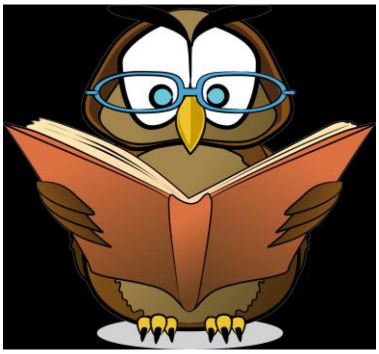


KNOWLEDGE

ST MARY'S SCIENCE DEPARTMENT:

APPLIED SCIENCE

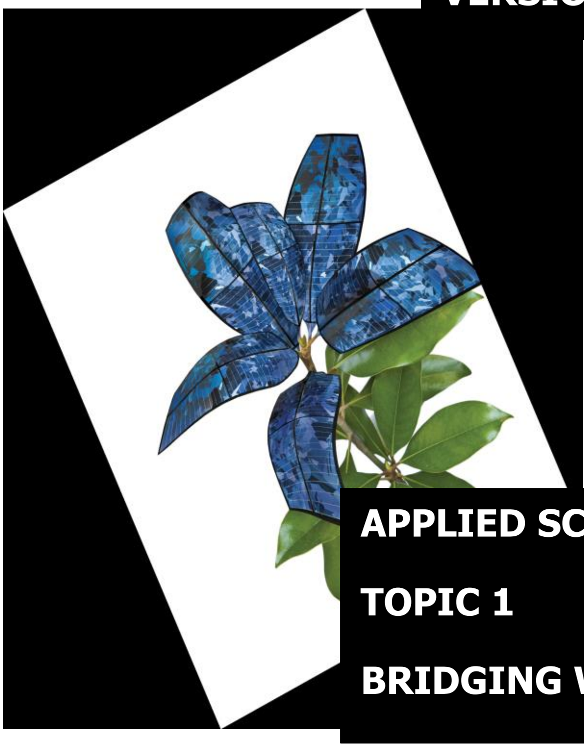
APPLIED SCIENCE BRIDGING COURSE

WEEK 1

Periodicity and Properties of Elements

NAME	
PHYSICS CLASS	

VERSION 1.1



APPLIED SCIENCE

TOPIC 1

BRIDGING WORK

THIS MUST BE BROUGHT TO LESSONS AT THE START OF YEAR 12.

WEEK 1

Contents

Overview

This bridging course will provide you with a mixture of information about Extended Certificate Applied Science and what to expect from the course, as well as key work to complete. Students who are expecting to study Extended Certificate Applied Science, and are likely to meet the entry requirements, must complete the bridging course fully and thoroughly, to the best of their ability.
You should complete all work on paper and keep it in a file, in an ordered way. You will submit it to your teacher in September.

All of the work will be reviewed, and selected work will be assessed, and you will be given feedback on it. This work will be signalled to you. If you do not have access to the internet, please contact the school and appropriate resources will be sent to you.

If you are thinking about studying Extended Certificate Applied Science, you should attempt this work to see whether or not you think studying a subject like this is right for you. If you later decide to study Extended Certificate Applied Science, you must ensure you complete this work in full.

This work should be completed after you have read and completed the Study Skills work that all of Year 12 should complete.

Introduction

Welcome to the Science Department at St Mary's Catholic School. I am delighted that you have chosen to study this qualification and hope you feel comfortable with your study environment, and that you value the support you will be given by your teaching staff. This book has been designed to inform you about your programme of study; including the ways you will be assessed and how to fully understand the grade you can achieve on completion.

BTEC courses do work differently to other subjects and you will be expected to work hard both in and out of your lesson to meet coursework deadlines. You will also be presented with many different opportunities to broaden your vocational learning, as this qualification contains a wide range of contemporary topics pertaining to Applied Science. A variety of assessment methods are used, ranging from external exams to course work. Additionally, this BTEC qualification has been designed with employers and representatives from higher education and professional bodies. In this way, the qualification is up to date and covers all of the knowledge, skills and attributes that are required across a range of technical science sectors.

I have high expectations of all students studying this BTEC qualification and all of the teachers in the department have researched and prepared in great depth to ensure that you receive quality teaching. Please use this book to get a feel of what the course is about and to ensure you are fully aware of what is expected of you. The Science Department has planned lots of innovative and creative ways to deliver the programme. This will help to promote your learning and successfully prepare you for the various assessments.

The Science Department has planned an exciting year ahead. We wish you the very best in your academic pathway and we look forward to working with you to achieve great success!

Course Overview

The BTEC Applied Science Extended Cerifiicate is a 2 year course comprising of 4 units.

Year 12
Unit 1 – 90 GLH
Unit 2 – 90 GLH
Year 13
Unit 3 – 120 GLH
Unit 10 – 60 GLH

GLH stands for Guideline Learning Hours, and represents the amount of time you will spend on each part of the course, and also how much each unit contributes to the final grade. You will see that Year 12 and 13 contribute equally to the final grade. In Year 12 Units, 1 and 2 are equally split and will take the same amount of time. In Year 13 Unit 3 contributes 66% of the mark and will take 66% of the time.

Each unit is awarded a grade (Distinction, Merit, Pass or Unclassified). Each grade is worth a certain number of marks.

	Unit 1*	**Unit 2**	**Unit 3***	**Unit 10**
Distinction	24	24	32	16
Merit	15	15	20	10
Pass	9	9	12	6
Unclassified	0	0	0	0

*These units are examined. For examined units marks may be awarded between these boundries. These marks are the minimum you need for a grade.

To arrive at your final grade your marks are added together. The final grade boundries are shown in the table.

Grade	**Points**	**A-Level Equivelent Grade**
Distinction*	90	A*
Distinction	74	A
Merit	52	C
Pass	36	E

To gain a pass you must achieve a pass in all units. If you fail a unit you will fail the course regardless of your score.

In units 2 and 10, a numbes of assignments contribute to each unit. Your grade for that unit will be the lowest grade you achieve. For example if you achieved a P, M, D, D in Unit 2, your grade for that unit would be a Pass.

Assessment Schedule

In Year 12 you will complete Unit 1 and Unit 2

Unit 2

This will be taught between the start of the year and Christmas.

You will complete 4 assignments

Assignment A – Titration and Colorimetry.
You will investigate methods to determine the pH of a solution. You will also use colorimetry to determine the concentration of a solution.

Assignment B – Calorimetry.
You will investigate different methods of measuring temperature. You will also use a cooling curve to determine the rate of cooling, and the melting point of different substances.

Assignment C – Chromatography.
You will use chromatography techniques to identify substances in a plant leaf extract and a mixture of amino acids.

Assignment D – Personal Review.
You will produce an assignment evaluating the personal and technical skills that you have developed over the unit.

This unit will be assessed by your teacher. All work must be submitted by the given deadline. A sample of assignments will be moderated by a Pearson moderator. Remember that you must pass the unit to pass the course, and that the grade awarded will be the lowest one you achieve.

You will be given interim and final deadlines for each assignment which must be met.

Unit 1

Unit 1 is made up of Biology, Chemistry and Physics theory, and will be taught by specialist teachers.
The level of difficulty is around that of A-level.

This will be taught between Christmas and May.

You will complete a Biology, Chemistry and Physics exam, each of which are worth 30 marks and last 40 minutes.

Biology

Chemistry

Physics

You are allowed to resit each exam once.

Aim

In this bridging course, we will outline the basic principles of the key topics covered in Unit 1.

In each topic, we will start by reviewing the understanding which you gained in GCSE Science and apply it to more advanced applications found in Extended Certificate Applied Science.

This is not a comprehensive overview of the Extended Certificate Applied Science specification, rather a taster on what is covered throughout the course.

This bridging course should give you an experience of the level you will be expected to study at, at the start of Year 12

Important

Please remember to look after your own wellbeing as you work through this bridging course.

Please take regular breaks as you go through this work.

This work should take approximately 5 hours, so should not be completed in one sitting.

Do not worry or panic if there is something challenging or which you do not understand at first. This is completely normal.

If you do not understand a concept after reviewing this work, please contact Mr. Turnbull on his school e-mail address.

WEEK 1: CHEMISTRY

TOPIC 1: THE ELECTRONIC STRUCTURE OF THE ATOM
SPEC CHECK

Specification	Learning Outcomes
Understand the electronic structure of atoms:	• Know that atoms have electron shells / energy levels which consist of subshells (s, p and d) • Know the number and type of subshells in the first four energy levels • Know how to place these subshells in order of increasing energy • Be able to write the electronic structure (configuration) of the first 36 elements of the periodic table using s, p and d subshell notation
Electronic orbitals	• Know that subshells contain electronic (atomic) orbitals • Know that an orbital is a region of space where an electron is likely to be found • Know that an orbital can hold up to 2 electrons • Know the shapes and orientation of s and p orbitals • Be able to interpret electron density plots for s and p orbitals • Know the number and type of orbitals in each subshell for the first 36 elements
Aufbau principle	• Understand that to predict the electronic configuration of an atom, electrons fill the shells, subshells and orbitals of lowest energy first • Understand that orbitals with the same energy must be filled singly before electrons are paired • Understand that when electrons are paired, the spins of the two electrons are opposite to each other, in order to reduce repulsion • Be able to use the electron-in-box model to show how electrons fill the orbitals in atoms of the first 36 elements
Bohr theory	• Understand that electrons occupy shells or energy levels, orbiting the nucleus of the atom • Understand that an electron can move from its ground state energy level to a higher energy level by absorption of a quantum of radiation

Read the information found in the key information to understand the concepts of this topic.

KEY INFORMATION

Bohr's model describes the shape in space where electrons can be found in atoms.

Main Energy Levels or Shells

Electrons orbit the nucleus in shells.
The further the shell, or main energy level, is from the nucleus, the higher its energy.
Each shell can hold a fixed number of electrons.

n	Shell	Number of Electrons
1	1st shell	2
2	2nd shell	8
3	3rd shell	18
4	4th shell	32

You can find the total number of electrons in an atom by using the atomic number.
This is defined as the number of protons in an atom, but for a neutral atom also equals the number of electrons.

Atomic Orbital

Each shell consists of atomic orbitals. These are regions in space where electrons may be found.
The larger the main energy level, the more orbitals it is made up from.
The first four types are called s, p, d and f orbitals.

An s-orbital is spherical in shape and has 2 electrons in it.
Each p-orbital is dumb-bell shaped and holds up to 2 electrons. Each p sub-shell has 3 p orbitals, p_x, p_y and p_z.

Writing Electron Configurations

Electron configurations are the arrangement of the electrons in atoms or ions.
Bohr's model recognises that the main shells are split into sub-shells and electron configurations should reflect that.

Element	Electron Configuration
B	$1s^2 2s^2 2p^1$
C	$1s^2 2s^2 2p^2$
N	$1s^2 2s^2 2p^3$
O	$1s^2 2s^2 2p^4$

In these diagrams, each box represents an orbital.
The arrows represent the electrons in the orbitals.

Energy →

			1s	2s	2px	2py	2pz
Lithium,	Li	$1s^2 2s^1$	↑↓	↑			
Beryllium,	Be	$1s^2 2s^2$	↑↓	↑↓			
Boron,	B	$1s^2 2s^2 2p^1$	↑↓	↑↓	↑		
Carbon,	C	$1s^2 2s^2 2p^2$	↑↓	↑↓	↑	↑	
Nitrogen,	N	$1s^2 2s^2 2p^3$	↑↓	↑↓	↑	↑	↑
Oxygen,	O	$1s^2 2s^2 2p^4$	↑↓	↑↓	↑↓	↑	↑
Fluorine,	F	$1s^2 2s^2 2p^5$	↑↓	↑↓	↑↓	↑↓	↑
Neon,	Ne	$1s^2 2s^2 2p^6$	↑↓	↑↓	↑↓	↑↓	↑↓

The direction of the arrows shows the **SPIN OF EACH ELECTRON.**
A pair always spin in opposite direction as it reduces the repulsion, they experience between each other.
Orbitals can be confused with sub-shells. Any orbital, regardless of the sub-shell it is in, can hold up to two electrons.

Rules for Arranging Electrons
1. Start at the lowest shell and add electrons one at a time to build up the configuration.
2. Fill each sub-level before starting on the next one.
Remember that the 4s orbital is lower in energy than the 3d when empty, but higher in energy when occupied.
3. Fill each orbital singly in a sub-level before pairing electrons.
4. Paired electrons have opposite spins, so they are shown as arrows pointing in opposite directions.

The ideas that electrons fill lower shells before the higher shells is called the **AUFBAU PRINCIPLE.**

REVISION SHEET

Highlight or underline the key information on the revision sheet to consolidate your understanding.

In this section you will revisit the structure of atoms, then see how that structure can result in different types of bonding. You will also learn about different types of interaction between molecules which determine the properties of materials. At the end of this section you will learn about important chemical quantities and how to calculate them from a properly balanced equation.

The Electronic Structure of Atoms

Definitions

Proton	a positively charged subatomic particle found in the nucleus of all atoms
Neutron	an uncharged subatomic particle found in the nucleus of all atoms except hydrogen
Electron	a negatively charged subatomic particle found orbiting the nucleus in all atoms
Bohr model	a simple atomic model with a small positively charged nucleus surrounded by defined rings of negatively charged electrons
Shell	an energy level occupied by electrons which surrounds the nucleus
Orbital	a subshell with a certain energy level, which can hold up to two electrons
Aufbau principle	electrons fill the lowest available energy shells first, and singly occupy orbitals in the same energy level before pairing up

The basic model of the atom has two main components – the positively charged nucleus, which contains positive **protons** and neutral **neutrons**, and the cloud of negatively charged **electrons** which encircle it. In all atoms, the number of protons equals the number of electrons. Protons and neutrons are far larger and heavier than electrons.

Mass number

$$^{23}_{11}\mathrm{Na}$$

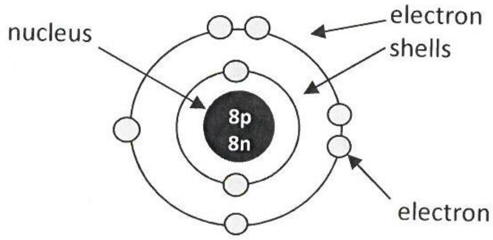
Atomic number

11 electrons
11 protons
12 neutrons

The number of electrons is equal to the number of protons. The number of protons is given by the **atomic number**.

The number of protons plus neutrons in an atom is given by the **mass number**.

In the **Bohr model** of the atom, electrons are arranged in **shells** around the nucleus, which have higher energy levels as you move further away from the centre. These shells of electrons orbit around the nucleus at fixed distances, rather like planets orbiting the sun. Each shell has a maximum number of electrons that can fit into it – once it is full, extra electrons must go into the next shell.



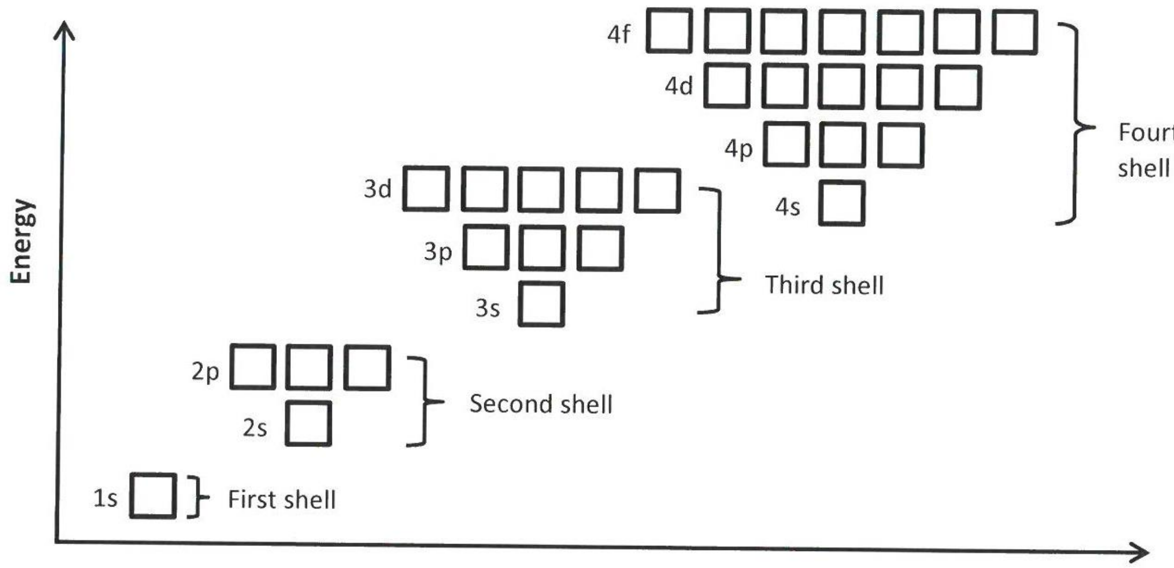
$^{16}_{8}O$
8 electrons
8 protons
8 neutrons

Shells and orbitals

Within a shell, not all electrons have the exact same energy – they are split into **subshells** which have different energies. Subshells are further divided into **orbitals**. Each orbital can hold **two** electrons with different **spins**.

Electron shell	Maximum number of electrons	Orbitals in shell
1	2	1 × s
2	8	1 × s, 3 × p
3	18	1 × s, 3 × p, 5 × d
4	32	1 × s, 3 × p, 5 × d, 7 × f
5	50	1 × s, 3 × p, 5 × d, 7 × f, 9 × g

Orbitals are filled according to the **Aufbau principle** – the lowest available energy level (the one closest to the nucleus) is filled first, before filling ones at higher levels.



Orbitals are labelled s, p, d or f, depending on which subshell they are in – p orbitals are higher in energy than the s orbitals in the same shells, d orbitals are higher than p orbitals, and so on.

Rules for filling orbitals

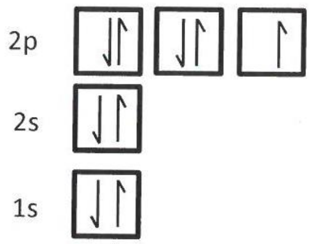
- ✓ Each box represents an orbital which can hold two electrons.
- ✓ An electron is represented as an arrow with half a head.
- ✓ The lowest energy levels are filled first.
- ✓ The 4s subshell is lower in energy than the 3d subshell, so is filled first (important for transition metals).
- ✓ All orbitals in a subshell are filled singly before electrons start to pair up (pairing up requires more energy due to repulsions between electrons in the same orbital).

By arranging electrons according to these rules, the most stable **electron configuration** is obtained.

Example

a) Use the box method and s/p/d notation to write out the electron configuration of fluorine.

Fluorine has nine electrons in total.



In s/p/d notation, this would be written as:

$$1s^2 2s^2 2p^5$$

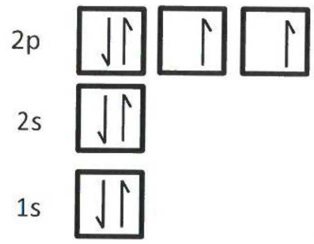
Number of electrons in the subshell

Number of the electron shell

Type of orbital in the subshell

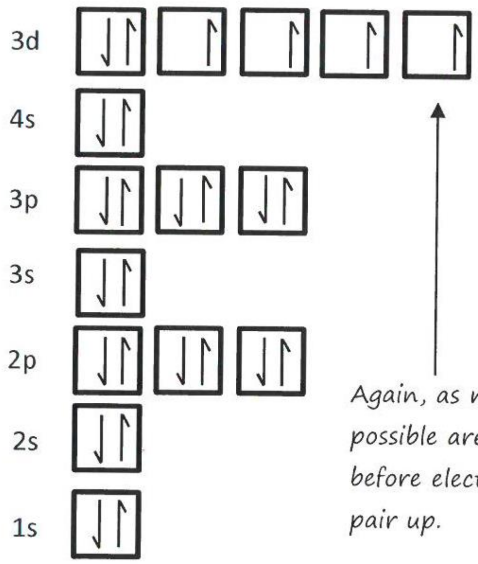
b) Use the box method and s/p/d notation to write out the electron configuration of oxygen.

Oxygen has eight electrons, and in s/p/d notation this is written as $1s^2 2s^2 2p^4$.



The 2p subshell can be filled with four electrons in two ways: with one doubly occupied orbital and two singly occupied orbitals, or two doubly occupied orbitals and one empty orbital. However, pairing up electrons is unfavourable, so the arrangement with the largest number of unpaired electrons is always used.

c) Use the box method to write out the electron configuration of iron.



Iron has 26 electrons, and in s/p/d notation this is written as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. Note that the 4s orbital is lower in energy than the 3d orbital, so it gets filled first.

Again, as many orbitals as possible are singly occupied before electrons start to pair up.

Credit: Zig Zag Resources Revision Guide Editions

VIDEO

To watch a video looking at this concept, please scan one of the following codes with your smartphone.

SCAN ME

Note

All rights to this video belong to the creator of the video.

This external third-party content – please check the contents are appropriate.

SELF ASSESSMENT

To practice your understanding, answer the following questions.

DO NOT WORRY IF YOU STRUGGLE AT FIRST.

The answers are found after the questions.

A1. How many electrons are there in:
A1.1 One atom of sulfur?

……

……

A1.2 A full p orbital?

……

……

A1.3 A full third electron shell.

……

……

A2. Draw the electron configuration of the following atoms using the box method and s/p/d notation.

A2.1 A nitrogen atom

A2.2 A vanadium atom

A3. Draw a Bohr-model diagram of the atomic structure of magnesium and label the components.

A4. What shell is the outer electron of a phosphorous atom in? Explain your answer.

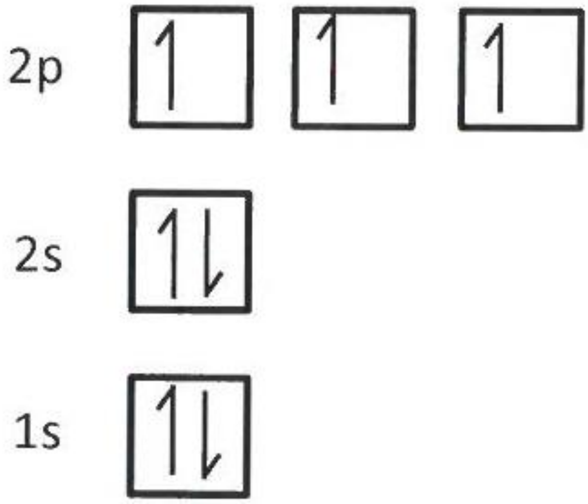
……

……

ANSWERS

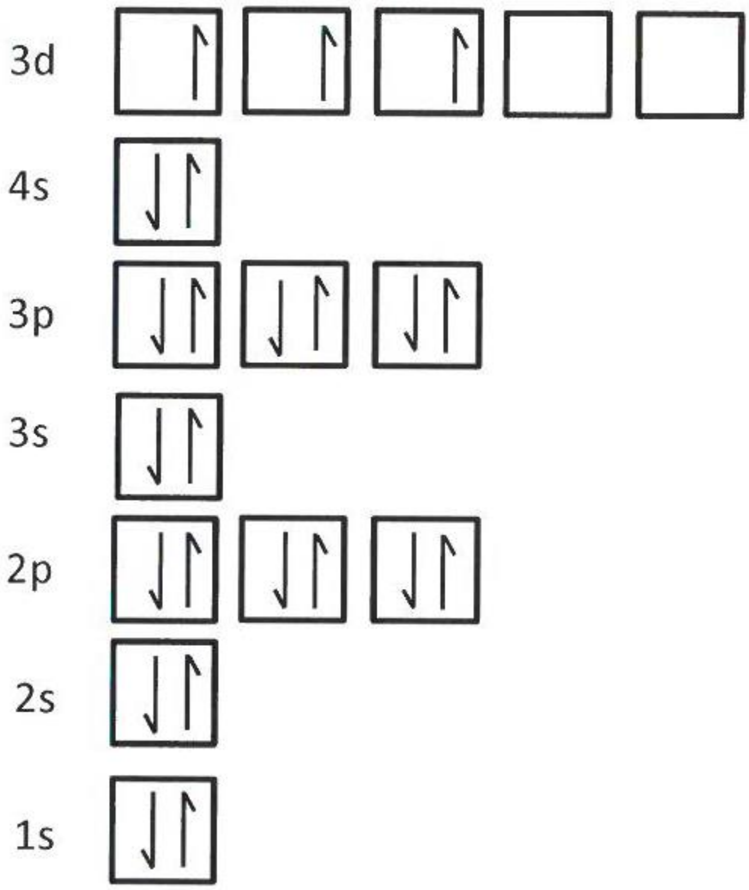
A1.
A1.1 16
A1.2 2
A1.3 18

A2.1



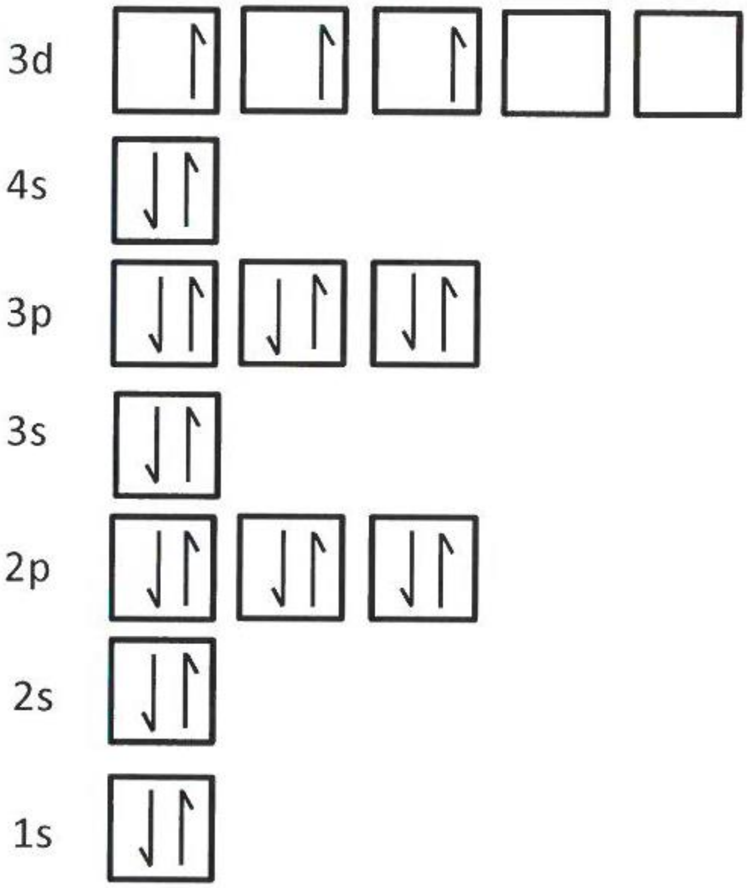
N = $1s^2 2s^2 2p^3$

A2.2 A vanadium atom



V = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

A3.



$V = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

A4. Phosphorus has 15 electrons, so the outer electron is in the third shell.

ASSESSMENT QUESTION

Please answer this assessment question on this topic in Applied Science.

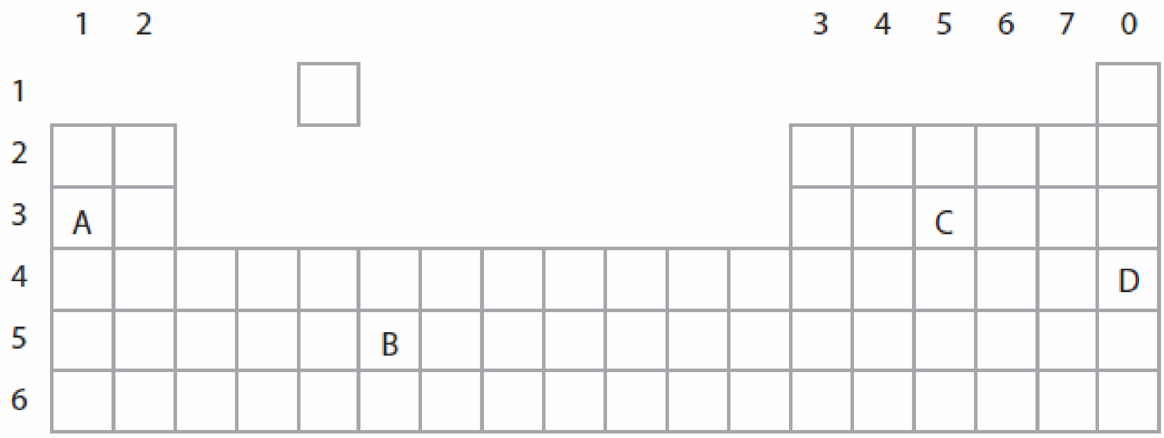
To submit this work, please scan in this sheet or take a picture of this work and send it to Mr. Turnbull via e-mail. If you cannot print out the sheet, please just answer the questions on a blank piece of paper.

This work will be formally assessed with feedback given.

1.

Chemists use the periodic table to predict the chemical properties of elements based on their location in the table.

Part of the periodic table is shown.



The letters, **A, B, C** and **D**, represent four different elements.

1.1 Which element, **A**, **B**, **C** or **D**, is in the d-block of the periodic table?

[1 Mark]

- ☐ **A**
- ☐ **B**
- ☐ **C**
- ☐ **D**

1.2 Which element, **A**, **B**, **C** or **D**, has the most stable electronic structure?

[1 Mark]

- ☐ **A**
- ☐ **B**
- ☐ **C**
- ☐ **D**

1.3 Which element, **A**, **B**, **C** or **D**, reacts violently with water?

[1 Mark]

☒ **A**

☒ **B**

☒ **C**

☒ **D**

1.4 An element has the electronic configuration $1s^2\ 2s^1$.
Identify which period the element is in.

[1 Mark]

....

....

1.5 Complete the electronic configuration for an atom of sodium.

[1 Mark]

$1s^2\ 2s^2$

....

Reference: BTEC Applied Science Specimen Materials 2

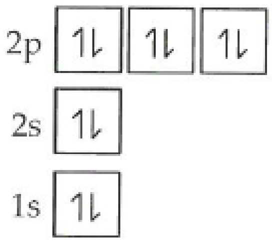
2. Electrons in atoms fit into shells made up of orbitals.

2.1 How many electrons fit in one orbital?

[1 Mark]

A	One	
B	Two	
C	Four	
D	Eight	

The electronic structure of a neon atom can be represented as:



2.2 In this way of writing an electronic structure, what does each box represent?

[1 Mark]

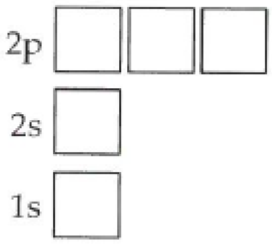
....

....

....

2.3 Write the electronic structure of an oxygen atom in this way.

[3 Marks]



Electronic configuration of atoms can be represented in different ways.
The electronic configuration of beryllium atom can be represented as $1s^2s^2$.
2.4 Write the electronic configuration of a nitrogen atom in this way.

[1 Mark]

....

....

....

2.5 In $1s^2$, what does 1 refer to?

[1 Mark]

A	Number of electrons	☐
B	Number of the shell	☐
C	Number of the orbital	☐
D	Number of atoms	☐

Reference: Zig Zag Educational Resources

TOPIC 2: IONIC BONDING
SPEC CHECK

Specification	Learning Outcomes
Understand ionic bonding:	• Be able to predict whether a compound has ionic bonding from its name or formula • Be able to predict the formula of an ionic compound from its elements (for groups 1, 2, 3, 6 and 7) • Understand how the physical properties of ionic substances, such as melting and boiling point, solubility and electrical conductivity, are affected by their bonding and structure
Strong electrostatic attraction between oppositely charged ions	• Understand that ionic bonding is a result of strong electrostatic attraction between positive and negative ions • Understand that electrostatic attraction between ions can occur in any direction • Understand that a giant ionic structure is a lattice of many ions held together by electrostatic attraction
Effects ionic radius and ionic charge have on the strength of ionic bonding	• Understand that the strength of electrostatic attraction increases with increasing ionic charge • Understand that the strength of electrostatic attraction decreases with increasing size of ionic radius • Understand why ionic radius increases down a group • Understand how ionic radius changes across a period (for groups 1, 2, 3, 5, 6 and 7) • Be able to predict differences in the strength of ionic bonding for different ionic compounds **(NB an understanding of polarisation is not required)**
Formation of ions in terms of electron loss or gain	• Be able to describe the formation of positive ions (cations) by the loss of electron(s) from metal atoms • Be able to describe the formation of negative ions (anions) by the gain of electron(s) by non-metal atoms • Be able to draw dot-and-cross diagrams for ionic compounds of groups 1, 2, 3, 6 and 7 elements, showing outer electrons and correct charges

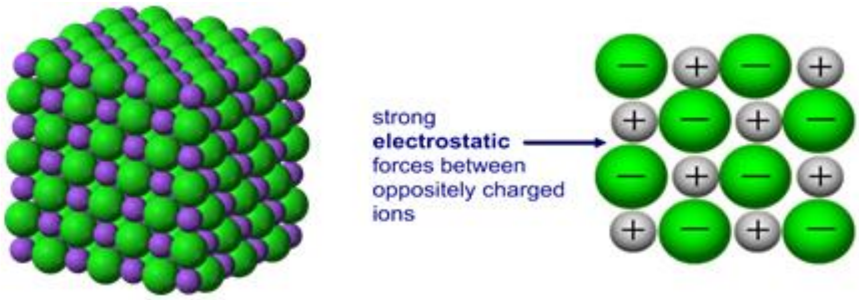
Read the information found in the key information to understand the concepts of this topic.

KEY INFORMATION

Ionic bonds are strong electrostatic attractions between positive and negative ions.
The ions in ionic bonds can be shown using electron configuration diagrams.

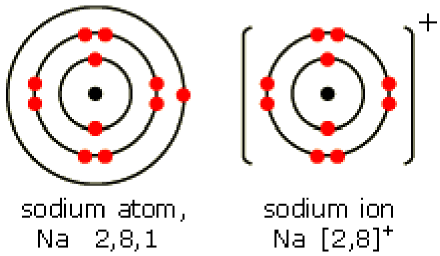
Ionic Compounds and Giant Ionic Structures

Ionic compounds have giant structures.
In a giant ionic structure, the ions are arranged in a regular, 3D pattern called a lattice.
The electrostatic forces between the ions act in all directions and keep the structure together.
The large number of these strong electrostatic attractions give ionic compounds high melting points.



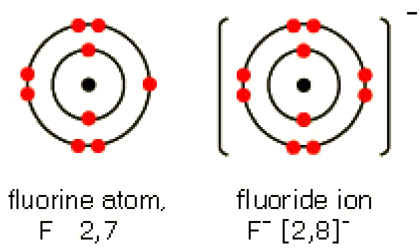
Positive Ions

Are generally formed by metal atoms losing electrons.
Have positive charge equal to the group number if formed from a group 1, 2 or 3 element.
Have different charges if formed from a transition metal (for example, Fe^{2+}, Fe^{3+}).
Can be represented in an electron configuration diagram, for example a Na^+ ion.
Are known as cations.



Negative Ions

Are generally formed by non-metal atoms gaining electrons from metal ions.
Have a negative charge equals to 8 minus the group number of the element.
Sometimes exist as polyatomic ions, such as CO_3^{2-}, SO_4^{2-}, NO_3^- and OH^-, whose charges should be learnt.
Can be represented in an electron configuration diagram, for example an F^- ion.
Are known as anions.



Strength of Ionic Bonds

To compare the relative strength of ionic bonds, the ionic charge and ionic radius have to be considered, sometimes called the charge/size ratio.

For instance, the ionic bonding in MgF_2 is much stronger than the ionic bonding in NaF. This is because the magnesium ion is smaller than the sodium ion and also has a greater charge. These factors increase the electrostatic attraction between the ions.

REVISION SHEET

Highlight or underline the key information on the revision sheet to consolidate your understanding.

Ionic bonding

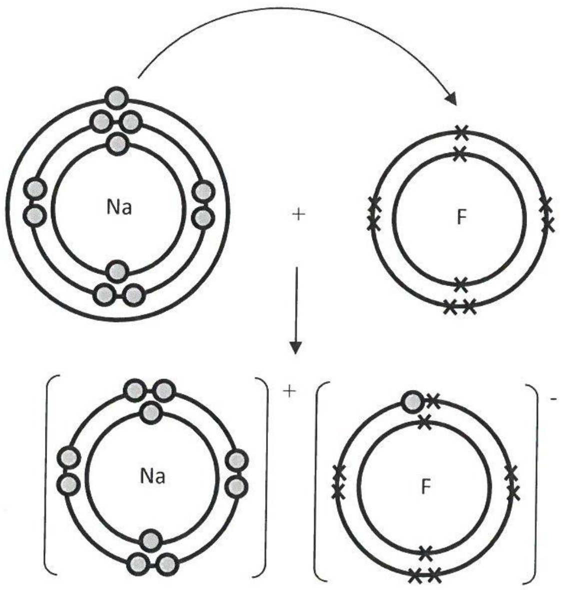
	Definitions
Electrostatic attraction	attractive forces between opposite charges
Noble gas configuration	a full shell of outer electrons; the most energetically stable arrangement
Giant ionic lattice	an arrangement of negative ions and positive ions in a regular 3D pattern

Ionic bonds are formed from the **electrostatic attraction** between oppositely charged ions.

- Ions are formed by the loss or gain of electrons.
- The atom losing electrons becomes positively charged (a **cation**).
- The atom gaining electrons becomes negatively charged (an **anion**).

By **exchanging** electrons, both elements are able to achieve a full outer shell of electrons, which is the most stable electronic configuration. This is also sometimes called a **noble gas configuration** because all noble gas atoms already have a full outer shell (which is the reason they are so unreactive).

Sodium has one outer electron, which it needs to lose to have a full outer shell



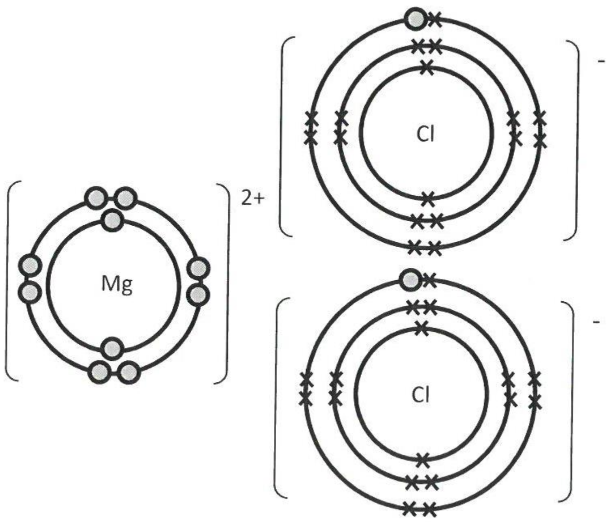
Fluorine has seven outer electrons, and needs one more to gain a full outer shell

Sodium gives one electron to fluorine and, therefore, becomes positively charged

Fluorine accepts one electron from sodium and, therefore, becomes negatively charged – both elements now have a full outer shell of electrons

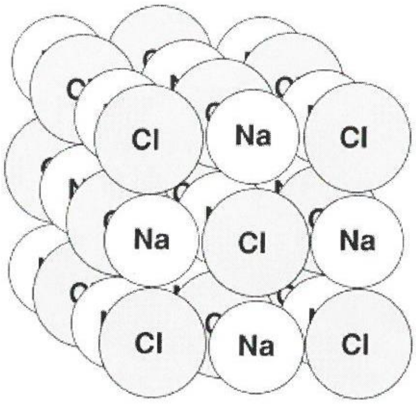
The ratio of positive ions to negative ions does not have to be 1:1, e.g. in $MgCl_2$.

Mg has two outer electrons, which it loses to become a 2+ ion

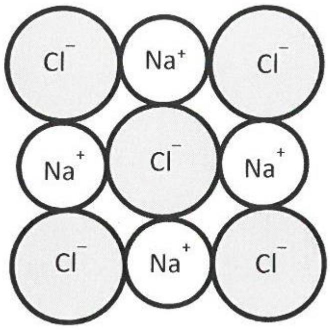


The opposing charges must cancel out, so for every Mg^{2+} ion there must be two Cl^- ions

Ionic bonds are most commonly observed between a metal and a non-metal, such as sodium and chlorine. Oppositely charged ions join together into a regularly ordered **giant ionic lattice**.



3D representation



2D representation

The strength of the ionic bond is affected by:

- ionic **charge**
- ionic **radius**

Smaller ions with higher charges result in stronger electrostatic attractions. For example:

- NaCl has stronger ionic bonding than KCl – even though sodium ions and potassium ions have the same charge, Na^+ is a smaller ion
- $MgCl_2$ has stronger bonding than NaCl – the ions are of a similar size, but magnesium ions have a charge of +2, whereas a sodium ion has a charge of +1

Credit: Zig Zag Resources Revision Guide Editions

VIDEO

To watch a video looking at this concept, please scan one of the following codes with your smartphone.

SCAN ME

Note

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SELF ASSESSMENT

To practice your understanding, answer the following questions.

DO NOT WORRY IF YOU STRUGGLE AT FIRST.

The answers are found after the questions.

A1.1 Define the term 'ionic bond' and describe how an ionic bond forms.

....

....

....

....

A2.1 Draw a dot-and-cross diagram to show the bonding in CaI_2.
Only draw the outer shells of the electrons.

A3.1 Will the bonding be stronger in NaF or NaCl? Explain your answer.

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....

A3.2 What about $CaCl_2$ and KCl?

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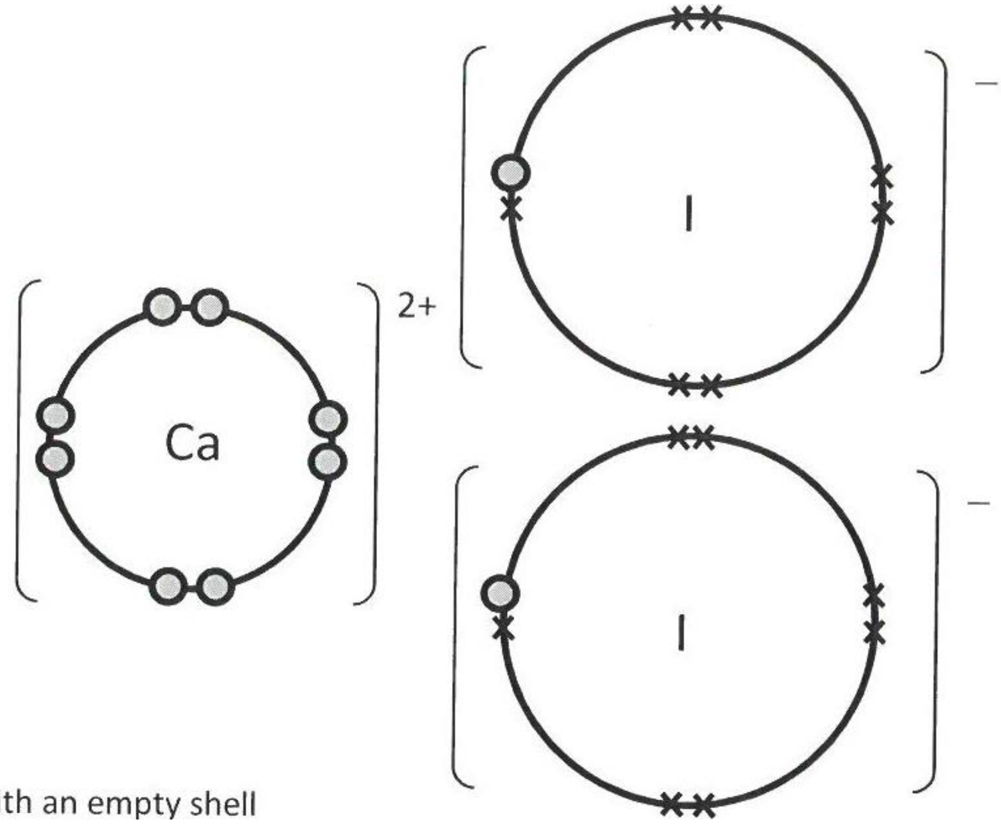
....

A4. Draw a 2D representation of the lattice structure of MgO.

ANSWERS

A1.1 An ionic bond is the electrostatic attraction between two opposing charges. Electrons are transferred from one atom to another to generate a pair of ions.

A2.1



Also accept Ca with an empty shell

A3.1 NaF has the stronger bonding – F^- has a smaller ionic radius than Cl^-, thus has a higher charge density and stronger electrostatic attraction.

A3.2 $CaCl_2$ has stronger bonding – $Ca2^+$ has a higher charge than Na^+, thus a higher charge density and stronger electrostatic attraction.

A4. Draw a 2D representation of the lattice structure of MgO.

ASSESSMENT QUESTION

Please answer this assessment question on this topic in Applied Science.

To submit this work, please scan in this sheet or take a picture of this work and send it to Mr. Turnbull via e-mail. If you cannot print out the sheet, please just answer the questions on a blank piece of paper.

This work will be formally assessed with feedback given.

1. Potassium chloride, KCl, is sometimes used as an alternative to sodium chloride, NaCl, in table salt.
Both of these compounds are ionic.
1.1 Describe how the bonding in KCl is formed.

[3 Marks]

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1.2 In KCl, ions are attracted towards each other. Why are the ions in KCL attracted to each other?

[2 Marks]

……

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1.3 Draw a dot and cross diagram, using outer electrons only, to show the bonding in KCl

[2 Marks]

Reference: Zig Zag Educational Resources

2. This question is about the properties of ions and ionic compounds.
Solid calcium carbonate, $CaCO_3$, has a giant ionic structure.

2.1 Draw a diagram (using dots or crosses) for a calcium **ion**.
Show **ALL** the electrons and the charge on the ion.

[1 Mark]

2.2 Complete the electronic configuration for a calcium **ion**.

[1 Mark]

$1s^2$

………

2.3 Would you expect a calcium ion to be bigger, smaller or the same size as a calcium atom?
Give **TWO** reasons to explain your answer.

[2 Marks]

………

………

………

………

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2.4 Explain why ionic compounds have relatively high melting temperatures.

[2 Marks]

………

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………

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Reference: EdExcel A-Level Chemistry Examination Resources

TOPIC 3: COVALENT BONDING

SPEC CHECK

Specification	Learning Outcomes
Understand covalent bonding:	• Be able to predict if a compound has covalent bonding from its name or formula • Understand how the physical properties of covalent substances, such as melting and boiling point, solubility and electrical conductivity, are affected by their bonding and structure (to include simple molecular and giant covalent structures)
Strong electrostatic attraction between two nuclei and the shared pair(s) of electrons between them	• Understand that covalent bonding involves the sharing of a pair of electrons between two atoms and that there is strong electrostatic attraction between the nuclei and the electrons being shared • Understand that electrostatic attraction between nuclei and the shared electrons is localised and in a specific direction • Understand that a giant covalent structure is a lattice of many atoms bonded covalently
Dot and cross diagrams to show electrons in simple covalent molecules, including those with multiple bonds and dative covalent (coordinate) bonds	• Be able to draw dot-and-cross diagrams for simple covalent molecules, showing outer electrons • Know that multiple bonds involve two or more pairs of electrons being shared • Know that a dative covalent (coordinate) bond is a covalent bond in which the pair of electrons being shared is donated by one atom
The relationship between bond lengths and bond strengths in covalent bonds	• Understand that as the number of shared pairs of electrons between two atoms increases, the bond length decreases • Understand that as the number of shared pairs of electrons between two atoms increases, the bond strength increases • Be able to represent covalent bonds in substances as 2D line diagrams
Tetrahedral basis of organic chemistry	• Understand that electron pairs in bonds repel each other in order to be as far apart as possible around a central atom • Know that carbon has four outer shell electrons so can form up to 4 single bonds/bonding electron pairs and will form tetrahedral shapes • Know the bond angle associated with 4 bonding electron pairs around a central carbon atom is 109.5° • Be able to represent covalent bonds in simple organic molecules as 3D line diagrams

Read the information found in the key information to understand the concepts of this topic.

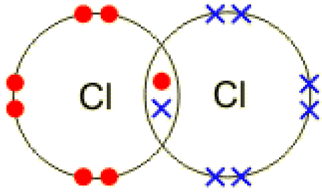
KEY INFORMATION

A covalent bond is an electrostatic attraction between a shared pair of electrons and the nuclei of the bonded atoms.

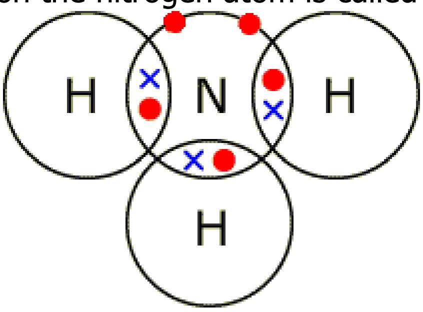
A covalent bond forms when atoms share a pair of electrons.
Generally each atom in the bond contributes to one electron to the pair, but a covalent bond consisting of an electron pair derived from one of the atoms is called a **DATIVE COVALENT (COORDINATE) BOND.**

Dot and Cross Diagrams

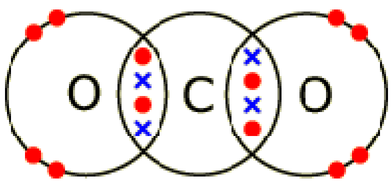
These are ways of showing the bonding between atoms, for instance in a chlorine molecules.
Dots represent electrons from one atom and crosses from the other.
The shared pair are in overlapping shells between the atoms – this is a covalent bond.



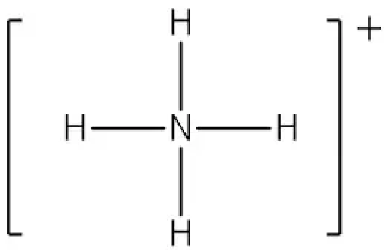
Ammonia consists of 3 hydrogen atoms bonded to a nitrogen atom. Each hydrogen atom has one 1 unpaired electron, so can form a single bond.
The nitrogen atoms have 3 unpaired electrons in its outer shell, so can form 3 bonds.
The pair of non-bonded electrons on the nitrogen atom is called a lone pair.



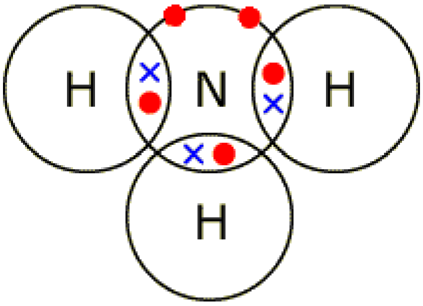
In carbon dioxide, each oxygen atom shares 2 pairs of electrons with the carbon atom.
Each oxygen forms a double bond to the carbon.



The ammonium ion forms from ammonia, NH_3 and a hydrogen ion, H^+.
The hydrogen ion has no electrons in its vacant orbital to form a bond, but the nitrogen atom in ammonia has a lone pair not involved in bonding. It uses this pair to bond to the hydrogen ion, forming a **DATIVE COVALENT BOND.**



Methane consists of 1 carbon atom bonded to 4 hydrogen atoms. The orbitals containing the electron pairs repel as far away as possible, forming a **TETRAHEDRAL** shape, common around carbon atoms in organic chemistry.



Strength of Covalent Bonds

Bond length and bond strength in covalent bonds are **INVERSELY RELATED**. This means that the shorter the covalent bond length, the greater the covalent bond strength.

REVISION SHEET

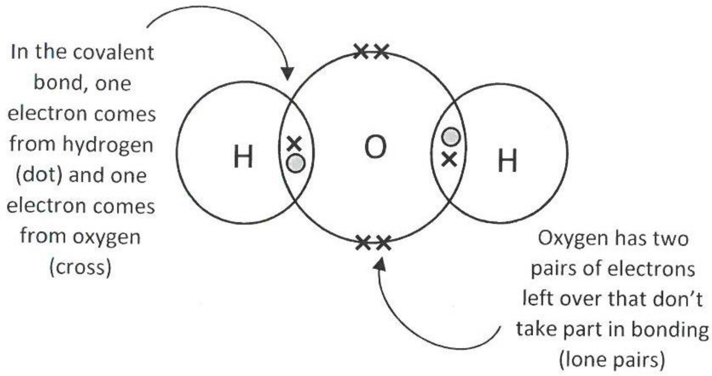
Highlight or underline the key information on the revision sheet to consolidate your understanding.

Covalent bonding

	Definitions
Covalent bond	a shared pair of electrons that is electrostatically attracted to both nuclei
Dative bond / coordinate bond	a covalent bond in which both electrons in the bonding pair were donated by one atom
Lone pair	pair of electrons not involved in covalent bonding

In contrast to ionic bonding, covalent bonding involves **sharing** of electrons rather than complete exchange from one atom to the other, and is most common between non-metals. The electrostatic attraction is between the positively charged nuclei and the negatively charged bonding electrons shared between them.

As in ionic bonding, the sharing of electrons is done to give each atom involved a full outer shell of electrons. In the example on the following page, both the hydrogen atom and the oxygen atom provide one electron each to form a bonding pair. Oxygen has six outer electrons and needs eight (two more) to have a full outer shell, so it can form covalent bonds with two hydrogen atoms.



Exam tip
In this example, only the outer shell of each electron has been drawn for simplicity. In an exam you may be asked to draw only the outermost electrons in your diagram, so make sure you check!

The stronger a covalent bond is, the shorter the bond length is. Stronger bonds are a result of the two bonding atoms having a large difference in electronegativity.

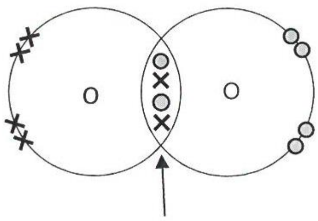
Bond	Bond strength (kJ mol^{-1})	Bond length (pm)
Cl–Cl	239	199
H–Cl	427	127
O–H	467	96

Bond strength **increases** / bond length **decreases**

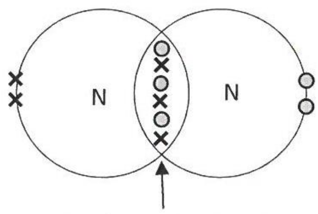
Multiple bonding

Sometimes, more than one pair of electrons must be shared in order to reach that noble gas configuration. One way to resolve this is to bond to more than one atom (as shown above); another is to share more than one pair of electrons and form a **multiple bond**.

Double bonds and triple bonds are stronger than single bonds and, therefore, are also shorter in length.



Double covalent bond
(two shared pairs of electrons)

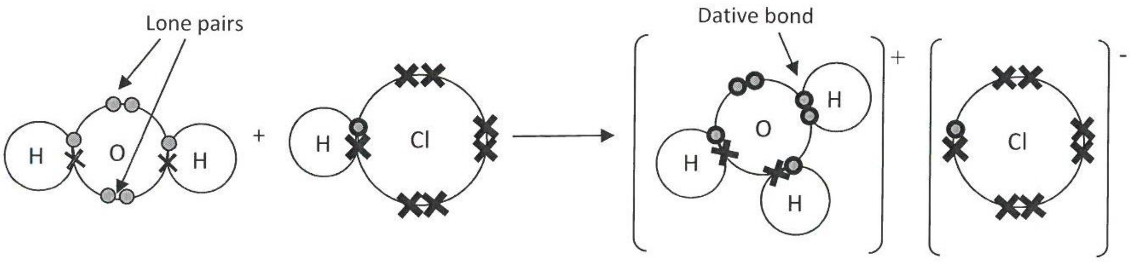


Triple covalent bond
(three shared pairs of electrons)

$O = O$ $\qquad$ $N \equiv N$

Dative bonding

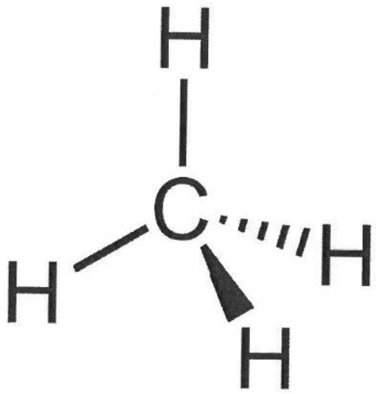
A **dative** covalent bond (or **coordinate bond**) occurs when **both** the electrons being shared have come from **the same** atom. This often happens with **lone pairs**, which are electron pairs that do not usually take part in bonding.



In the example above, hydrogen chloride is dissolved in water to form hydrochloric acid. The oxygen in water uses one of its lone pairs of electrons to bond to another hydrogen atom to form a positive **hydroxonium ion**. All three O–H bonds are identical in terms of their properties, and all three hydrogens are equally likely to be removed.

Shapes of organic molecules

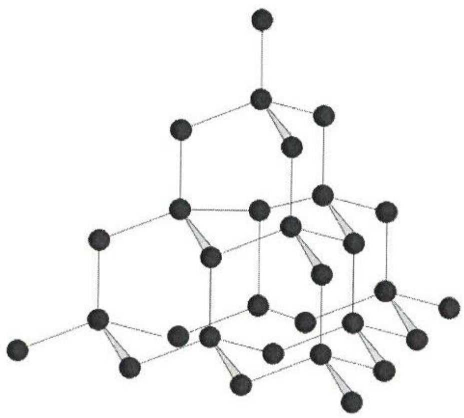
Each pair of electrons around an atom repels the others. This repulsion is what gives molecules their 3D shape, as bonding electrons arrange themselves to minimise repulsions. Carbon needs four electrons to achieve a full outer shell so forms four covalent bonds with other atoms. The four bonds surrounding a carbon atom all repel each other and sit as far away from each other as possible, forming a **tetrahedral** shape with an angle of 109.5° between each bond. This is the structure seen in diamond and in the organic molecules that are important for life.



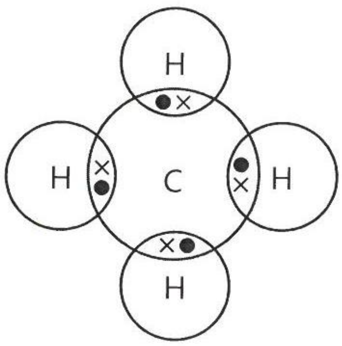
Exam tip

This is one example of how molecules are represented in 3D. A straight line represents a bond in the plane of the page, a black wedge represents a bond pointing out from the plane towards you, and a dashed wedge represents a bond pointing away from you.

Covalent structures can be giant, like the infinitely repeating pattern of carbon atoms in diamond, or small molecules, such as methane (CH_4). Despite having the same types of bonds, these two types of material have very different physical properties.



Giant covalent structure (diamond)



Simple covalent molecule (methane)

Credit: Zig Zag Resources Revision Guide Editions

VIDEO

To watch a video looking at this concept, please scan one of the following codes with your smartphone.

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SELF ASSESSMENT

To practice your understanding, answer the following questions.

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The answers are found after the questions.

A1. Draw a dot and cross diagram to show the bonding in an ammonium ion ($NH4^{-}$) ion.
What is special about one of the bonds?

A2. Explain why diatomic oxygen (O_2) forms a double covalent bond but diatomic fluorine does not.

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A3. Explain why chloroform ($CHCl_3$) is not a flat molecule.

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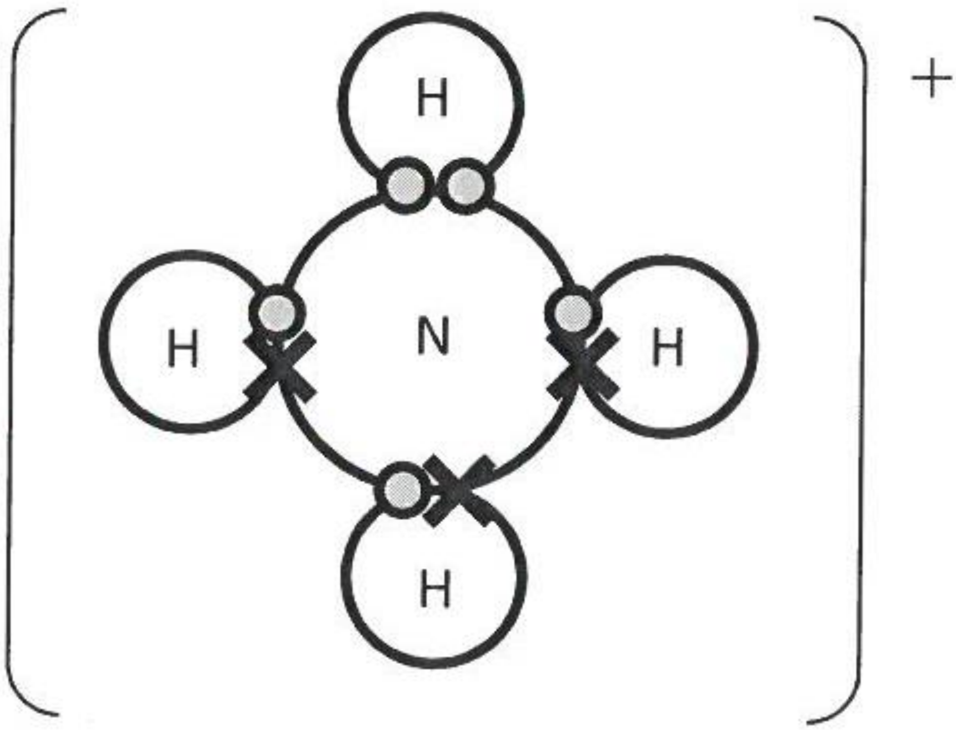
ANSWERS

A1.



One of the bonds is a dative/coordinate covalent bond – both electrons in the bonding pair come from the N atom.

A2.

An oxygen atom has six electrons in its outer shell with two unpaired electrons, so forms a double bond (two pairs of shared electrons) with another oxygen atom to obtain a full outer shell. An atom of fluorine has seven outer electrons with one unpaired electron, so only forms a single bond.

A3.

Each carbon must form four covalent bonds which arrange themselves in a 3D tetrahedral shape to minimise repulsion between bonding electron pairs.
The separation between each bond is 109.5^{o}.

ASSESSMENT QUESTION

Please answer this assessment question on this topic in Applied Science.

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This work will be formally assessed with feedback given.

1. Ethane and ethene are compounds which contain covalent bonds.
Ethane has the formula C_2H_6
1.1 Draw ethane using a dot and cross diagram

[3 Marks]

Ethen contains as many carbon atoms as ethane but has a double bond between the carbon atoms which is stronger.
1.2 Write the formula of ethane.

[1 Mark]

....

1.3 Describe the bonding of ethane in terms of
Carbon-carbon bond length compared to ethene.
Geometry around the carbon atoms.
The number of co-ordinate bonds

[3 Marks]

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Reference: Zig Zag Educational Resources

2.1 Explain how the atoms are held together by the covalent bond in a molecule of hydrogen.

[1 Mark]

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....

....

Draw the dot and cross diagrams for

1.2 methane, CH_4

[1 Mark]

1.3 ethene, $CH_2{=}CH_2$

[1 Mark]

1.4 nitrogen, N_2

[1 Mark]

1.5 the ammonium ion, NH_4^+

[1 Mark]

Silicon exists in a giant covalent lattice.
1.6 The electrical conductivity of pure silicon is very low.
Explain why this is so in terms of the bonding.

[2 Marks]

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1.7 Explain the high melting temperature of silicon in terms of the bonding.

[2 Marks]

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Reference: EdExcel A-Level Chemistry Examination Resources

EXTENSION

To further extend your understanding in the Applied Science course to prepare yourself for Year 12 course, read over information on the following topics:

Metallic Bonding

Intermolecular forces

London forces

Dipole-dipole interactions

Hydrogen bonds

DATASHEET

The Periodic Table of Elements

Key

relative atomic mass **atomic symbol** name atomic (proton) number

1 (1)	2 (2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	0 (8) (18)
							1.0 **H** hydrogen 1										4.0 **He** helium 2
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12											27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	95.9 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La*** lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111							

Elements with atomic numbers 112-116 have been reported but not fully authenticated

* Lanthanide series	140 **Ce** cerium 58	141 **Pr** praseodymium 59	144 **Nd** neodymium 60	[147] **Pm** promethium 61	150 **Sm** samarium 62	152 **Eu** europium 63	157 **Gd** gadolinium 64	159 **Tb** terbium 65	163 **Dy** dysprosium 66	165 **Ho** holmium 67	167 **Er** erbium 68	169 **Tm** thulium 69	173 **Yb** ytterbium 70	175 **Lu** lutetium 71
* Actinide series	232 **Th** thorium 90	[231] **Pa** protactinium 91	238 **U** uranium 92	[237] **Np** neptunium 93	[242] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[245] **Bk** berkelium 97	[251] **Cf** californium 98	[254] **Es** einsteinium 99	[253] **Fm** fermium 100	[256] **Md** mendelevium 101	[254] **No** nobelium 102	[257] **Lr** lawrencium 103

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Only constructive and reasoned feedback will be considered.