CAMBRIDGE
UNIVERSITY PRESS

Chemistry

For the IB Diploma

Brighter Thinking
Better Learning

Chapter 7

The covalent model

Covalent bonding

Covalent bond the electrostatic attraction between a shared pair of electrons and the nuclei of the atoms making up the bond.

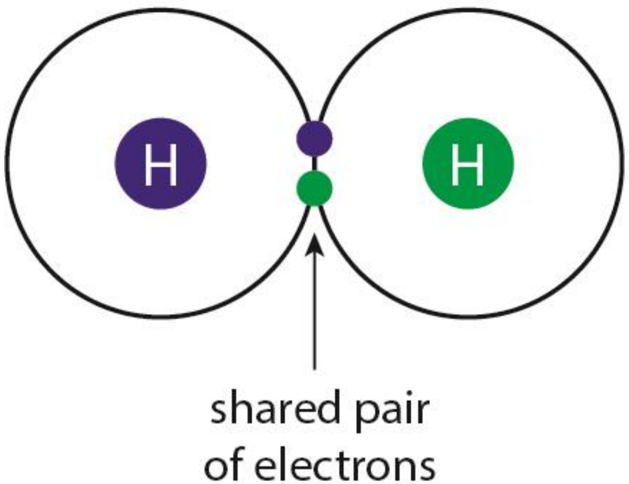

Figure 7.1: A pair of electrons.

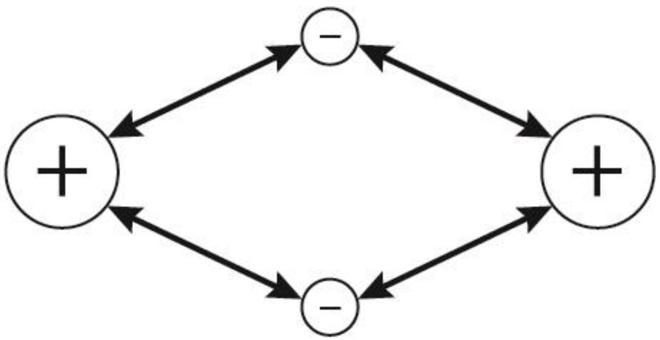

Figure 7.2: The positively charged nuclei attract the negatively charged electrons in the bond.

The Octet Rule

In most molecules, all atoms have eight electrons in their outer (valence) shell – this is called the octet rule. The outer atoms will always have an octet (unless it is hydrogen), but this is not always the case with the central atom.

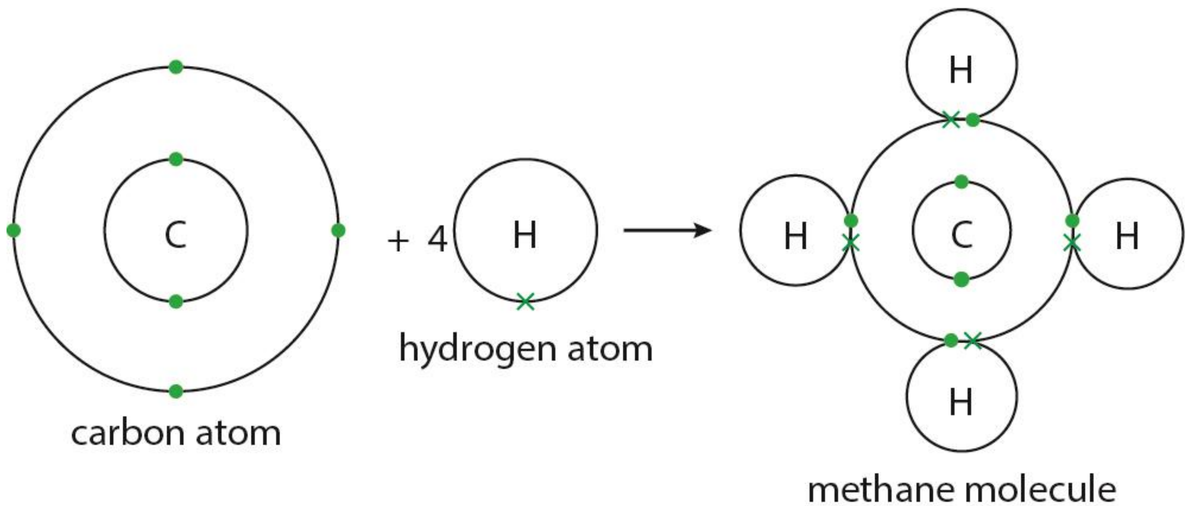

Figure 7.3: The covalent bonding in CH_4.

Coordinate covalent bonds (dative covalent bonds)

This is a type of covalent bond where *both electrons come from the same atom*.

Once it has been formed, it is identical to an *ordinary* covalent bond.

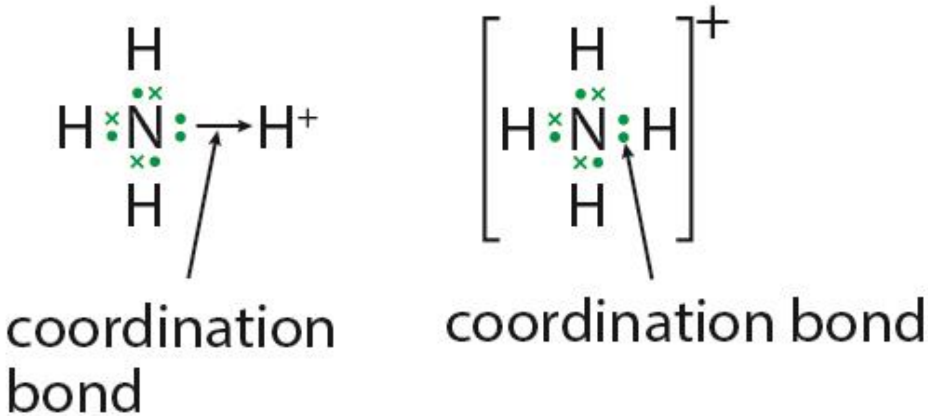

Figure 7.4: The ammonium ion **a** with the coordination bond shown and **b** with no distinction between the types of bonds.

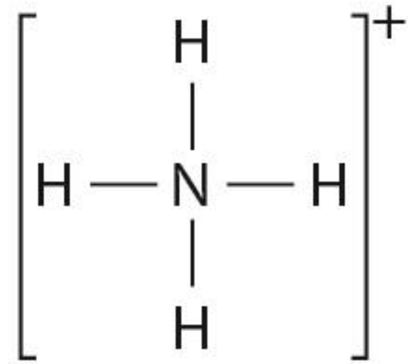

Figure 7.5: Formation of a coordination bond in NH_4^+.

Lewis (electron dot) structures

A set of guidelines for working out Lewis formulas is

- Start with the outer atoms – the outer atom(s) is (are) usually the more electronegative (see below) atom (except when hydrogen is also present in the molecule).
- All outer atoms (except, of course, hydrogen which should have two) have eight electrons in their outer shell. This is achieved by forming single bonds, double bonds, triple bonds and/or having + or − charges.
- The outer atom will form the maximum number of bonds it can (e.g., oxygen will form double bonds wherever possible) up to a maximum limited by the fact that the central atom can form a maximum of four bonds.
- The central atom will usually have eight electrons in its outer shell. It cannot have more than eight but it can have fewer than eight.

Valence shell electron-pair repulsion (VSEPR) theory

Pairs of electrons (electron domains) in the valence (outer) shell of the central atom repel each other and will, therefore, take up positions in space to minimise these repulsions – to be as far apart in space as possible.

How lone pairs of electrons affect the shape

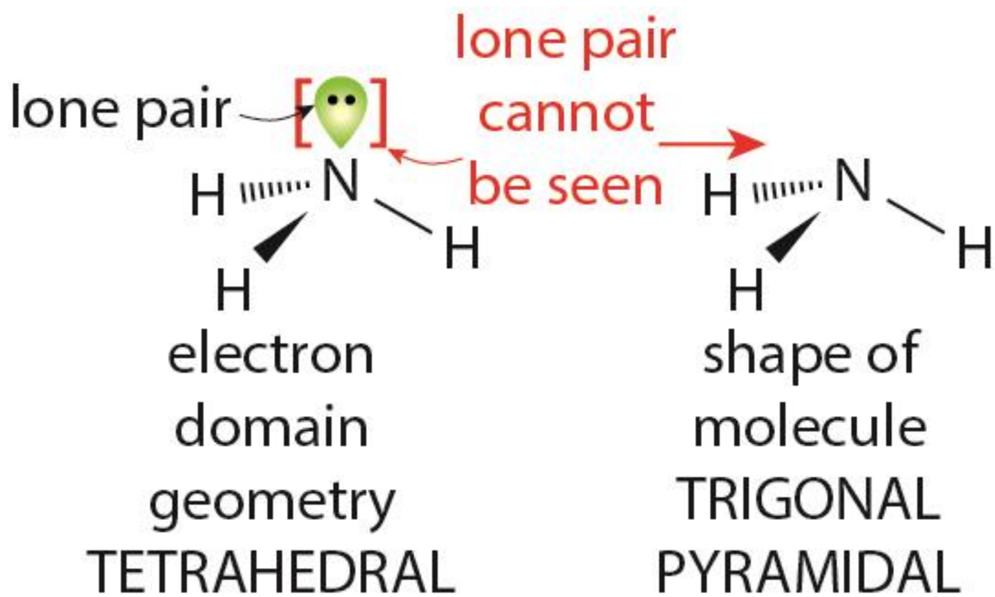

Figure 7.6: The electron domain geometry of the NH_3 molecule is tetrahedral, but the shape of the molecule is trigonal pyramidal.

Intermolecular forces

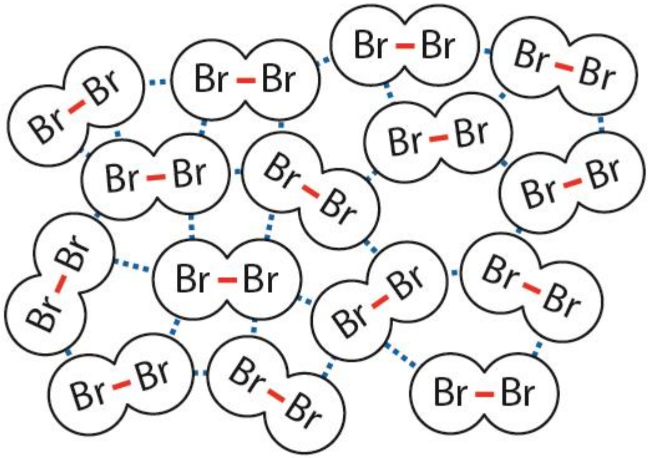

Figure 7.7: Intermolecular and intramolecular forces in liquid bromine. Intramolecular forces (covalent bonds) are shown in red and intermolecular forces are the dashed lines in blue.

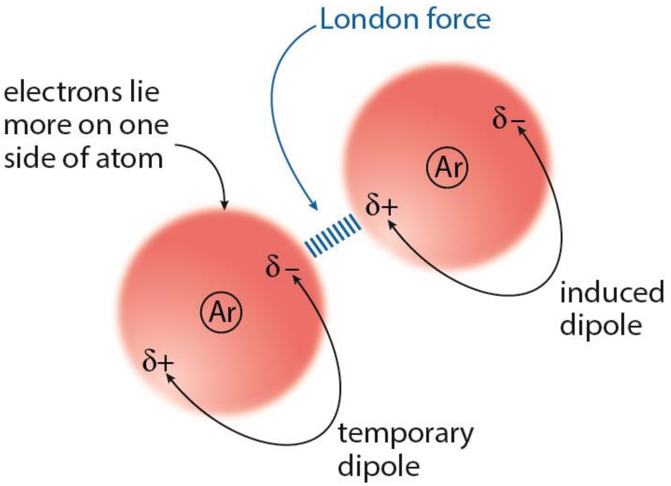

Figure 7.8: The origin of London forces.

Boiling points of halogens vs noble gases

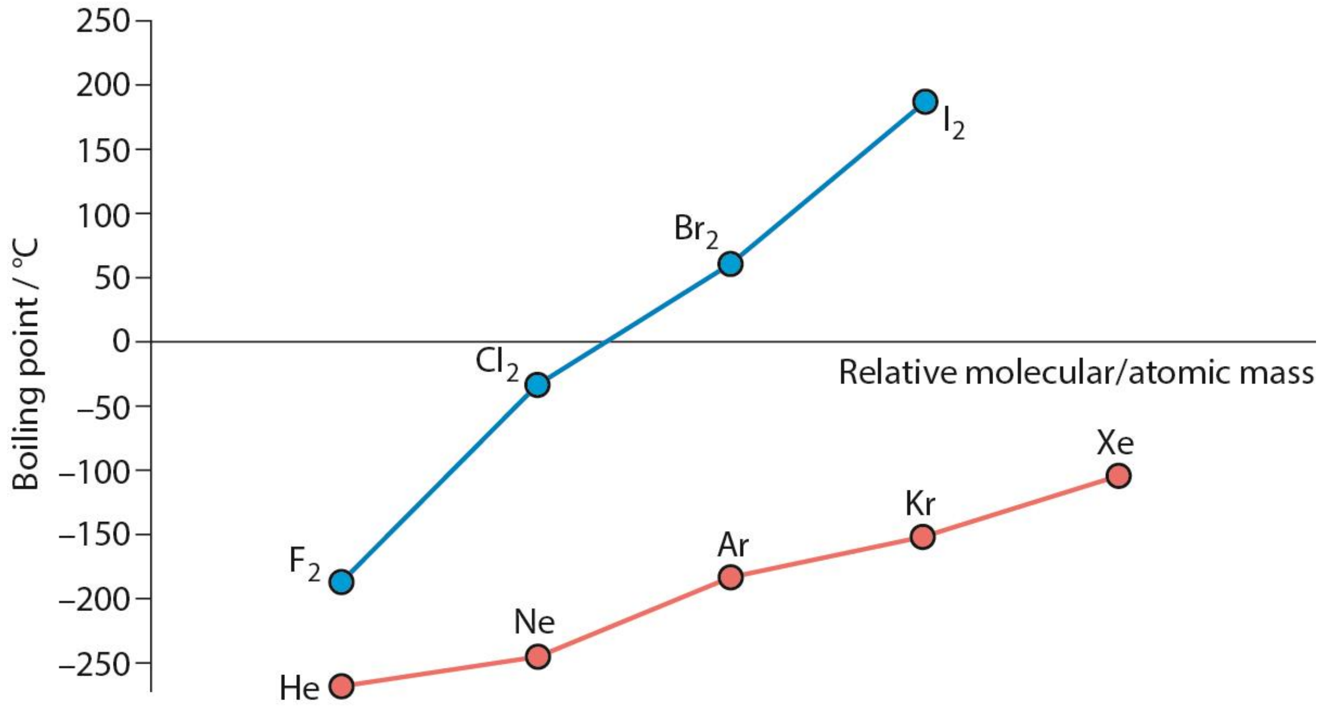

Figure 7.9: Variation in boiling points of elements in Groups 17 and 18.

Diagram showing how permanent dipole permanent–dipole interactions are formed

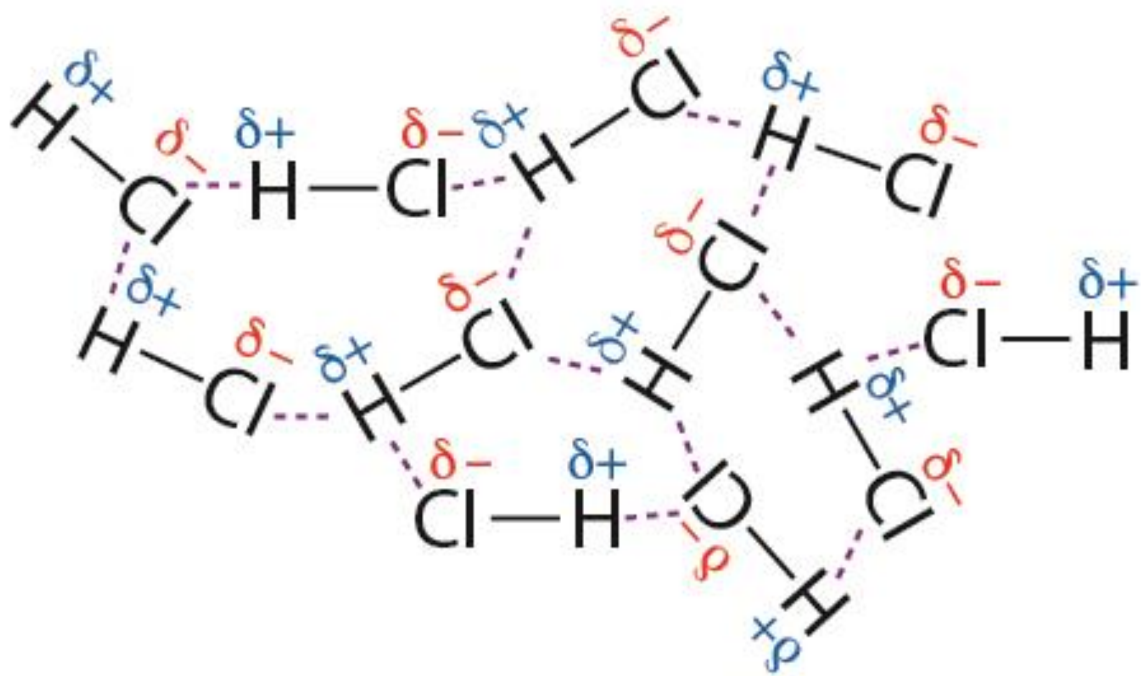

Figure 7.10: Permanent dipole–permanent dipole interactions exist between molecules. These are shown in purple.

Boiling points graph to show hydrogen bonding is present in water

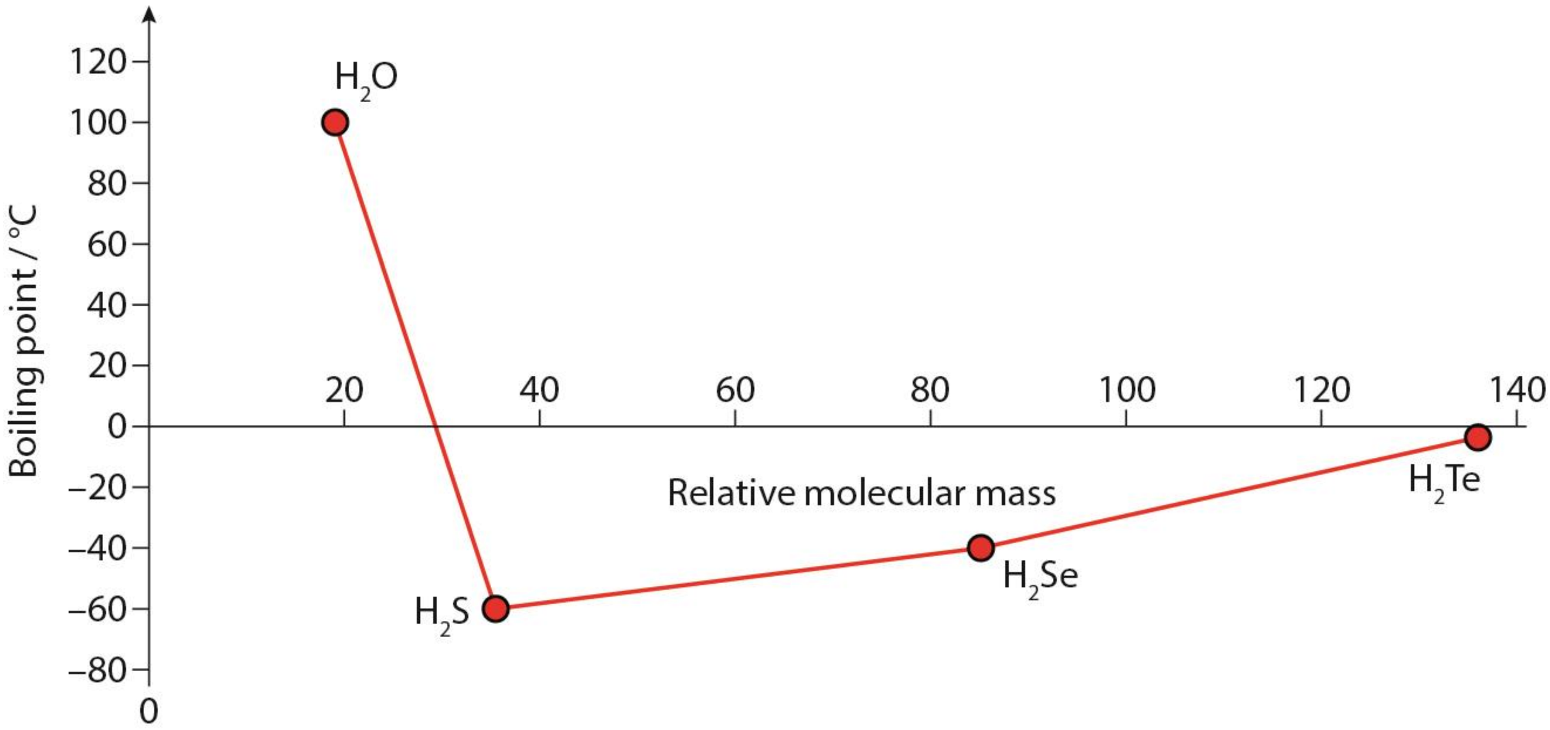

Figure 7.11: Boiling points of Group 16 hydrides.

Diagrams to show how hydrogen bonds are formed

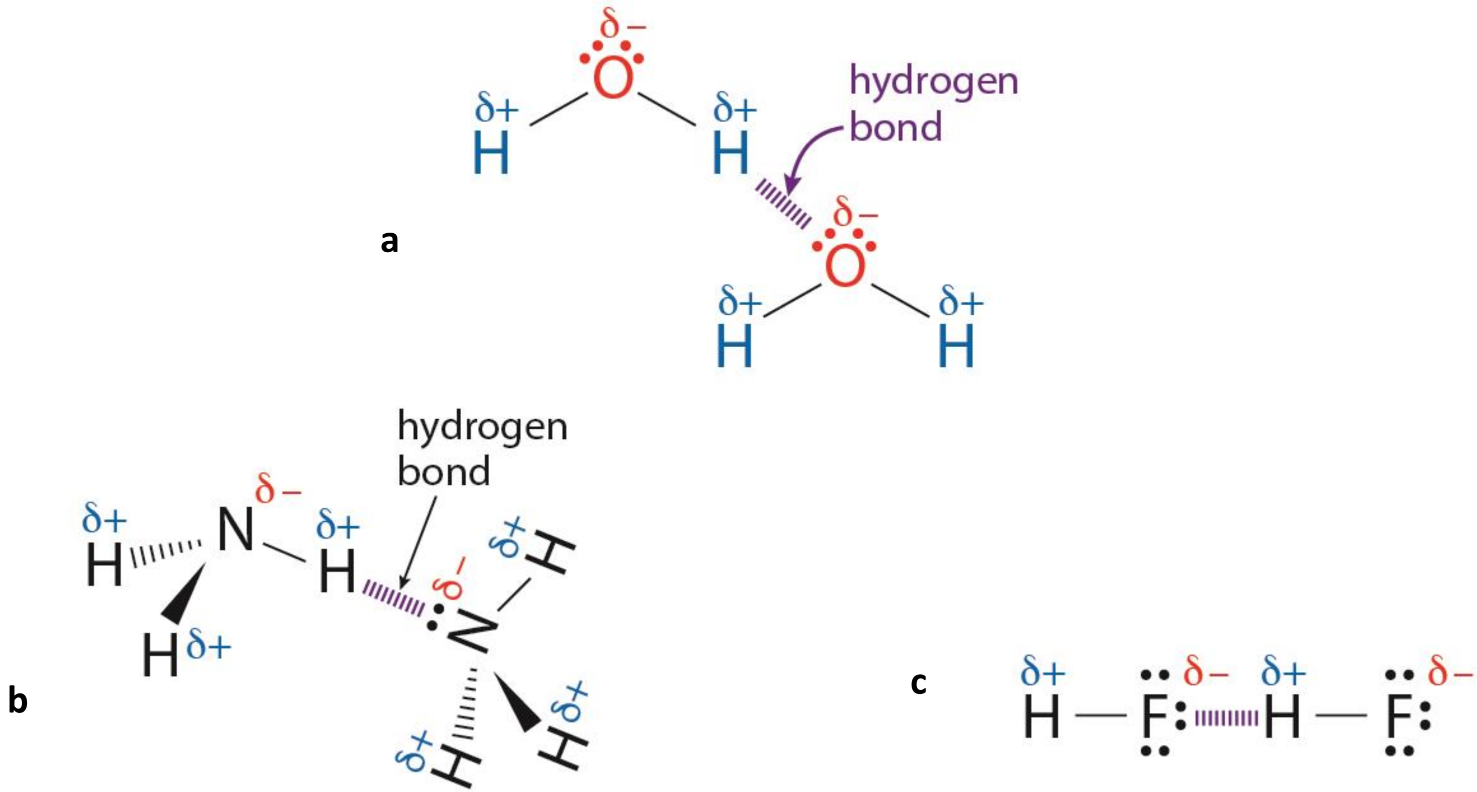

Figure 7.12: Hydrogen bonds between **a** water molecules, **b** ammonia (NH_3) molecules and **c** HF molecules. The hydrogen bond is the dashed purple line between the lone pair of a δ^- oxygen, nitrogen or fluorine on one molecule and the δ^+ hydrogen on the other molecule.