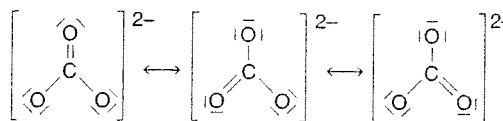
DELOCALIZATION OF ELECTRONS

Resonance structures can also be explained by the delocalization of electrons. For example, in the ethanoate ion the carbon atom and the two oxygen atoms each have a p orbital containing one electron after the σ bonds have been formed. Instead of forming just one double bond between the carbon atom and one of the oxygen atoms the electrons can delocalize over all three atoms. This is energetically more favourable than forming just one double bond.

Delocalization can occur whenever alternate double and single bonds occur between carbon atoms. The delocalization energy in benzene is about 150 kJ mol^{-1}, which explains why the benzene ring is so resistant to addition reactions.
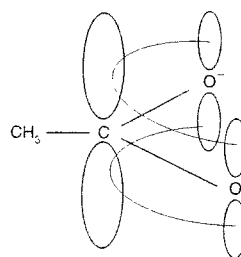
ethanoate ion shown as
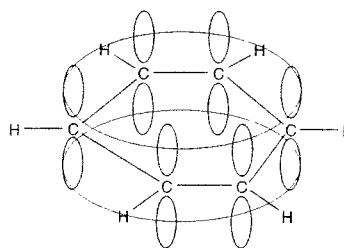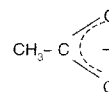
benzene ring shown as

IB QUESTIONS – BONDING

1. Which compound contains both covalent and ionic bonds?

   A. sodium carbonate, Na_2CO_3
   B. magnesium bromide, $MgBr_2$
   C. dichloromethane, CH_2Cl_2
   D. ethanoic acid, CH_3COOH

2. Which pair of elements is most likely to form a covalently bonded compound?

   A. Li and Cl   C. Ca and S
   B. P and O   D. Zn and Br

3. Given the following electronegativities,

   H: 2.2   N: 3.0   O: 3.5   F: 4.0

   which bond would be the most polar?

   A. O–H in H_2O C. N–O in NO_2
   B. N–F in NF_2 D. N–H in NH_3

4. What is the correct Lewis structure for methanal?

   A. $H:C:::O:H$ C. $\begin{array}{l} H \\ \ddot{C}::\ddot{O}: \\ H \end{array}$

   B. $\begin{array}{l} H \\ H:C::O: \end{array}$ D. $\begin{array}{l} :\ddot{C}\ddot{O}:H \\ H \end{array}$

5. When CH_4, NH_3, H_2O, are arranged in order of **increasing** bond angle, what is the correct order?

   A. CH_4, NH_3, H_2O C. NH_3, CH_4, H_2O
   B. NH_3, H_2O, CH_4 D. H_2O, NH_3, CH_4

6. When the H–N–H bond angles in the species NH_2^-, NH_3, NH_4^+ are arranged in order of increasing bond angle (smallest bond angle first), which order is correct?

   A. $NH_2^- < NH_3 < NH_4^+$ C. $NH_3 < NH_2^- < NH_4^+$
   B. $NH_4^+ < NH_3 < NH_2^-$ D. $NH_2^- < NH_4^+ < NH_3$

7. In which of the following pairs does the second substance have the lower boiling point?

   A. F_2, Cl_2 C. C_2H_6, C_3H_8
   B. H_2O, H_2S D. CH_3OCH_3, CH_3CH_2OH

8. In which of the following substances would hydrogen bonding be expected to occur?

   I. CH_4

   II. CH_3COOH

   III. CH_3OCH_3

   A. II only   C. I and III only
   B. I and III only   D. I, II and III

9. Which one of the following statements is correct?

   A. The energy absorbed when liquid ammonia boils is used to overcome the covalent bonds **within** the ammonia molecule.
   B. The energy absorbed when solid phosphorus (P_4) melts is used to overcome the ionic bonds **between** the phosphorus molecules.
   C. The energy absorbed when sodium chloride dissolves in water is used to form ions.
   D. The energy absorbed when copper metal melts is used to overcome the non-directional metallic bonds between the copper atoms.

10. A solid has a melting point of 1440 °C. It conducts heat and electricity. It does not dissolve in water or in organic solvents. The bond between the particles is most likely to be

    A. covalent.   C. ionic.
    B. dipole:dipole.   D. metallic.

(HL)

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11. What are the types of hybridization of the carbon atoms in the compound

    $H_2ClC–CH_2–COOH$?
        1   2   3

	_1_	_2_	_3_		_1_	_2_	_3_
A.	sp^2	sp^2	sp^2	C.	sp^3	sp^3	sp^2
B.	sp^3	sp^2	sp	D.	sp^3	sp^3	sp

12. Which molecule or ion does not have a tetrahedral shape?

    A. XeF_4 C. BF_4^-
    B. $SiCl_4$ D. NH_4^+

13. When the substances below are arranged in order of increasing carbon–carbon bond length (shortest bond first), what is the correct order?

    I. H_2CCH_2 II. H_3CCH_3 III. ⬡

    A. I < II < III   C. II < I < III
    B. I < III < II   D. III < II < I

14. Which of the following species is considered to involve sp^3 hybridization?

    I. BCl_3 II. CH_4 III. NH_3

    A. I only   C. I and III only
    B. II only   D. II and III only

15. How many π bonds are present in CO_2?

    A. One   C. Three
    B. Two   D. Four

16. When the following substances are arranged in order of increasing melting point (lowest melting point first) the correct order is

    A. $CH_3CH_2CH_3$, CH_3COCH_3, $CH_3CH_2CH_2OH$
    B. $CH_3CH_2CH_3$, $CH_3CH_2CH_2OH$, CH_3COCH_3
    C. CH_3COCH_3, $CH_3CH_2CH_2OH$, $CH_3CH_2CH_3$
    D. $CH_3CH_2CH_2OH$, $CH_3CH_2CH_3$, CH_3COCH_3