Structure 2.2

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The octet rule



The octet rule states that atoms bond together in order to achieve a full valence shell containing 8 electrons.

The octet rule

| 2 He 4.00 | 1s ² | Noble gases are stable because they have full |
|--------------------------|---|---|
| 10 Ne 20.18 | 1s ² 2s ² 2p ⁶ | valence shells. |
| 18 Ar 39.95 | 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ | 6 |
| 36 Kr 83.90 | 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ | ⁶ 3d ¹⁰ 4s ² 4p ⁶ |

Atoms can gain a full valence shell by either sharing electrons (covalent bonding) or by transferring electrons (ionic bonding).

The octet rule

Covalent bonding (CH₄)

Ionic bonding (NaCl)

H: C: H H: H







Covalent bonding (sharing of electrons)

Bromine, Br₂ Molecular Oxygen, O₂ Br:Br:Br: O::O

The octet rule

In covalent bonding, atoms share electrons to achieve a full valence shell.



The octet rule **Ionic bonding (transfer of electrons)**



Both ions have the electron configuration of a noble gas.



The octet rule

Exceptions to the octet rule

H:H He:

:CI:Be:CI:

• •

 \bullet \bullet





The octet rule

Exceptions to the octet rule

| | Number of electrons in valence shell | Example |
|-----------------------------|--------------------------------------|-------------------|
| Hydrogen | 2 | H ₂ |
| Helium | 2 | Не |
| Beryllium | 4 (incomplete octet) | BeCl ₂ |
| Boron | 6 (incomplete octet) | BF ₃ |
| Period 3 elements (S, P) | More than 8 (expanded octet) | SF ₆ |

Covalent bonding

Covalent bonding

Covalent bonding occurs between non-metal elements.

| | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---------------------------|---------------------------|------------------------------------|---------------------------|---------------------------|---------------------------|---------------------------|---------------------------|---------------------------|---------------------------|---------------------------|---------------------------|----------------------------|----------------------------|----------------------------|----------------------------|----------------------------|----------------------------|
| 1 | 1 H 1.01 | | | | | | Me | etals | | | | | | | | | | 2 He 4.00 |
| 2 | 3 Li 6.94 | 4 Be 9.01 | | | | | No | n-m | etals | S | | | 5 B 10.81 | 6 C 12.01 | 7 N 14.01 | 8 O 16.00 | 9 F 19.00 | 10 Ne 20.18 |
| 3 | 11 Na 22.99 | 12 Mg 24.31 | | | | | Me | etallo | oids | | | | 13 Al 26.98 | 14 Si 28.09 | 15 P 30.97 | 16 S 32.07 | 17 Cl 35.45 | 18 Ar 39.95 |
| 4 | 19 K 39.10 | 20 Ca 40.08 | 21 Sc 44.96 | 22 Ti 47.87 | 23 V 50.94 | 24 Cr 52.00 | 25 Mn 54.94 | 26 Fe 55.85 | 27 Co 58.93 | 28 Ni 58.69 | 29 Cu 63.55 | 30 Zn 65.38 | 31 Ga 69.72 | 32 Ge 72.63 | 33 As 74.92 | 34 Se 78.96 | 35 Br 79.90 | 36 Kr 83.90 |
| 5 | 37 Rb 85.47 | 38 Sr 87.62 | 39 Y 88.91 | 40 Zr 91.22 | 41 Nb 92.91 | 42 Mo 95.96 | 43 Tc (98) | 44 Ru 101.07 | 45 Rh 102.91 | 46 Pd 106.42 | 47 Ag 107.87 | 48 Cd 112.41 | 49 In 114.82 | 50 Sn 118.71 | 51 Sb 121.76 | 52 Te 127.60 | 53 I 126.90 | 54 Xe 131.29 |
| 6 | 55 Cs 132.91 | 56 Ba 137.33 | 57 † La 138.91 | 72 Hf 178.49 | 73 Ta 180.95 | 74 W 183.84 | 75 Re 186.21 | 76 Os 190.23 | 77 Ir 192.22 | 78 Pt 195.08 | 79 Au 196.97 | 80 Hg 200.59 | 81 Tl 204.38 | 82 Pb 207.20 | 83 Bi 208.98 | 84 Po (209) | 85 At (210) | 86 Rn (222) |
| 7 | 87 Fr (223) | 88 Ra (226) | 89 ‡ Ac (227) | 104 Rf (267) | 105 Db (268) | 106 Sg (269) | 107 Bh (270) | 108 Hs (269) | 109 Mt (278) | 110 Ds (281) | 111 Rg (281) | 112 Cn (285) | 113 Uut (286) | 114 Uuq (289) | 115 Uup (288) | 116 Uuh (293) | 117 Uus (294) | 118 Uuo (294) |
| | | | t | 58 Ce 140.12 | 59 Pr 140.91 | 60 Nd 144.24 | 61 Pm (145) | 62 Sm 150.36 | 63 Eu 151.96 | 64 Gd 157.25 | 65 Tb 158.93 | 66 Dy 162.50 | 67 Ho 164.93 | 68 Er 167.26 | 69 Tm 168.93 | 70 Yb 173.05 | 71 Lu 174.97 | |
| | | | ŧ | Th 232.04 | Pa 231.04 | U 238.03 | Np (237) | Pu (244) | Am (243) | Cm (247) | Bk (247) | Cf (251) | Es (252) | Fm (257) | Md (258) | No (259) | (262) | |



Covalent bonding

Covalent bonding occurs between non-metal elements.

| Difference in electronegativity | Type of bonding |
|---------------------------------|--------------------|
| 0-0.4 | non-polar covalent |
| 0.5-1.7 | polar covalent |
| ≥ 1.8 | ionic |

Covalent bonds can be classified as non-polar covalent or polar covalent depending on the difference in electronegativity between the bonded atoms. Tutorials for IB Chemistry

Covalent bonding **Covalent bonds exist between the atoms in** molecular compounds such as CO_2 , CH_4 and H_2O_2 . They also exist between the atoms in giant covalent substances such as silicon dioxide and diamond.













Covalent bonding

shared pair of electrons



H-H H:H

electrostatic attraction

A covalent bond is the electrostatic attraction between positive nuclei and shared pairs of electrons.



| | Number of shared | C to C bond | C to C bond |
|--------|------------------|----------------------------------|------------------------------|
| Bond | electrons | strength (kJ mol ⁻¹) | length (10 ⁻¹² m) |
| Single | 2 | 347 | 153 |
| Double | 4 | 614 | 134 |
| Triple | 6 | 839 | 120 |

Covalent bonding

Covalent bonding occurs between non-metal elements. A covalent bond is the electrostatic attraction between positive nuclei and shared pairs of electrons. **Covalent bonds can be polar or non-polar depending** on the difference in electronegativity between the atoms.

Single covalent bonds are weaker and longer than double or triple covalent bonds which are stronger and shorter.

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Lewis formulas



Lewis formulas

Lewis formulas (or Lewis structures) show all the valence electrons in a molecule; the bonding electrons and the lone pairs of electrons.



Lewis structures

- 1) Count the total number of valence electrons in all the atoms in the molecule.
- 2) Determine the number of electrons needed for each atom to achieve an octet.
- 3) Subtract 1 from 2 to get the number of bonding electrons in the molecule.
- 4) Add electrons to each atom until it has an octet.
- 5) Count the total number of valence electrons, it should be equal to the number in part 1.



Lewis formulas

Methane (CH₄)

- 1) 4 + (4 × 1) = 8 valence electrons
- 2) 8 + (4 × 2) = 16 electrons needed to complete each atom's octet
- 3) 16 8 = 8 bonding electrons
- 4) Complete each atom's octet
- 5) Total number of electrons = 8



Dichloromethane (CH₂Cl₂) 1) $4 + (2 \times 1) + (2 \times 7) = 20$ valence electrons 2) $8 + (2 \times 2) + (2 \times 8) = 28$ electrons needed to complete each atom's octet 3) 28 - 20 = 8 bonding electrons 4) Complete each atom's octet

Lewis formulas

5) Total number of electrons = 20

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Lewis formulas

Ammonia (NH₃)

- 1) 5 + (3 × 1) = 8 valence electrons
- 2) 8 + (3 × 2) = 14 electrons needed to complete each atom's octet
- 3) 14 8 = 6 bonding electrons
- 4) Complete each atom's octet
- 5) Total number of electrons = 8



Ethene (C_2H_4) 1) $(2 \times 4) + (4 \times 1) = 12$ valence electrons

2) $(2 \times 8) + (4 \times 2) = 24$ electrons needed to complete each atom's octet

Lewis formulas

- 3) 24 12 = 12 bonding electrons
- **Complete each atom's octet** 4)

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5) Total number of electrons = 12 H:C:C:H



Coordination bonds



Coordination bonds

In the formation of a single covalent bond, each atom contributes one electron to the bond.



In a coordination bond, one atom contributes both the bonding electrons to the bond.

In a coordination bond, one atom contributes both the bonding electrons to the bond.

Coordination bonds

Once a coordination bond is formed, it is identical to a regular covalent bond.

Hydronium ion (H_3O^+)

[H-Ö-H]⁺



Ammonium ion (NH_4^+)

Coordination bonds

 Al_2Cl_6 – the dimer formed between two molecules of $AlCl_3$.



VSEPR theory

VSEPR theory

- Valence shell electron pair repulsion theory is used to predict the geometry (shape) of molecules.
- Electron pairs (bonds or lone pairs) repel each other and spread apart as far as possible.

greatest repulsion

least repulsion

lone pair – lone pair > lone pair – bonding pair > bonding pair – bonding pair

 The term electron domain is used to refer to bonds or lone pairs of electrons around an atom in a molecule.





Single bonds, double bonds, triple bonds and lone pairs of electrons count as one electron domain.



4 electron domains around the carbon atom (4 bonding domains)







 $O \equiv C \equiv C$

4 electron domains around the oxygen atom (2 bonding domains, 2 lone pairs)

2 electron domains around the carbon atom (2 bonding domains)

VSEPR theory





| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|-----------------|-------------|----------------|
| domains | domains | pairs | geometry | geometry | angle |
| 4 | 4 | 0 | tetrahedral | tetrahedral | 109.5 ° |

VSEPR theory

NH₃





| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|-----------------|-----------------------|----------------|
| domains | domains | pairs | geometry | geometry | angle |
| 4 | 3 | 1 | tetrahedral | trigonal pyramidal | 107.8 ° |

VSEPR theory



| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|-----------------|-----------|----------------|
| domains | domains | pairs | geometry | geometry | angle |
| 4 | 2 | 2 | tetrahedral | bent | 104.5 ° |

Water has a bent shape due to the extra repulsion from the two lone pairs of electrons on the oxygen atom.

VSEPR theory





| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|--------------------|--------------------|--------------|
| domains | domains | pairs | geometry | geometry | angle |
| 3 | 3 | 0 | trigonal planar | trigonal planar | 120 ° |







| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|-----------------|-----------|-------|
| domains | domains | pairs | geometry | geometry | angle |
| 3 | 2 | 1 | trigonal planar | bent | <120° |

VSEPR theory





| electron | bonding | lone | electron domain | molecular | bond |
|----------|---------|-------|-----------------|-----------|--------------|
| domains | domains | pairs | geometry | geometry | angle |
| 2 | 2 | 0 | linear | linear | 180 ° |
VSEPR theory

| Electron domains | Bonding domains | Lone pairs | Electron domain geometry | Molecular geometry | Bond angle | Example |
|---------------------|--------------------|---------------|--------------------------|-----------------------|----------------|------------------|
| 4 | 4 | 0 | tetrahedral | tetrahedral | 109.5° | CH ₄ |
| 4 | 3 | 1 | tetrahedral | trigonal pyramidal | 107.8° | NH ₃ |
| 4 | 2 | 2 | tetrahedral | bent / v-shaped | 104.5 ° | H ₂ O |
| 3 | 3 | 0 | trigonal planar | trigonal planar | 120 ° | BF ₃ |
| 3 | 2 | 1 | trigonal planar | bent / v-shaped | < 120° | SO ₂ |
| 2 | 2 | 0 | linear | linear | 180 ° | CO2 |

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Polar and non-polar covalent bonds

Covalent bonds can be classified as polar or non-polar depending on the difference in electronegativity between the bonding atoms.

| Difference in electronegativity | Polar or non-polar covalent bond | Example |
|---------------------------------|---|---------|
| 0 | non-polar (pure) covalent bond | CI-CI |
| 0.1–0.4 | non-polar (weakly polar) covalent bond | C-H |
| 0.5-1.7 | polar covalent bond | C-F |



Polar and non-polar bonds



Increasing difference in electronegativity

Polar and non-polar bonds Covalent bonds can be classified as polar or non-polar depending on the difference in electronegativity between the bonding atoms.

Polar covalent bonds have a bond dipole.





Non-polar bonds

C - HCI-CI $O \equiv O$ 3.4 3.4 3.2 2.2 3.2 2.6 $\Delta EN = 0$ $\Delta EN = 0$ $\Delta EN = 0.4$



Polar bonds

C-F H-CIC=O2.2 3.2 4.0 2.6 2.6 3.2 **ΔEN = 1.0 ΔEN = 1.4** $\Delta EN = 0.6$ $\delta + \delta_{-}$ $\delta +$ δ- $\delta +$ δ- $C \equiv O$ H-CI



Polar bonds





H-N

Polar bonds



 $\delta + \delta -$ H-Ō

H–F

| UC | MSJChem torials for IB Chemist | | ir and non-r | oolar bonds |
|----|-----------------------------------|----------|---------------------|-------------|
| | Non-polar (weakly pol | r ar) | | |
| | covalent | | Polar covalent | Ionic |
| | 0 | 0.4 | | 1.8 |
| | More or le | ess U | nequal sharing | Formation |
| | equal sharin | g of | of electrons | of ions |
| | electrons | 3 | Z 1 Z 2 1 Z | Z 1 Z |
| | | | - 0 + 0 - 0 + 0 - 0 | |

MSJChem Tworks for IB Chemistry

Polar and non-polar molecules



The polarity of a molecule depends on two factors:

The presence of polar bonds in the molecule.
The geometry of the molecule.
Polar molecules have a net dipole moment.





Polar and non-polar molecules



Non-polar molecules



MSJChem Tutorials for IB Chemistry Non-polar molecules Non-polar bonds (CCl_4)



 CCl_4 is a non-polar molecule; the bond polarities cancel out, therefore, it has no net dipole moment.



Polar molecules

Polar molecule (CH₂Cl₂)



 CH_2Cl_2 is a polar molecule; the bond polarities do not cancel out therefore it has a net dipole moment.

Polar molecules



MSJChem Tutorials for IB Chemistry Polar and non-polar molecules Tetrahedral molecules with the same type of atom bonded to the central atom are non-polar (CH_4 , CCl_4). Tetrahedral molecules with different atoms bonded to the central atom are usually polar (CH_3CI , CH_3OH). Trigonal planar molecules with the same type of atom bonded to the central atom are non-polar (BF_3) . Trigonal pyramidal molecules are usually polar if they contain polar bonds (NH₃). Bent molecules are usually polar if they contain polar bonds (H_2O) .

Polar and non-polar molecules Tutorials for IB Chemistry Linear molecules with the same atom bonded to the central atom are non-polar (CO₂). Linear molecules with different atoms bonded to the central atom are usually polar (HCN). Diatomic molecules with the same atom bonded together are non-polar (H₂, Cl₂, N₂). **Diatomic molecules with different atoms bonded** together are usually polar (HCl, HF).

MSJChem Tườnals for IB Chemistry

Properties of molecular compounds and covalent network structures





Tutorials for IB Chemistry Properties of covalent substances

| Compound | Molar mass g mol ⁻¹ | Melting point (°C) | Boiling point (°C) | State at room temperature (25 °C) |
|------------------|-----------------------------------|-----------------------|-----------------------|--------------------------------------|
| CH ₄ | 16.04 | -183 | -162 | Gas |
| CO ₂ | 44.01 | sublimes | s at -78.5 | Gas |
| CCl ₄ | 153.81 | -22.9 | 76.7 | Liquid |
| NH ₃ | 17.04 | -77.7 | -33.3 | Gas |
| H ₂ O | 18.02 | 0 | 100 | Liquid |

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Properties of covalent substances

Covalent network substances have high melting and boiling points.



Diamond M.P. 3550 °C B.P. 4830 °C Silicon M.P. 1414 °C B.P. 3265 °C



Oxygen ator

Silicon dioxide M.P. 1713 °C B.P. 2950 °C

Tutorials for IB Chemistry Properties of covalent substances

| Property | Molecular elements / compounds | Covalent network substances |
|----------------------------|---|---|
| Solubility | Polar molecules are soluble in polar solvents Non-polar molecules are soluble in non-polar solvents | Insoluble in both polar and non-polar solvents |
| Electrical conductivity | Poor electrical conductors (do not have free moving ions or delocalised electrons) | Poor electrical conductors (do not have free moving ions or delocalised electrons) |

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Allotropes of carbon



Tutorials for IB Chemistry

Allotropes are different forms of the same element in the same physical state. Carbon has 4 allotropes – graphite, diamond, Fullerene C₆₀ and graphene.















Graphite has a layered structure. The layers are held together by weak intermolecular forces – the layers can slide over one another. Each carbon atom is bonded to 3 other carbon atoms. The bond angle is 120°, trigonal planar. **Good conductor of electricity** (delocalised electrons).



Diamond

Diamond has a covalent network structure. High melting point, high boiling point and very hard (strong covalent bonds between atoms). Each carbon is bonded to 4 other carbon atoms. Bond angle is 109.5°, tetrahedral. Does not conduct electricity (no delocalised electrons).





Structure consists of 12 pentagons and 20 hexagons. Each carbon atom is bonded to 3 other carbon atoms. Shows some electrical conductivity.





Very thin (one layer thick) but also very strong. Each carbon atom is bonded to 3 other carbon atoms. **Bond angle between carbon** atoms is 120°, trigonal planar. Very good electrical and thermal conductivity.



| Allotrope | Graphite | Diamond | Fullerene C ₆₀ | Graphene |
|-------------------------|---------------------------------------|--|---------------------------------------|---|
| Structure | Layered (weak IMF between layers) | Covalent network | 12 pentagons 20 hexagons | Layered structure |
| C-C bonding | Each C atom bonded to 3 other C atoms | Each C atom bonded to 4 other C atoms | Each C atom bonded to 3 other C atoms | Each C atom bonded to 3 other C atoms |
| C-C bond angle | 120° trigonal planar | 109.5° tetrahedral | 120° trigonal planar | 120° trigonal planar |
| Electrical conductivity | High | None | Medium | Very high |
| Thermal conductivity | Low | Very good | Low | Very good |
| Delocalised electrons | Yes | No | Yes | Yes |
| Properties | Very soft | Very hard | Very light and strong | High melting point, flexible, stronger than steel |

Internolecular forces



Intermolecular forces

Intermolecular forces are forces between molecules that determine physical properties such as the melting point and boiling point of a substance.



London dispersion forces (weakest) Dipole-dipole forces Hydrogen bonding (strongest)



Intermolecular forces

- Intermolecular forces are forces that exist between molecules.
- They influence the physical properties of a substance such as melting and boiling point.
- They are weaker than bonds between atoms such as covalent bonds.
 - London dispersion forces (weakest) Dipole-dipole forces Hydrogen bonding (strongest)





Intermolecular forces

| Type of IMF | Type of molecule | Energy (kJ mol ⁻¹) |
|-----------------------------|---|--------------------------------|
| London dispersion forces | All types of molecules | 0.05-40 |
| Dipole-dipole | Polar molecules | 5-25 |
| Hydrogen bonding | Molecules with O-H, N-H and H-F bonds | 10-40 |

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London dispersion forces are caused by the movement of electrons within an atom or molecule. The constant motion of electrons within an atom or molecule can cause a temporary (instantaneous) dipole.



MSJChem Tutorials for IB Chemistry London dispersion forces

| | Molar mass (g mol ⁻¹) | Boiling point (°C) |
|-----------------|-----------------------------------|---------------------------|
| F ₂ | 30.8 | -188 |
| Cl ₂ | 70.9 | -34.0 |
| Br ₂ | 160 | 58.0 |
| I ₂ | 254 | 193 |

As molar mass increases, the strength of the London dispersion forces between the molecules also increases. This results in an increased boiling point.


Dipole-dipole forces

Dipole-dipole forces occur between polar molecules (molecules that have a net dipole moment).

The dipole-dipole force is the electrostatic attraction between the partial positive charge on one molecule and the partial negative charge on another.



Dipole-dipole forces

| Compound | Molar mass (g mol ⁻¹) | Dipole moment (D) | Boiling point (K) |
|----------------------------------|--------------------------------------|----------------------|----------------------|
| CH ₃ OCH ₃ | 46.07 | 1.30 | 248 |
| CH ₃ Cl | 50.48 | 1.87 | 249 |
| CH₃CHO | 44.05 | 2.69 | 294 |
| CH ₃ CN | 41.05 | 3.92 | 355 |

Hydrogen bonding occurs when a hydrogen atom is bonded to either a nitrogen, oxygen or fluorine atom.

Hydrogen bonding



Hydrogen bonding



The hydrogen bond is between the partial positive charge on the hydrogen atom and a lone pair of electrons on the oxygen atom. Water has a much higher boiling point compared to other molecules with similar molar masses.

Hydrogen bonding



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| Type of molecule | Intermolecular forces | Examples |
|---|--|--|
| Non-polar | London dispersion forces | $\begin{array}{cccc} Cl_2 & H_2 & N_2 & O_2 \\ & CH_4 & CCl_4 \end{array}$ |
| Polar | London dispersion forces Dipole-dipole forces | HCI HCN CH ₃ CI CH ₃ CHO |
| Molecules with H bonded to N, O or F | London dispersion forces Dipole-dipole forces Hydrogen bonding | $H_2O NH_3 HF$ C_2H_5OH CH_3COOH |

Solubility and intermolecular forces

A polar solvent is a liquid composed of polar molecules. A non-polar solvent is a liquid composed of non-polar molecules. Molecules.

| Polar solvents | Non-polar solvents |
|---|---|
| Water H ₂ O | Hexane C ₆ H ₁₄ |
| Methanol CH ₃ OH | Octane C ₈ H ₁₈ |
| Ethanol C ₂ H ₅ OH | Benzene C ₆ H ₆ |
| Propanone CH ₃ CH ₂ CHO | Methylbenzene C ₆ H ₅ CH ₃ |
| Ethanoic acid CH ₃ COOH | Carbon tetrachloride CCl ₄ |

MSJChem Tutorials for IB Chemistry Polar and non-polar solvents

- Polar substances are soluble in polar solvents (miscible). Non-polar substances are soluble in non-polar solvents. The phrase 'like dissolves like' is useful to remember. Examples:
- Hexane is soluble in octane (both non-polar).
- Methanol is soluble in water (both polar).
 Water is known as the universal solvent because of its ability to dissolve so many different substances.



lon-dipole forces

Ion-dipole forces occur between water molecules and ions in aqueous solutions.



When an ionic compound dissolves in water, ion-dipole forces occur between the ions and the oppositely-charged ends of the water molecules. The water molecules surround the ion forming a hydration shell.



lon-dipole forces

Ion-dipole forces occur between water molecules and ions in aqueous solutions.



When an ionic compound dissolves in water, ion-dipole forces occur between the ions and the oppositely-charged ends of the water molecules. The water molecules surround the ion forming a hydration shell.



Hydrogen bonds

Hydrogen bonds occur between water molecules and polar molecules such as alcohols.



Methanol is soluble in water because it is able to form hydrogen bonds with water molecules.

Ammonia, NH_3 , is also able to form hydrogen bonds with water molecules.



MSJChemistry London dispersion forces London dispersion forces occur between non-polar molecules.



$$\begin{array}{cccc} H & O \\ H - C - O - C - (CH_2)_{16}CH_3 \\ O \\ H - C - O - C - (CH_2)_{16}CH_3 \\ O \\ H - C - O - C - (CH_2)_{16}CH_3 \\ H \end{array}$$

Hexane, a non-polar molecule, is soluble in oil. Hexane is insoluble in polar solvents.





Most ionic compounds are soluble in water because of its polar nature forming ion-dipole forces. Polar substances are soluble in water because they are able to form hydrogen bonds with water molecules.

Non-polar substances are soluble in non-polar solvents (oil and hexane are miscible) because of the London dispersion forces that occur between the molecules.

Types of bonding and electrical conductivity For a substance to conduct electricity it must have one of the following:

- Delocalised electrons electrons that are free to move within the structure.
- Free moving (mobile) ions ions that are free to move in a solution or in a liquid.







MSJChem Tutorials for IB Chemistry Bonding and electrical conductivity

Molecular elements and compounds are poor conductors of electricity.



They do not have delocalised electrons (electrons are localised in covalent bonds between atoms).

MSJChem Bonding and electrical conductivity Giant covalent substances are poor conductors of electricity.



They do not have delocalised electrons (electrons are localised in covalent bonds between atoms).

Tutorials for IB Chemistry Bonding and electrical conductivity



MSJChem Tentorials for IB Chemistry Bonding and electrical conductivity Molecules with delocalised electrons are not good conductors of electricity.



Benzene C₆H₆





Bonding and electrical conductivity Ionic compounds conduct electricity only when melted or dissolved in water.



When solid, the ions are held in fixed positions. When melted or dissolved, the ions are free to move and can carry an electric current.

MSJChem Tutorials for IB Chemistry Bonding and electrical conductivity

Metallic substances (metals) are good electrical conductors.



The delocalised electrons within the metallic structure are free to move and conduct electricity.

MSJChem Tutorials for IB Chemistry Bonding and electrical conductivity

| Type of bonding | Electrical conductivity | Reason |
|--------------------|---|---|
| Covalent | Poor electrical conductivity | No delocalised electrons |
| lonic | Only when melted or dissolved in water | lons are free to move when melted or dissolved |
| Metallic | Good electrical conductivity | Delocalised electrons within the metallic structure |

Chromatography



Chromatography

Chromatography is a separation technique used to separate a mixture of solutes in a solvent.

A small sample of the mixture is spotted near the bottom of a piece of filter paper (known as the origin).

The filter paper is suspended in a solvent with the spot above the level of the solvent.

Chromatography

As the solvent rises up the filter paper by capillary action, the components of the mixture will distribute themselves between two phases – the stationary phase (the filter paper) and the mobile phase (the solvent).

This distribution is determined by the strength of the intermolecular forces experienced by the components of the mixture in each phase.

If the component forms stronger intermolecular forces with the stationary phase it will not travel as far up the paper.

Conversely, a component that forms stronger intermolecular forces with the mobile phase will travel further up the paper.





- The components of the mixture can be identified by calculating the retardation factor, R_f.
- The retardation factor is calculated by dividing the distance moved by the component by the total distance moved by the solvent (known as the solvent front).

$$R_{f} = \frac{distance moved by component}{distance moved by solvent front}$$

 This R_f value can be compared to the R_f values of known substances and the components of the mixture can be identified.

Chromatography

Calculating the retardation factor, R_f



In the chromatogram shown, component B travels 3 cm and the solvent front is 10 cm.

The retardation factor is:

$$R_f = \frac{3 \ cm}{10 \ cm} = 0.3$$