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Structure 2.2

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The octet rule

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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.

2 **He** 4.00	$1s^2$
10 **Ne** 20.18	$1s^2\ 2s^2\ 2p^6$
18 **Ar** 39.95	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
36 **Kr** 83.90	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

Noble gases are stable because they have full valence shells.

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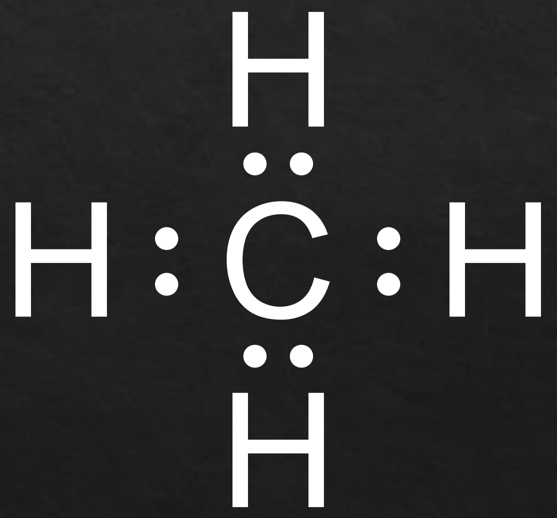
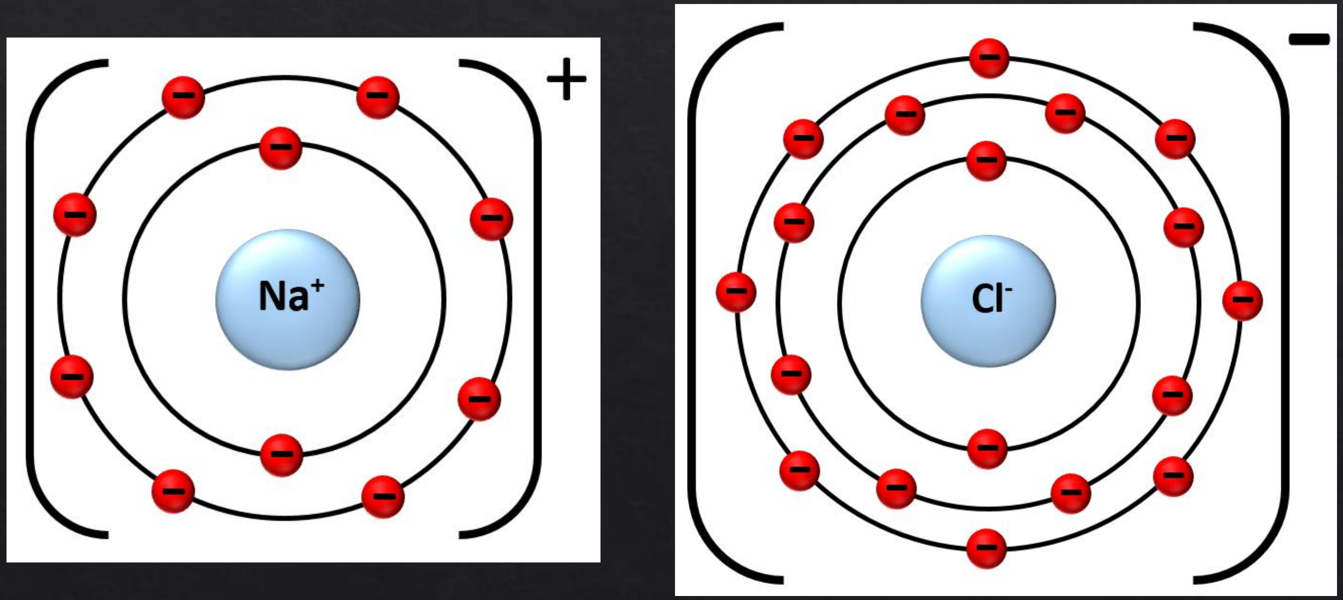
The octet rule

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

Covalent bonding (CH_4)



Ionic bonding (NaCl)



The octet rule

Covalent bonding (sharing of electrons)

Bromine, Br_2

Molecular Oxygen, O_2

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

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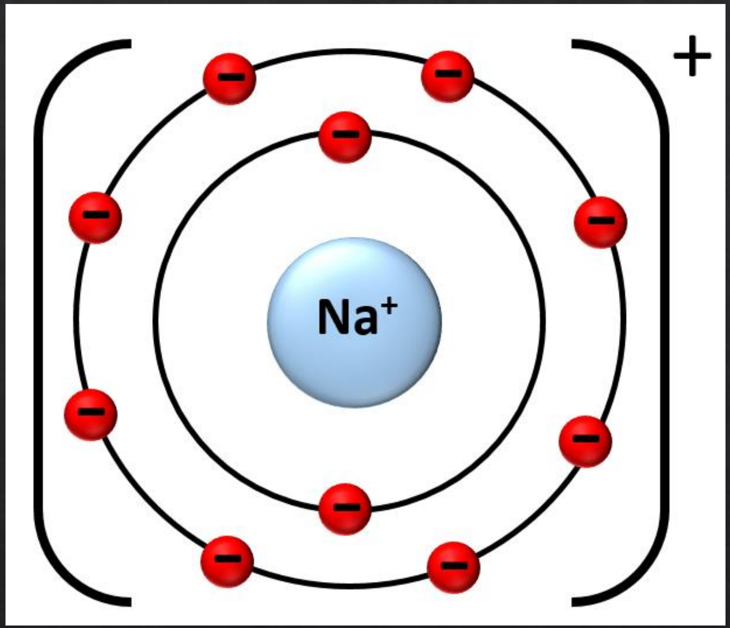
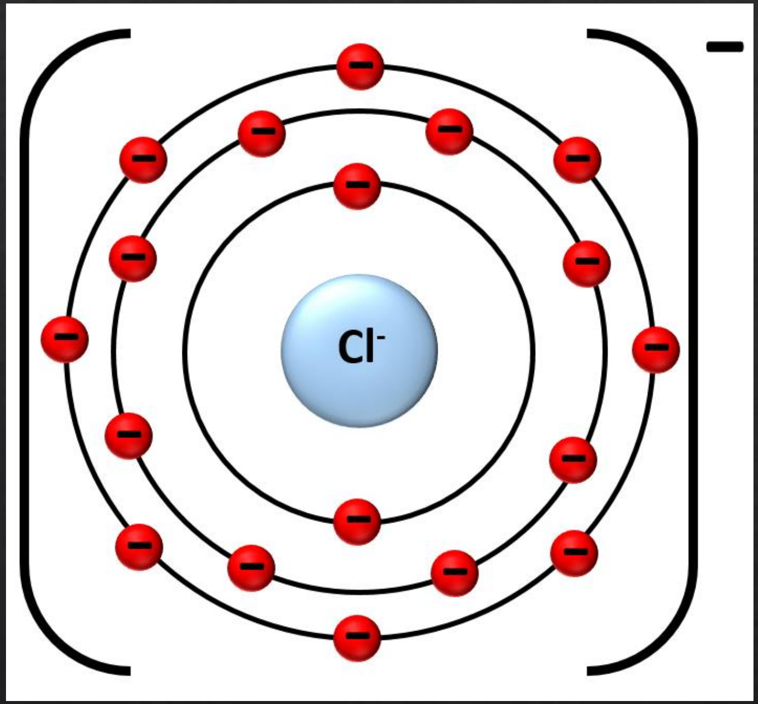
The octet rule

Ionic bonding (transfer of electrons)

Na^+



$1s^2\ 2s^2\ 2p^6$



Cl^-

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Both ions have the electron configuration of a noble gas.

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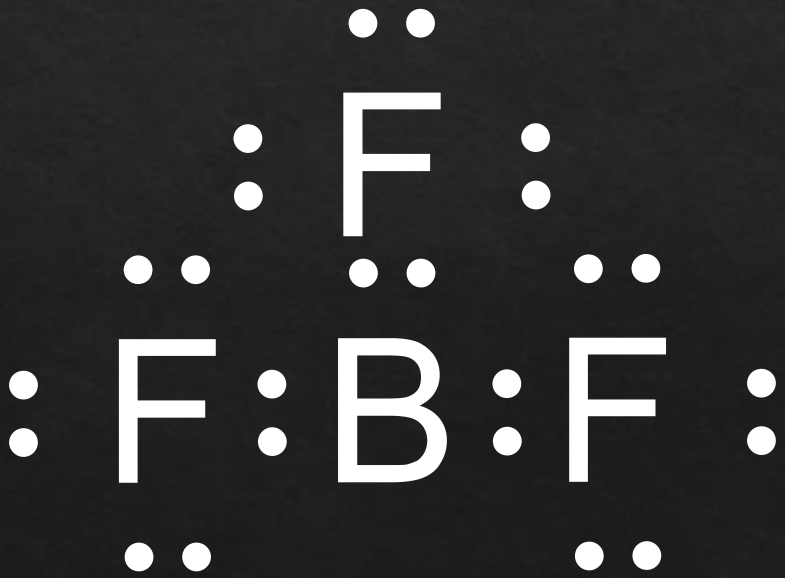
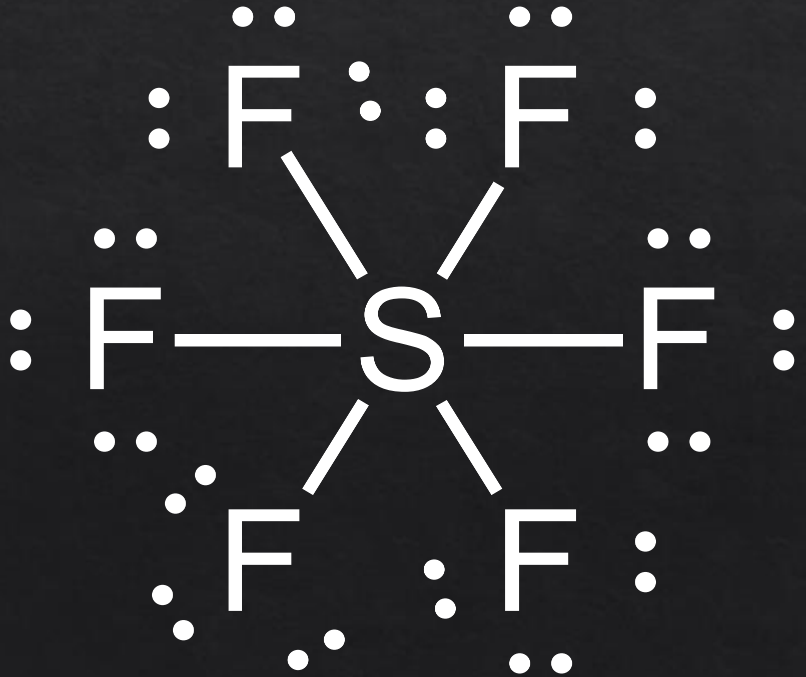
The octet rule

Exceptions to the octet rule





The octet rule

Exceptions to the octet rule

	Number of electrons in valence shell	Example
Hydrogen	2	H_2
Helium	2	He
Beryllium	4 (incomplete octet)	$BeCl_2$
Boron	6 (incomplete octet)	BF_3
Period 3 elements (S, P)	More than 8 (expanded octet)	SF_6

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Covalent bonding

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Covalent bonding

Covalent bonding occurs between non-metal elements.

- Metals
- Non-metals
- Metalloids

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											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											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)
			†	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	
			‡	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)	

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Covalent bonding

Covalent bonding occurs between non-metal elements.

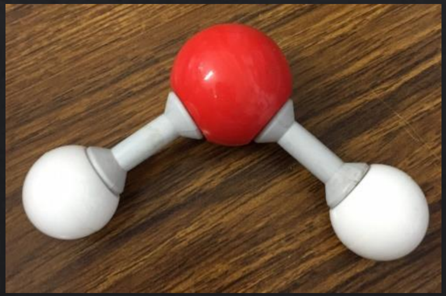
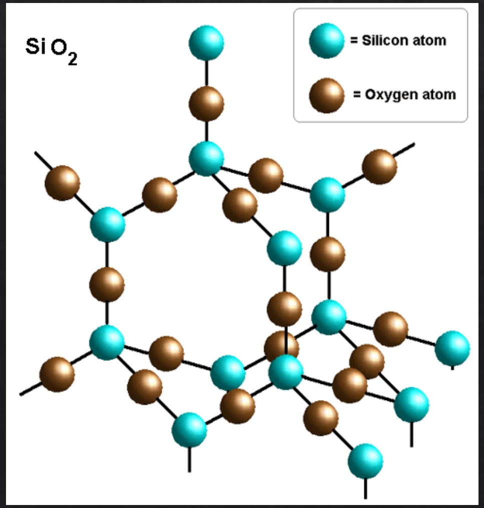
Difference in electronegativity	Type of bonding
0—0.4	non-polar covalent
0.5—1.7	polar covalent
≥1.8	ionic

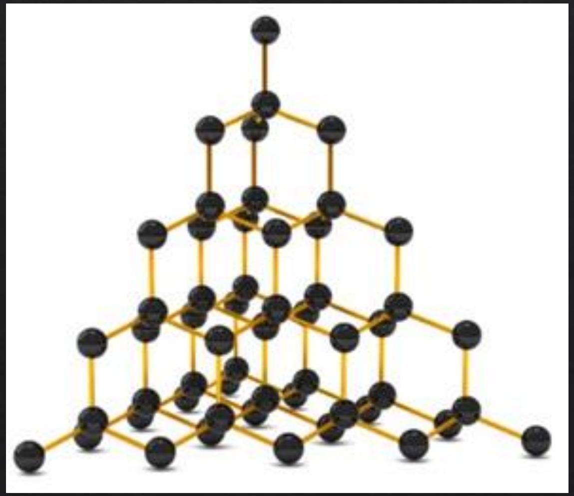
Covalent bonds can be classified as non-polar covalent or polar covalent depending on the difference in electronegativity between the bonded atoms.

Covalent bonding

Covalent bonds exist between the atoms in molecular compounds such as CO_2, CH_4 and H_2O. They also exist between the atoms in giant covalent substances such as silicon dioxide and diamond.





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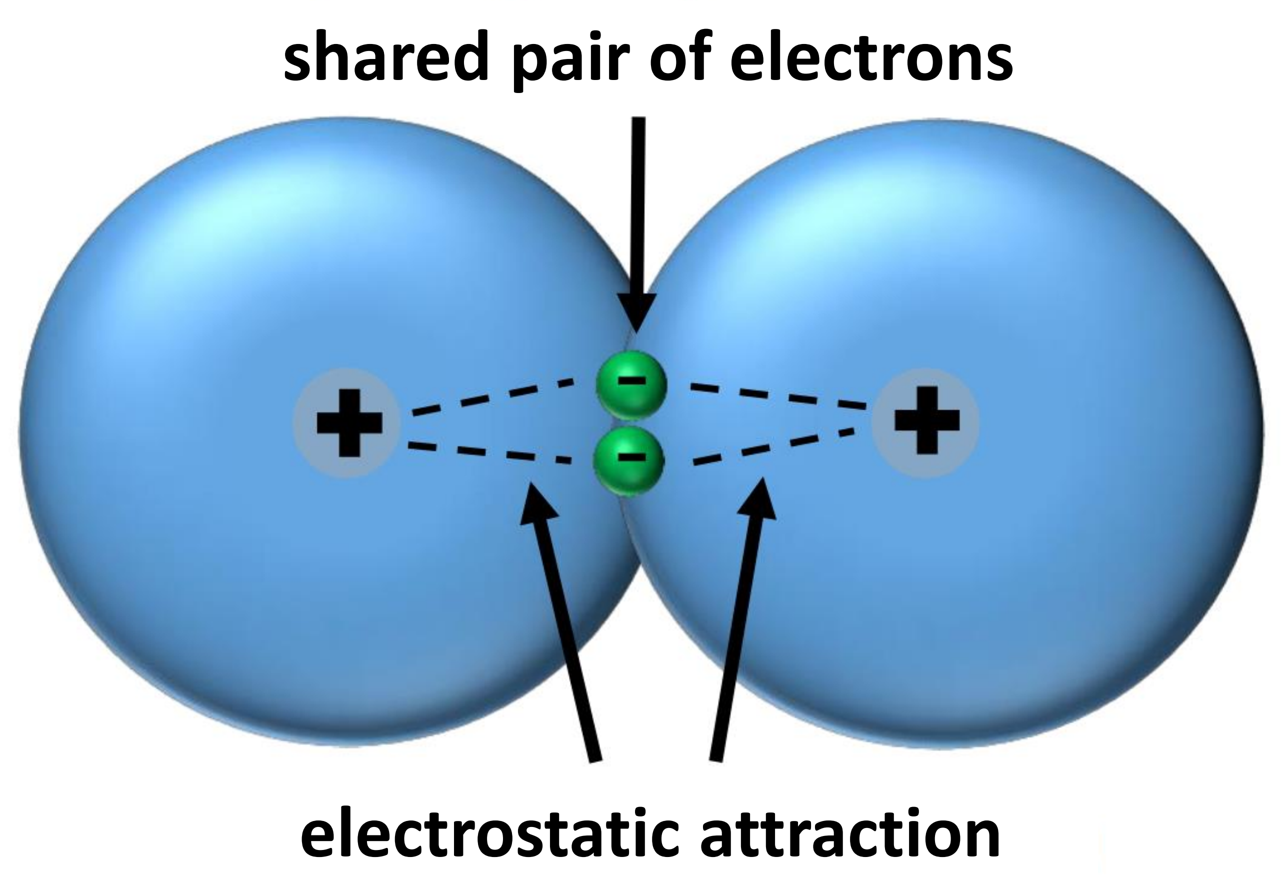
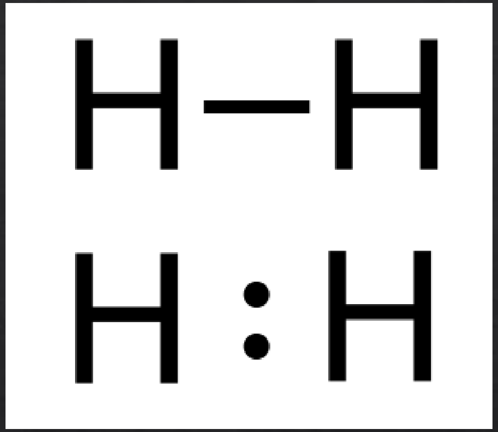
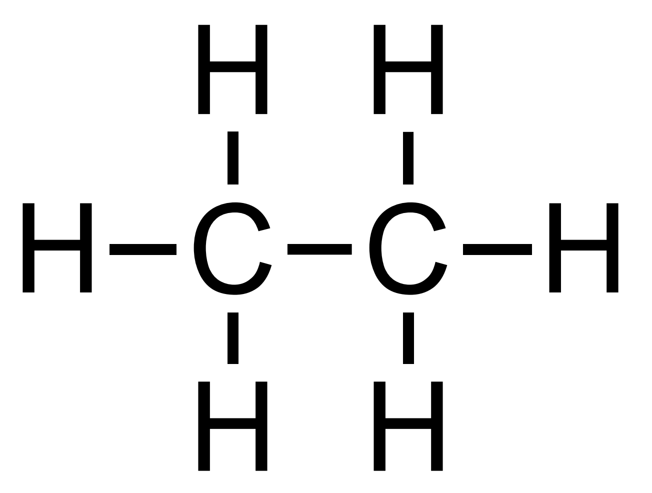
Covalent bonding

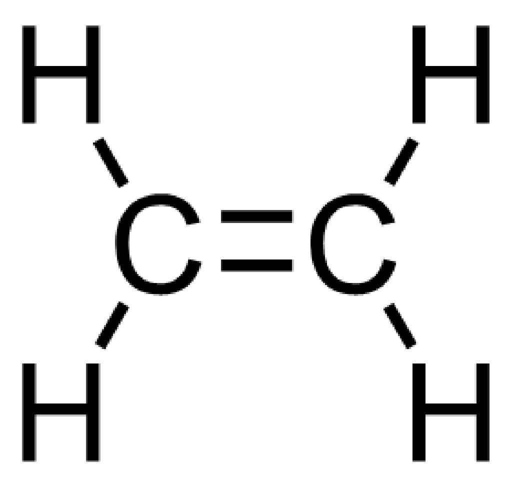


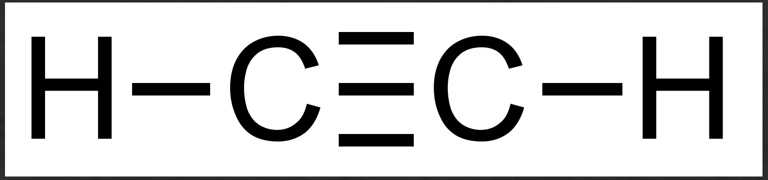


A covalent bond is the electrostatic attraction between positive nuclei and shared pairs of electrons.

Covalent bonding







Bond	Number of shared electrons	C to C bond strength (kJ mol^{-1})	C to C bond 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

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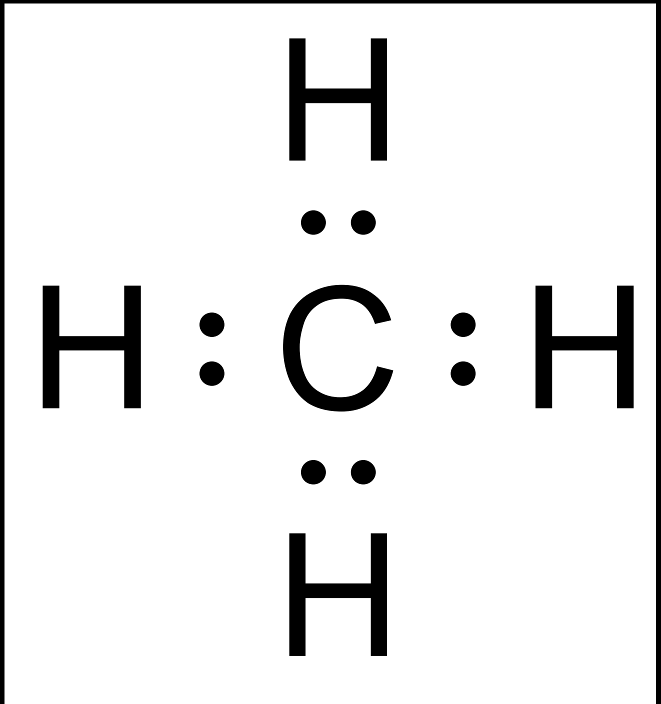
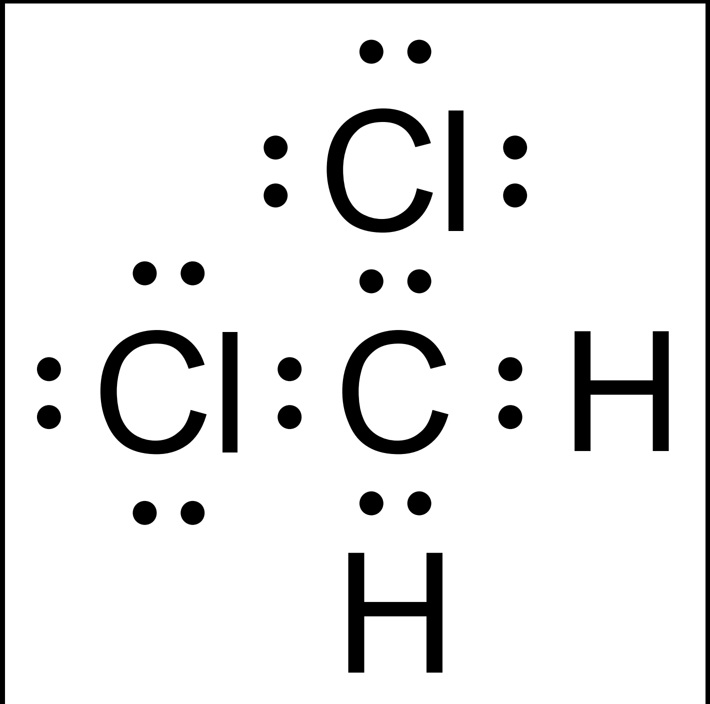
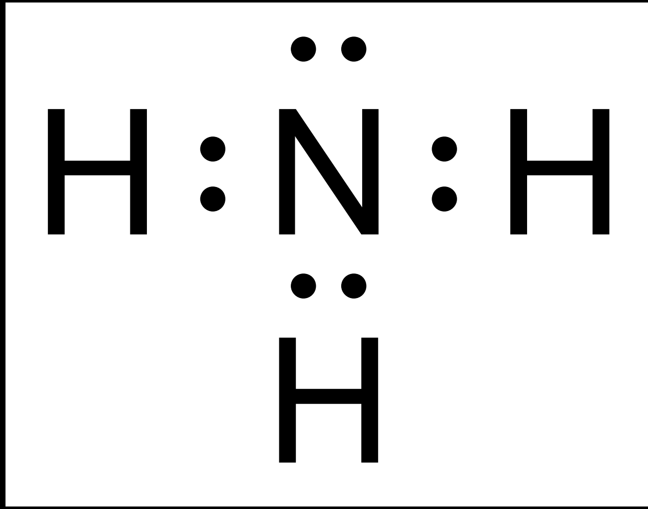
Lewis formulas

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







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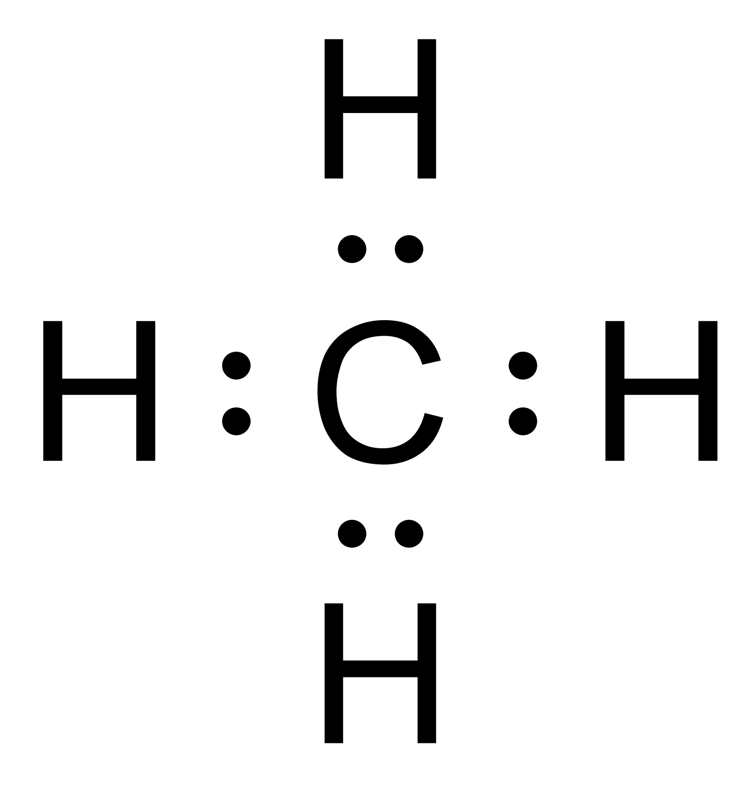
Lewis structures

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 (CH_4)

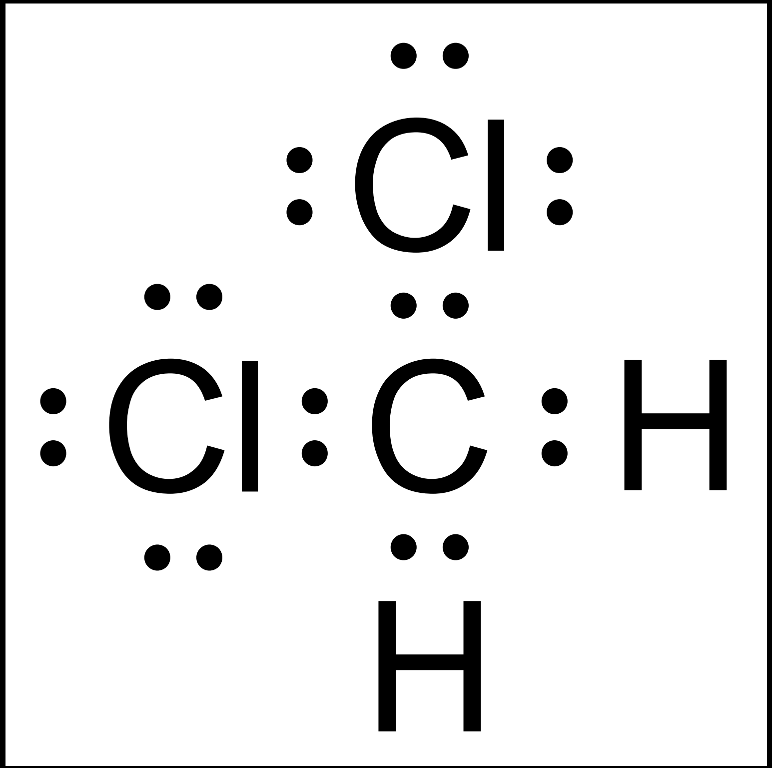
1) 4 + (4 × 1) = 8 valence electrons
2) 8 + (4 × 2) = 16 electrons needed to complete each atom's octet
3) 16 – 8 = 8 bonding electrons
4) Complete each atom's octet
5) Total number of electrons = 8



Lewis formulas

Dichloromethane (CH_2Cl_2)

1) 4 + (2 × 1) + (2 × 7) = 20 valence electrons
2) 8 + (2 × 2) + (2 × 8) = 28 electrons needed to complete each atom's octet
3) 28 – 20 = 8 bonding electrons
4) Complete each atom's octet
5) Total number of electrons = 20



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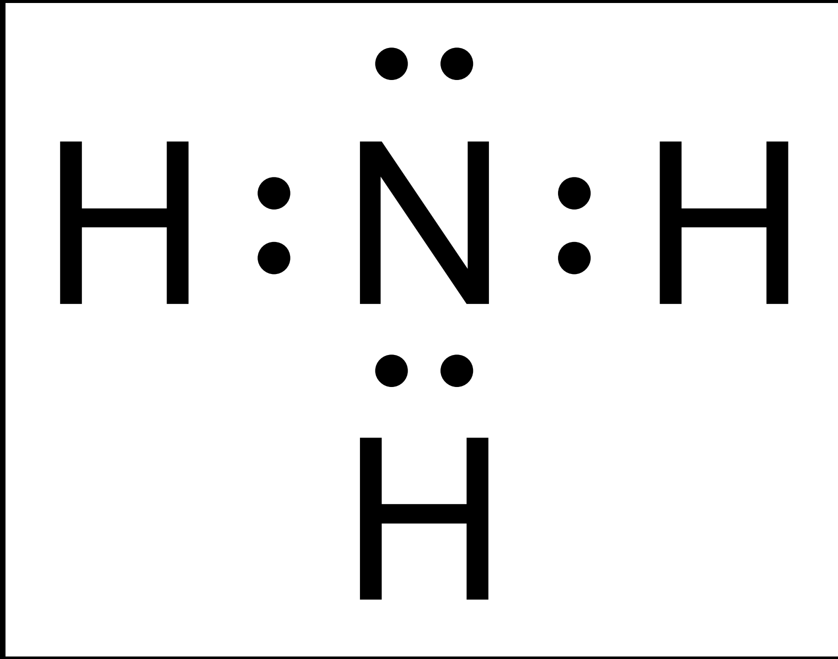
Lewis formulas

Ammonia (NH_3)

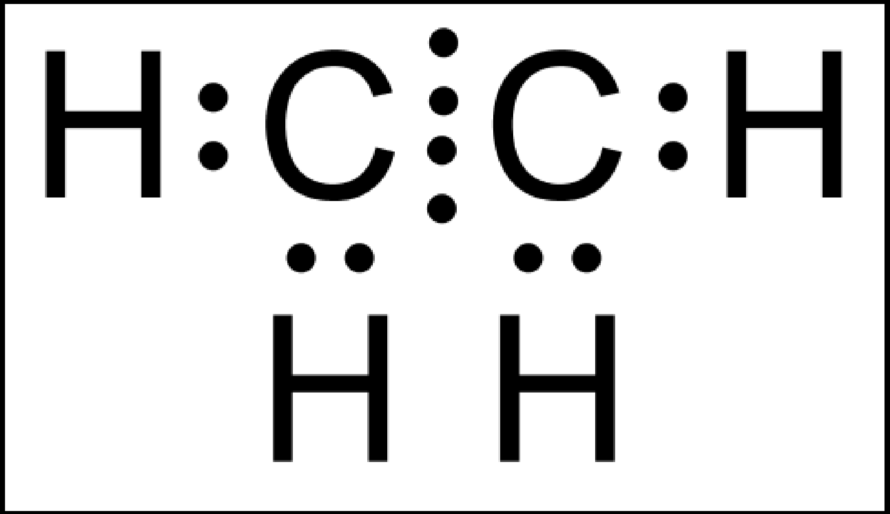
1) $5 + (3 \times 1) = 8$ valence electrons
2) $8 + (3 \times 2) = 14$ electrons needed to complete each atom's octet
3) $14 – 8 = 6$ bonding electrons
4) Complete each atom's octet
5) Total number of electrons = 8



Lewis formulas

Ethene (C_2H_4)

1) (2 × 4) + (4 × 1) = 12 valence electrons
2) (2 × 8) + (4 × 2) = 24 electrons needed to complete each atom's octet
3) 24 – 12 = 12 bonding electrons
4) Complete each atom's octet
5) Total number of electrons = 12



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Coordination bonds

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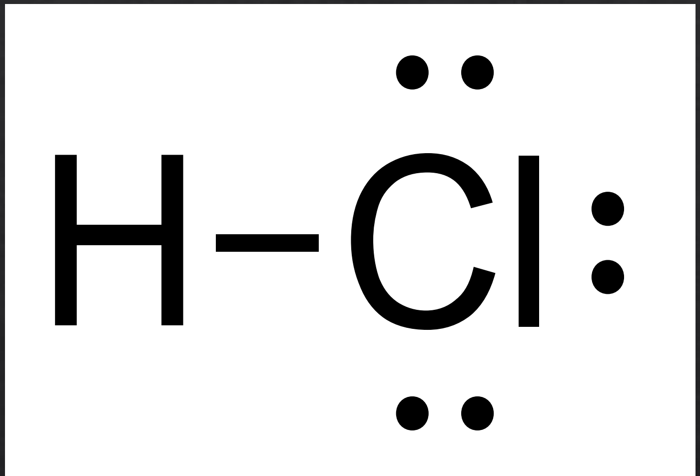
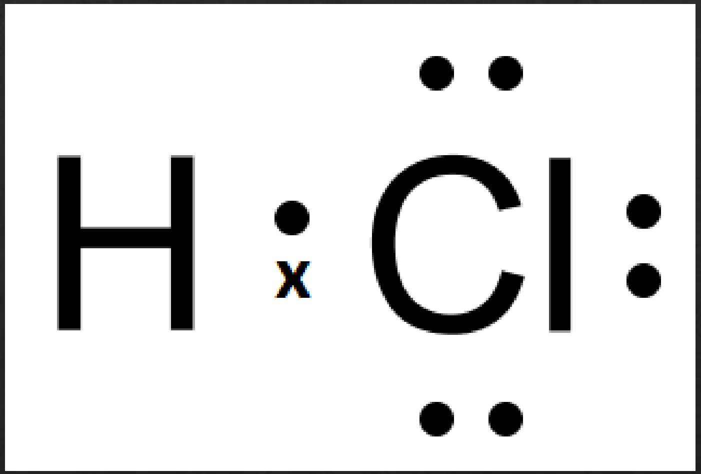
Coordination bonds

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

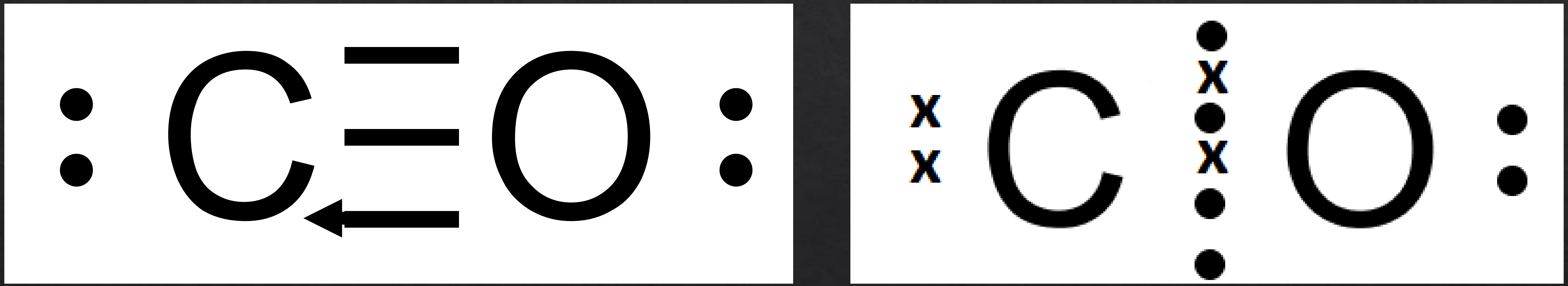




In a coordination bond, one atom contributes both the bonding electrons to the bond.

Coordination bonds

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



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

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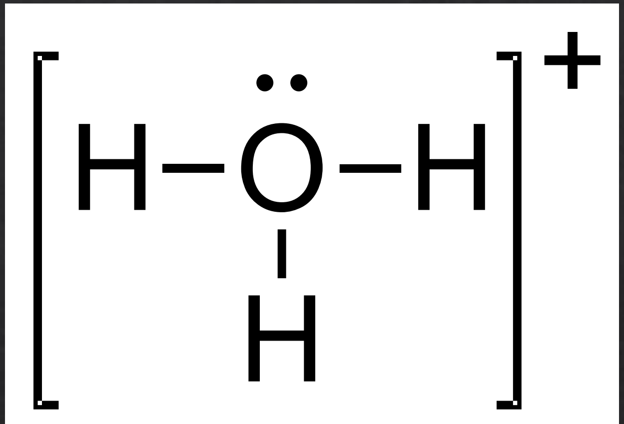
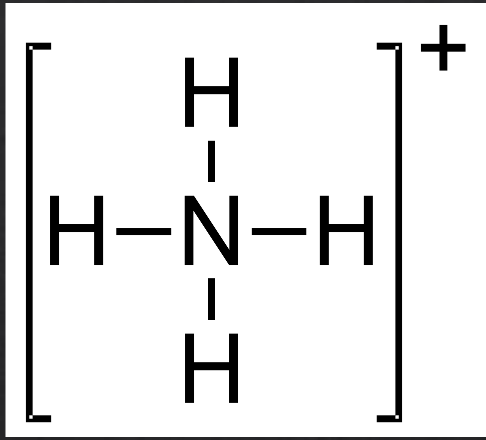
Coordination bonds

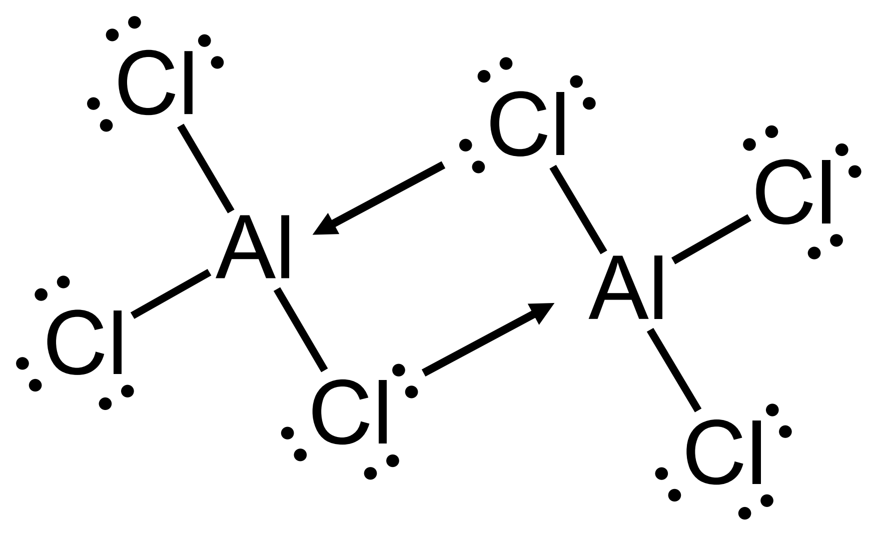
Hydronium ion (H_3O^+)



Ammonium ion (NH_4^+)



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



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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.

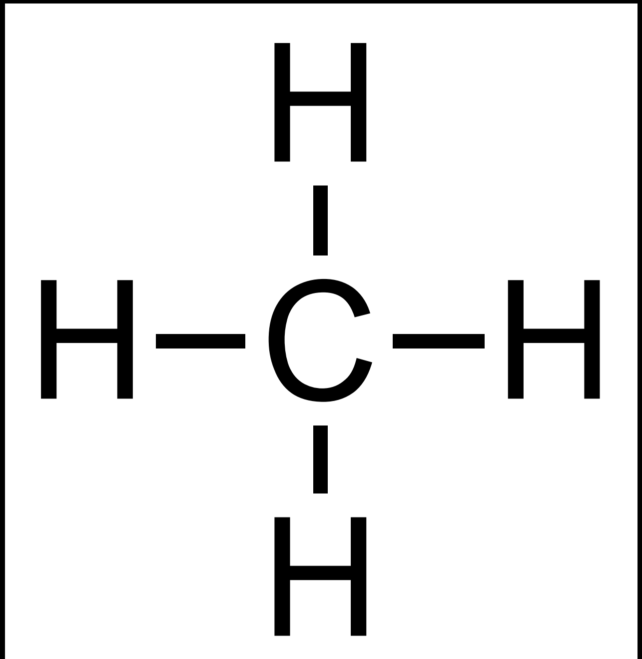
greatest repulsion → least repulsion

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.

VSEPR theory

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)

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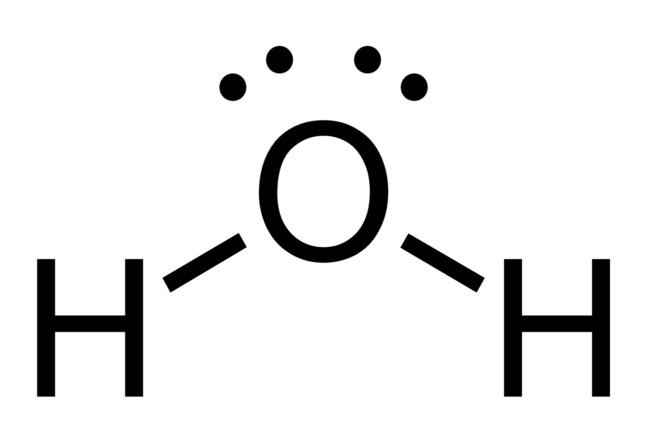
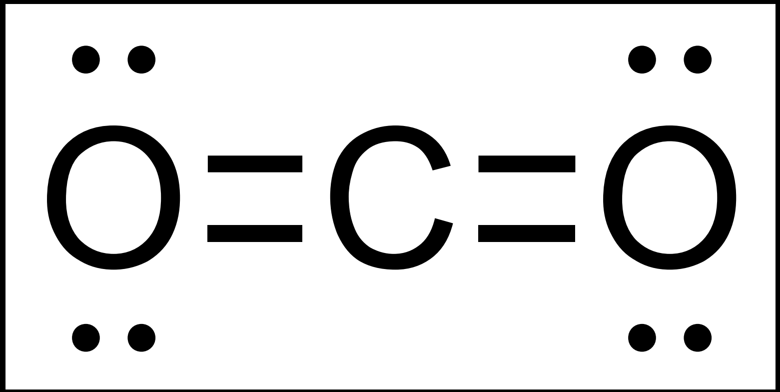
VSEPR theory





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

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

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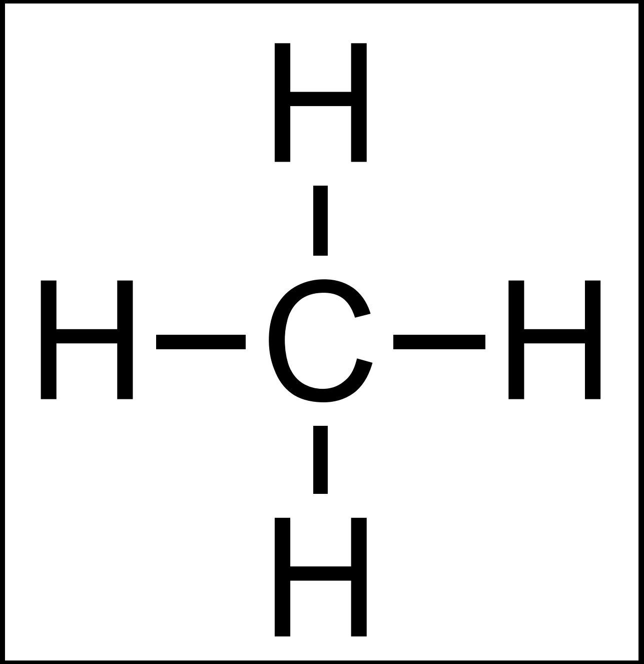
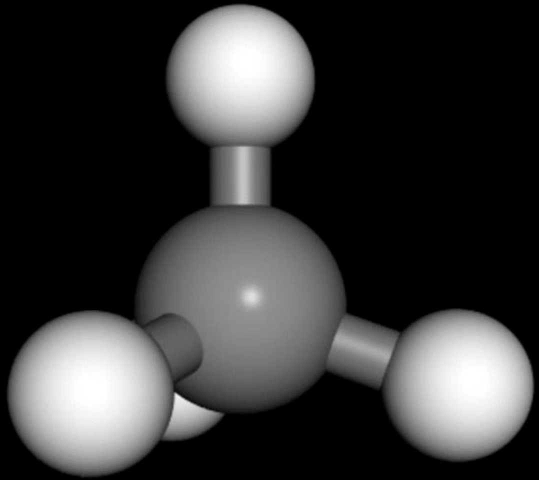
VSEPR theory

CH_4





electron domains	bonding domains	lone pairs	electron domain geometry	molecular geometry	bond angle
4	4	0	tetrahedral	tetrahedral	109.5º

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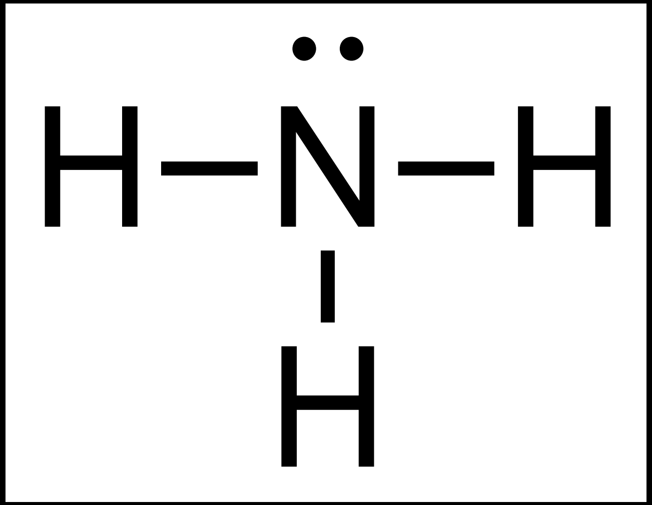
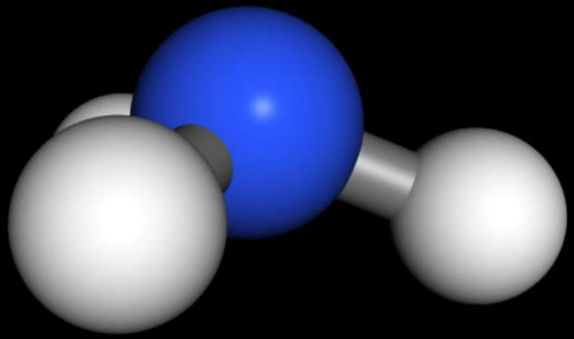
VSEPR theory

NH_3





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

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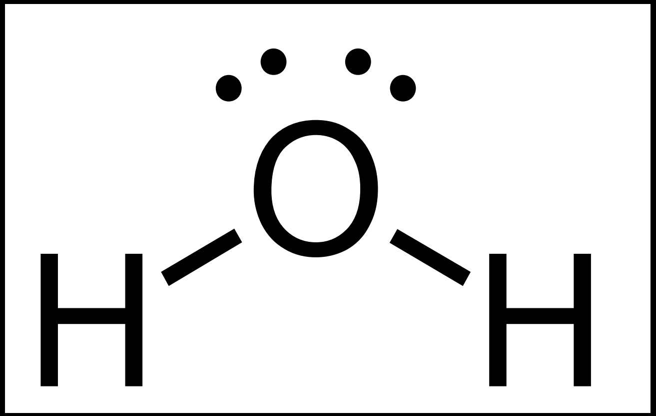
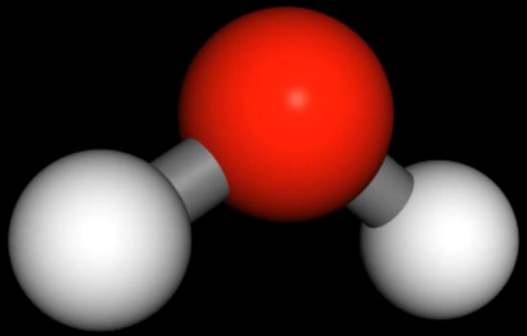
VSEPR theory

H_2O



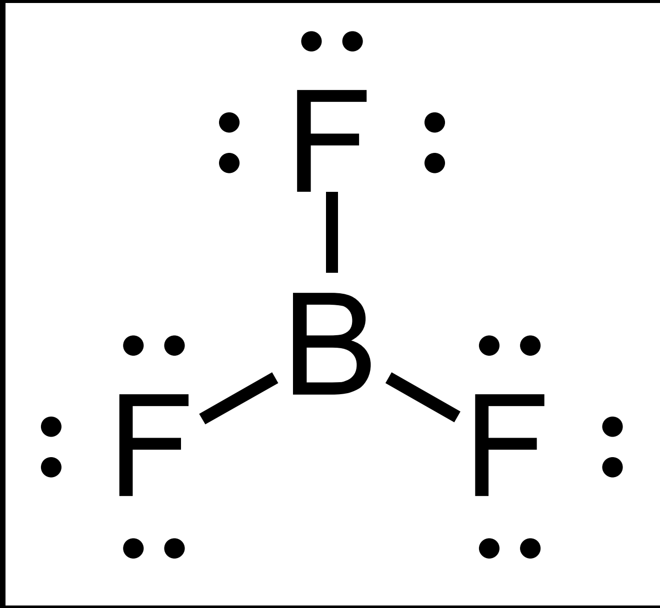


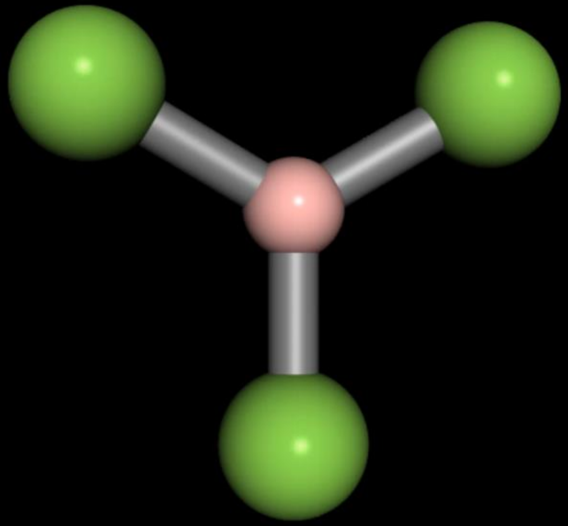
electron domains	bonding domains	lone pairs	electron domain geometry	molecular geometry	bond angle
4	2	2	tetrahedral	bent	104.5°

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

VSEPR theory

BF_3





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

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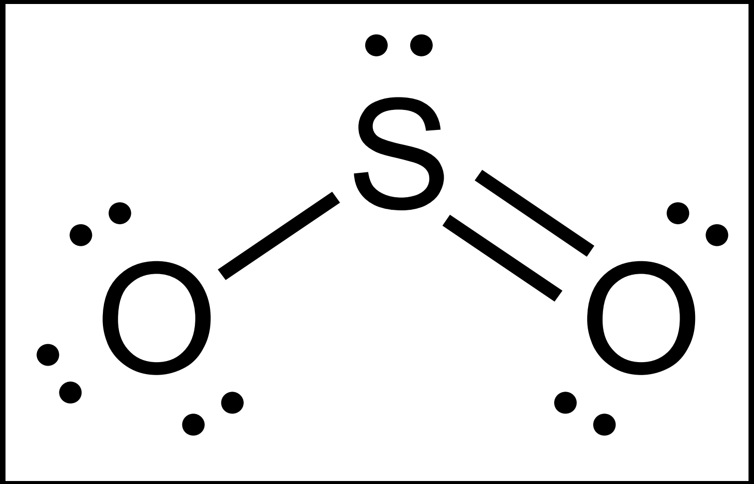
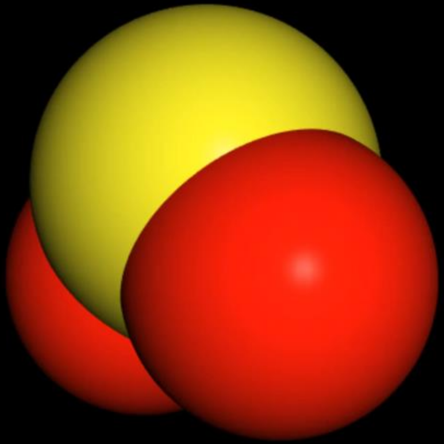
VSEPR theory

SO_2





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

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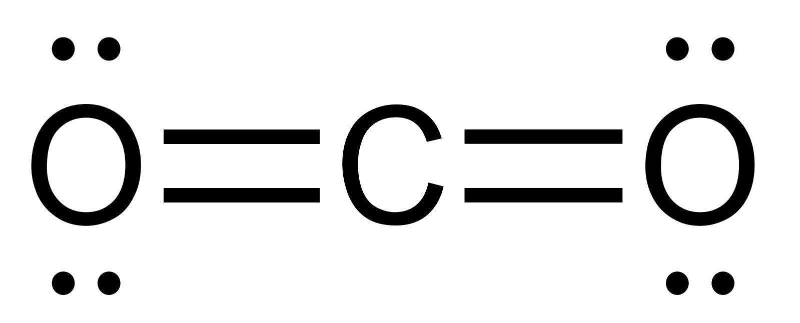
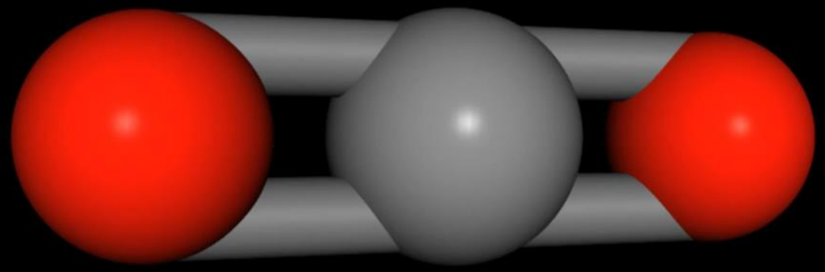
VSEPR theory

CO_2





electron domains	bonding domains	lone pairs	electron domain geometry	molecular geometry	bond angle
2	2	0	linear	linear	180º

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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_2O
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°	CO_2

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Polar and non-polar covalent bonds

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.

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

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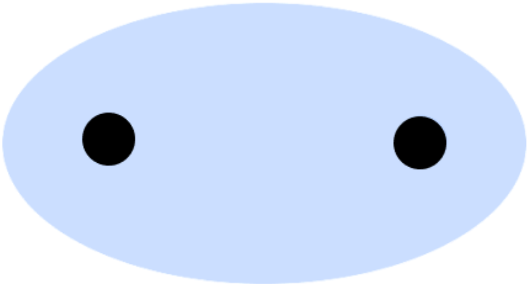
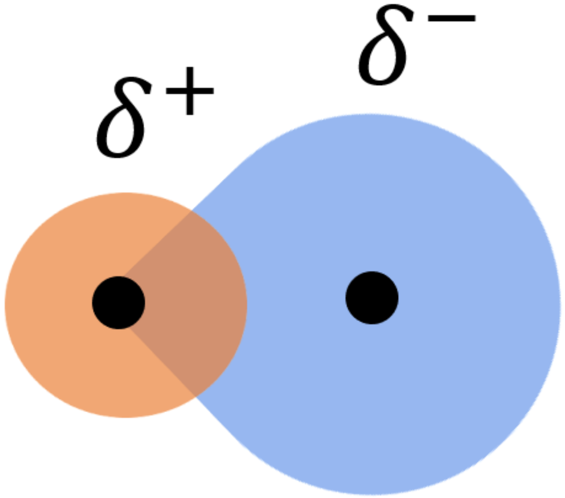
Polar and non-polar bonds

Non-polar covalent bond



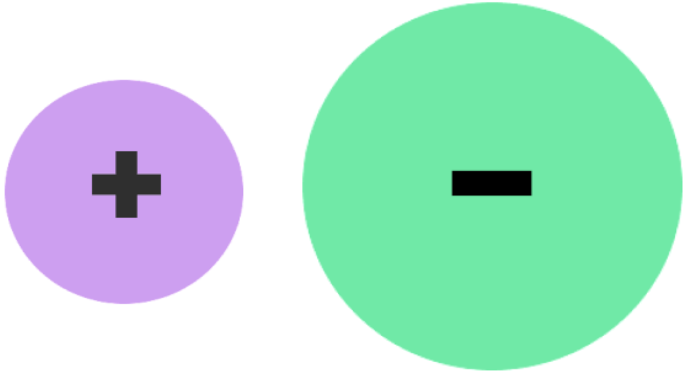
Equal sharing of electrons

Polar covalent bond



Unequal sharing of electrons

Ionic bond



No sharing of electrons in bond

Increasing difference in electronegativity

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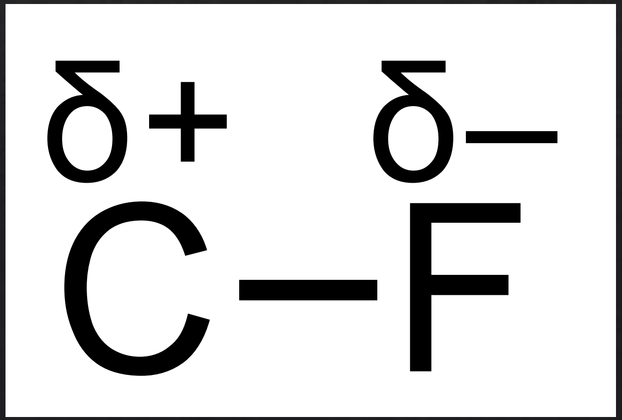
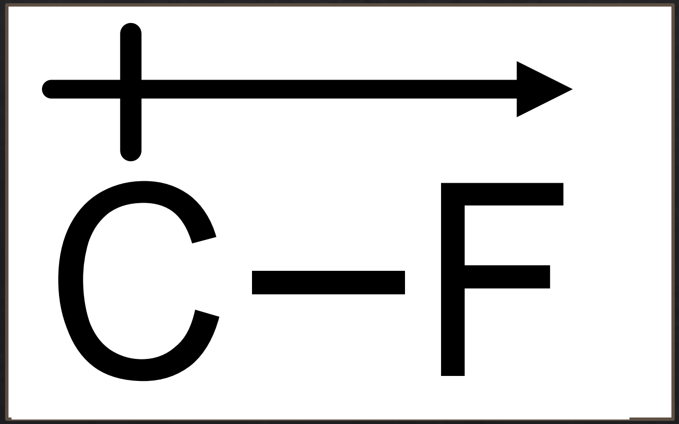
Polar and non-polar bonds

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.





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Non-polar bonds

Cl-Cl	O=O	C–H
3.2 3.2	3.4 3.4	2.6 2.2
ΔEN = 0	ΔEN = 0	ΔEN = 0.4

Polar bonds

C=O	H–Cl	C–F
2.6 3.2	2.2 3.2	2.6 4.0
ΔEN = 0.6	ΔEN = 1.0	ΔEN = 1.4

$\overset{\delta+}{C}=\overset{\delta-}{O}$

$\overset{\delta+}{H}-\overset{\delta-}{Cl}$

$\overset{\delta+}{C}-\overset{\delta-}{F}$

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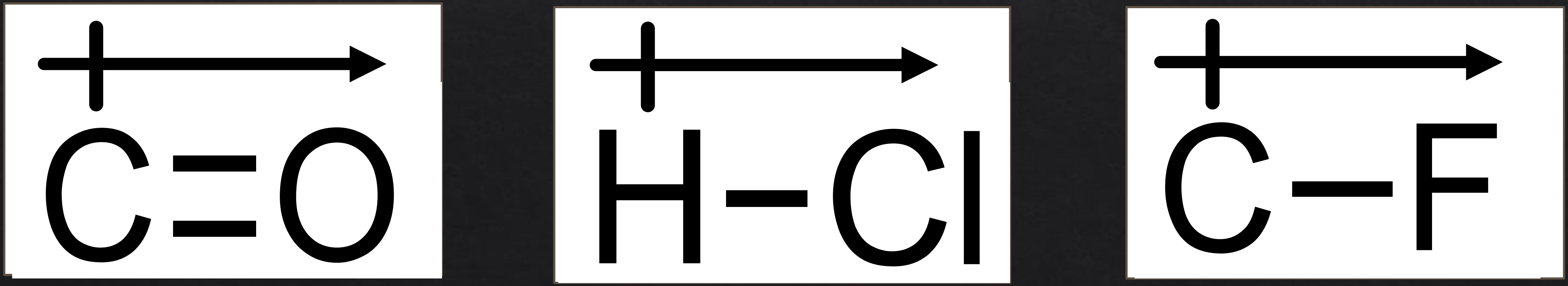
Polar bonds

C=O	H–Cl	C–F
2.6 3.2	2.2 3.2	2.6 4.0
ΔEN = 0.6	ΔEN = 1.0	ΔEN = 1.4



Polar bonds

H–N	H–O	H–F
2.2 3.0	2.2 3.4	2.2 4.0
ΔEN = 0.8	ΔEN = 1.2	ΔEN = 1.8

$\overset{\delta+}{H}-\overset{\delta-}{N}$ $\overset{\delta+}{H}-\overset{\delta-}{O}$ $\overset{\delta+}{H}-\overset{\delta-}{F}$

Polar and non-polar bonds

Non-polar (weakly polar) covalent | **Polar covalent** | **Ionic**

More or less equal sharing of electrons | **Unequal sharing of electrons** | **Formation of ions**

Cl-Cl C–H $\overset{\delta+}{C}=\overset{\delta-}{O}$ $\overset{\delta+}{H}-\overset{\delta-}{Cl}$ $\overset{\delta+}{C}-\overset{\delta-}{F}$ $\overset{\delta+}{H}-\overset{\delta-}{F}$

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Polar and non-polar molecules

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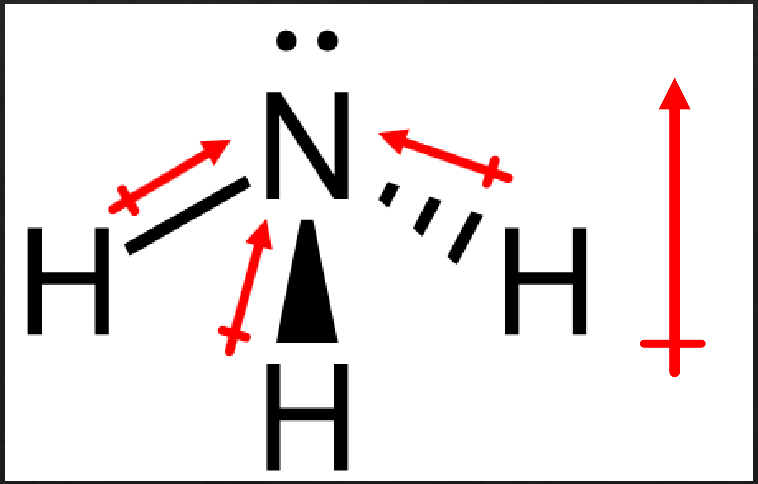
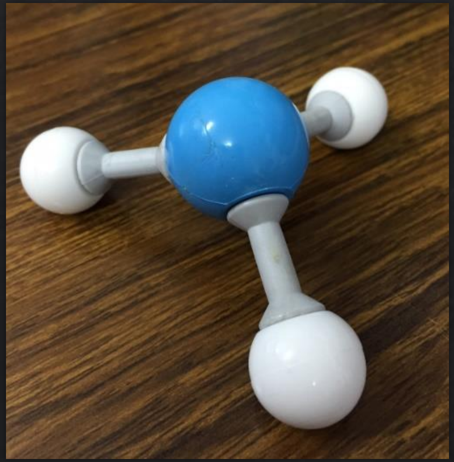
Polar and non-polar molecules

The polarity of a molecule depends on two factors:

1. The presence of polar bonds in the molecule.
2. The geometry of the molecule.

Polar molecules have a net dipole moment.





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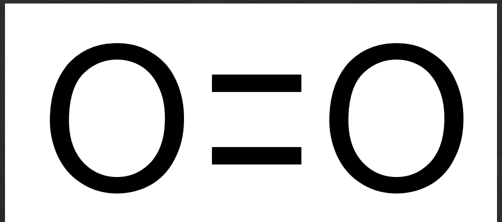
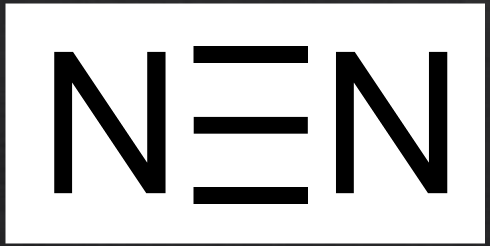
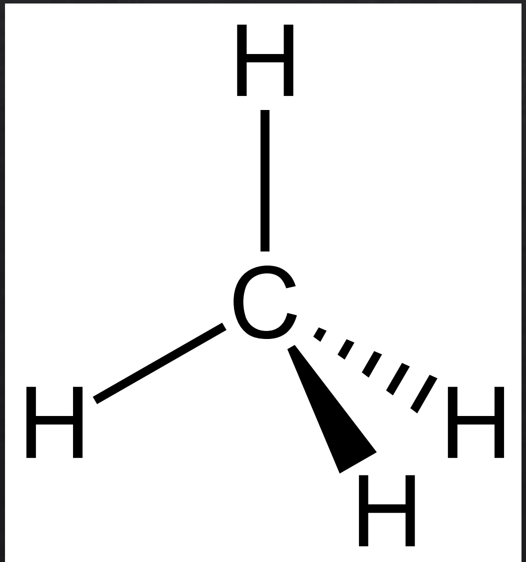
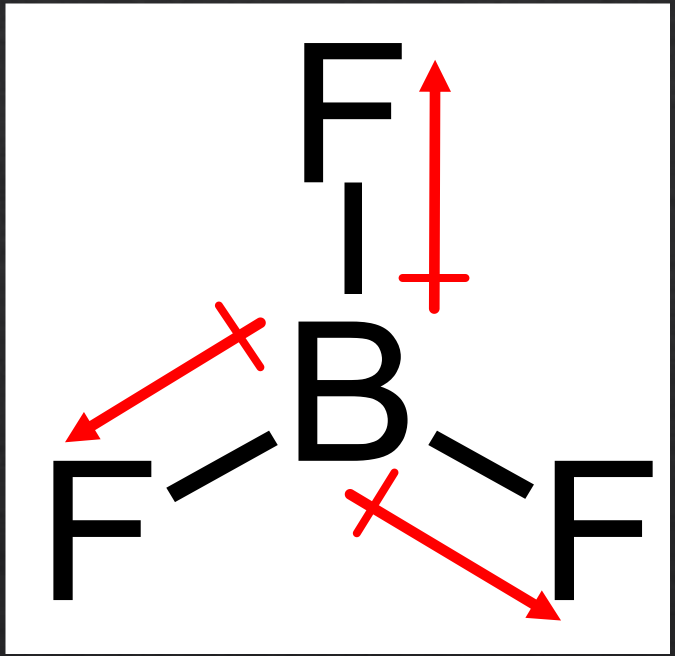
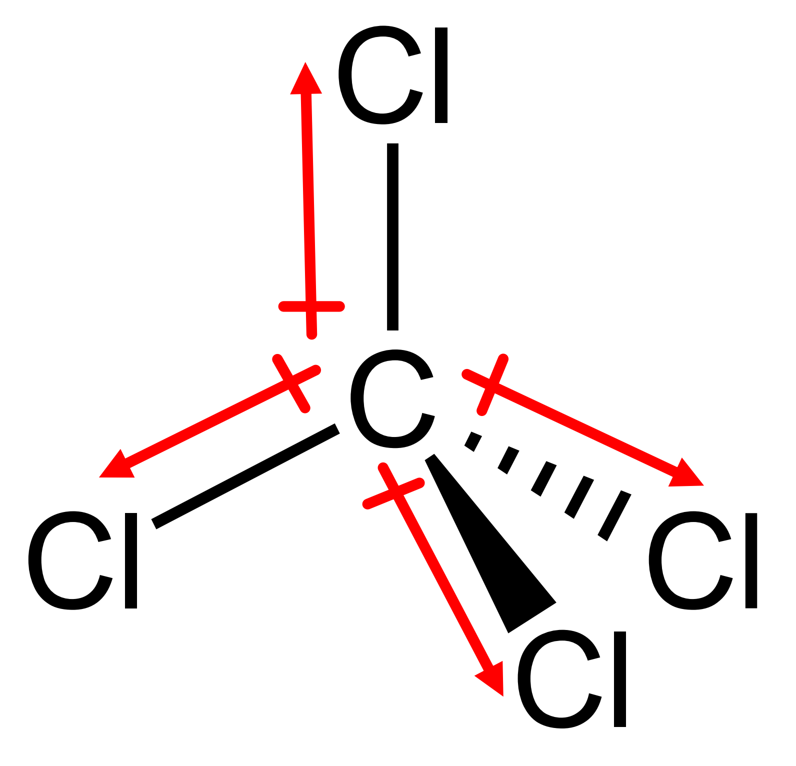
Non-polar molecules

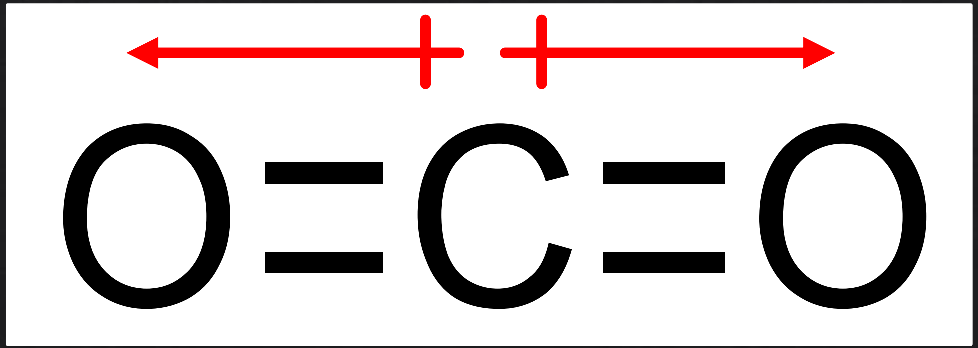












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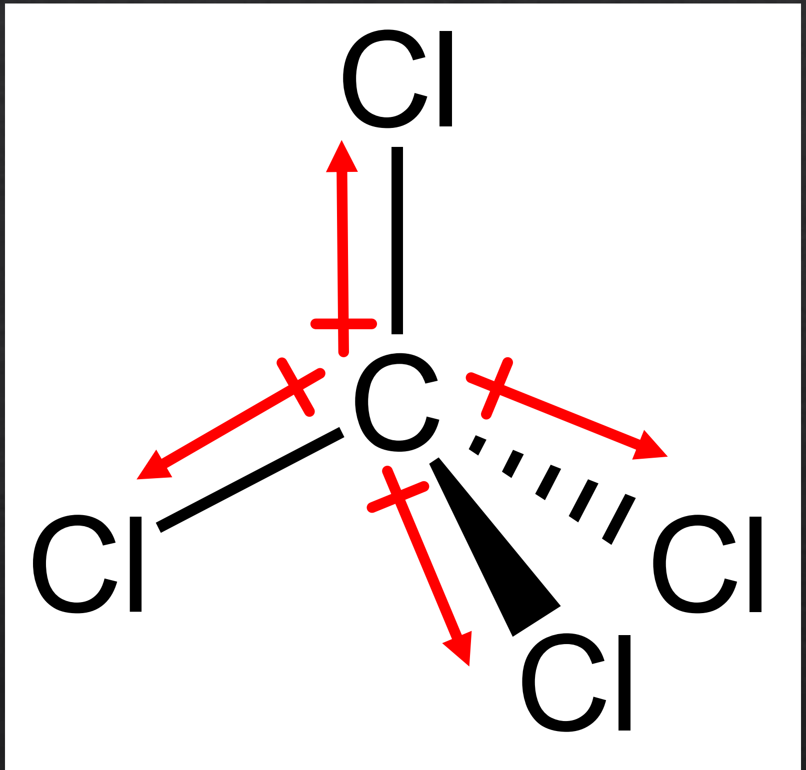
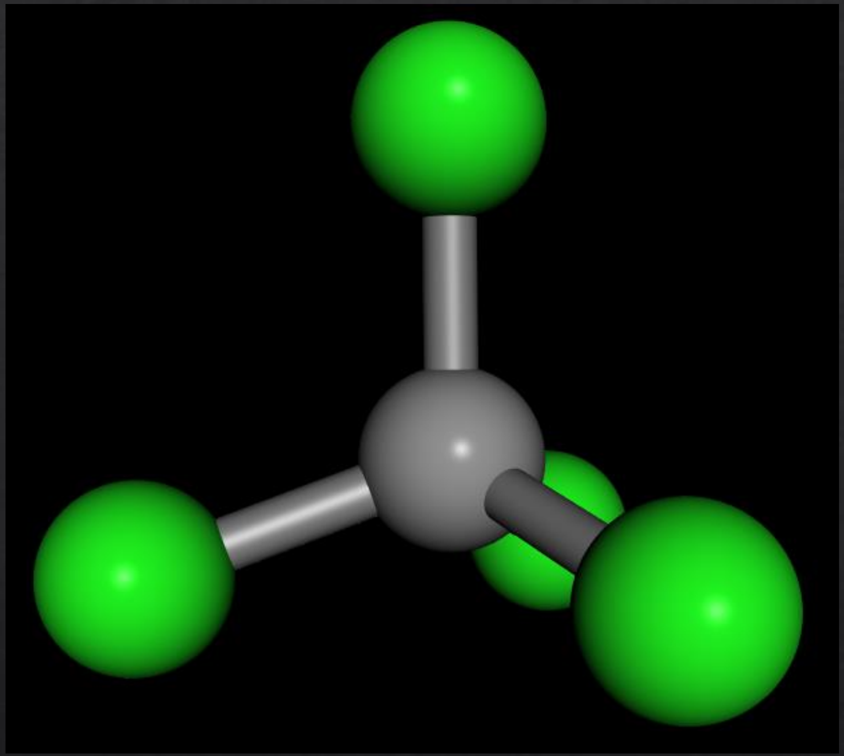
Non-polar molecules

Non-polar molecule with polar bonds (CCl_4)

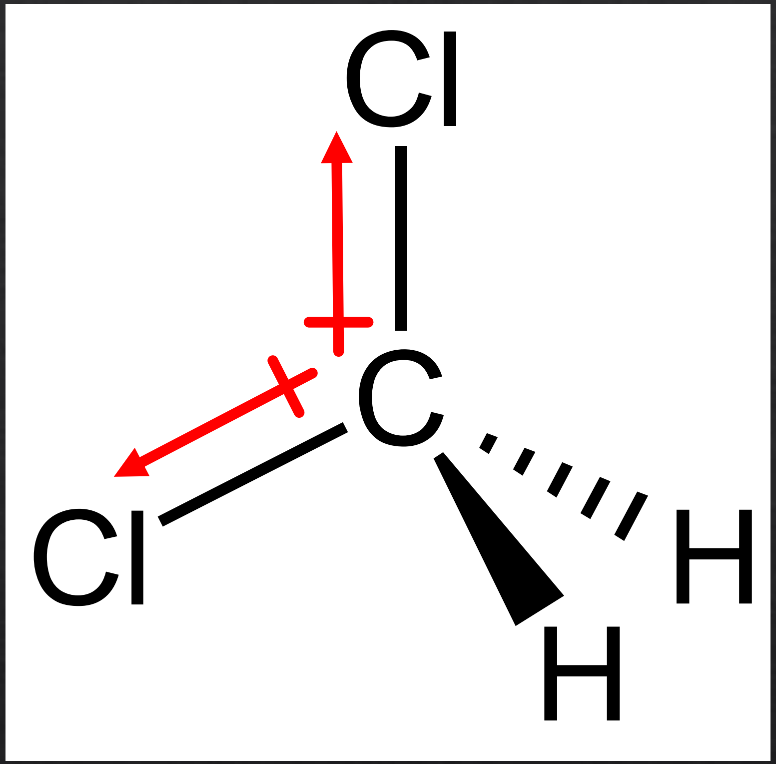


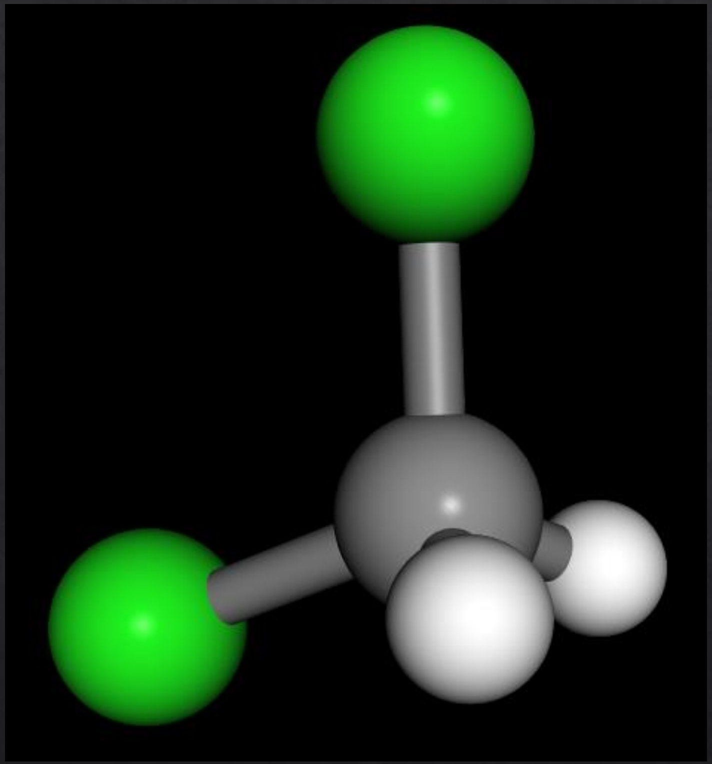


CCl_4 is a non-polar molecule; the bond polarities cancel out, therefore, it has no net dipole moment.

Polar molecules

Polar molecule (CH_2Cl_2)





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

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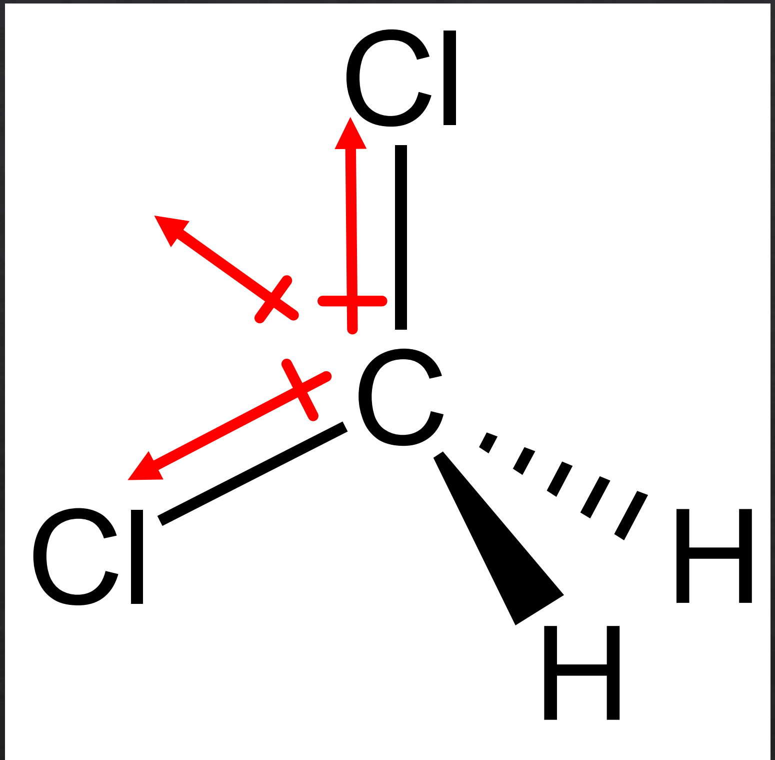
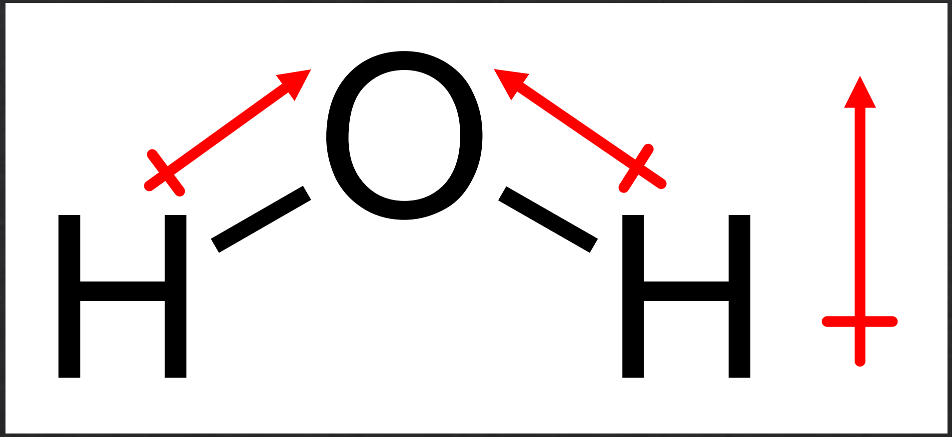
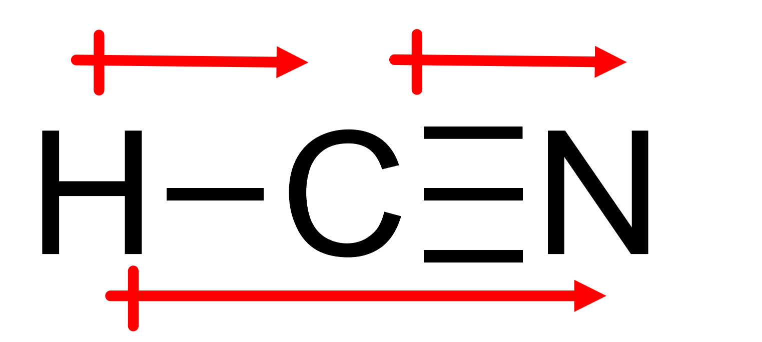
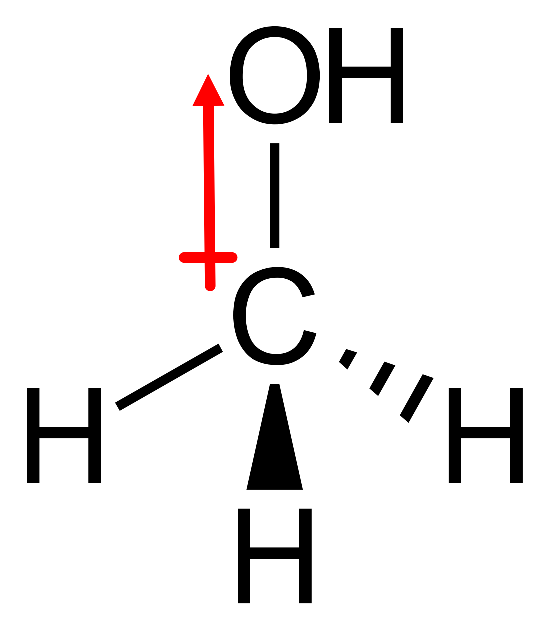
Polar molecules









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_3Cl, 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 (NH_3).

Bent molecules are usually polar if they contain polar bonds (H_2O).

Polar and non-polar molecules

Linear molecules with the same atom bonded to the central atom are non-polar (CO_2).

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 (H_2, Cl_2, N_2).

Diatomic molecules with different atoms bonded together are usually polar (HCl, HF).

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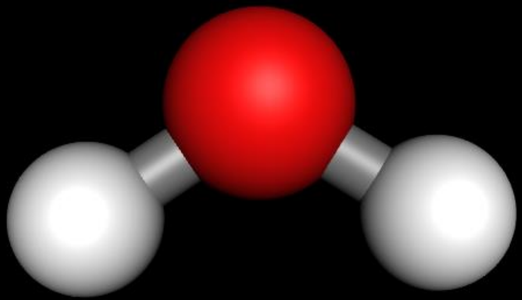
Properties of molecular compounds and giant covalent structures

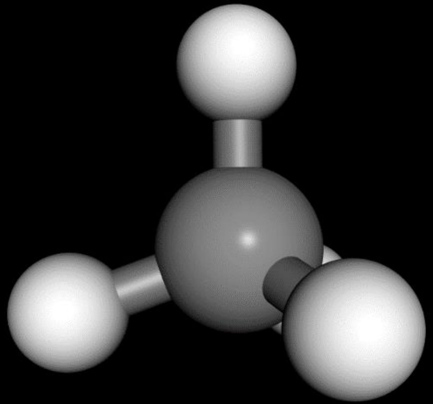
Properties of covalent substances

Covalent substances can exist as individual molecules or as giant covalent (network covalent) structures.

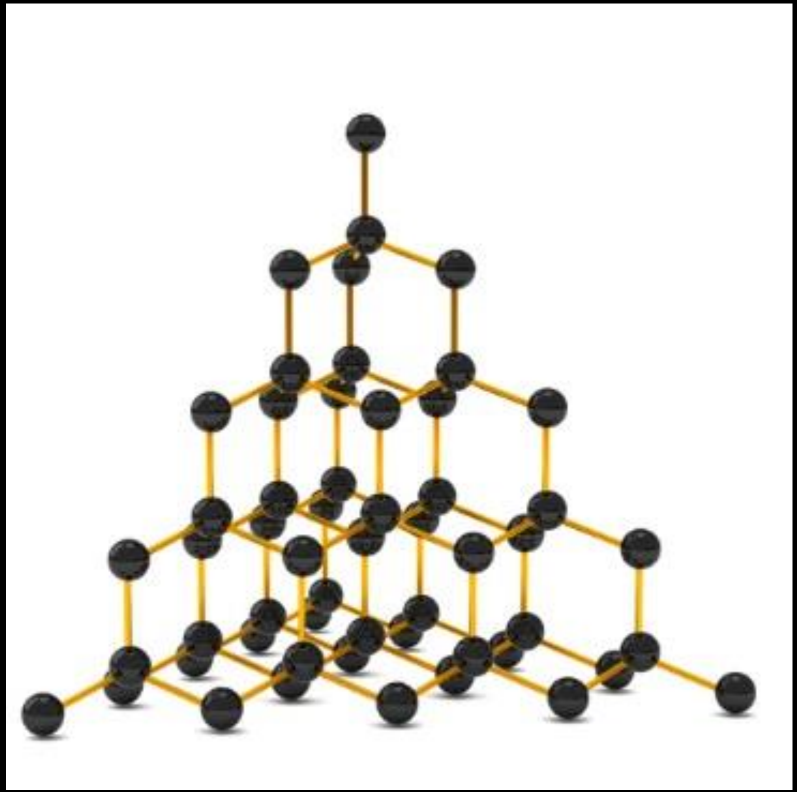
H_2O



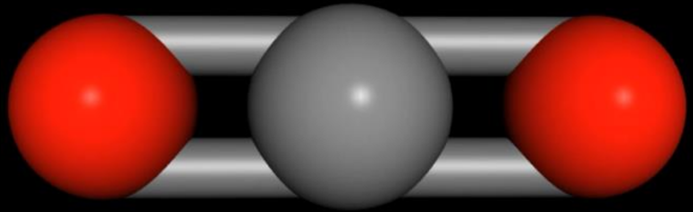
CH_4



Diamond



CO_2



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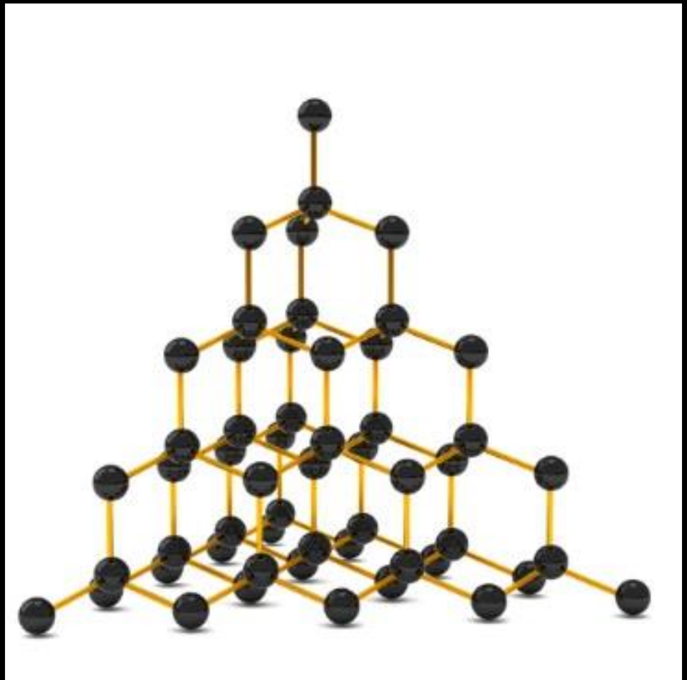
Properties of covalent substances

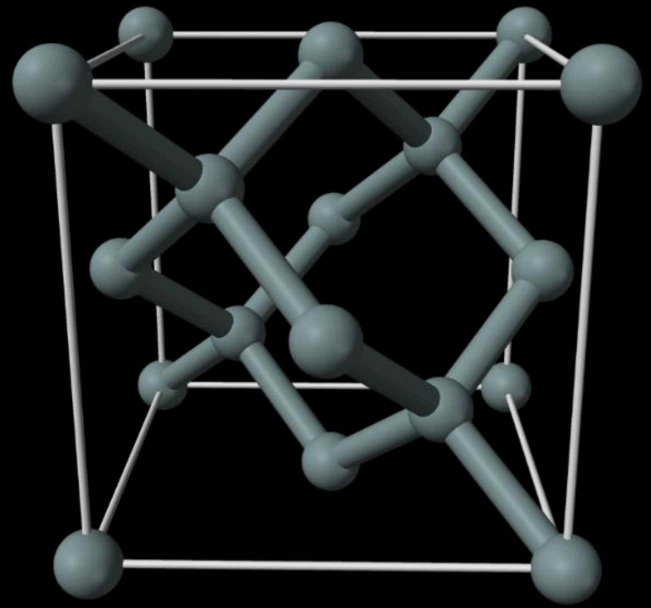
Compound	Molar mass g mol^{-1}	Melting point (°C)	Boiling point (°C)	State at room temperature (25 °C)
CH_4	16.04	-183	-162	Gas
CO_2	44.01	sublimes at -78.5		Gas
CCl_4	153.81	-22.9	76.7	Liquid
NH_3	17.04	-77.7	-33.3	Gas
H_2O	18.02	0	100	Liquid

Properties of covalent substances

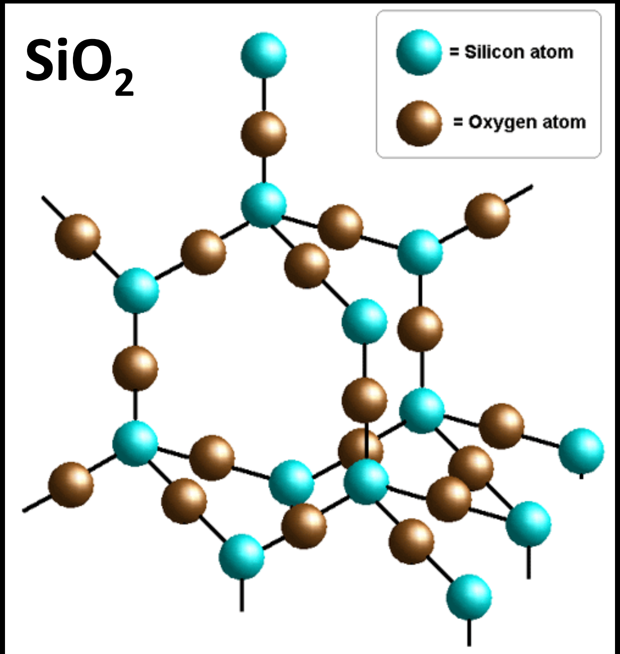
Giant covalent substances have high melting and boiling points.



Diamond
M.P. 3550 °C
B.P. 4830 °C



Silicon
M.P. 1414 °C
B.P. 3265 °C



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

Properties of covalent substances

Property	Molecular elements / compounds	Giant covalent structures
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 (don’t have free moving ions or delocalised electrons)	Poor electrical conductors (don’t have free moving ions or delocalised electrons)

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Allotropes of carbon

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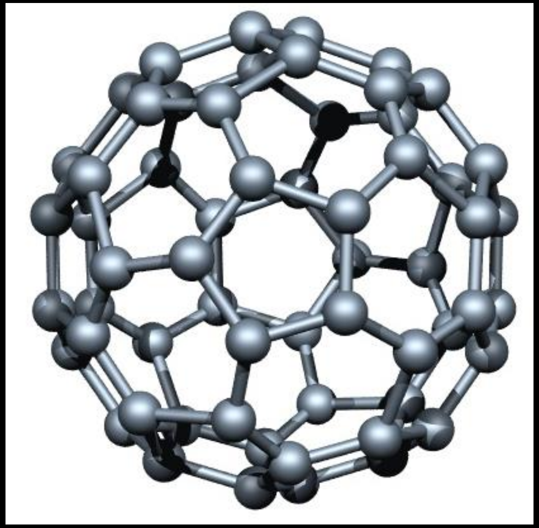
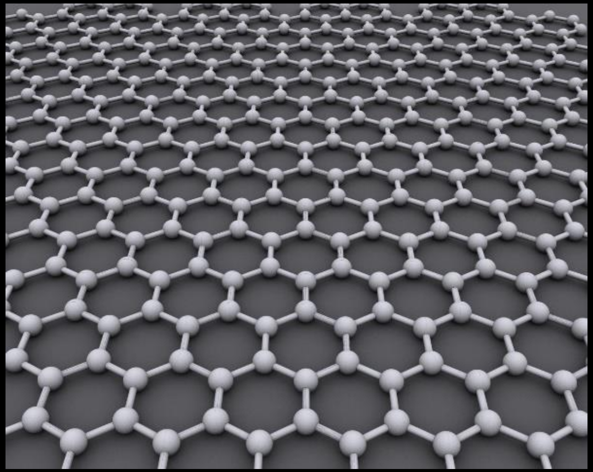
Allotropes of carbon

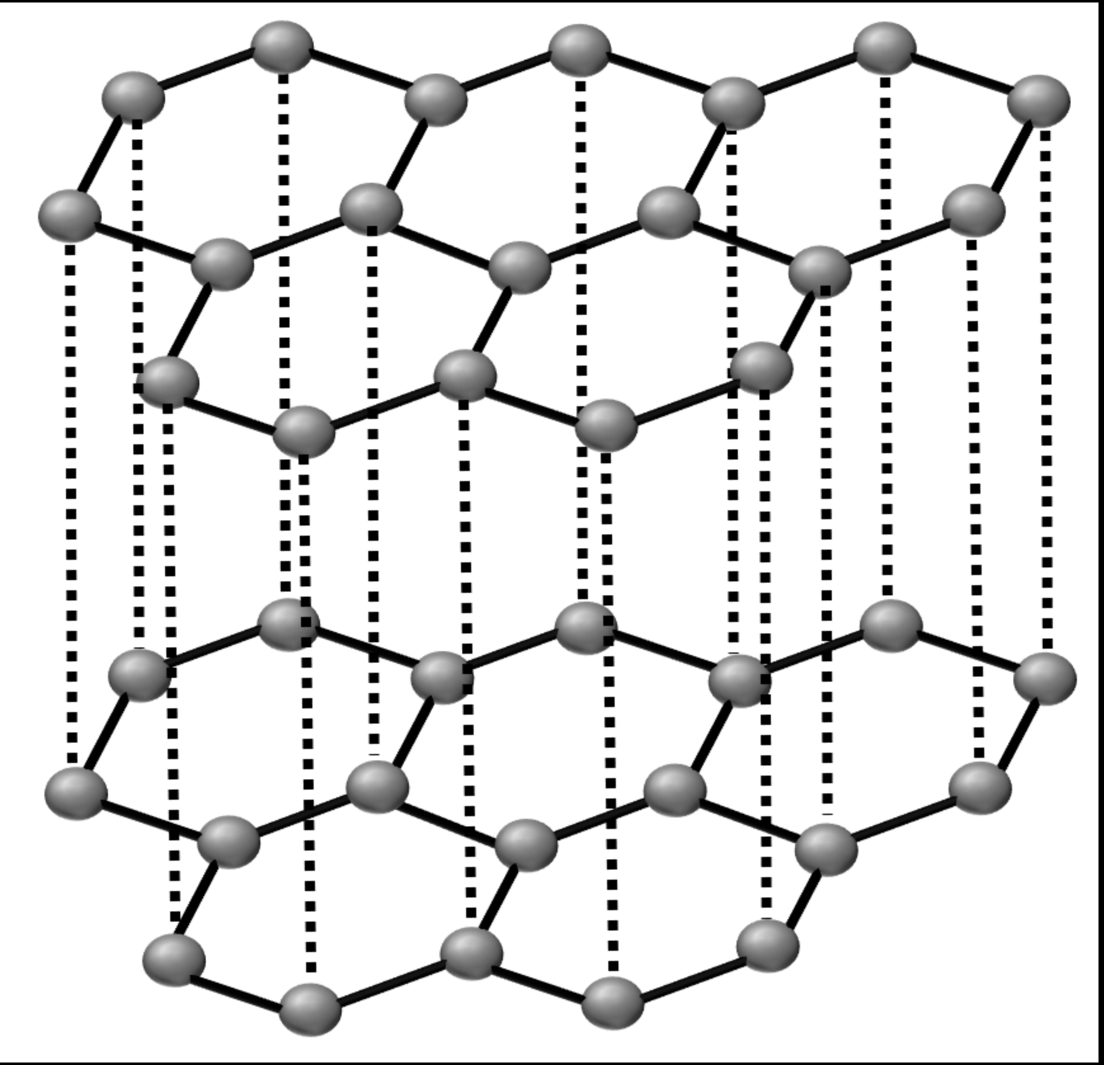
Allotropes are different forms of the same element in the same physical state.

Carbon has 4 allotropes – graphite, diamond, Fullerene C_{60} and graphene.





Graphite



Graphite has a layered structure.

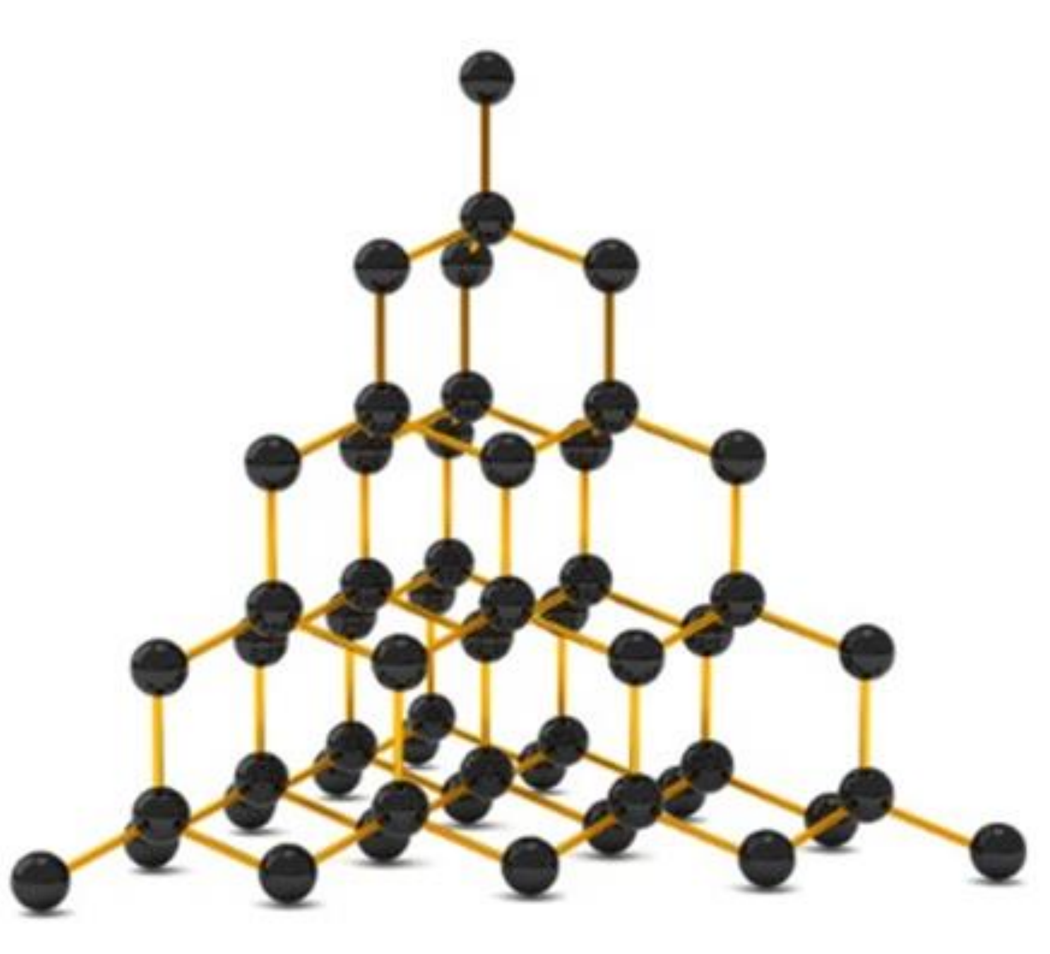
The layers are held together by weak intermolecular forces – the layers can slide over one another.

Each carbon atom is bonded to 3 other carbon atoms.

The bond angle is 120^{o}, trigonal planar.

Good conductor of electricity (delocalised electrons).

Diamond



Diamond has a giant covalent 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).

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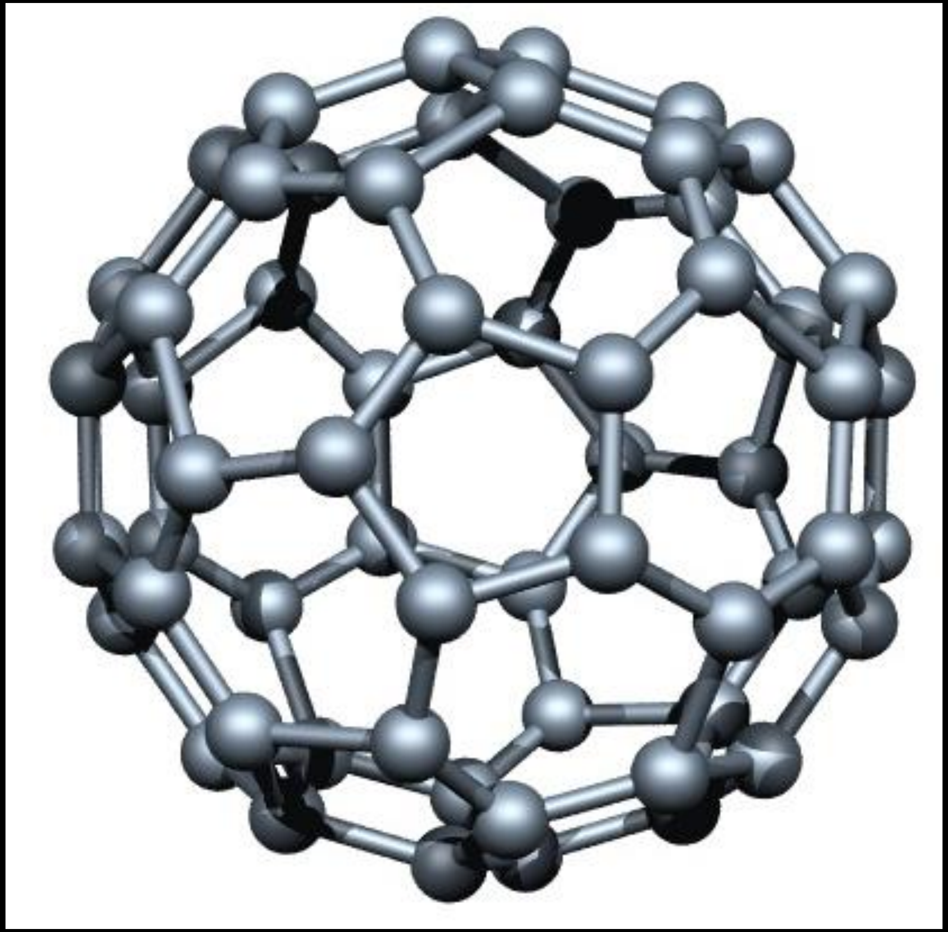
Fullerene C_{60}



Structure consists of 12 pentagons and 20 hexagons.

Each carbon atom is bonded to 3 other carbon atoms.

Shows some electrical conductivity.

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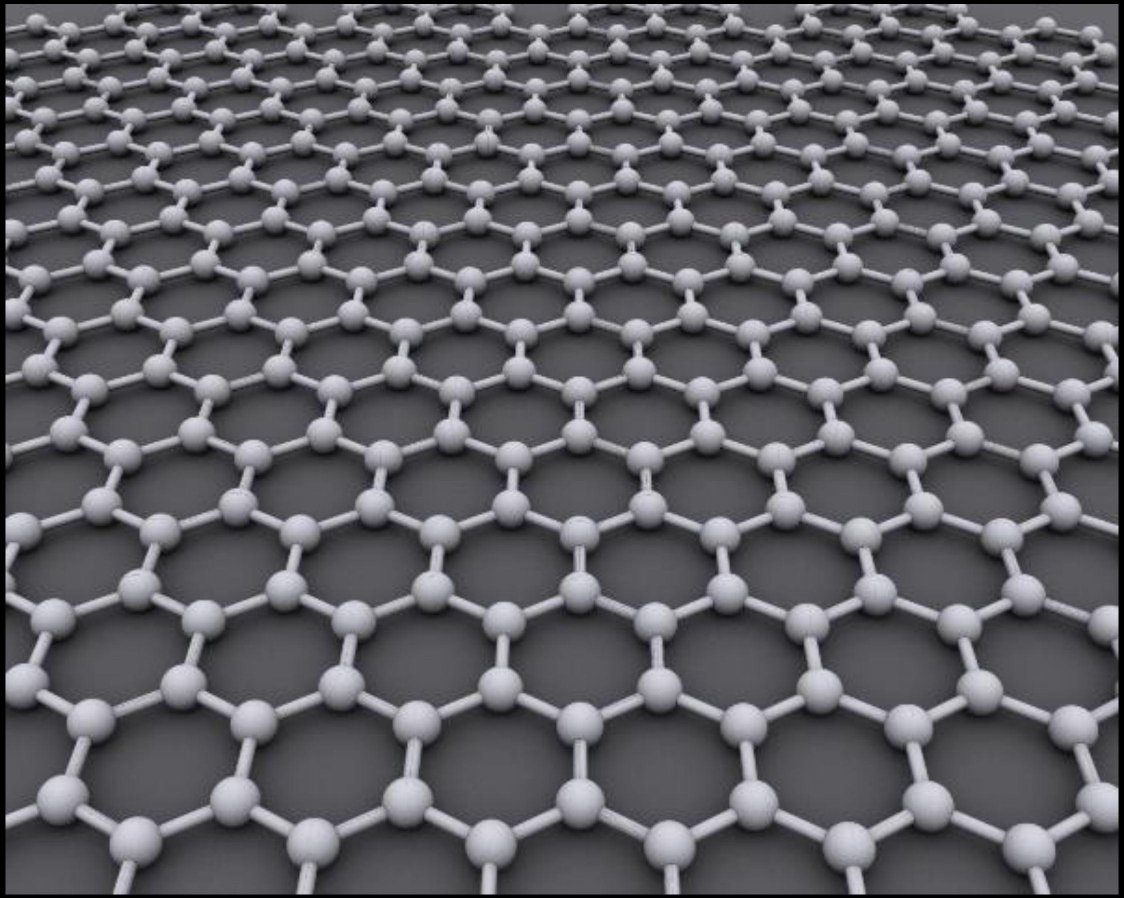
Graphene



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.

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Summary

Allotrope	Graphite	Diamond	Fullerene C_{60}	Graphene
Structure	Layered (weak IMF between layers)	Giant covalent	12 pentagons 20 hexagons	Honeycomb 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^{o} trigonal planar	109.5^{o} tetrahedral	120^{o} trigonal planar	120^{o} trigonal planar
Electrical conductivity	High	None	Medium	Very high
Thermal conductivity	Low	Very good	Low	Very good (higher than diamond)
Delocalised electrons	Yes	No	Yes	Yes
Properties	Very soft	Very hard	Very light and strong	High melting point, flexible, stronger than steel

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Intermolecular forces

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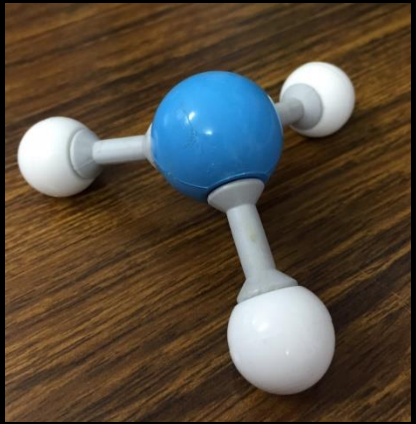
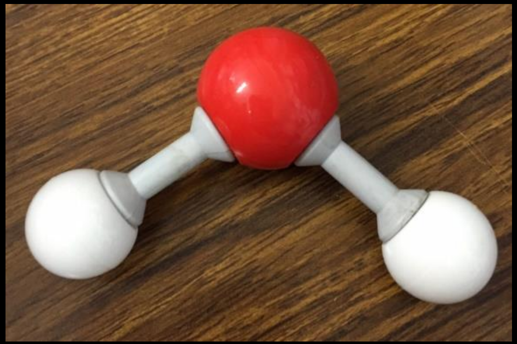
Intermolecular forces

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)

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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)

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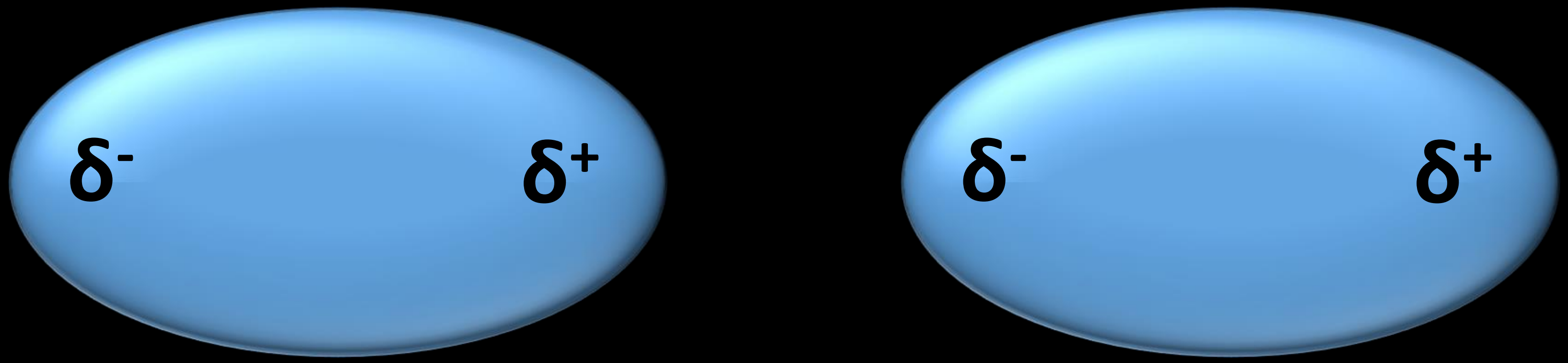
Intermolecular forces

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

London dispersion forces

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.



instantaneous dipole induced dipole

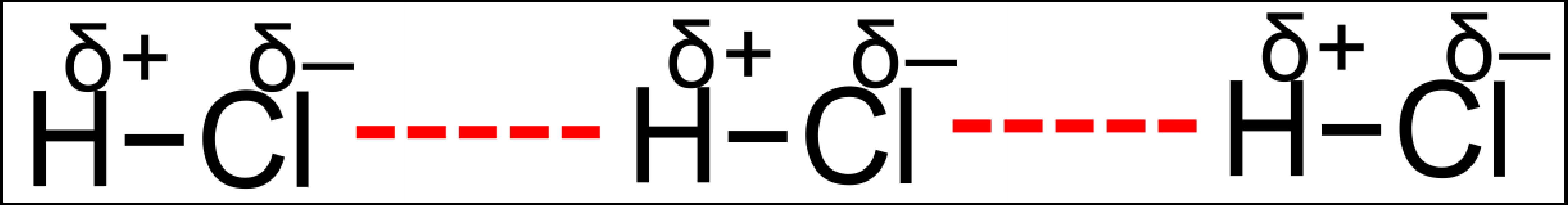
London dispersion forces

	Molar mass ($g\ mol^{-1}$)	Boiling point (^{o}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.

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Dipole-dipole forces

Compound	Molar mass (g mol^{-1})	Dipole moment (D)	Boiling point (K)
CH_3OCH_3	46.07	1.30	248
CH_3Cl	50.48	1.87	249
CH_3CHO	44.05	2.69	294
CH_3CN	41.05	3.92	355

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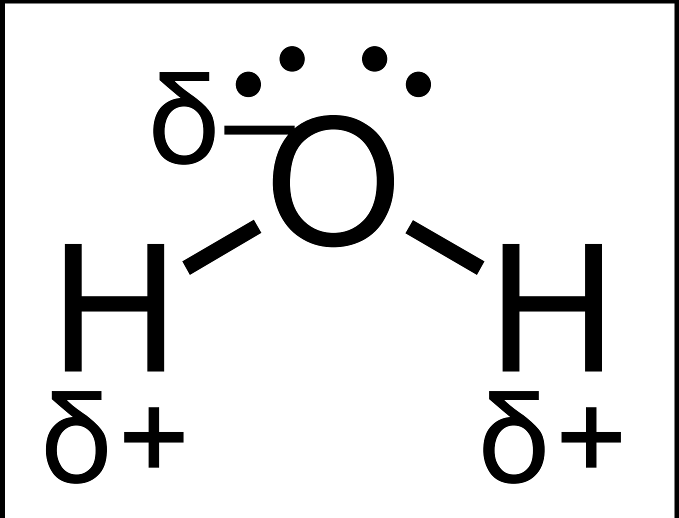
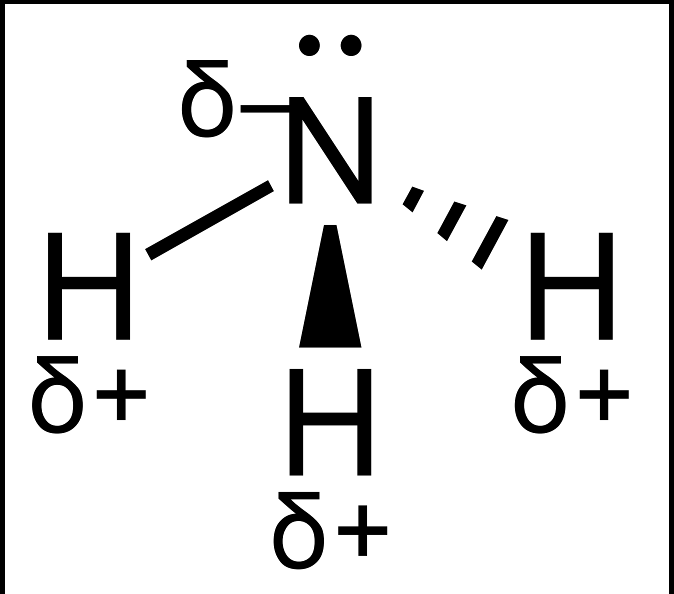
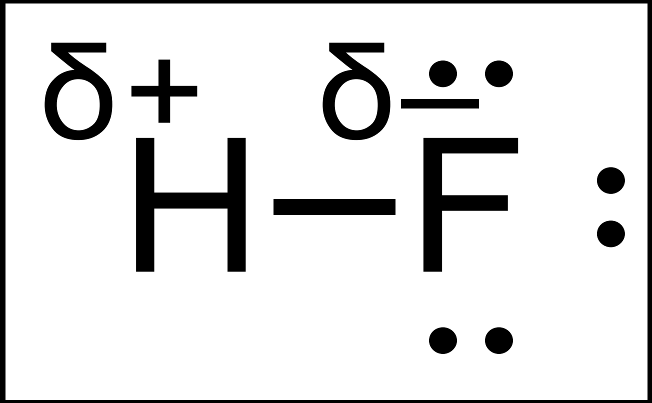
Hydrogen bonding

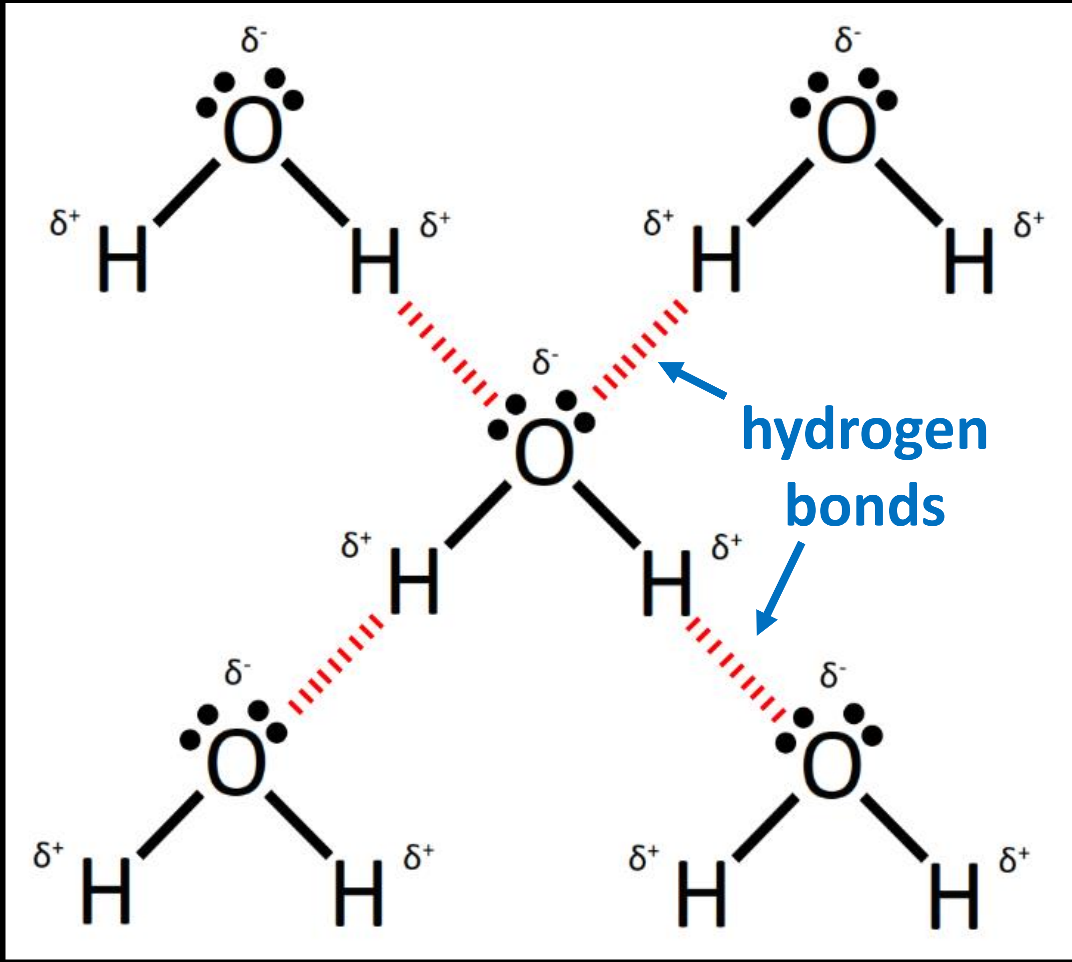
Hydrogen bonding occurs when a hydrogen atom is bonded to either a nitrogen, oxygen or fluorine atom.







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.

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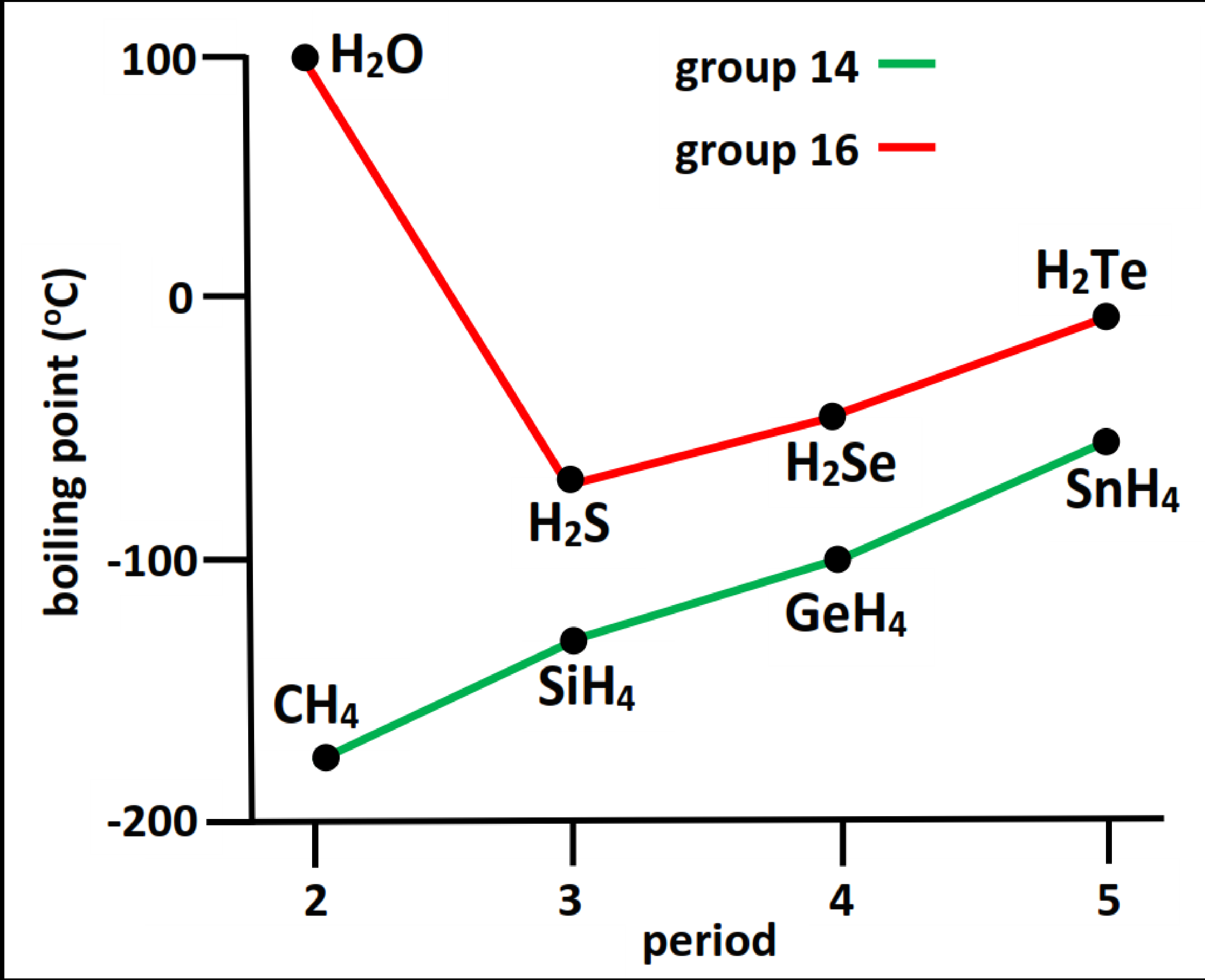
Hydrogen bonding



Intermolecular forces

Type of molecule	Intermolecular forces	Examples
Non-polar	London dispersion forces	Cl_2 H_2 N_2 O_2 CH_4 CCl_4
Polar	London dispersion forces Dipole-dipole forces	HCl HCN CH_3Cl CH_3CHO
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

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Solubility and intermolecular forces

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Polar and non-polar solvents

A polar solvent is a liquid composed of polar molecules.

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

Polar solvents	Non-polar solvents
Water H_2O	Hexane C_6H_{14}
Methanol CH_3OH	Octane C_8H_{18}
Ethanol C_2H_5OH	Benzene C_6H_6
Propanone CH_3CH_2CHO	Methylbenzene $C_6H_5CH_3$
Ethanoic acid CH_3COOH	Carbon tetrachloride CCl_4

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.

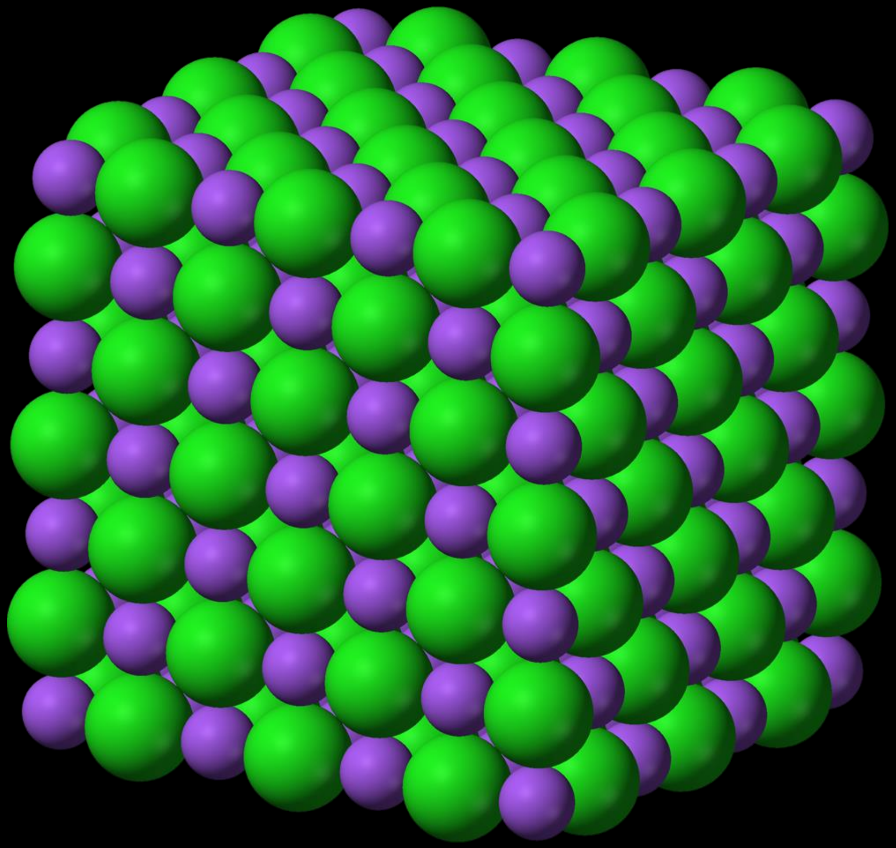
Examples:

- 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.

Ion-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.

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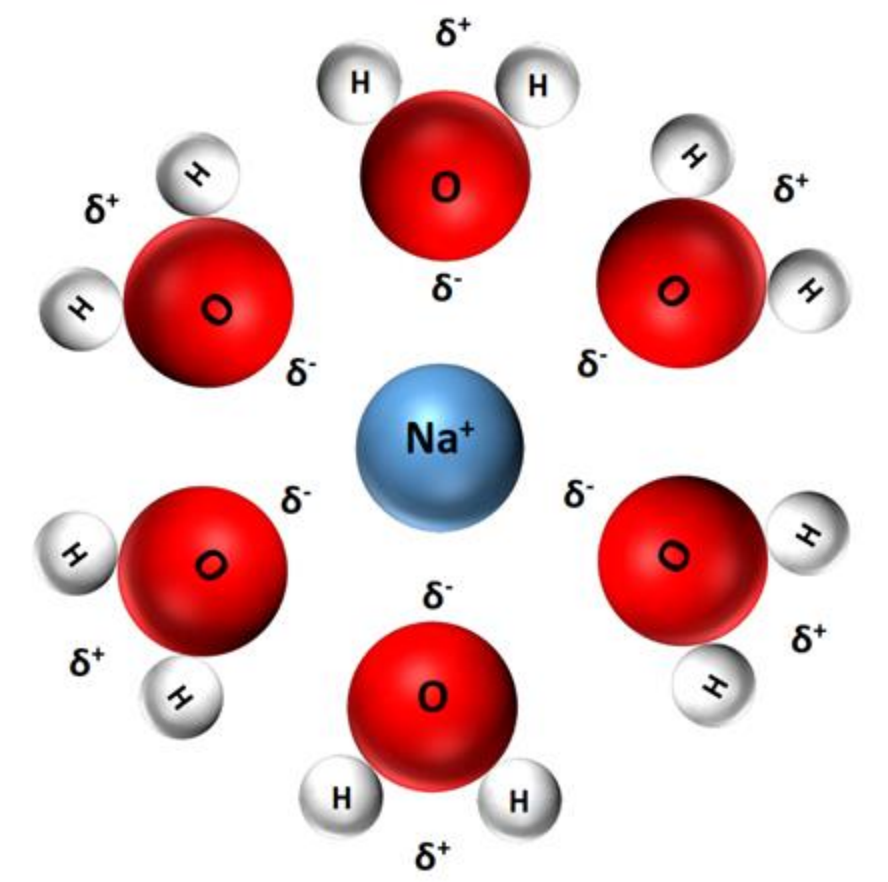
Ion-dipole forces

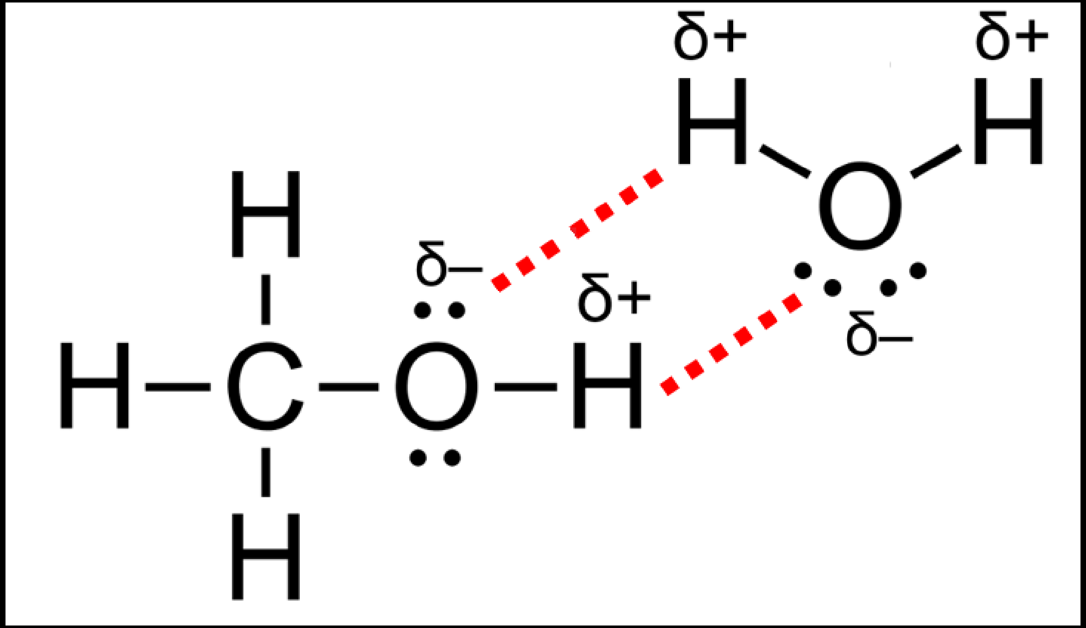
Ion-dipole forces occur between water molecules and ions in aqueous solutions.



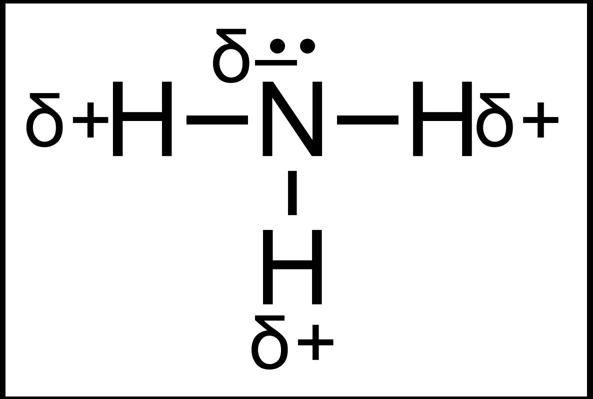
When an ionic compound dissolves in water, ion-dipole forces occur between the ions and the oppositely-charged ends of the water molecules.
The water molecules surround the ion forming a hydration shell.

Hydrogen bonds

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



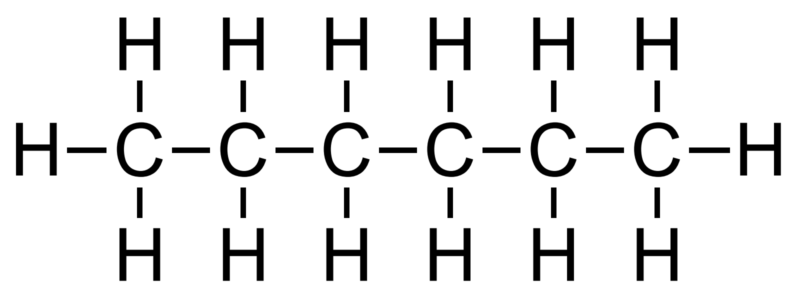
Methanol is soluble in water because it is able to form hydrogen bonds with water molecules.

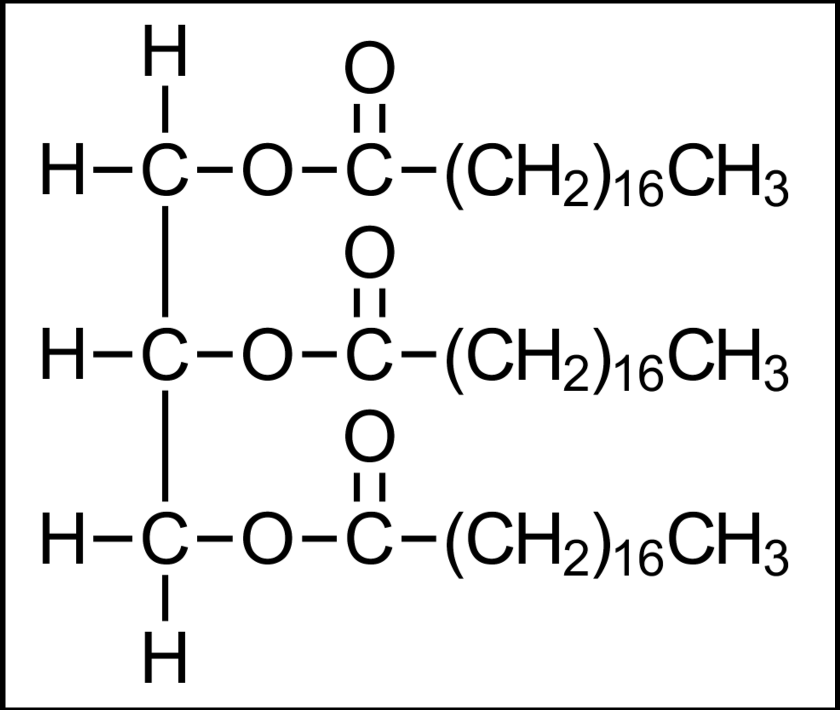


Ammonia, NH_3, is also able to form hydrogen bonds with water molecules.

London dispersion forces

London dispersion forces occur between non-polar molecules.





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

Hexane is insoluble in polar solvents.

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Summary

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.

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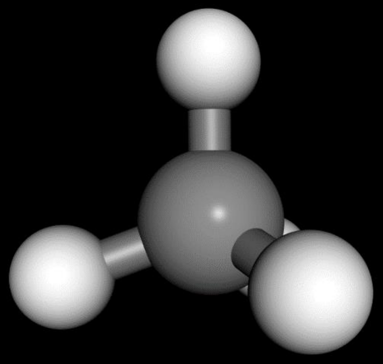
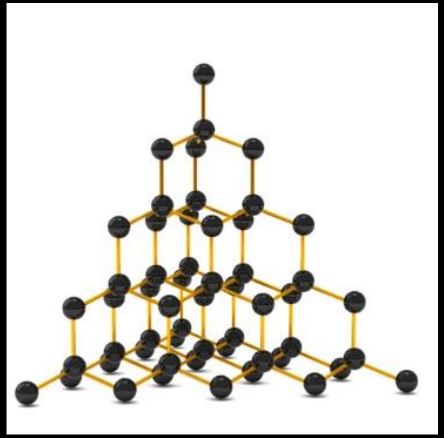
Types of bonding and electrical conductivity

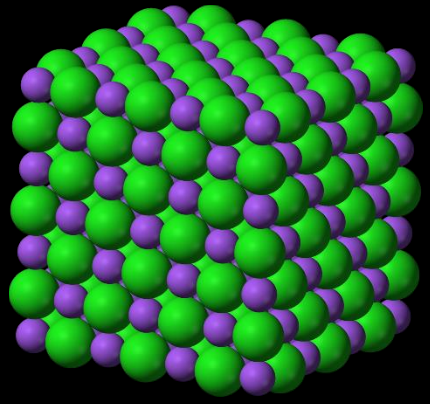
Bonding and electrical conductivity

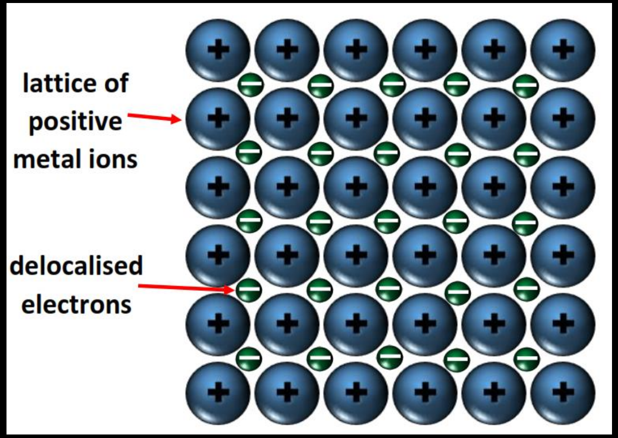
For a substance to conduct electricity it must have one of the following:

- 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.



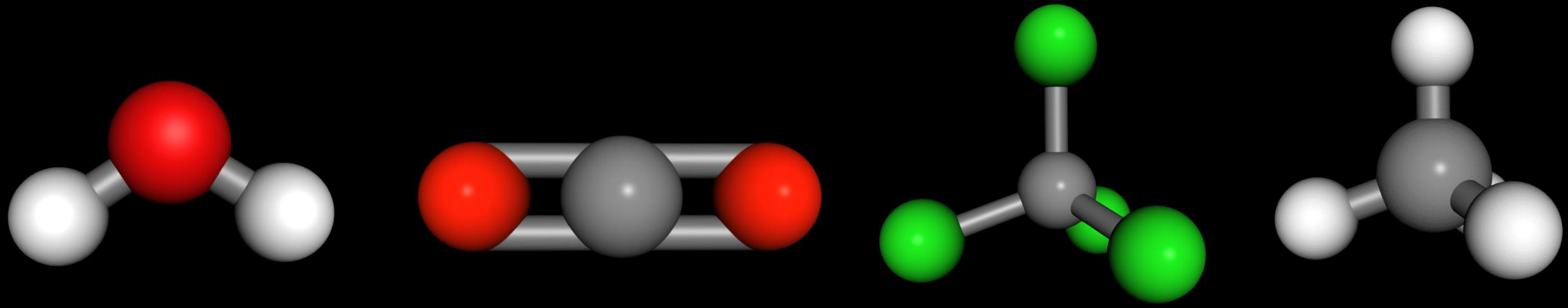






Bonding and electrical conductivity

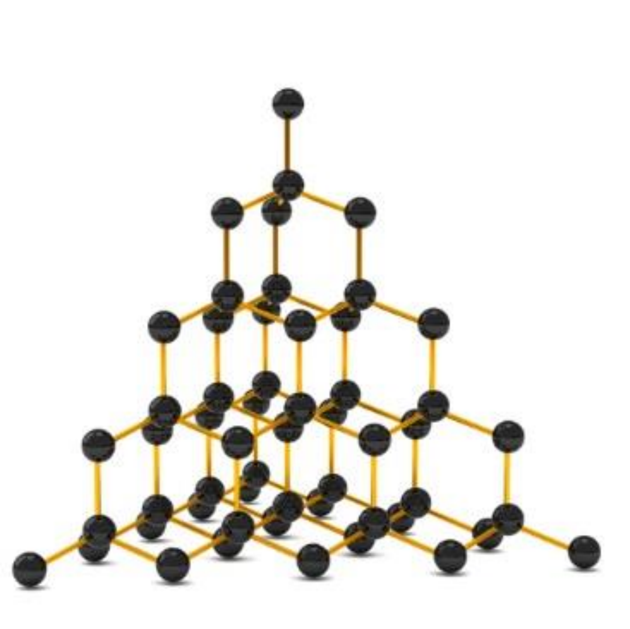
Molecular elements and compounds are poor conductors of electricity.

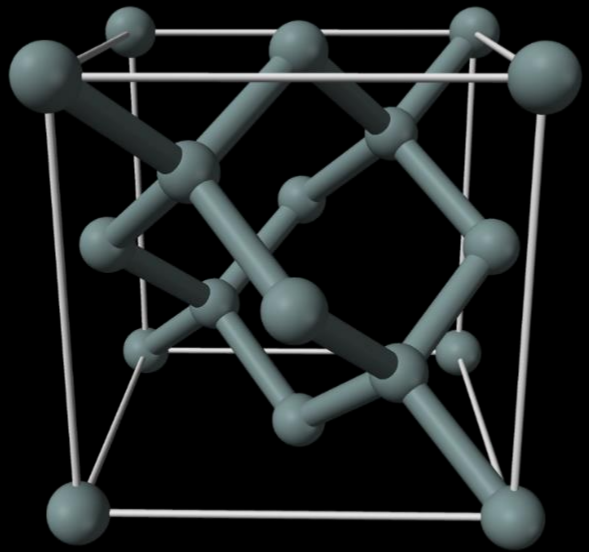


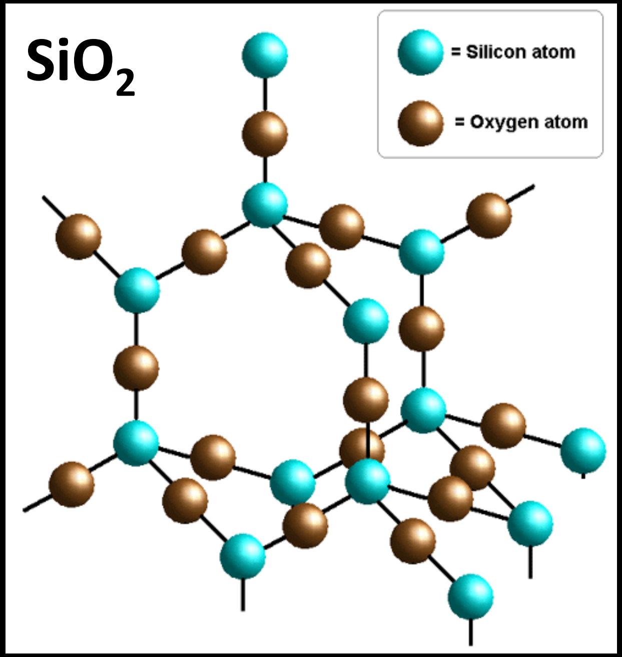
They do not have delocalised electrons (electrons are localised in covalent bonds between atoms).

Bonding and electrical conductivity

Giant covalent substances are poor conductors of electricity.

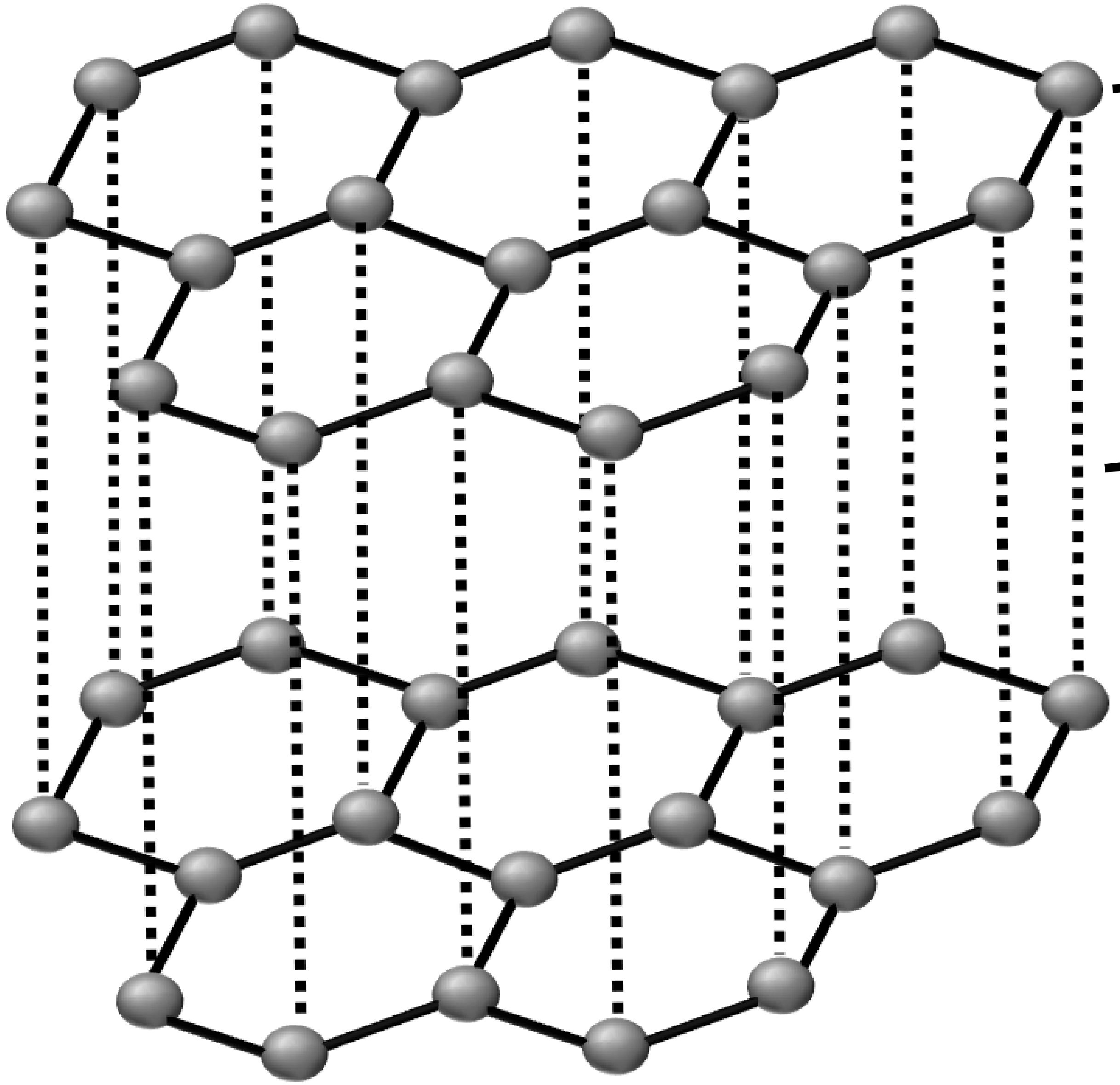






They do not have delocalised electrons (electrons are localised in covalent bonds between atoms).

Bonding and electrical conductivity



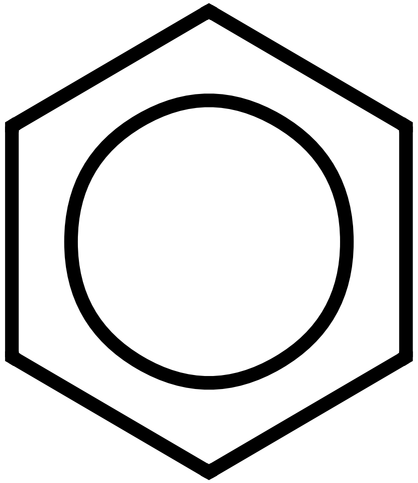
Layers of carbon atoms bonded by covalent bonds.

London dispersion forces between the layers.

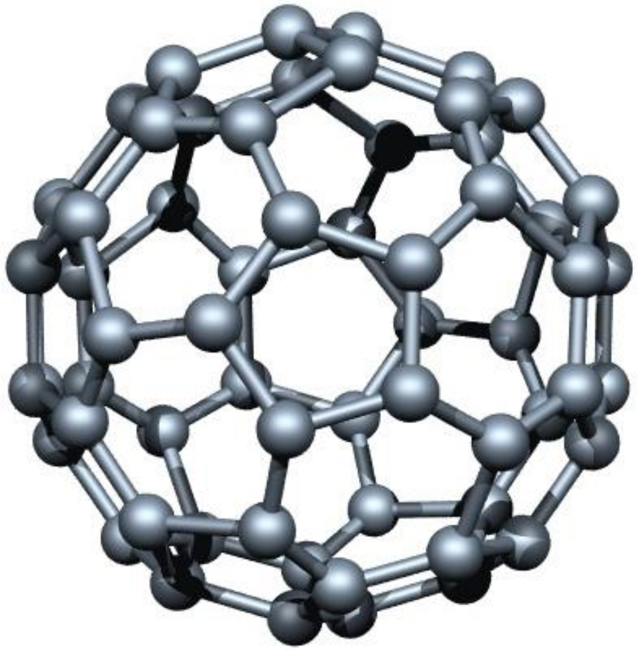
Delocalised electrons are free to move and conduct electricity.

Bonding and electrical conductivity

Molecules with delocalised electrons are not good conductors of electricity.



Benzene C_6H_6



Fullerene C_{60}

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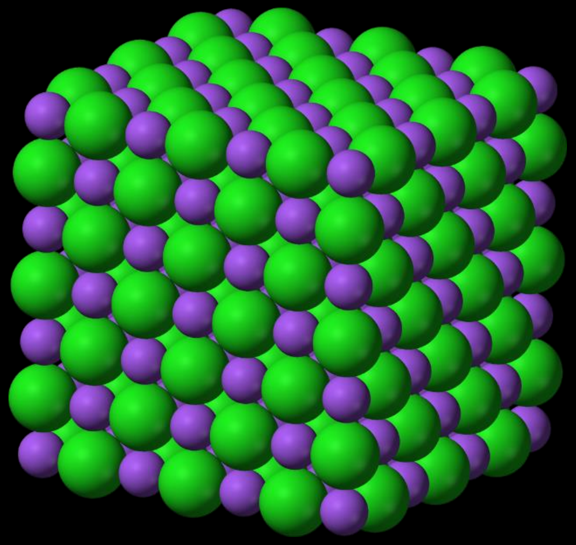
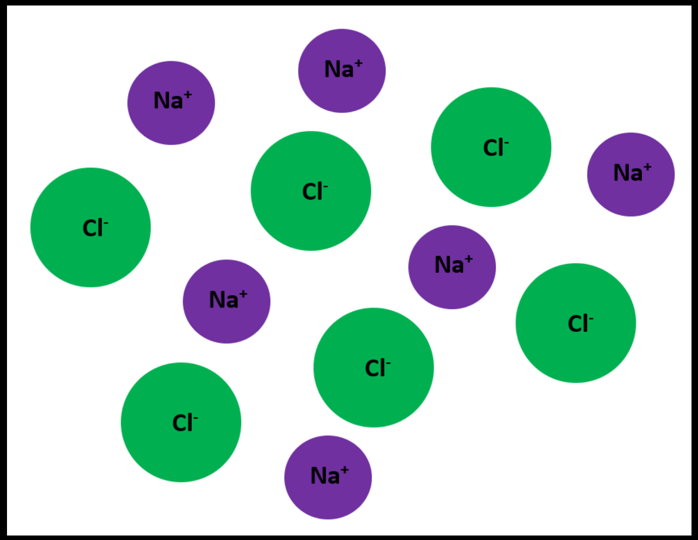
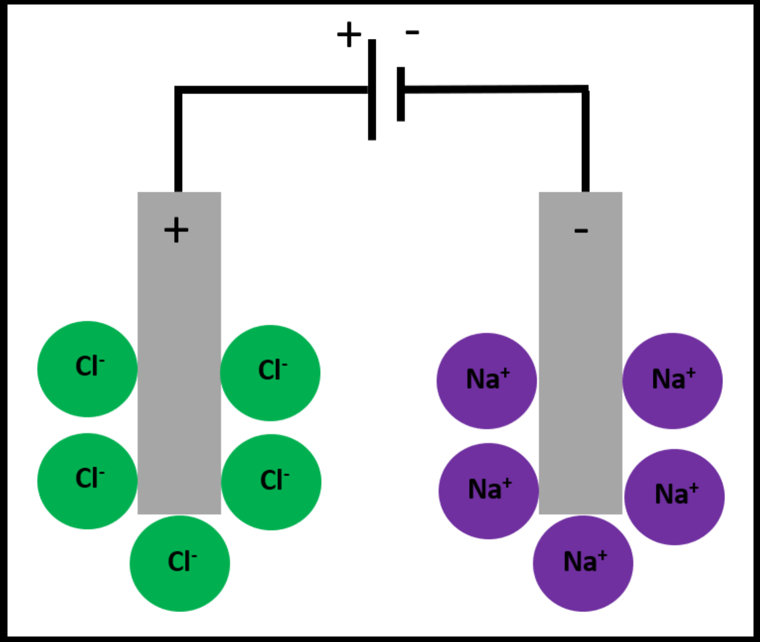
Bonding and electrical conductivity

Ionic compounds conduct electricity only when melted or dissolved in water.





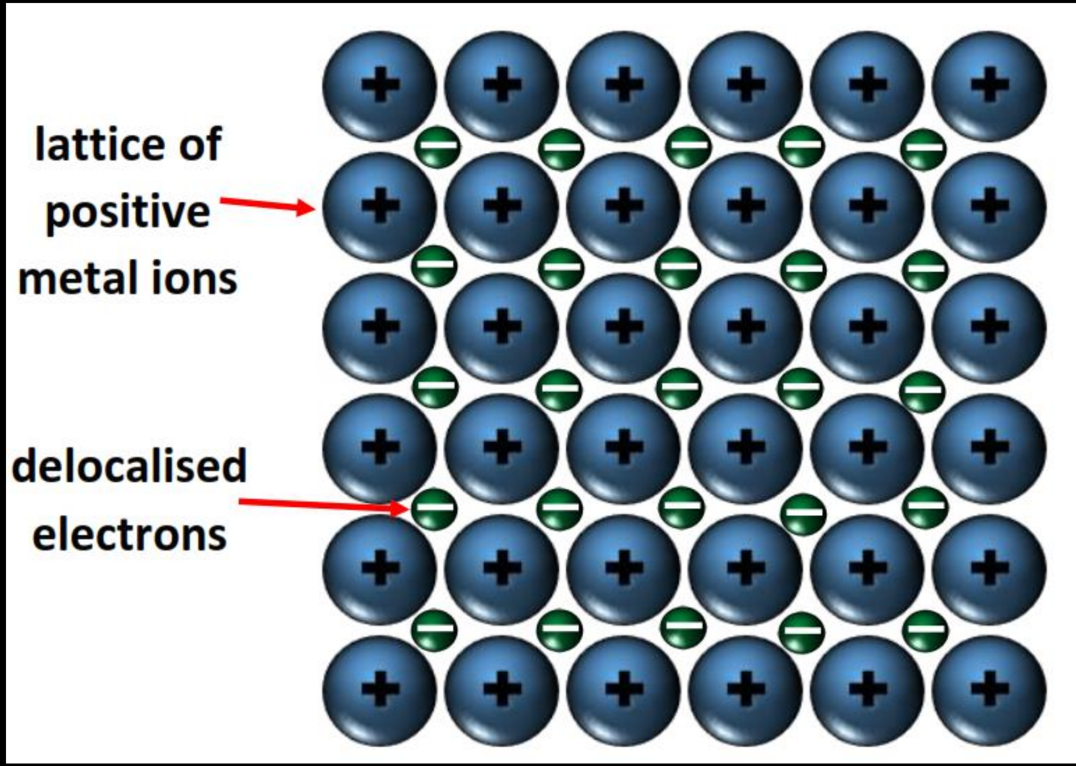


When solid, the ions are held in fixed positions.

When melted or dissolved, the ions are free to move and can carry an electric current.

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.

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Bonding and electrical conductivity

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

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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.

Chromatography

- 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.

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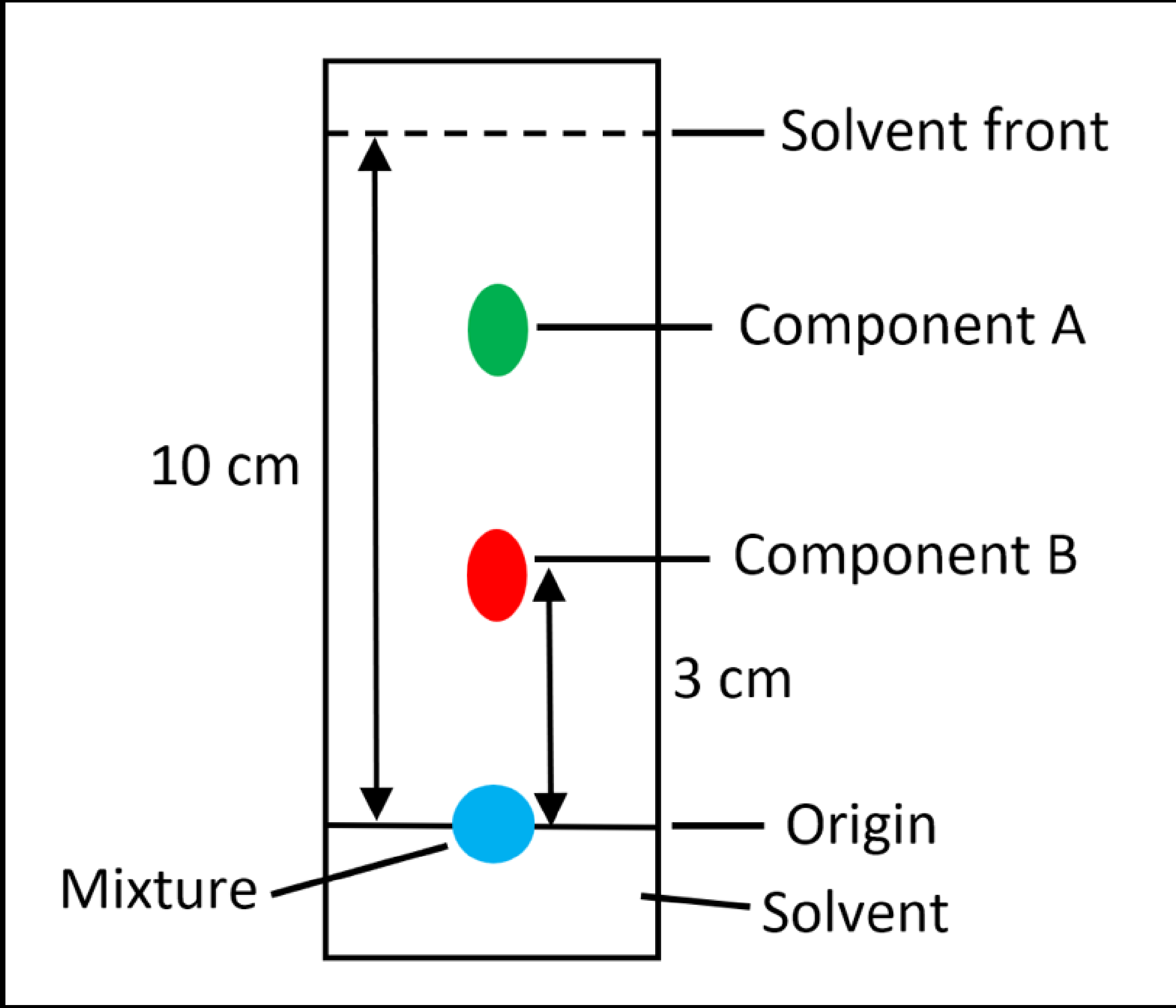
Chromatography

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:

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