

SESSION 3: CHEMICAL BONDING

Key Concepts

In this session we will focus on summarising what you need to know about:

- Bonding
- Covalent bonding
- Electronegativity in covalent bonding
- Shapes of molecules
- Polar covalent substances
- Bond energy and length

X-planation

Bonding

When atoms combine, the electrons involved are the **valence electrons**.

Valence electrons are the electrons on the highest energy level or in the outer shell of an electron.

The number of electrons in the atom of an element is equal to the group number of the element.

When atoms bond, the valence electrons are either shared (covalent bonding) or transferred between atoms of elements (ionic bonding).

The number of bonds an atom can form is referred to as the **valency** of that atom. For groups I-IV, the valency is the same as the group number, but for group V-VIII, the group number is equal to 8 minus the group number.

Covalent Bonding

Covalent bonds form between non-metals.

They involve sharing of electrons, i.e. two atoms share an electron pair.

To form a bond, the two atoms must approach each other in such a way that their half-filled orbitals overlap. If the overlapping orbitals each contain one electron with opposite spin, the electrons can pair, and the shared electron pair is called the **bonding pair**. The other pairs of electrons that are not shared are called **lone pairs**. Covalent bonds can be single (one e⁻ pair shared), double (2 e⁻ pairs shared) and triple bonds (3 e⁻ pairs shared).

Electronegativity in Covalent Bonding

Electronegativity is a measure of the tendency of an atom to attract the shared electron pair towards itself. On the Periodic Table of Elements, electronegativity increases from left to right in a period, and decreases from top to bottom in a group. In a **polar covalent bond** there is a difference in electronegativity between the bonded atoms, e.g. HCl. The bonded pair is attracted closer to the Cl than it is to the H because Cl is more electronegative than H. The molecule is a **polar molecule**. **The H side of the molecule has a slight positive charge and the Cl side of the molecule has a slight negative charge.**

If the electron pair is shared equally between the two atoms because they have the same electronegativity, e.g. H₂, O₂ or N₂, the bond is non-polar. This forms a non-polar molecule.

Shapes of Molecules are determined by the number of bond pairs and lone pairs. The shapes of molecules affect boiling point, melting point, physical properties and intermolecular forces between molecules.

Linear molecules - H_2 , N_2 , CO_2 , HCl (mostly when 2 atoms bond)

Angular - Three atoms that are not arranged linearly, two shared electron pairs, e.g. H_2O , H_2S

Pyramidal - Four atoms, three shared electron pairs, one lone pair, e.g. NH_3

Tetrahedral - Five atoms, four shared electron pairs, e.g. CH_4 , CCl_4

Ionic Bonding - Involves a transfer of electrons. One atom loses an electron – electron donor – attains a positive charge (called a cation). Another gains an electron – electron acceptor – attains a negative charge (called an anion). The substance that loses electrons is oxidised (reducing agent); the substance that gains electrons is reduced (oxidising agent)

Polar covalent substances

Depending on the shape of the molecule and the electronegativity difference of the constituent atoms, some molecules are more polar than others.

A pure covalent bond can be considered as one which exists between two identical atoms. The atoms exert equal forces of attraction on the bonded pair of electrons and

the molecule is non-polar.

In the event of a diatomic molecule composed of two different atoms with different electronegativities, the bonded pair is not equally shared. As the electronegativity difference increases, the dipole nature of the molecule increases.

As a result of such unequal sharing of electrons, a bond can be considered to be partly ionic and partly covalent.

If the electronegativity difference is greater than 1,7, the compound is more ionic than covalent, and if the difference is less than 1,7, the compound is more covalent than ionic.

Bond energy and length

The strength of bonds within a molecule is of great significance in chemical reactions.

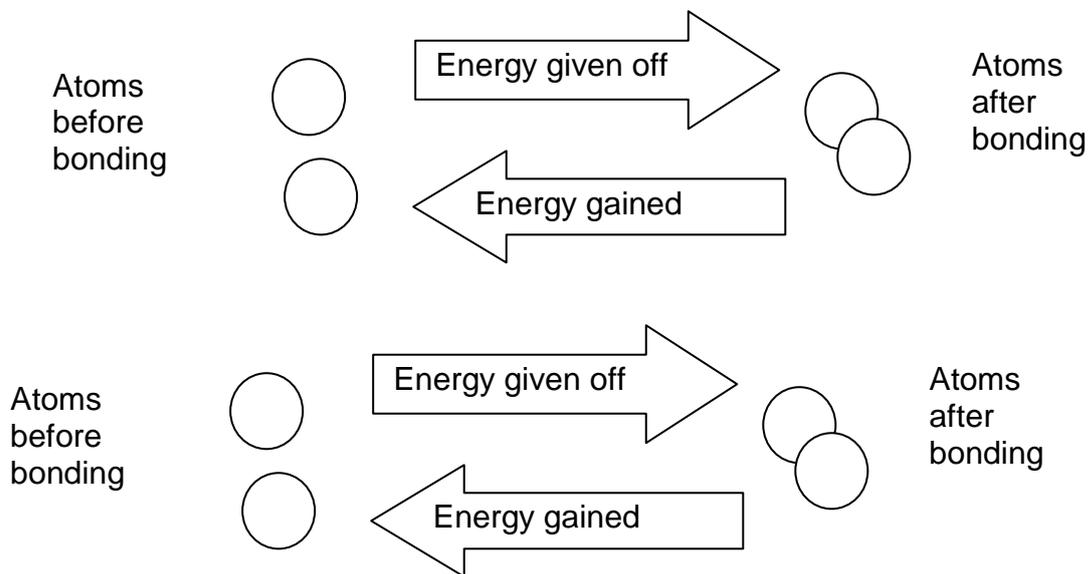
What is meant by **bond strength**? Bond strength relates to the energy needed to break bonds between atoms in a molecule. This is indicated by the bond energy - the amount of energy required to “break” the bond. It is used to overcome the force of attraction between the atoms. This is the same as the energy given off when two atoms bond.

The relationship between strength of bond between two chemically bonded atoms and the length of the bond between them: the length of a bond determines the strength of the bond. The **bond length** is the distance between the centers of the bond atom pair. The stronger the bond, the smaller the bond length.

Effect of the **size of the atoms** on the strength of the bond: small atoms form stronger bonds as the attraction towards the nucleus of a small atom is greater. The atomic radius decreases towards the right in a period, and increases towards the bottom of a group.

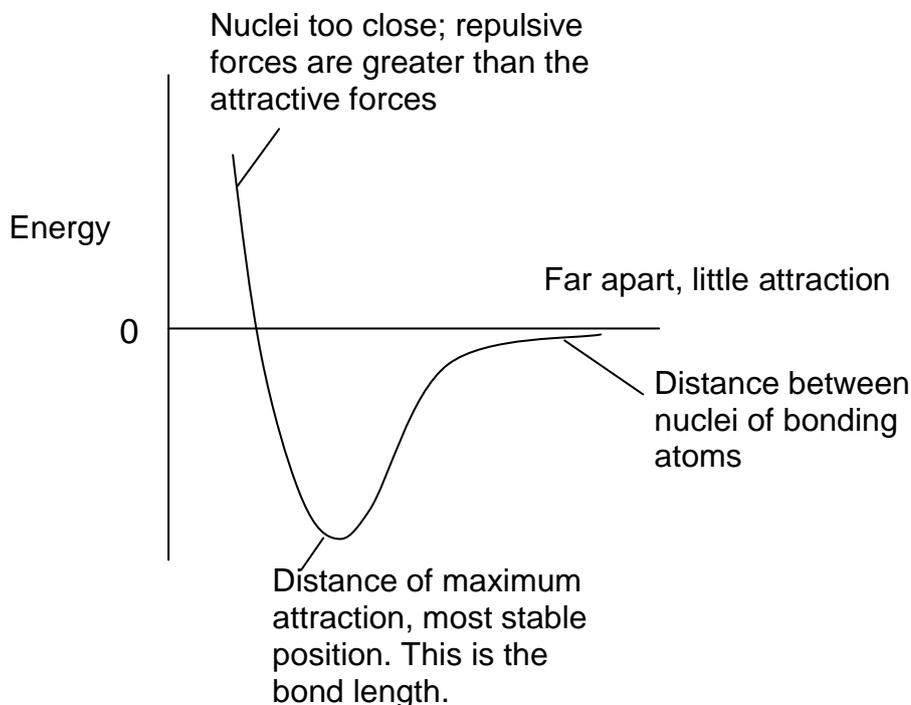
The effect of the **number of bonds** or bond order (single, double, triple) on the strength of the bond: multiple bonds are stronger than single bonds.

When bonds are broken, energy is required to overcome the maximum attraction between the bonding atoms, and the potential energy levels increase.



Potential energy against distance graph for two bonding atoms

The graph below shows the effect of the distance between atoms on the stability of the bond.



When the atoms approach each other, the energy levels decrease. Attractive forces increase when atoms approach each other, repulsive forces increase when atoms move too close together.

At the bond length position the atoms are most stable; repulsion and attractive forces are equal.

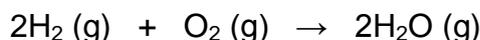
Bond energy is measured in $\text{kJ}\cdot\text{mol}^{-1}$ and shows the energy used to make or break the bond during a reaction. Bond energy indicates the strength of bonds.

When the sum of the energy required to break the bonds is greater than that released to make new bonds, the net energy value will be positive. This indicates an exothermic reaction.

Compare the given data.

BOND PAIR	BOND ENERGY ($\text{kJ}\cdot\text{mol}^{-1}$)	BOND LENGTH (pm)	ATOMIC RADII (pm)
H-H	436	75	120
H-F	570	92	120 (H) 140(F)
O=O	497	148	145
C-H	435	109	120 (H)
C=O	803	123	145 (O)
H-O	464	96	120 (H) 145 (O)
C-C	347	154	
C-O	358	143	145 (O)

Example: Use the equation and answer questions below.



Determine the energy required to break the bonds in the equation above.

Separating an H-H bond requires $436 \text{ kJ}\cdot\text{mol}^{-1}$; therefore, two molecules require $2 \times 436 \text{ kJ}\cdot\text{mol}^{-1}$.

Separating an O=O bond requires $497 \text{ kJ}\cdot\text{mol}^{-1}$

Total energy required = $2369 \text{ kJ}\cdot\text{mol}^{-1}$

Determine the energy given off when making new bonds.

In 2 molecules of H_2O four O-H bonds are formed, giving off $4 \times 464 \text{ kJ}\cdot\text{mol}^{-1}$

Total energy given off = $1856 \text{ kJ}\cdot\text{mol}^{-1}$

The energy difference (**heat of reaction**) = energy required – energy given off

$$2369 - 1856 = -513 \text{ kJ}\cdot\text{mol}^{-1}$$

A positive answer indicates an endothermic reaction (products have more energy than reactants).

A negative answer indicates an exothermic reaction (products have less energy than the reactants).

The energy of the bonds (enthalpy) is shown by a symbol H, the difference in bond energy of the reactants and products is ΔH , also known as heat of the reaction.

The amount of substance is given by the relative atomic mass.

The **relative atomic mass** (A_r) has no unit; it only indicates how many times a substance is heavier than $\frac{1}{12}$ of a ${}^{12}_6\text{C}$ atom.

Formulae of substances can be used to calculate **relative formula mass** (M_r). It is also known as **relative molecular mass**.

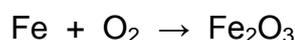
M_r is the sum of the relative atomic masses of all the atoms in a molecule; it has no unit.

However, **molecular of formula mass** is measured in $\text{g}\cdot\text{mol}^{-1}$

X-ample Questions

Question 1

Corrugated iron sheets need to be painted regularly to avoid rusting. The following shows how flakes form:



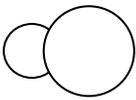
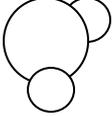
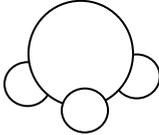
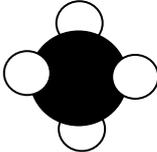
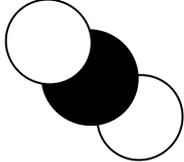
- 1.1 Give the oxidation number for each of the atoms in the reaction. (4)
- 1.2 Which substance is reduced? (2)
- 1.3 Name the reducing agent (2)
- 1.4 Give the Lewis structure of oxygen (2)
- 1.5 Fe^{2+} can form a dative covalent bond with carbon monoxide. Explain how this dative bond is formed. (4)
- 1.6 Draw the Lewis diagram of a carbon dioxide molecule. What is the shape of the molecule? (4)

Question 2

- 2.1 What is the difference between the type of bonding in Cl_2 and MgCl_2 ? (4)
- 2.2 Determine the oxidation number of chlorine in each of the following substances:
 - 2.2.1 Cl_2O
 - 2.2.2 ClO_2
 - 2.2.3 Cl_2O_6
 - 2.2.4 Cl_2O_7 (4)

Question 3: MOLECULAR SHAPE and POLARITY

Complete the following table:

					
Molecular shape					
Formula of substance with this shape					
Polarity of bond					
Polarity of molecule					

Question 4

Study the values given in the table below and answer the questions that follow.

BOND	BOND ENERGY kJ per mole of bonds
N - N	158
N \equiv N	945
N - H	389
O - H	464
O = O	497
C = C	612
N = O	631
H - H	436
C - H	435
C - C	347

- 4.1 Identify the most unstable bond. Give a reason for your answer (3)
- 4.2 What is the general relationship between the size of the atoms and the bond energy? (2)
- 4.3 Which bond is the strongest? Explain. (3)
- 4.4 Explain the difference in bond energy between
- a.) N \equiv N and N - N
- b.) C = C and C - C (4)

Question 5

- 5.1 Chemical bonding results in the increased stability of a substance. True / false. Explain. (3)
- 5.2 The chemical bond with the greatest bond energy is:
a. C – F
b. C – O
c. C – N
d. C – C
Give a possible reason for your answer (3)
- 5.3 For the reaction $2 \text{O}_3(\text{g}) \rightarrow 3 \text{O}_2(\text{g})$, do the number of moles of molecules increase, decrease or remain the same as the reaction proceeds? (2)
- 5.4 Study the reactions below.
i. $\text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2 \text{e}^-$
ii. $4 \text{H}^+(\text{aq}) + \text{O}_2(\text{g}) + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O}(\text{l})$
Indicate the oxidation numbers of each substance, then identify the substances that are reduced and oxidised (8)

X-ercise

Question 1

- 1.1 Use the following examples to explain the term “chemical bond”:
a. H_2 (4)
b. H_2O (4)
- 1.2 Explain why two H atoms form an H_2 molecule (2)
- 1.3 Explain why He does not form He_2 (2)
- 1.4 Describe a chemical bond as a shared pair of electrons (3)
- 1.5 Which “bonding rule” best explains each of the following situations:
a. formation of diatomic molecules, e.g. F_2 , N_2 (1)
b. Inert gases, e.g. Ne, Ar do not form chemical bonds (1)
c. formation of multiple bonds in a molecule, e.g. H_2S , NO_2 (1)
d. the existence of NH_4^+ and H_3O^+ ions (1)

Question 2

- 2.1 Represent the following atoms using Lewis diagrams:
- a. Chlorine (1)
 - b. Helium (1)
- 2.2 Draw Lewis diagrams for the following simple molecules:
- a. F_2 (2)
 - b. H_2O (2)
 - c. NH_3 (2)
- 2.3 Draw the Lewis diagram of the formation of substances where a molecule donates a lone pair of electrons to a molecule or ion with vacant orbitals in the valence shell.
- a. NH_4^+ (ammonia) (2)
 - b. The substance that donates the electron pair is called (1)
 - c. The substance that accepts the electron pair is called (1)

Question 3



- 3.1 the energy required to break the bonds (3)
- 3.2 the energy needed to form new bonds (3)
- 3.3 the total change in energy for the reaction (1)
- 3.4 Is the reaction endothermic or exothermic? Explain. (2)