

# CAIE Chemistry A-level

## 3: Chemical Bonding Definitions

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## Definitions and Concepts for CAIE Chemistry A-level Chemical Bonding

**Atomic radius:** The radius of an atom. The distance from the nucleus to the outermost electrons.

**Boiling temperature:** The temperature at which a substance changes from a liquid state to a gaseous state.

**Bond angle:** The angle between two bonds that originate from the same atom.

**Bond enthalpy:** The energy required to break one mole of the stated bond in a gaseous state, under standard conditions.

**Bond length:** Internuclear distance between two covalently bonded atoms.

**Covalent bond:** The strong electrostatic attraction between two nuclei and the shared pair of electrons between them. Polar covalent bonds occur when there is an asymmetric electron distribution within the covalent bond due to difference in electronegativities.

**Covalent substance:** A substance that is made up of atoms that are covalently bonded to each other.

**Dative covalent bonding:** Occurs when one atom donates both electrons in a bond. For example, in  $^+NH_4$  the nitrogen atom supplies both electrons for one of the covalent bonds with a hydrogen atom.

**Dot-and-cross diagram:** Diagrams used to model the bonding that occurs in a simple molecule. The shells of an atom are drawn as circles, with crosses or dots marked on the circles to represent the electrons. The circles overlap when there is a covalent bond. The electrons from one atom are drawn as dots, and the electrons from a different atom are drawn as crosses.

**Electronegativity:** A measure of the ability of an atom to attract a bonding pair of electrons within a covalent bond. The Pauling scale is often used, with fluorine being the most electronegative element and caesium and francium being the least electronegative elements.

**Hybridisation:** The combining of two or more atomic orbitals to form new 'hybrid' orbitals which are different to the original orbitals.





**Hydrogen bonding:** An interaction between a hydrogen atom and an electronegative atom, commonly nitrogen, fluorine or oxygen. The slightly positive hydrogen is attracted to the lone pair on the electronegative atom. Hydrogen bonds are stronger than van der Waals and dipole-dipole forces but weaker than ionic and covalent bonds.

**Intermolecular forces:** The forces which exist between molecules. The strength of the intermolecular forces impact physical properties like boiling/melting point.

**Ionic bond:** Strong electrostatic attraction between two oppositely charged ions. The strength of attraction depends on the relative sizes and charges of ions.

**Ionic charge:** The electrical charge of an ion caused by the gain (negative charge) or loss (positive charge) of electrons. The magnitude of the charge is related to how many electrons have been lost or gained as electrons have a relative charge of -1.

**Ionic compound:** A compound made up of anions and cations, held together by ionic bonds, which arise due to the electrostatic attraction between oppositely charged ions. These structures are neutral overall.

**Ionic radius:** The radius of an ion. It is the distance between the nucleus and the outermost electron of the ion.

**Linear:** The shape of a molecule when the central atom has 2 bonding pairs and no lone pairs of electrons.

**Melting temperature:** The melting point of a substance is the temperature at which it changes from solid state to liquid state.

**Metallic bonding:** Strong electrostatic attraction between positive metal ions and the sea of delocalised electrons that surround them.

**Non-linear/bent:** The shape of a molecule when the central atom has 2 bonding pairs and 2 lone pairs of electrons.

**Nuclear charge:** Total charge of all the protons in the nucleus. It has the same value as the atomic number. Increases as you go across the periodic table.

**Octahedral:** The shape of a molecule with 6 bonding pairs of electrons and a bond angle of  $90^\circ$ . This is often found in metal complexes E.g.  $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ .

**Permanent dipole-dipole forces:** When molecules with polar covalent bonds interact with dipoles in other molecules dipole-dipole intermolecular forces are produced between the molecules. These intermolecular forces are generally stronger than van der Waals forces but weaker than hydrogen bonding.





**Polar bond:** A covalent bond between two atoms in which the electrons in the bond are unevenly distributed. This causes a slight charge difference which induces a dipole in the molecule.

**Polarisability:** The ability of a molecule to induce a dipole in another molecule.

**Shielding:** A decrease in the nuclear attraction experienced by an outer shell electron caused by electron-electron repulsion between the outer shell electron and electrons from adjacent shells.

**Square planar:** One of the possible shapes of a molecule with 4 bonding pairs and a bond angle of  $90^\circ$ . Platinum and nickel metal complexes form square planar complexes. E.g,  $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ .

**Surface tension:** As a result of water's ability to form hydrogen bonds between its molecules, water molecules on the surface do not have water molecules above them. This causes them to bond more strongly to the water molecules around them. This causes a 'film' of water at the surface.

**Tetrahedral:** The shape of a molecule when the central atom has 4 bonding pairs.

**Trigonal bipyramidal:** The shape of a molecule when the central atom has 5 bonding pairs.

**Trigonal planar:** The shape of a molecule when the central atom has 3 bonding pairs.

**Van der Waals:** Also known as induced dipole–dipole, dispersion and London forces, van der Waals forces exist between all molecules. They arise due to fluctuations of electron density within a nonpolar molecule. These fluctuations may temporarily cause an uneven electron distribution, producing an instantaneous dipole. This dipole can induce a dipole in another molecule, and so on.

**VSEPR:** Valence shell electron pair repulsion theory which is a method of predicting the shape of a molecule by the number of bonding and non bonding pairs of electrons.

**$\pi$ -bond:** A type of covalent bond formed when adjacent p orbitals overlap above and below the  $\sigma$  single carbon bond. Pi ( $\pi$ ) bonds can't be rotated. As  $\pi$ -bonds have low bond enthalpy, alkenes are more reactive than alkanes.

**$\sigma$ -bond:** A type of covalent bond formed by direct overlap of orbitals between the bonding atoms to form a single carbon bond. These have a high bond enthalpy.

