COVALENT MOLECULES

Molecules can be defined as the smallest part of a pure substance that can exist separately. Therefore, there are several classes of molecules:

1. Monatomic molecules – consist of one atom e.g. helium, He
2. Diatomic molecules – consist of two atoms e.g. oxygen, O₂, and carbon monoxide, CO
3. Triatomic molecules – consist of three atoms e.g. ozone, O₃, and water, H₂O
4. Tetra-atomic molecules – consist of four atoms e.g. phosphorous, P₄, and ammonia, NH₃

Note that all molecules are composed of non-metals. Some are comprised of one type of element, in which case they are known as elemental molecules. Others contain more than one type of element, making them molecular compounds.

As well as gaining electrons to become ions, non-metal atoms can achieve stability by sharing electron pairs with other non-metal atoms. A stable octet is obtained by this process. The resulting bond is called a covalent bond. The number of electron pairs that are shared will determine whether the bond is a single, double or triple covalent bond.

Lewis electron dot structures can be used to show the formation of these bonds.

![Lewis dot structures](image)

Structural formulae can also be used to show the bonding arrangements in molecules. The electron pair is represented by a bond line.

![Structural formulae](image)
In order to determine the chemical formula of a covalent compound, it is necessary to consider the number of electrons that are required by each atom to form a stable octet (or pair, as is the case for hydrogen).

**EXAMPLE**

1. For hydrogen and oxygen – oxygen requires two electrons to achieve a stable octet, however, a hydrogen atom has only one electron to share. Therefore, two hydrogen atoms are required, H₂O.

2. For carbon and oxygen – carbon requires four electrons to achieve a stable octet, however, each oxygen atom only requires two. Therefore, two oxygen atoms are necessary, CO₂.

When writing the formula, the element with the lowest Group number is written first. If two elements are in the same Group, the lower member of the Group is written first.

The following rules are used to name covalent compounds:

1. Name the element with the lowest group number (i.e. is furthest *left* on the Periodic Table) first.

2. Use a prefix to denote the number of atoms in each element (the prefix ‘mono’ is not used for the first element).

3. Change the suffix of the second element to ‘ide’.

<table>
<thead>
<tr>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
</tr>
</thead>
<tbody>
<tr>
<td>mono</td>
<td>di</td>
<td>tri</td>
<td>tetra</td>
<td>pent</td>
<td>hex</td>
<td>hept</td>
<td>oct</td>
<td>non</td>
<td>dec</td>
</tr>
</tbody>
</table>

**EXAMPLE**

1. N₂O₄ – dinitrogen tetroxide

2. SO₃ – sulphur trioxide

**QUESTION 24**

Determine the chemical formulae for the following molecular compounds:

(a) Boron trichloride

(b) Dihydrogen sulphide

(c) Carbon tetrahydride

(d) Phosphorous tribromide

(e) Sulphur dichloride
QUESTION 25
Name the following molecular compounds:

(a) HF
(b) BrCl
(c) NI₃
(d) H₃P
(e) CBr₄

QUESTION 26
Draw Lewis electron dot structures for the following ionic and covalent compounds:

(a) Aluminium chloride
(b) Nitrogen trihydride
(c) Calcium oxide
(d) Ethene, C₂H₄

QUESTION 27
Both hydrogen and sodium react violently with chlorine. With the aid of electron dot diagrams, compare and contrast the changes which occur at the atomic level in these two reactions, and the types of bonding in the resulting compounds.

Solution

(6 marks)
PHYSICAL AND CHEMICAL CHANGE

Syllabus:

- Identify the differences between physical and chemical change in terms of rearrangement of particles.
- Summarise the differences between the boiling and electrolysis of water as an example of the difference between physical and chemical change.

Physical changes do not lead to the formation of new chemical substances. Examples include changes of state, filtration, evaporation and change of shape or size. These processes generally involve very small amounts of energy.

Chemical changes form new chemical substances. Examples include burning, neutralising and decomposing. These processes involve larger quantities of energy.

A useful way to compare physical and chemical changes is by analysing the boiling versus the electrolysis of water. When water boils, the forces that exist between the molecules (intermolecular) are broken and the water changes state from a liquid to a gas. The chemical structure of the molecule does not change and no new chemical substances are formed. Therefore, the process is termed a physical change.

When an electric current is run through water, the energy is sufficient to break the forces within the water molecules (intramolecular) and create new ones. Therefore, new chemical substances are formed. This process, known as electrolysis, is an example of a chemical change. Chemical changes can be denoted by a chemical equation:

\[2\text{H}_2\text{O} (l) \rightarrow 2\text{H}_2 (g) + \text{O}_2 (g)\]
QUESTION 28
Heating water to boil it simply separates its molecules BUT electrolysis separates its atoms to form the elements. Which statement best explains this difference?

A Water is ionic and the ions move to opposite electrodes.
B Electrolysis provides much more energy per molecule than boiling.
C The polar water molecules respond to the applied electric field.
D Only electrolysis can break the hydrogen bonds in water.

QUESTION 29
It has been stated that the first industrial revolution was based on energy stored in steam and that the next might be based on hydrogen, also obtained from water, but storing about 7 times as much energy.

Describe a **first-hand** experiment you have performed to produce hydrogen from water AND explain why much more energy is required for this than to convert water to steam.

(4 marks)
Syllabus:

- Identify light, heat and electricity as the common forms of energy that may be released or absorbed during the decomposition or synthesis of substances and identify examples of these changes occurring in everyday life.

- Explain that the amount of energy needed to separate atoms in a compound is an indication of the strength of the attraction, or bond, between them.

- Plan and safely perform a first-hand investigation to show the decomposition of a carbonate by heat, using appropriate tests to identify carbon dioxide and the oxide as the products of the reaction.

Compounds can be decomposed and synthesised using various forms of energy, including heat, light and electricity. The diagram below shows the equipment that can be used to decompose carbonate compounds. For example:

\[ \text{MgCO}_3 (s) \rightarrow \text{MgO} (s) + \text{CO}_2 (g) \]

In order to confirm that the decomposition has been successful, it is necessary to test the products. Carbon dioxide can be identified by its reaction with limewater, creating a milky product. Magnesium oxide is a basic substance that can be identified using acid/base indicators.

UV light decomposes silver salts such as silver chloride. The white compound turns grey when exposed to UV light which indicates the formation of metallic silver. Chlorine vapour is released in the process.

\[ 2\text{AgCl} (s) \rightarrow 2\text{Ag} (s) + \text{Cl}_2 (g) \]
Many synthesis reactions are initiated by heat energy, such as the production of ammonia and the synthesis of iron (III) chloride.

\[
\text{N}_2 (g) + 3\text{H}_2 (g) \rightarrow 2\text{NH}_3 (g)
\]

\[
2\text{Fe} (s) + 3\text{Cl}_2 (g) \rightarrow 2\text{FeCl}_3 (s)
\]

Others occur in the presence of light and electricity.

\[
\text{H}_2 (g) + \text{Cl}_2 (g) \rightarrow 2\text{HCl} (g)
\]

\[
\text{N}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{NO} (g)
\]

The amount of energy required to separate atoms in a compound is an indication of the strength of the bonds present. For example, given that more energy is required to extract magnesium from its oxide than zinc, we can conclude that the ionic bonds between magnesium and oxide ions are stronger than those between zinc and oxide. In terms of covalent bonds, multiple bonds are stronger than single bonds. Molecules such as carbon dioxide and nitrogen require more energy to separate their individual atoms than water and hydrogen gas.

**QUESTION 30**

A student was asked to carry out a *first-hand* investigation on the effect of light on silver chloride.

(a) Describe a procedure that the student could follow.

(b) Identify the independent variable and the observations that the students should expect.

(c) Identify one risk posed in this investigation and describe how it can be minimised.
QUESTION 31
The flow chart below sets out one proposed method for removing carbon dioxide from the atmosphere, or from the waste products of combustion.

(a) Write a balanced ionic equation for the formation of calcium carbonate in Step 2; and show that this step regenerates sodium hydroxide solution for Step 1.

(b) With the aid of a balanced equation, describe the reaction in which calcium carbonate is decomposed in Step 3.

QUESTION 32
A student placed a piece of burning sodium in a gas jar containing of chlorine. The sodium burnt with a bright yellow flame forming clouds of a white, crystalline smoke.

(a) State whether energy was released or absorbed in this reaction. Justify your answer.

(b) Identify the white, crystalline substance that was synthesised in this reaction. Use an equation to support your answer.
STRUCTURE AND BONDING

**Syllabus:**

- Describe the physical properties used to classify compounds as ionic or covalent molecular or covalent network.

- Explain the relationship between the properties of conductivity and hardness and the structure of ionic, covalent molecular and covalent network structures.

Compounds can be classified into three groups or lattice types based on their physical properties.

<table>
<thead>
<tr>
<th>Type of compound</th>
<th>Ionic</th>
<th>Covalent molecular</th>
<th>Covalent network</th>
</tr>
</thead>
<tbody>
<tr>
<td>Particles forming the lattice</td>
<td>cations and anions</td>
<td>molecules</td>
<td>atoms</td>
</tr>
<tr>
<td>Forces holding the particles in the lattice</td>
<td>ionic bonds</td>
<td>intermolecular forces</td>
<td>covalent bonds</td>
</tr>
<tr>
<td>Melting point</td>
<td>high</td>
<td>low</td>
<td>very high</td>
</tr>
<tr>
<td>Other properties</td>
<td>hard, brittle</td>
<td>soft, brittle</td>
<td>very hard, brittle</td>
</tr>
</tbody>
</table>

**QUESTION 33**
Classify the following substances into one of the three lattice groups described above:

(a) silicon nitride
(b) barium iodide
(c) phosphorous
(d) iron chloride
Syllabus:

- **Distinguish between metallic, ionic and covalent bonds.**
- **Describe metals as three-dimensional lattices of ions in a sea of electrons.**
- **Describe ionic compounds in terms of repeating three-dimensional lattices of ions.**
- **Explain why the formula for an ionic compound is an empirical formula.**
- **Identify common elements that exist as molecules or as covalent lattices.**

A lattice describes the geometric arrangement of particles in a crystalline solid. The physical properties of lattices allow us to make inferences about their structure.

**METALLIC LATTICES**

A metallic lattice is comprised of positive metal ions that are surrounded by delocalised electrons that are free to move. The electrons are referred to as ‘delocalised’ as they have been lost from the valence shell of the metal atoms and now belong to the lattice as a whole. The attraction between the positive metal ions and the electrons is called a metallic bond, which is considered a strong chemical bond.

Metallic lattices have the following properties:

- High melting and boiling points – due to strong forces between particles in the lattice.
- High electrical and thermal conductivity – due to mobile electrons.
- High density – due to tightly packed metal atoms throughout the lattice.
- High lustre – light photons are absorbed and released by the mobile electrons.
- High malleability and ductility – the layers of ions within the lattice slide over each other when forces are applied and establish new bonds.

**IONIC LATTICES**

Ionic lattices consist of a three-dimensional arrangement of cations and anions. The oppositely charged ions are attracted to each other, forming ionic bonds, which are strong chemical bonds.

Ionic lattices have the following properties:

- High melting points – due to the strong ionic bonds throughout the lattice.
- Brittle – shearing forces cause ions of similar charge to come in close contact, leading to strong repulsion, causing the crystals to shatter.
• Non-conductors of electricity – as solids, ionic lattices do not have free electrons or free ions to carry charge. However, if a crystal is melted or dissolved, the ions are free to move, which makes them good electrical conductors.

Ionic compounds are represented by what is known as an empirical formula. The empirical formula of a compound represents its atomic or ionic composition expressed as a simple whole number ratio. The formula for sodium chloride, \( \text{NaCl} \), indicates that the ratio of sodium to chloride ions is 1:1.

**COVALENT NETWORK LATTICES**

Covalent network compounds are composed of a three dimensional array of atoms linked by strong covalent bonds. Examples include silicon dioxide and diamond.

Covalent network lattices have the following properties:

• Very high melting points – due to strong covalent bonds throughout the lattice.

• Non-conductors of electricity – due to an absence of free electrons or ions (except for graphite). Highly insoluble and so do not conduct in solutions.

• Very hard – due to the extensiveness of the strong covalent bonds. Diamond is the hardest natural substance.

**COVALENT MOLECULAR LATTICES**

Covalent molecules that are held in place by intermolecular forces to form a lattice structure. Examples include ice, dry ice (solid carbon dioxide), sulphur, iodine and phosphorous.

Covalent molecular lattices have the following properties:

• Low melting points – due to the weak intermolecular forces, which require little energy to break.

• Non-conductors of electricity – due to an absence of mobile electrons or ions.

• Soft and brittle – due to the weak intermolecular forces which make them.