Covalent Bonding Map

AQA A-level Chemistry · 3.1.3 Bonding · Covalent bonding

1 The covalent bond 3.1.3.2

- Non-metal atoms **share** a pair of electrons
- Both **nuclei attract** the shared pair between them
- A strong electrostatic attraction
- Attraction to the pair outweighs nuclear repulsion
- Giant covalent solids: very high melting temperatures

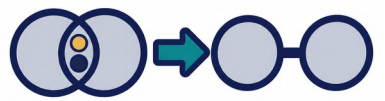
2 Dot-and-cross diagrams 3.1.3.2

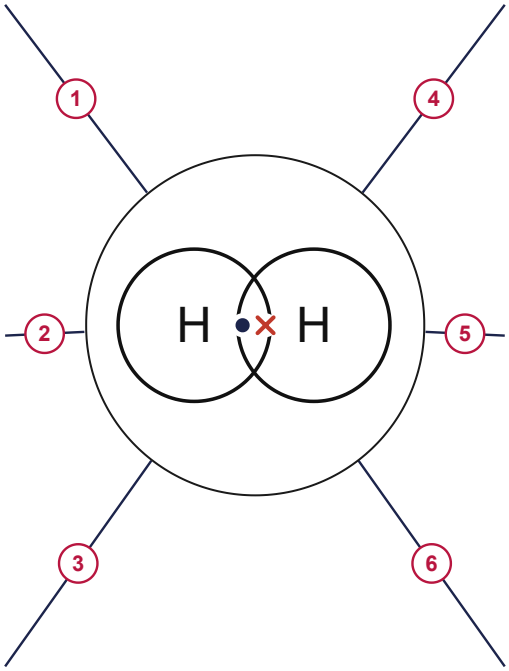
- Outer shells overlap; shared pairs in the overlap
- Dots for one atom, crosses for the other
- **Lone pairs**: unshared outer pairs
- Usually 8 round each atom (2 for H)
- Double bond: 2 pairs (O=C=O); triple: 3 (N≡N)



3 Bonds as lines 3.1.3.2

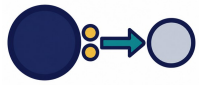
- Each covalent bond: **a line**, one shared pair
- Single H–Cl; double O=C=O; triple N≡N
- **Displayed formula**: every bond drawn
- Ethene: C=C plus four C–H
- Lone pairs are not lines





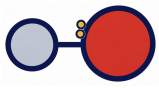
4 Co-ordinate bonds 3.1.3.2

- Dative covalent: both electrons from **one atom**
- Donor: a **lone pair**; acceptor: an **empty orbital**
- Drawn as an **arrow** from the donor: $H_3N \rightarrow H^+$
- NH_4^+ and H_3O^+: brackets and + round the ion
- Al_2Cl_6: two Cl→Al co-ordinate bonds



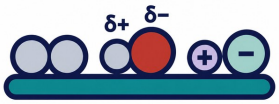
5 Electronegativity 3.1.3.6

- Attraction of an atom for a **bonding pair**
- Pauling scale: F 4.0 the highest
- **Increases across** a period: more protons
- **Decreases down** a group: larger, more shielded
- O 3.5, N 3.0, Cl 3.0, C 2.5, H 2.1



6 Polarity and the continuum 3.1.3.6

- Different electronegativities: a **polar** bond
- δ− on the more electronegative atom
- Difference 0: non-polar; small: polar covalent
- Large difference: **ionic**, electron transferred
- Ionic and covalent: ends of one **continuum**



Electronegativity difference and bond type

Pauling values; the boundaries are gradual, not sharp

Bond	Electronegativities	Difference	Partial charges	Bond type
Cl–Cl	3.0, 3.0	0	none	non-polar covalent
C–H	2.5, 2.1	0.4	almost none	almost non-polar
H–Cl	2.1, 3.0	0.9	Hδ+ Clδ−	polar covalent
O–H	3.5, 2.1	1.4	Oδ− Hδ+	polar covalent
Al–Cl	1.5, 3.0	1.5	Alδ+ Clδ−	polar, ionic character
Na–Cl	0.9, 3.0	2.1	Na^+ Cl^-	ionic

Worked examples

Drawing a molecule

- Count outer electrons of each atom
- Share until each has 8 (H: 2)
- Add the lone pairs last

Spotting a co-ordinate bond

- Is there a lone pair (N, O, Cl)?
- Is there an empty orbital (H^+, B, Al)?
- Arrow from the lone pair to it

Comparing bonds

- One line per shared pair
- Bigger difference: more polar
- Very big difference: ionic

Polar bonds do not always make polar molecules: the shape decides (Shapes of molecules and ions).

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