This is your booklet for the term.

Look after it.

If you need to take it away to do prep, it is your responsibility to bring it to the next lesson.

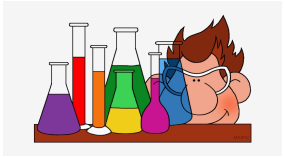
If you fail to bring your booklet to two consecutive lessons you'll be given another, your tutor contacted, an email sent home and a charge will be added to the bill.

If you manage to lose it and are unable to do your prep then it is your responsibility to download another copy of from **FROG**.

<u>The Departmental Website</u>.

You can log-on to the departmental website from anywhere by navigating to ellesmereonline.com and clicking

1. Pupils and Students
2. Science Link
3. Chemistry
4. AQA GCSE



9.04 – Atomic Structure

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

9.04 – Atomic Structure

Syllabus

- The relative electrical charges of the particles in atoms are:

Name of particle	Relative charge
Proton	+1
Neutron	0
Electron	–1

- Students should be able to use the nuclear model to describe atoms.
- Atoms are very small, having a radius of about 0.1 nm (1×10^{-10} m).

- The radius of a nucleus is less than 1/10 000 of that of the atom (about 1×10^{-14} m).
- Almost all of the mass of an atom is in the nucleus. The relative masses of protons, neutrons and electrons are:

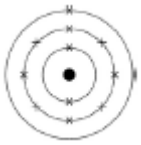
Name of particle	Relative mass
Proton	1
Neutron	1
Electron	Very small

- In an atom, the number of electrons is equal to the number of protons in the nucleus.
- Atoms have no overall electrical charge.
- The number of protons in an atom of an element is its atomic number.
- All atoms of a particular element have the same number of protons.
- Atoms of different elements have different numbers of protons.
- The sum of the protons and neutrons in an atom is its mass number.
- Atoms can be represented as shown in this example:

(Mass number) 23
(Atomic number) 11 Na

- Students should be able to calculate the numbers of protons, neutrons and electrons in an atom or ion, given its atomic number and mass number.
- Students should be able to relate size and scale of atoms to objects in the physical world.
- Use SI units and the prefix nano.

- Atoms of the same element can have different numbers of neutrons and are called ***isotopes***
- The electrons in an atom occupy the lowest available energy levels (innermost available shells).



- The electronic structure of an atom can be represented by numbers or by a diagram. For example, the electronic structure of Sodium is 2,8,1 or showing two electrons in the lowest energy level, eight in the second energy level and one in the third energy level
- Students should be able to represent the electronic structures of the first twenty elements

- The relative atomic mass of an element is an ***average value*** that takes account of the ***abundance*** of the isotopes of the element.
- Students should be able to calculate the relative atomic mass of an element given the percentage abundance of its isotopes.
- Before the discovery of the electron, atoms were thought to be tiny spheres that couldn't be divided.
- The discovery of the electron led to the plum pudding model of the atom. The plum pudding model suggested that the atom is a ball of positive charge with negative electrons embedded in it.
- The results from the alpha particle scattering experiment led to the conclusion that the mass of an atom was concentrated at the centre (nucleus) and that the nucleus was charged. This nuclear model replaced the plum pudding model.
- Niels Bohr adapted the nuclear model by suggesting that electrons orbit the nucleus at specific distances. The theoretical calculations of Bohr agreed with experimental observations.
- Later experiments led to the idea that the positive charge of any nucleus could be subdivided into a whole number of smaller particles, each particle having the same amount of positive charge. The name proton was given to these particles.
- The experimental work of James Chadwick provided the evidence to show the existence of neutrons within the nucleus. This was about 20 years after the nucleus became an accepted scientific idea.

Students should be able to describe:

why the new evidence from the scattering experiment led to a change in the atomic model
the difference between the plum pudding model of the atom and the nuclear model of the atom.

{Details of experimental work supporting the Bohr and Chadwick are not required.}

Atomic Structure.

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Atoms are small. Really small. This piece of paper only weighs about 5g.

It will contain over 300,000,000,000,000,000,000,000 atoms!

And yet there are ***sub-atomic particles*** – particles smaller than atoms.

Some are found in the centre (or ***nucleus***) of the atom.

Some are found around the outside in ***shells***.

Sub-atomic particle	Position	Charge	Mass
Proton	Nucleus	+1	1
Neutron	Nucleus	0	1
Electron	Shells	-1	1/1836 (0)

Q1. ***Nucleons*** are particles found in the nucleus. Which two sub-atomic particles are nucleons?

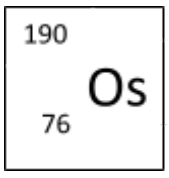
Q2. Why do you think:

a) Particles with no charge were named neutrons?

b) Particles with positive charge were named protons?

Q3 Which type of sub-atomic particles flow through wire in electrical circuits?

Atomic & Mass Numbers.



Every common element has its own one or two letter symbol.

It's important to write them correctly, with only one capital letter, or we could easily be confused.

Os = Osmium (element) **OS =Oxygen (O) and Sulphur (S) combined**

Each element also has two numbers. The smaller number is the ***ATOMIC*** (or proton) number.

It tells you where in the Periodic Table to find the element and also how many protons there are in the nucleus.

Atoms are neutral so there must be as many electrons & protons for the + and – charges to cancel.

The ***MASS*** (or nucleon) number tells you how many sub-atomic particles are in the nucleus (the heavy ones).

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Q1 Which element has 8 protons?

Q2 Which element has 32 protons?

Q3 Which element has 10 electrons?

Q4 How many electrons and protons does Carbon have?

Q5 How many protons and neutrons does Helium have?

Q6 How many protons and neutrons does Hydrogen have?

Q7 If an element has 13 protons and 14 neutrons: i) What are its Atomic and Mass numbers?

ii) Which element is it?

iii) How many electrons must it have?

Q8. If an element has a Mass number of 7 and has 3 electrons:

i) How many protons must it have?

ii) Which element is it?

iii) How many neutrons does it have?

Q9. Find Sodium on the Periodic Table. How many protons, neutrons & electrons does it have?

Q10 Find Xenon on the Periodic Table. How many protons, neutrons & electrons does it have?

Isotopes.

If the number of protons in an atom changed it would have a different Atomic number and would be a different element – but that would be Physics rather than Chemistry!

And to have a different number of electrons it would need a different number of protons.

But the number of neutrons can be different without changing either.

1 H 1	2 H 1	3 H 1

There are three slightly different versions of Hydrogen. These are called **Isotopes**.

All Hydrogen's isotopes react in exactly the same way.

Q1. What is the difference between the three isotopes of Hydrogen?

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Q2. Oxygen is element 8, it has stable isotopes with 8, 9 and 10 neutrons.

Complete the element cards for the three isotopes.

O 8	O 8	O 8

Q3. Uranium has major isotopes. One is very stable.

One can be made into Nuclear weapons. All Uranium atoms have 92 protons.

One isotope has 143 neutrons, the other has 3 more. Complete the element cards.

U 92	U 92

Electron Structure.

Electrons are found in the ***shells*** around the outside of the atom.

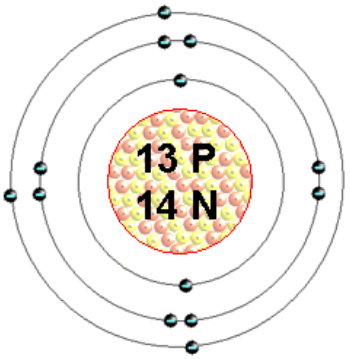
But not all the shells are the same size and so they can't all accept the same number of electrons.

- The ***inner shell*** is small and can take only ***2***.
- The ***second*** is larger and can take ***8***.
- The ***third*** can also take ***8***

- The shells are filled from the centre.
- One shell must be full before any electrons can be placed in another shell

So, since Aluminium has an Atomic Number of 13, it must have 13 electrons.

They will be arranged as ***(2, 8, 3)*** – two in the inner shell, eight in the second, the rest in the third.



Aluminium: $1s^2 2s^2 2p^6 3s^2 3p^1$

Q1 What is the electronic structure of Carbon?

Q2 What is the electronic structure of Oxygen?

Q3 What is the electronic structure of Sodium?

This lesson involves …

*Practice
*Perseverance
*Resilience

Q4 Which element would have the electronic structure (2,1)?

Q5 Which element would have the electronic structure (2,7)?

Q6 Which element would have the electronic structure (2,8.7)?

Q7 An element has 10 protons. What will its electronic structure be?

Q8 An element has 18 protons. What will its electronic structure be?

Q9 An element has 14 protons. What will its electronic structure be?

This lesson involves …

*Practice
*Perseverance
*Resilience

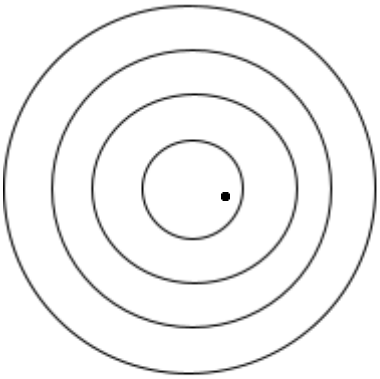
Filling Shell Diagrams

When filling shell diagrams, we must fill the shells from the centre of the atom.

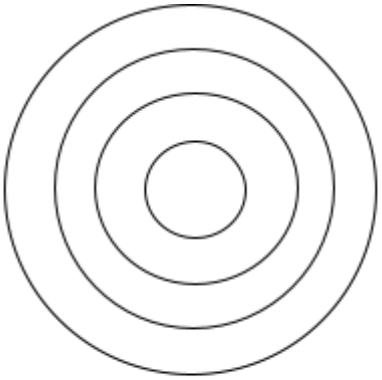
In the second and third shells it is a good idea to place electrons singly until there are 4 single electrons and then to pair them until the shell is full with 8.

❖ ***Task***

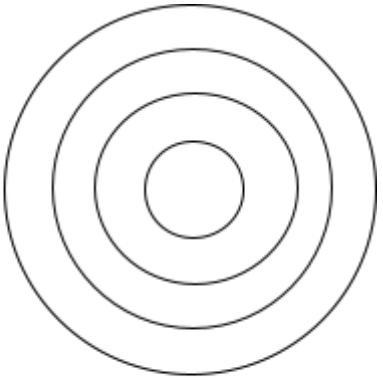
Complete the shell diagrams below for the first eighteen elements.



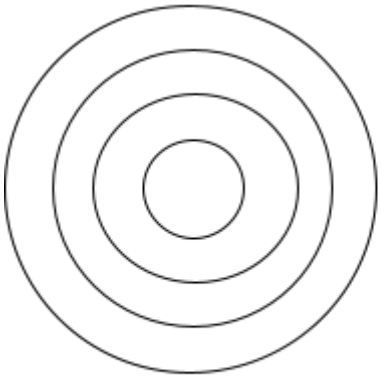
Hydrogen – $_{1}H$



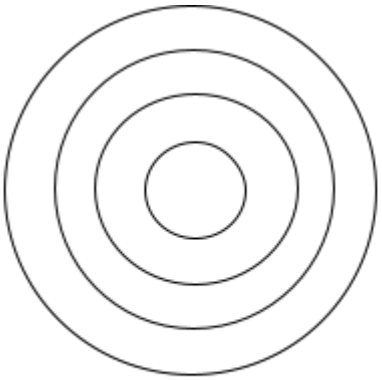
Boron – $_{5}B$



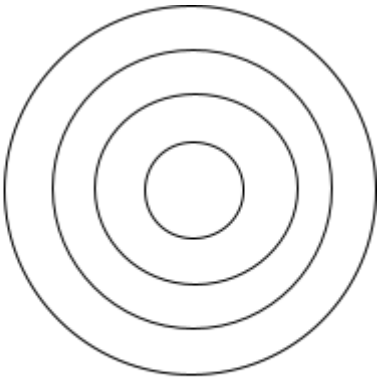
Oxygen – $_{8}O$



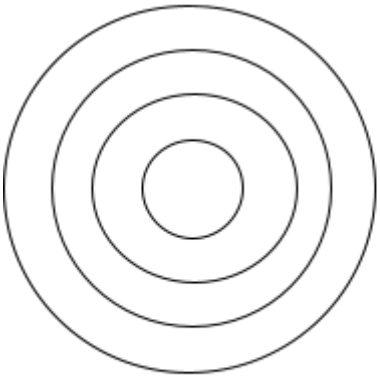
Fluorine– $_{9}F$



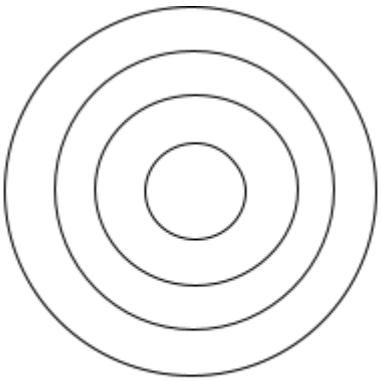
Helium – $_{2}He$



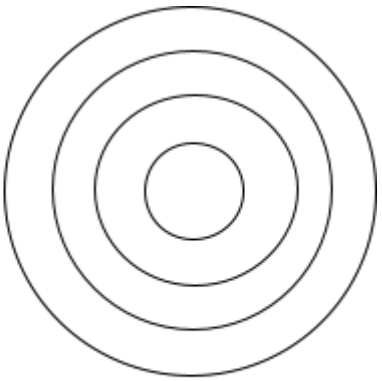
Lithium – $_{3}Li$



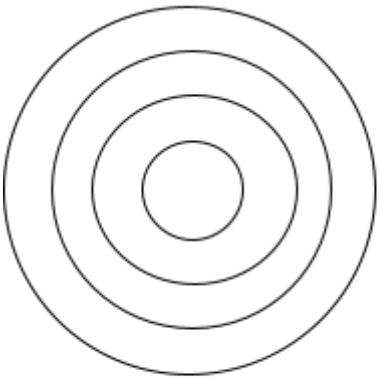
Nitrogen – $_{7}N$

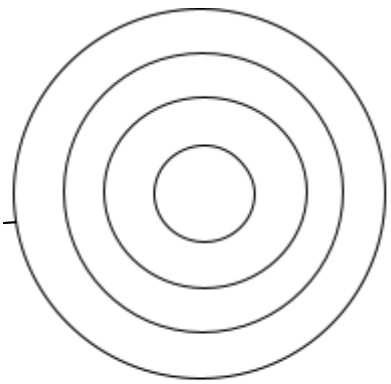


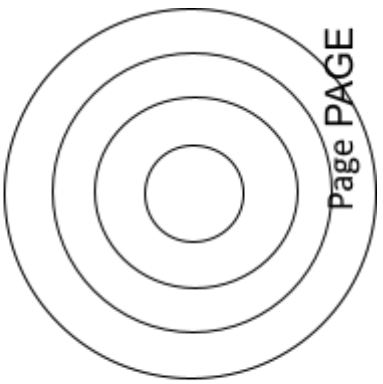
Carbon – $_{6}C$

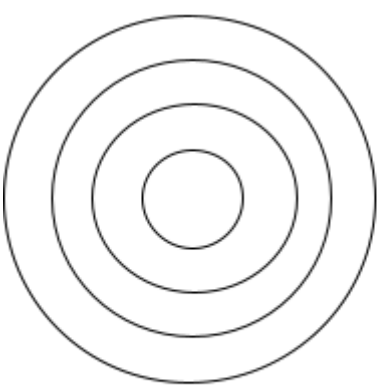


Beryllium– $_{4}Be$

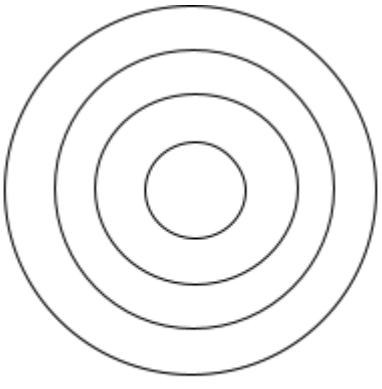




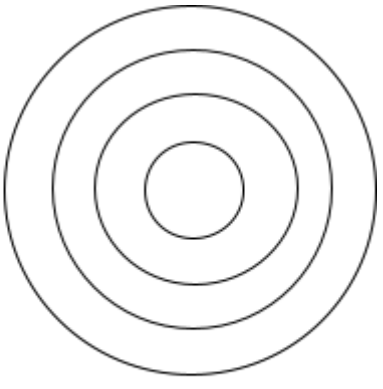




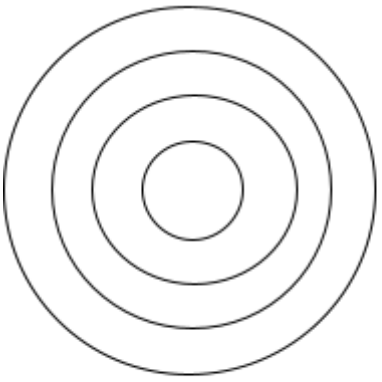
Aluminium – $_{13}$Al



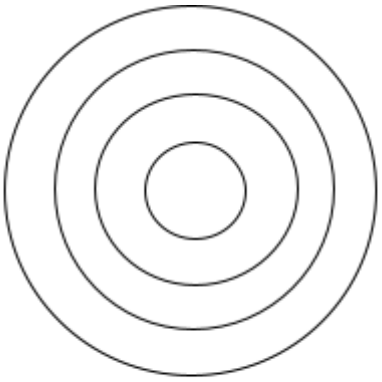
Silicon– $_{14}$Si



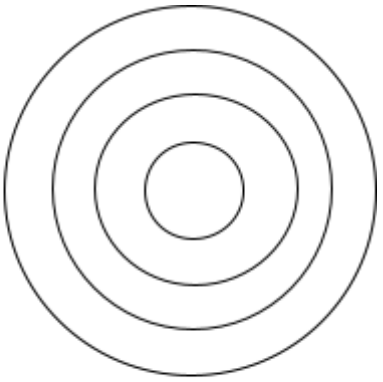
Magnesium – $_{12}$Mg



Sodium – $_{11}$Na

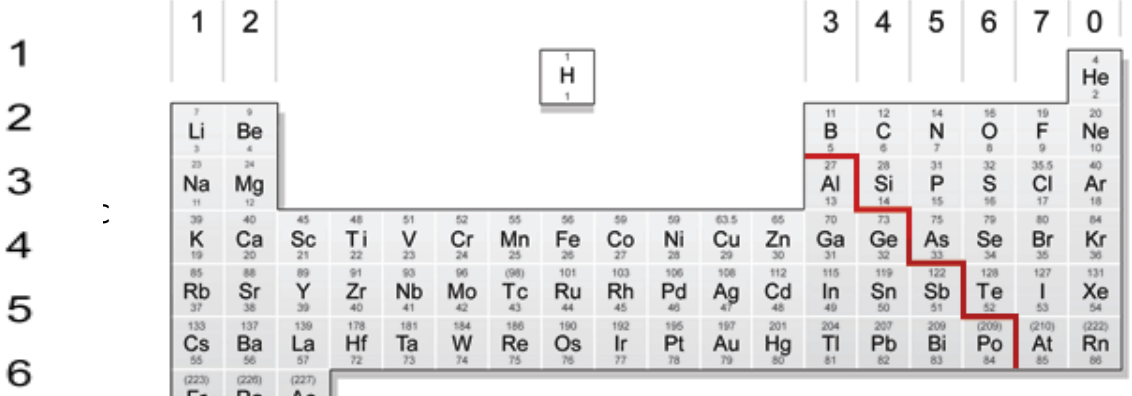


Chlorine – $_{17}$Cl



Neon – $_{10}$Ne

Electronic Structure & the Periodic Table



llabus

Page PAGE *

MERGEFORMAT 102

The periodic table is divided into ***vertical*** columns called ***Groups***.

And ***horizontal*** rows called ***Periods***.

If we look at the electronic structure of the Group 1 elements we see:

$_{3}Li$ => (2,1) $_{11}Na$ => (2,8,1) $_{19}K$=> (2,8,8,1)

- The number of electrons in the outside shells is the same as the Group number.
- The number of shells is the same as the Period number

So, is an atom has the electronic structure (2,8,7) – it must be in group 7, and Period 3 – it is Chlorine

❖ **Task** Complete the table below

This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Electronic Structure	Group	Period	Element
2.8.7	7	3	Chlorine
2.1			
2.8.8.2			
2.8.3			
	6	2	
	5	3	
			Beryllium

Questions

Potassium has electronic configuration 2.8.8.1

Q1. Explain how you know that it is in Group 1

Q2.Explain how you know it is in Period 4

Q3. What would the electronic structure of an element in Group 3 and Period 2 be?

Q4. What would the electronic structure of an element in Group 5 and Period 3 be?

Q5. What would the electronic structure of an element in Group 7 and Period 2 be?

Q6. What would the electronic structure of an element in Group 2 and Period 4 be?

Atomic Mass

An atom of ^{35}Cl has a mass of 35 amu (atomic mass units)

An atom of ^{37}Cl has a mass of 37 amu

On the Periodic Table an atom of Cl has a mass of 35.5 amu

But it isn't possible to have ½ a neutron so this must be an average mass – but not a simple average because the simple mean of 35 and 37 would be 36 amu.

How do we get to 35.5?

When we study Chlorine we find that 25% of the atoms are ^{37}Cl and the other 75% are ^{35}Cl.

So, in a sample of 100 Chlorine atoms the total mass would be:

(25x37) + (75x35) = 3550

And the average mass would be 3550/100 = 35.5

Q1. Find the average mass of Lithium which is found to be 7.6% ^{6}Li and 92.4% ^{7}Li

Q2. Find the average mass of Potassium which is found to be 93.25% ^{39}K and 6.75% ^{41}K

Q3. Find the average mass of Copper which is found to be 69.2% ^{63}Cu and 30.8% ^{65}Cu

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Q4. Find the average mass of Magnesium found to be 79% ^{24}Mg, 10% ^{25}Mg and 11% ^{26}Mg

History of the Atom

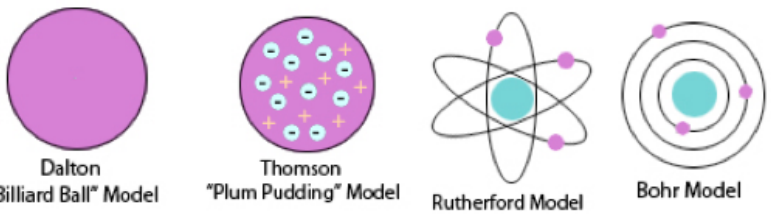
Science works by developing "models" (ideas about how things work) and then testing to see if the model matches the way things behave in real life.

The idea of atoms is very old but no one really believed it until two hundred years ago.

Eventually, most Scientists came round to the idea that everything was made out of solid particles that couldn't be broken down into anything simpler.

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

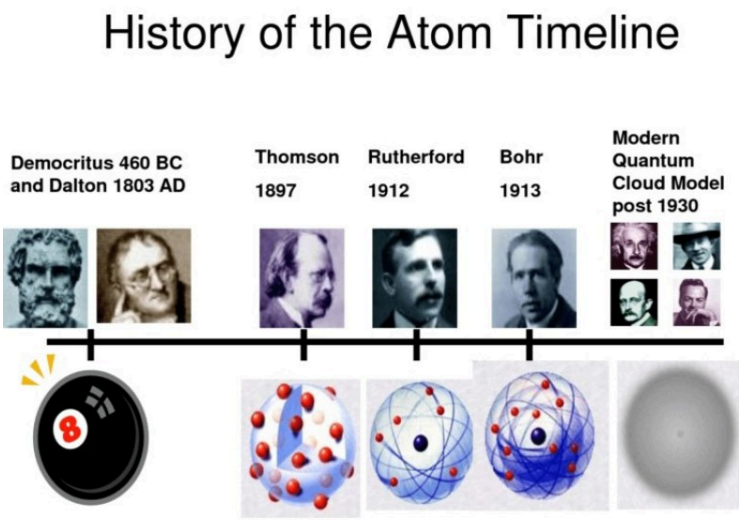


However, eventually it was found that negative particles much smaller than atoms existed. A new model was needed. And then another. And another…..

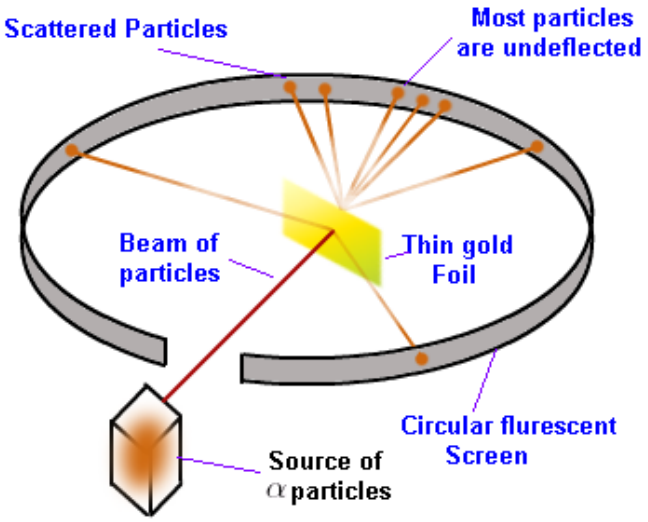
1. Thomson decided that if electrons were negative this meant that the rest of the atom was positive – why did he think this?

2. Thomson believed that the electrons would be scattered throughout the atom, as far apart as possible. Why?

3. Describe what Ernest Rutherford did that proved Thomson had to be wrong?



Rutherford Scattering



This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

1. Which particles was Rutherford scattering?

2. If Thomson's model had been correct, what would have happened to these particles?

3. How did the experiment show that most of an atom is empty space?

4. How did the experiment suggest that the nucleus was likely to be positive?

5. What did Nils Bohr do to alter the model?

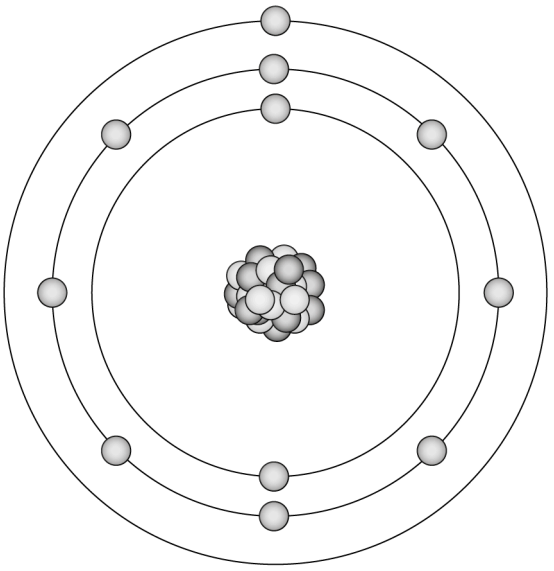
6. What did James Chadwick discover that needed adding to the model?

Question — Model of an atom

The model of an atom has developed since the time when atoms were thought to be tiny spheres that could not be divided to the simple model we use today.

A diagram of a model of a sodium atom that we use today is shown in **Figure 2**.

Figure 2



One of the early models of an atom is the plum pudding model.

Compare the plum pudding model of the atom with the simple model of the atom we use today.

[4 marks]

Mark scheme

Answers	Mark
Level 2: Scientifically relevant features are identified; the way(s) in which they are similar/different is made clear and (where appropriate) the magnitude of the similarity/difference is noted.	3–4
Level 1: Relevant features are identified and differences noted.	1–2
Indicative content Similarities ● Both models refer to a positive charge in the atom. ● Both models contain electrons. ● Both models have electrons each with a negative charge. Differences ● The plum pudding model states that atoms are a ball of positive charge. ● Today's model refers to the atom having a nucleus ● Which consists of protons each with a positive charge ● And neutrons that are neutral. ● The plum pudding model states that electrons are embedded in the ball of positive charge. ● Today's model states that electrons are orbiting the nucleus ● Some distance away. Credit may be given for diagrams showing the models of the atom.	

Full-marks answer

Both the plum pudding model and today's model of an atom refer to an area of positive charge and negatively charged electrons in an atom. However, the plum pudding model assumes that atoms are a ball of positive charge whereas the modern model has a central nucleus that contains protons and neutrons. The protons have a positive charge and the neutrons are neutral. The plum pudding model assumes that the electrons are inside the ball of positive charge, like raisins in a pudding, whereas the modern model assumes that the electrons are moving around the nucleus in orbits.

Partial answer and commentary

The plum pudding model shows atoms as a pudding with electrons in it, like raisins. The model of an atom that we use today has a nucleus with protons and neutrons and electrons moving around

The student has included some of the relevant points about the two models of atoms but has omitted important points, such as the charges on the subatomic particles. There is no attempt to identify the similarities and differences in the two models and both of these would be needed to score a higher mark, so this is a Level 1 answer. To move to Level 2, the student would need to identify at least one similarity and one difference between the models.
1 out of 4 marks awarded.

PES/MPC

This lesson involves ...

*Practice
*Perseverance

1 This question is about the element beryllium.

(a) Use words from the box to complete the sentences about beryllium.

Each word may be used once, more than once or not at all.

(7)

1 Boron is an element in Group 3 of the Periodic Table.

An atom of boron can be represented as $^{11}_{5}B$

9 This question is about bromine and some of its compounds.

(a) Atoms of bromine can be represented as ^{79}Br and ^{81}Br

(i) State the number of protons, neutrons and electrons in an atom of ^{79}Br

(2)

Protons .....

Neutrons .....

Electrons .....

1 The table shows the numbers of particles in two atoms, L and M.

	Atom L	Atom M
number of electrons	6	6
number of neutrons	8	6
number of protons	6	6

(ii) What n

.....

(iii) Why dc

(a) Which particles are present in the nuclei of both atoms?

(1)

.....

- ☐ **A** electrons and neutrons
- ☐ **B** electrons and protons
- ☐ **C** neutrons and protons
- ☐ **D** neutrons, protons and electrons

.....

(iv) The rel
more a

Sugges
numbe

(b) (i) The atomic number of atom L is .....

(d) Use information from the table to explain why atoms L and M are isotopes of the same element.

(2)

.....

.....

.....

.....

(d) An atom has atomic number 8 and mass number 18.

How many protons, neutrons and electrons does this atom contain?

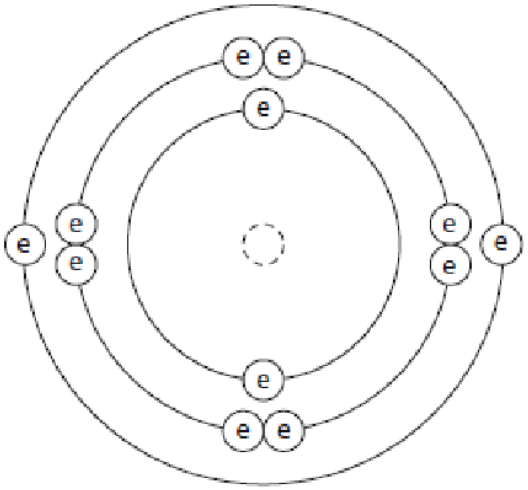
(2)

protons .....

neutrons .....

electrons .....

3 The diagram shows the electronic configuration of an atom of element X.



(a) (i) How many protons does the nucleus of the atom contain?

(1)

.....

(ii) Which group of the Periodic Table contains element X?

Give a reason for your choice.

(2)

.....

.....

.....

.....

(iii) Give the formula of the ion formed by element X in its compounds.

(1)

.....

(b) Element X has three isotopes.

The table gives the mass number of each isotope and its percentage abundance in a sample of element X.

Mass number	Percentage abundance (%)
24	79.0
25	10.0
26	11.0

This lesson involves ...

*Practice
*Perseverance
*Resilience

Calculate the relative atomic mass (A_r) of element X.

Give your answer to one decimal place.

(3)

relative atomic mass of X = ...

1 Neon is an element with atomic number 10.

(a) Which sub-atomic particles are present in the nucleus of a neon atom?

(1)

- ☒ **A** electrons and neutrons
- ☒ **B** electrons and protons
- ☒ **C** electrons and neutrons and protons
- ☒ **D** neutrons and protons

(b) Use words from the box to complete the sentences about the particles in a neon atom.

Each word may be used once, more than once or not at all.

(3)

electrons	neutrons	nuclei	protons

The particles with the smallest mass are .....

An atom of neon has no overall charge because it contains equal numbers

of ..... and .....

The chemical properties of neon depend on the number of

..... in the outer shell.

(c) What is the electronic configuration of a neon atom?

(1)

- ☒ **A** 2.8
- ☒ **B** 2.2.6
- ☒ **C** 2.8.8
- ☒ **D** 2.8.8.2

(d) Neon has two main isotopes that can be represented as ^{20}Ne and ^{22}Ne.

(i) Explain, with reference to sub-atomic particles, what is meant by the term **isotopes.**

(2)

(ii) The relative atomic mass of neon is 20.2

How does this information support the fact that a sample of neon contains more ^{20}Ne than ^{22}Ne?

(1)

This lesson involves …

*Practice
*Perseverance
*Resilience

(e) Neon belongs to the family of noble gases and is inert.

(i) What is meant by the term **inert**?

(1)

(ii) Why are noble gases inert?

(1)

(Total for Question 1 = 10 marks)

1 The table shows the numbers of protons, neutrons and electrons in some atoms and ions.

Atom or ion	Protons	Neutrons	Electrons
P	6	8	6
Q	5	6	5
R	9	10	10
S	3	4	2
T	6	6	6

This lesson involves …

*Automaticity
*Speed & Accuracy

(a) (i) Which particles have the same mass?

(1)

☒ **A** electrons and protons

☒ **B** electrons and neutrons

☒ **C** neutrons and protons

☒ **D** electrons, neutrons and protons

(ii) What is the atomic number of P?

(1)

☒ **A** 6

☒ **B** 8

☒ **C** 12

☒ **D** 14

(iii) What is the mass number of Q?

(1)

☒ **A** 5

☒ **B** 6

☒ **C** 10

☒ **D** 11

(b) Which group of the Periodic Table contains element T?

(1)

.....

(c) (i) Which two letters represent isotopes of the same element?

(1)

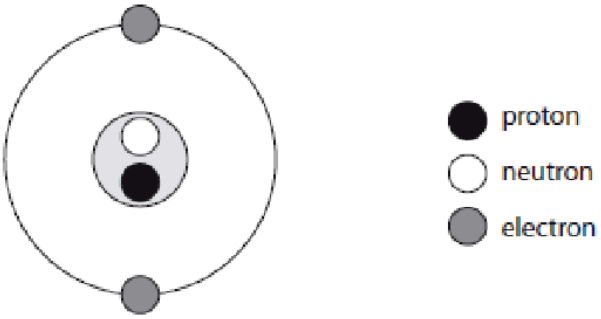
..... and .....

(ii) Which letter represents a positive ion?

(1)

.....

(d) The diagram shows the arrangement of particles in another ion.



This lesson involves ...

*Practice
*Perseverance
*Resilience

How does the diagram show that this ion has a negative charge?

(1)

.....

.....

(Total for Question 1 = 7 marks)

1 An atom of an element has an atomic number of 6 and a mass number of 12.

(a) Using this information, complete the table to show the numbers of protons, neutrons and electrons in one atom of this element.

(2)

number of protons	
number of neutrons	
number of electrons	

This lesson involves ...

*Automaticity
*Speed & Accuracy

(b) The Periodic Table shows the positions of five elements, J, Q, T, X and Z.

The letters do **not** represent the symbols for the elements.

Period	1	2						Group					3	4	5	6	7	0
1																		
2	J																	Q
3	T																	
4													X		Z			
5																		
6																		

(i) How many electrons are there in the outer shell of an atom of X?

(1)

.....

(ii) There are 31 protons in an atom of X.

Using this information, explain how many protons there are in an atom of Z.

(2)

.....

.....

.....

.....

(iii) What is the electronic configuration of an atom of Q?

(1)

.....

(iv) State one similarity and one difference between the electronic configurations of atoms of J and T.

(2)

similarity .....

difference .....

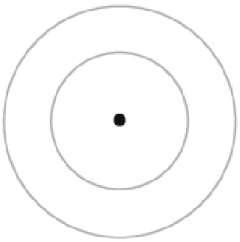
(Total for Question 1 = 8 marks)

This lesson involves …

*Practice
*Perseverance
*Resilience

(c) Complete the diagram to show the electronic configuration of an atom of fluorine, using x to represent an electron.

(1)



(Total for Question 1 = 8 marks)

1 Use the the Periodic Table on page 2 to answer this question.

(a) (i) **The symbol for silver is**

(1)

☐ **A** Ag ☐ **B** As ☐ **C** S ☐ **D** Si

Syllabus

2 Bromine is an element in Group 7 of the Periodic Table.

(a) What is the name given to the Group 7 elements?

(1)

☐ **A** alkali metals ☐ **B** alkaline earth metals ☐ **C** halogens ☐ **D** noble gases

(b) The symbols of two isotopes of bromine are $^{79}_{35}Br$ and $^{81}_{35}Br$.

(i) State what is meant by the term **isotopes**.

(2)

.....

.....

.....

.....

(ii) Complete the table to show the number of protons, neutrons and electrons in one atom of $^{79}_{35}Br$ and in one atom of $^{81}_{35}Br$.

(3)

Isotope	Number of protons	Number of neutrons	Number of electrons
$^{79}_{35}Br$			
$^{81}_{35}Br$			

This lesson involves ...

*Automaticity
*Speed & Accuracy

ys the percentage composition of a sample of magnesium.

Isotope	^{24}Mg	^{25}Mg	^{26}Mg
Percentage (%)	78.6	10.1	11.3

Calculate the relative atomic mass of magnesium.

Give your answer to one decimal place.

(3)

This lesson involves ...

*Automaticity
*Speed & Accuracy

This lesson involves ...

*Practice
*Perseverance
*Resilience

(d) The table shows the percentage composition by mass of a sample of silicon.

Isotope	^{28}Si	^{29}Si	^{30}Si
Percentage (%)	92.2	4.70	3.10

Calculate the relative atomic mass of this sample of silicon.

Give your answer to one decimal place.

(2)

This lesson involves ...

*Practice
*Perseverance
*Resilience

relative atomic mass = ...

9.05 – Periodic Table

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

9.05 – Periodic Table

Syllabus

- The elements in the Periodic Table are arranged in order of ***atomic (proton) number*** and so that elements with ***similar properties*** are in columns, known as ***groups***.
- The table is called a Periodic Table because similar properties occur at ***regular intervals***.
- Elements in the same ***group*** in the Periodic Table have the ***same number of electrons in their outer shell*** (outer electrons) and this gives them ***similar chemical properties***.

Students should be able to:

- explain how the ***position*** of an element in the Periodic Table is related to the ***arrangement of electrons in its atoms*** and hence to its ***atomic number***
- ***predict*** possible reactions and probable ***reactivity*** of elements from their positions in the Periodic Table.
- Elements that react to form positive ions are metals.
- Elements that do not form positive ions are non-metals.
- The majority of elements are metals.
- Metals are found to the left and towards the bottom of the Periodic Table. Non-metals are found towards the right and top of the Periodic Table.

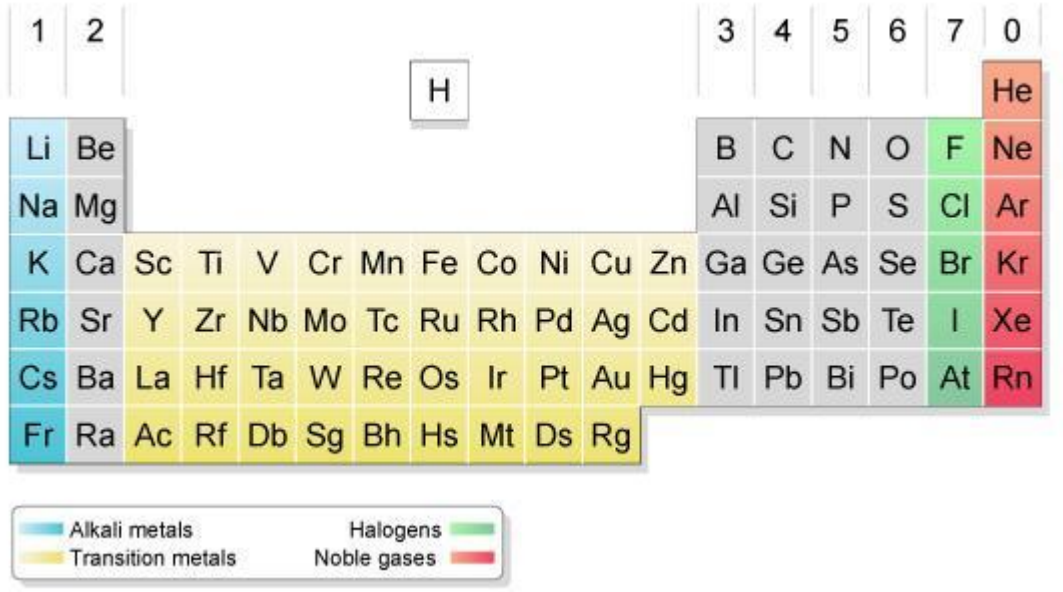
Students should be able to:

- explain the differences between metals and non-metals on the basis of their characteristic physical and chemical properties.
- explain how atomic structure of metals & non-metals relates to their position in the Periodic Table
- explain how the reactions of elements are related to the arrangement of electrons in their atoms and hence to their atomic number.

- Before the discovery of protons, neutrons and electrons, scientists attempted to classify the elements by arranging them in order of their atomic weights.
- The early Periodic Tables were incomplete and some elements were placed in inappropriate groups if the strict order of atomic weights was followed.
- Mendeleev overcame some of the problems by leaving gaps for elements that he thought had not been discovered and in some places changed the order based on atomic weights.
- Elements with properties predicted by Mendeleev were discovered and filled the gaps.
- Knowledge of isotopes made it possible to explain why the order based on atomic weights was not always correct.

Students should be able to describe these steps in the development of the Periodic Table.

The Periodic Table



The Periodic Table is divided in ***Groups*** – vertical columns of elements that (generally) share ***properties***. In other words, an element's properties should be very similar to the properties of the elements above it and below it on the table.

The entire middle block of the Periodic Table (Transition Metals) are very similar so we don't give them individual Group Numbers

The numbers are based on the electronic structures of its elements.

Q1. Write down the electronic structures of the first three Group 1 elements (Lithium, Sodium and Potassium).

Q2. Write down the electronic structures of the first three Group 2 elements (Beryllium, Magnesium and Calcium).

Q3. Write down the electronic structures of the first two Group 6 elements (Oxygen, and Sulphur).

Q4. Write down the electronic structures of the first two Group 7 elements (Fluorine, and Chlorine).

The Periodic Table

Q5. How does the Electronic Structure of an element tell us which Group to place it in?

Q6. All Group 2 elements have electronic structures that have two electrons in the outer shell. Why might all these elements behave in a similar way?

This lesson involves ...

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Q7. Hydrogen is element 1. Its electronic structure is (1). Group 1 elements are dangerously reactive metals.

i) Give one reason why Hydrogen is sometimes placed in Group 1.

ii) Give a reason why we generally do not put it in Group 1.

Q8. The Electronic Structure of Group 0 elements are He (2), Ne (2,8), Ar (2,8,8). This group used to be called Group 8 but is now called Group 0.

i) Suggest a reason why the name was changed.

ii) Suggest why the name Group 0 was chosen.

Q9. There are only two Period 1 elements H (1) and He (2).
Period 2 elements have electronic structures (2,1), (2,2), (2,3), (2,4), (2,5), (2,6), (2,7) & (2,8).
Period 3 elements have (2,8,1), (2,8,2), (2,8,3), (2,8,4), (2,8,5), (2,8,6), (2,8,7) & (2,8,8).

How does the electronic structure of an element determine which period it is placed in?

The Periodic Table

Because of the way we arrange elements in the Periodic Table elements that need to lose electrons to gain electronic stability end up on the left.

Elements that need to gain end up on the right.

So the table is divided between metals and non-metals. But the dividing line is a not clear-cut. Some elements have metallic and non-metallic physical properties.

Q1. Divide these physical properties into Metal or Non-Metal properties.

Ductile	
Brittle	
Low Melting-point	
Malleable	
Strong	
Insulating	
High Boiling Point	
Shiny	
Conductive	

Metals

Non-Metals

Q2. Connect the physical property to its definition

Ductile	Allows electricity or heat to pass through.
Malleable	Makes a ringing noise when struck.
Lustrous	Can be rolled or hammered into shape without breaking.
Sonorous	Can be stretched into a wire.
Brittle	Has a shiny surface (when freshly made).
Conductive	Is easily broken if dropped or hammered.

The Periodic Table

Semi-metals or ***Metalloids*** have some physical properties that suggest they are metals and some that suggest they are non-metals.

If this is the case we generally look at their ***chemical properties*** (how they react and what they react with) to decide if they are metal or non-metal.

* ***Metal Oxides*** neutralise acids (are ***Bases***) . Most dissolve to form ***Alkalis*** *

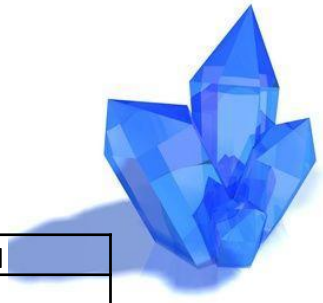
* ***Non-metal Oxides*** neutralise alkalis (are ***Acidic***) Most dissolve to form ***Acids*** *

Experiment 1.

Copper Oxide (CuO) is a metal oxide. It should form an alkali but ***cannot*** dissolve.

- Measure 30 cm^3 of ***Sulphuric Acid*** (H_2SO_4) and place in a 250cm^3 beaker.
- Test its ***pH*** with ***Universal Indicator paper***.
- Add a small spatula of ***Copper Oxide*** (CuO) and stir. When the powder dissolves measure the pH again.
- Continue adding Copper Oxide (CuO) and measuring pH until no more will dissolve.
- It may be necessary to gently warm the mixture. **DO NOT BOIL!!!!!!!.**
- Add two more spatulas and stir vigorously. Re-measure the pH

Results



Spatulas of CuO	pH
0	
1	
2	
3	
4	
5	
6	

This lesson involves ...

*Collaborative
*Concerned for Society
*Confident

You may extend this tab
necessary

Page PAGE * MERGEFORMA

Q1. What evidence is there that Copper Oxide (CuO) is a ***base***?

Q2. What evidence is there that Copper Oxide **<u>cannot</u>** form an ***alkali***?

Q3 When a Base neutralises an acid a salt is formed. Which salt? Where is it?

Q4. The symbol equation for this reaction doesn't need balancing. Why?

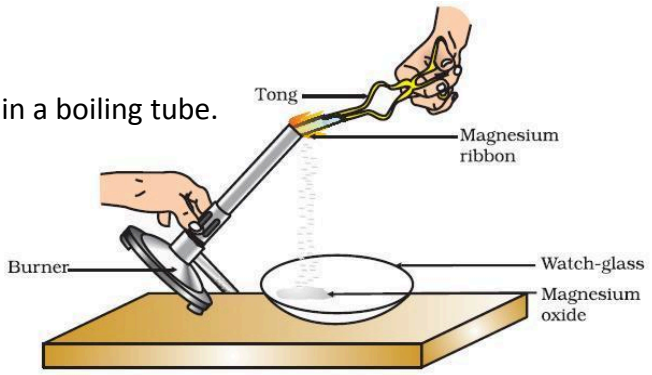
$$CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$$

Q5. Write the word equation for the reaction

Metal Oxides and Hydroxides

Experiment 2.

- Burn a short piece of Magnesium.
- **DO NOT LOOK DIRECTLY AT THE FLAME!**
- Collect the ash and drop into 10cm^3 of water in a boiling tube.
- Add a few drops of Universal Indicator.
- Place a bung in the top and shake vigorously.



Results:

The white ash formed is _____.

It dissolves in water to form an **acid/alkali.** **(Delete one alternative)**

Q1. Write a word equation for burning Magnesium.

Q2. Balance the symbol equation for this reaction:

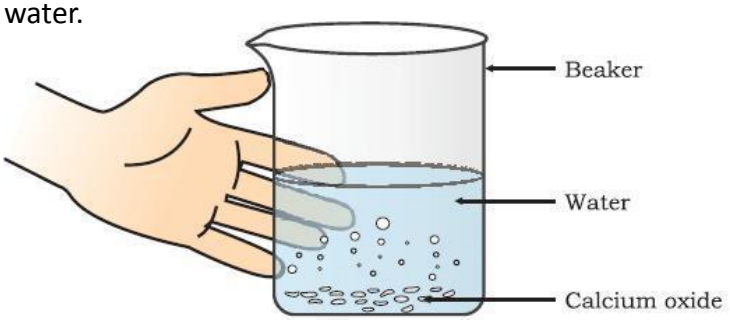
$$Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$$

This lesson involves …

*Collaborative
*Concerned for Society
*Confident

Experiment 3.

- Strongly heat some Calcium Carbonate (Limestone) for at least 5 minutes.
- Drop the hot Limestone chip in 25cm^3 of water.
- Add some Universal Indicator.



Results.

The Calcium Carbonate thermally decomposes (splits up when heated) to form Calcium Oxide and Carbon Dioxide.

Q1. Write a word equation for this reaction.

The Calcium Oxide reacts with water to form Calcium Hydroxide which dissolves.

Q2. What evidence is there that Calcium Hydroxide is an alkali?

This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Q3. The symbol equation is $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$

Write a word equation for the reaction.

Electrons and Reactivity

Group 0 elements are known as the ***Noble Gases*** because they (almost) never react with other elements.

The ***electronic structures*** of the first three Noble Gases are He (2), Ne (2,8), Ar (2,8,8).

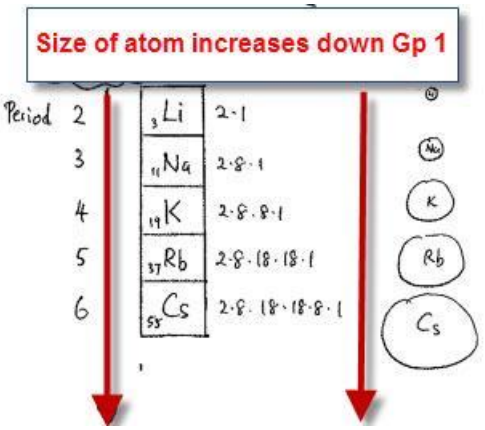
Q1. How do the electronic structures explain why Noble Gases don't react?

Metals generally lose electrons to become ***positive*** ions (***Cations***).

Non-metals gain electrons to become negative ions (***Anions***).

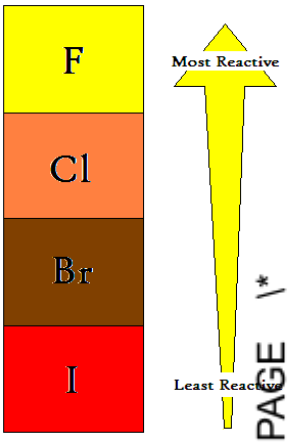
Group 1 and Group 2 elements become more reactive as we go down the group from small to large atoms.

Q2. Suggest why it might be easier for metals to lose electrons as their size increases.



Q3. Non-metals become more reactive as we go up the group.

Suggest why smaller atoms might find it easier to gain electrons.



This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Lavoisier tried to list the elements known in his day in some sort of logical order – but too few had been discovered and some elements weren't elements at all. Decades later Dobereiner discovered his "Triads"

1. How many elements in a triad?

2. Why did Dobereiner put these elements together in triads?

3. Newlands came up with the "law of Octaves". How many elements did he have in each octave? (Careful!)

4. Mendeleev put most of his elements in order of mass – why did he reverse the order of some?

5. Why did he leave some gaps in his table?

Extended Response Questions

Question — The periodic table

Mendeleev arranged the elements in the periodic table in order of increasing atomic weights. **Figure 3** shows part of the periodic table.

Figure 3.

1	2											3	4	5	6	7	0
							1 H hydrogen 1										4 He helium 2
7 Li lithium 3	9 Be beryllium 4		Key: relative atomic mass, **atomic symbol**, name, atomic (proton) number									11 B boron 5	12 C carbon 6	14 N nitrogen 7	16 O oxygen 8	19 F flourine 9	20 Ne neon 10
23 Na sodium 11	24 Mg magnesium 12											27 Al aluminium 13	28 Si silicon 14	31 P phosphorus 15	32 S sulfur 16	35.5 Cl chlorine 17	40 Ar argon 18
39 K potassium 19	40 Ca calcium 20	45 Sc scandium 21	48 Ti titanium 22	51 V vanadium 23	52 Cr chromium 24	55 Mn manganese 25	56 Fe iron 26	59 Co cobalt 27	59 Ni nickel 28	63.5 Cu copper 29	65 Zn zinc 30	70 Ga gallium 31	73 Ge germanium 32	75 As arsenic 33	79 Se selenium 34	80 Br bromine 35	84 Kr krypton 36

Explain why argon with relative atomic mass 40 is placed before potassium with relative atomic mass 39.

Your answer should refer to the atomic structures of argon and potassium.

[6 marks]

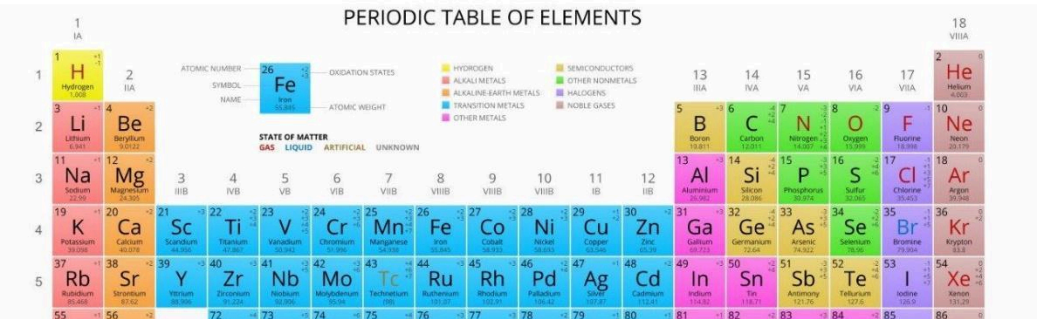
Mark scheme

Answers	Mark
Level 3: Relevant points (reasons/causes) are identified, given in detail and logically linked to form a clear account.	5–6
Level 2: Relevant points (reasons/causes) are identified, and there are attempts at logically linking. The resulting account is not fully clear.	3–4
Level 1: Points are identified and stated simply, but their relevance is not clear and there is no attempt at logical linking.	1–2
No relevant content	0
Indicative content ● The modern periodic table has elements arranged in order of increasing atomic number / number of protons. ● Protons had not been discovered when Mendeleev published his periodic table. ● Argon atoms have a higher relative atomic mass than potassium atoms as they have 18 protons and 22 neutrons. ● Potassium atoms have 19 protons but (only) 20 neutrons. ● Electronic structure provides additional evidence for the different positions in the periodic table. ● Argon has electronic structure 2.8.8. ● There are electrons in three shells so it is in Period 3. ● The outer shell is full of electrons so it is in Group 0. ● Potassium has electronic structure 2.8.8.1. ● There are electrons in four shells so it is in Period 4. ● The outer shell has one electron so it is in Group 1. Credit may be given for diagrams showing the electronic structures.	

Question — The periodic table

The Modern Periodic Table of elements is arranged according to the structure of the atoms in each element. It is also arranged by chemical properties. Chemical properties and atomic structure are linked.

PES/MPC



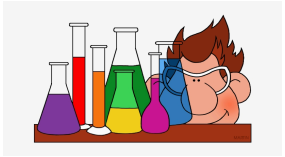
Explain how the table is arranged in groups and periods according to the atomic structure of atoms.

Explain how chemical properties are linked to the arrangement of elements in the table, and are also linked to atomic structure

[6 marks]

Mark scheme

Answers	Mark
Level 3: Most relevant points (reasons/causes) are identified, and logically linked to form a clear account.	5–6
Level 2: Some relevant points are identified, and there are attempts at logically linking.	3–4
Level 1: Some points are identified and stated simply, but there is little or no attempt at logical linking.	1–2
No relevant content	0
Indicative content ● The modern periodic table has elements arranged in order of increasing atomic number / number of protons. ● Metals are on the left and bottom of the table ● Non metals are on the right and at the top. ● Elements in the same group share similar chemical properties ● Elements in the same group have the same number of electrons in their outer shells ● So, they react to gain/lose or share electrons in the same way. ● Elements in the same period **do not** share similar chemical properties ● Elements in the same group have the same number of shells ● But they react to gain/lose or share electrons differently (because they have different numbers of outer-she electrons). Credit may be given for diagrams showing the electronic structures.	



3 The diagram shows a section of the Periodic Table and the symbols for the first 20 elements.

																	H	He
Li	Be										B	C	N	O	F	Ne	
Na	Mg										Al	Si	P	S	Cl	Ar	
K	Ca																

This lesson involves …

*Practice
*Perseverance
*Resilience

(a) (i) What name is given to a horizontal row of elements such as Na to Ar?

(1)

.....

(ii) Name two metals in the row Na to Ar.

(1)

..... and .....

(iii) Which is the least reactive element in the row Na to Ar?

Explain your answer.

(2)

least reactive element .....

explanation .....

.....

.....

(b) State, in terms of electronic configurations, why the elements in the column Li to K have similar chemical properties.

(1)

.....

.....

(c) (i) Which element has atomic number 6?

(1)

.....

(ii) Which element has atoms with an electronic configuration of 2.8.6?

(1)

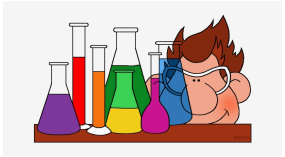
.....

PES/M

3 The diagram shows the elements in Period 3 of the Periodic Table.

Na	Mg	Al	Si	P	S	Cl	Ar

(a) (i) Identify an element in Period 3 that forms a basic oxide.



2 The diagram shows the positions of some elements in part of the Periodic Table.

							He
Li		B		N		F	
	Mg		Si		S		Ar

(a) How many periods and groups are shown in this diagram?

(1)

	Periods	Groups
☐ **A**	2	4
☐ **B**	3	4
☐ **C**	2	8
☐ **D**	3	8

(b) How many elements shown in the diagram are noble gases?

(1)

☐ **A** 1

☐ **B** 2

☐ **C** 3

☐ **D** 4

This lesson involves …

*Practice
*Perseverance
*Resilience

(c) What is the formula of the compound formed between magnesium and fluorine?

(1)

☐ **A** MgF

☐ **B** Mg_2F

☐ **C** MgF_2

☐ **D** Mg_2F_2

4 The Periodic Table consists of elements arranged in order of increasing atomic number.

(a) (i) What is meant by the term **atomic number**?

(1)

☒ **A** the number of neutrons in an atom

☐ **B** the number of protons in an atom

☐ **C** the number of protons plus the number of electrons in an atom

☐ **D** the number of protons plus the number of neutrons in an atom

(b) The elements in the periods of the Periodic Table show a change in properties across the period from left to right.

The table gives information about some elements in Period 2.

	Li	Be	B	C	N	F
melting point	low	high	high			
structure		giant	giant			molecular
acid-base character of the oxide		basic	acidic			

Complete the table by writing one of these words in each blank space.

- high or low for melting point
- giant or molecular for structure
- acidic or basic for acid-base character of the oxide

(3)

1 Use the Periodic Table on page 2 to help you answer this question.

Give the name or symbol of

(a) the element in group 3 and period 4.

(1)

.....

(b) an element in period 3 that is a good conductor of electricity.

(1)

.....

(c) the element in group 7 that is the most reactive.

(1)

.....

PES

(d) the element in group 5 that is present in a molecule of ammonia.

(1)

.....

(e) an element with an atom containing 8 electrons in its outer shell.

9.07 Ionic Bonding

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

9.07 Ionic Bonding

Syllabus

- When a metal atom reacts with a non-metal atom electrons in the outer shell of the metal atom are transferred.
- Metal atoms lose electrons to become positively charged ions.
- Non-metal atoms gain electrons to become negatively charged ions.
- The charge on the ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 relates to the group number of the element in the Periodic Table.
- The ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 have the electronic structure of a noble gas (Group 0).
- Students should be able to

1. work out the charge on the ions of metals and non-metals from the group number of the element, limited to the metals in Groups 1 and 2, and non-metals in Groups 6 and 7.

- An ionic compound is a giant structure of ions. Ionic compounds are held together by strong ***electrostatic*** forces of attraction between oppositely charged ions. These forces act in all directions in the lattice and this is called ***ionic bonding***.
- The electron transfer during the formation of an ionic compound can be represented by a dot and cross diagram, eg for Sodium Chloride.

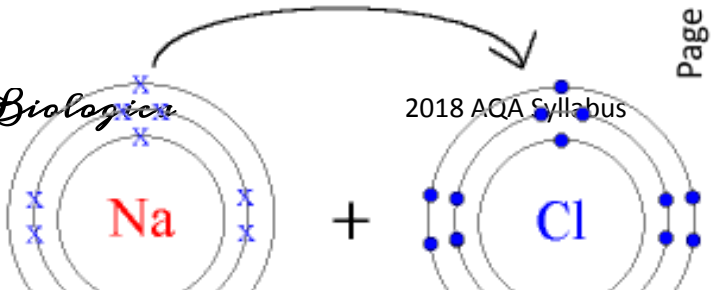
- Students should be :

1. draw ***dot and cross diagrams*** for ionic compounds formed by metals in Groups 1 and 2 with non-metals in Groups 6 and 7.
2. able to deduce a compound is ionic from a diagram of its structure in one of the specified forms
3. able to describe the ***limitations*** of using dot and cross, ball and stick, two and three-dimensional diagrams to represent a ***giant ionic structure***
4. able to work out the ***empirical formula*** of an ionic compound from a given model or diagram that shows the ions in the structure.
5. familiar with the structure of Sodium Chloride but do not need to know the structures of other ionic compounds.

- Ionic compounds have regular structures (giant ionic lattices) in which there are strong electrostatic forces of attraction in all directions between oppositely charged ions.
- These compounds have high melting points and high boiling points because of the large amounts of energy needed to break the many strong bonds.
- When melted or dissolved in water, ionic compounds conduct electricity because the ions are free to move and so charge can flow.
- Knowledge of the structures of specific ionic compounds other than Sodium Chloride is ***not required***.

Ionic Bonding.

Covalent bonding usually happens between two **non-metals.**



But if a **metal** combines with a **non-metal** they combine differently. This is **Ionic** Bonding.

Metal atoms have almost empty outer shells. They lose their outer electrons, becoming positive (**Cations**.)

Non-metal atoms have almost full outer shells. They gain these electrons, becoming negative **Anions**.

Positive cations and negative anions **attract** each other – this is the **Ionic bond**.

In the example on the right, the Sodium atom has an electronic structure of (2.8.1). It **loses** 1 electron to become a **1+** Sodium ion.

The Chlori***n***e atom has electronic structure (2.8.7). It **gains** the electron to become a -1 Chlori***d***e ion.

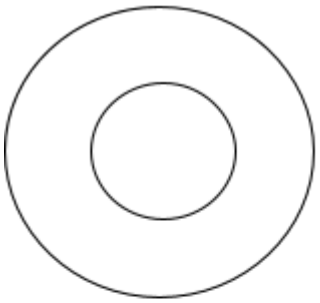
❖ Task

Draw the electronic structures of each pair of atoms. Draw an arrow to show how the electrons move and what the charge on the ions formed would be.

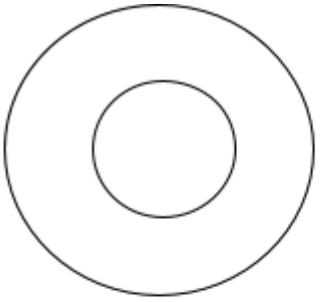
Q1

This lesson involves …

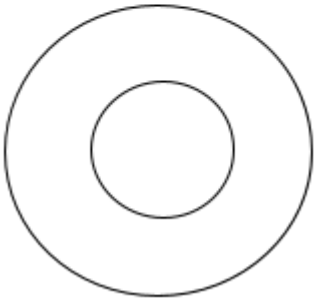
*Logical Thinking
*Precision
*Multi-step problem-solving



Lithium atom

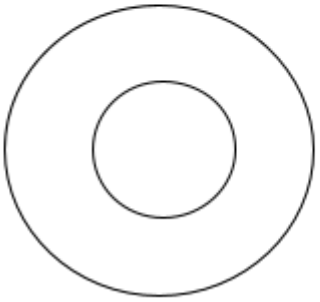


Fluorine atom



Charge?

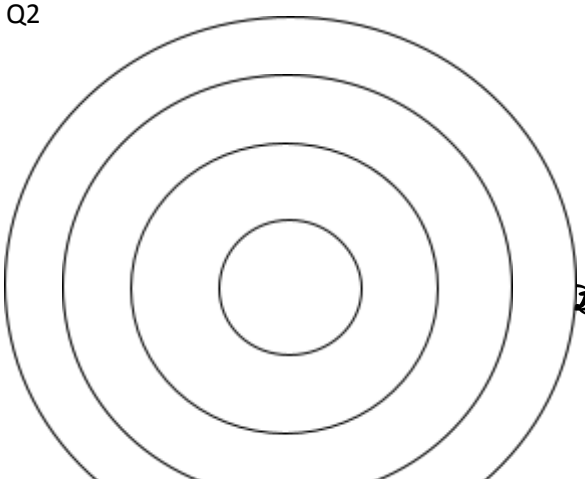
Lithium ion

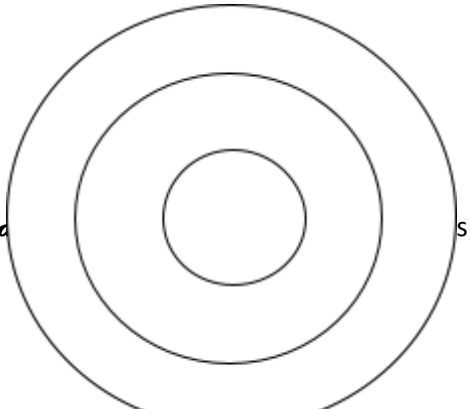


Charge?

_____? ion

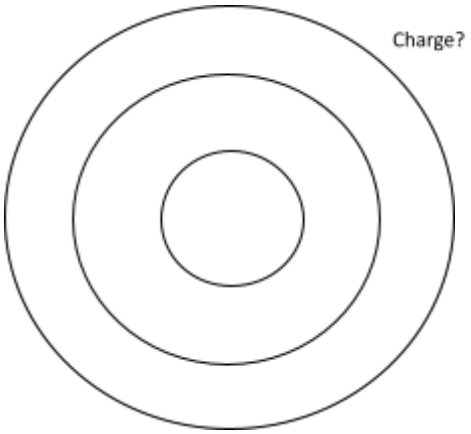
Q2

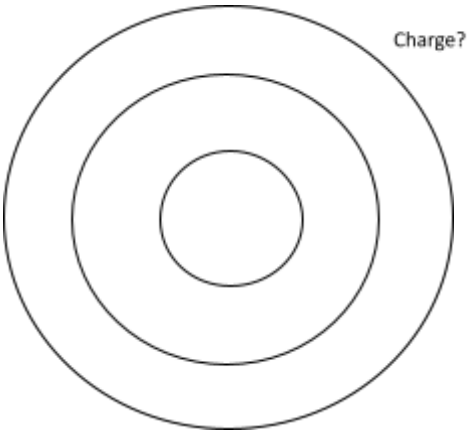




Calcium atom

Oxygen atom





Calcium ion

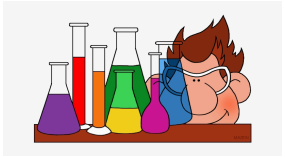
_____? ion

Mismatched Elements

Sometimes one metal atom hasn't enough electrons to fill the outer shell of the non-metal it is bonding to.

Sometimes one non-metal atom hasn't enough room in its outer shell to accept all the electrons the metal atom needs to lose.

How do we cope with that?



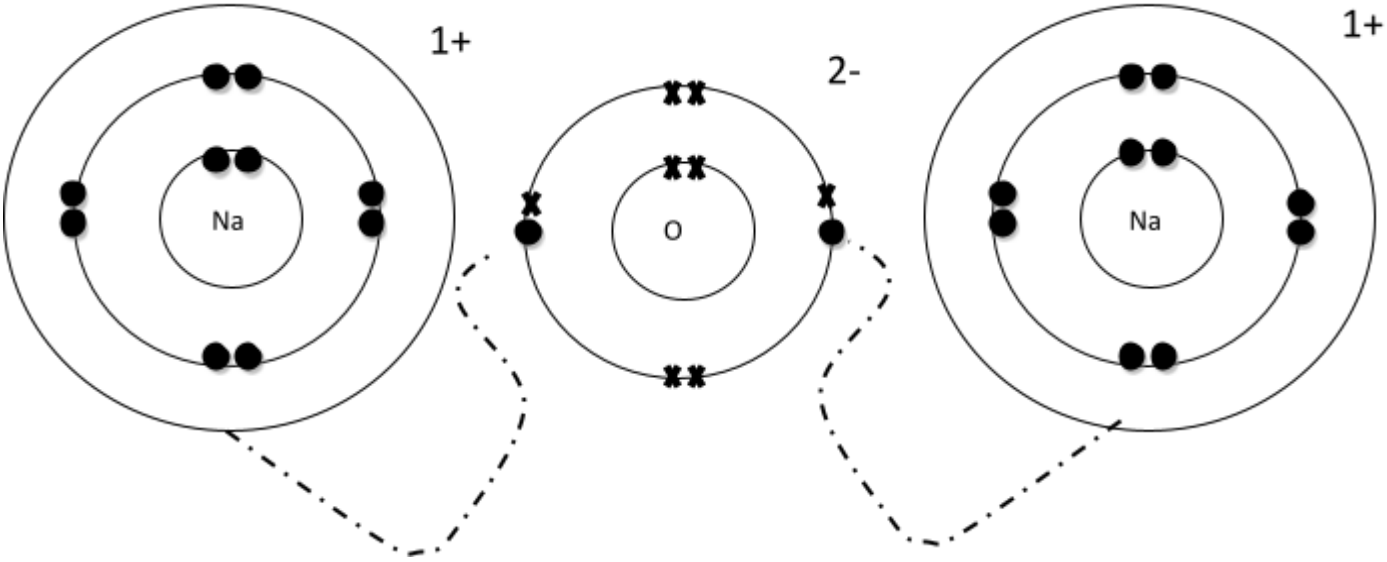
Example.

Sodium Oxide is made from Sodium and Oxygen atoms. First write the electronic structures.

$_{11}$Na => (2.8.1) This atom needs to lose **1 electron** to become **1+**

$_{8}$O => (2.6) This atom needs to **gain 2 electrons** to become **2-**

It takes two 1+ ions to cancel the charge of a 2- ion. So two Sodium ions bond to each Oxide ion.



❖ Task

- Write down the electronic structures of **Magnesium** and **Fluorine**.
- Work out how many atoms of each element are needed to form **Magnesium Fluoride.**
- Draw a **dot-cross diagram** to show the Ionic Bonding in the compound.

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

❖ Task

- Write down the electronic structures of **Potassium** and **Sulphur**.
- Work out how many atoms of each element are needed to form **Potassium Sulphide.**
- Draw a **dot-cross diagram** to show the Ionic Bonding in the compound.

❖ Task

- Write down the electronic structures of **Calcium** and **Chlorine**.
- Work out how many atoms of each element are needed to form **Calcium Chloride.**
- Draw a **dot-cross diagram** to show the Ionic Bonding in the compound.

❖ Task

- Write down the electronic structures of **Aluminium** and **Fluorine**.
- Work out how many atoms of each element are needed to form **Aluminium Fluoride**.
- Draw a **dot-cross diagram** to show the Ionic Bonding in the compound.

❖ Task

- Write down the electronic structures of **Aluminium** and **Oxygen**.
- Work out how many atoms of each element are needed to form **Aluminium Oxide.**
- Draw a **dot-cross diagram** to show the Ionic Bonding in the compound.

This lesson involves …

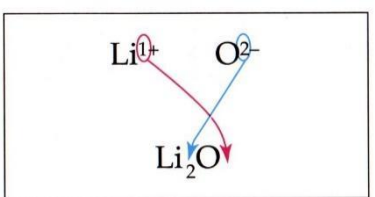
*Logical Thinking
*Precision
*Multi-step problem-solving

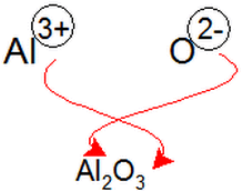
Writing Ionic Formulas

When the charge of ions match we use one of each in the formula

Eg Na^{+} & Cl^{-} → NaCl Ca^{2+} & O^{2-} → CaO Ga^{3+} & P^{3-} → GaP

When they don't match we "Swap and Drop"

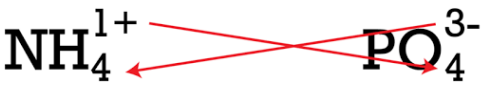




If we need to use more than one complicated ion (one with more than one atom in it) we use brackets

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving



$(NH_4)_3PO_4$

Question 1.

1. Find the following elements on the Periodic Table and use it to work out the charges of their ions.

a	Calcium	b	Oxide (from Oxygen)	c	Lithium
d	Aluminium	e	Flouride (from Fluorine)	f	Magnesium
g	Chloride	h	Bromide	I	Iodide

j Sodium k Gallium l Sulphide

2. Use your answers to Q1 to write formulas for:

a Calcium Oxide b Magnesium Fluoride

c Lithium Oxide d Magnesium Sulphide

e Aluminium Chloride f Sodium Iodide

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

g Gallium Bromide h Magnesium Bromide

Compound ions

Some ions are unpredictable so you will be told their charges

Copper (II) ions = Cu^{2+} **Iron (II) ions = Fe^{2+}** **Iron (III) ions = Fe^{3+} etc**

And some you just have to learn

Positive ion (Cation)	Formula	Negative ion (Anion)	Formula
Ammonium	NH_4^+	Carbonate	CO_3^{2-}
		Sulphate	SO_4^{2-}
		Nitrate	NO_3^-



		Hydroxide	OH^-

Hydrogen ions are H^+, Zinc ions are Zn^{2+}, Silver ions are Ag^+

Question

Write formulas for

a) Copper (II) Chloride

b) Iron (II) Oxide

c) Iron (III) Oxide

d) Calcium Carbonate

e) Calcium Hydroxide

f) Copper (II) Sulphate

g) Ammonium Fluoride

h) Ammonium Hydroxide

i) Ammonium Nitrate

j) Ammonium Carbonate

Giant Ionic Lattice.

We know an Ionic bond is an attraction from a Cation (+) to an Anion (-) and vice versa.

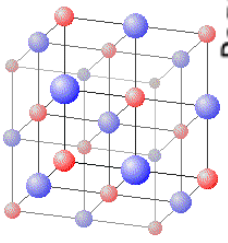
But Positive ions attract all the negative ions nearby, and the negative ions attract all the positive ions.

This leads to an arrangement that alternates from + to – and can be as large as you like. We call this an **Ionic Lattice.**

Ionic lattice structure

But this pattern continues in all three directions so that the real pattern is more like the bottom picture than the top.

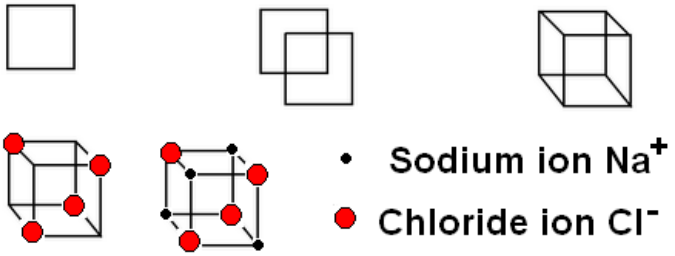
Since this is so difficult to draw we tend to space out the ions and draw the attractions between near-neighbours



Caution

Don’t confuse a Giant Ionic Lattice with a Giant Ironic Lettuce

Q1 Using the “spaced-out” style of ionic lattice draw 2 Sodium Chloride (NaCl) lattices with a key to show which ions are which. The guide below may he



Draw your own lattice

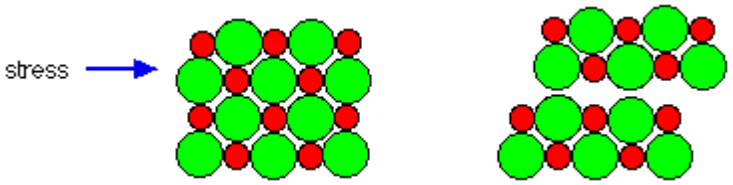
1. Start with a single vertical square.
2. Add another.
3. Join them.
4. Add ions of one colour in opposite corners.
5. Add ions of another colour in the remaining corners.
6. Make a key.

Q2 Why do you think we don’t show **all** the attractive forces between positive and negative ions?

Q3 Why does Magnesium Oxide take up the **same** lattice structure even though its ions are 2^+ and 2^-?

Ionic properties

If you hit an ionic crystal with a hammer and chisel (apply some stress) it would break.



Q1. Why can’t the crystal hold together in the new arrangement?

Q2. Substances that break easily when hit would be described as _____.

Q3. Ionic bonds often take a great deal of energy to break. What does this tell you about the temperature at which Ionic substances melt?

Sodium Chloride Crystal Structure

Ionic properties

Electrical properties.

Solid Ionic substances **do not conduct** electricity.

They are **Insulators**.

Q1. Why?

Q2. If we heat up (most) Ionic compounds enough they will eventually conduct. How?

This lesson involves …

*Logical Thinking
*Precision
*Multi-step problem-solving

Q3. Magnesium Oxide but can be used as an electrical insulator up to very high temperatures. Why is it different to Sodium Chloride in this regard?

Q4. Why do soluble ionic compounds conduct electricity when they've been dissolved?

Extended Response Questions

Question **Ionic bonding**

Table 2 gives some information about the elements lithium and fluorine.

Table 2

	Lithium	**Fluorine**
Metal or non-metal	Metal	Non-metal
Symbol	Li	F
Number of electrons in one atom	3	9

Lithium fluoride is an ionic compound.

Describe, in terms of electrons, how lithium atoms react with fluorine atoms to form lithium fluoride.

You may include diagrams to show the electronic structures of the atoms and ions.

[4 marks]

Mark scheme

Answers	Mark
Level 2: Scientifically relevant facts, events or processes are identified and given detail to form an accurate account.	3–4
Level 1: Facts, events or processes are identified and simply stated but their relevance is not clear.	1–2
No relevant content	0
Indicative content ● Electronic structure of lithium is 2.1 or diagram to show this. ● Lithium atoms lose electrons. ● One electron lost per atom. ● Electronic structure of lithium ion is 2 or diagram to show this. ● Formula of ion is Li^+. ● Electronic structure of fluorine is 2.7 or diagram to show this. ● Fluorine atoms gain electrons. ● One electron gained per atom. ● Electronic structure of fluoride ion is 2.8 or diagram to show this. ● Formula of ion is F^-. ● Formula of lithium fluoride is LiF.	

(c) (i) The electronic configurations of atoms of sodium and chlorine are

Na 2.8.1

Cl 2.8.7

Describe the changes in the electronic configurations of sodium and chlorine when these atoms form sodium chloride.

(3)

This lesson involves ...

*Practice
*Perseverance
*Resilience

(d) Solid lithium nitride conducts electricity and is used in batteries.

Why would you expect solid lithium nitride **not** to conduct electricity?

(1)

8 When lithium is burned in air, the two compounds lithium oxide (Li_2O) and lithium nitride (Li_3N) are formed.

Both compounds are ionic and their ions can be represented by dot and cross diagrams.

The dot and cross diagram for the ions in lithium oxide is

6 Potassium sulfide is an ionic compound.

(a) Complete the table to show the arrangement of electrons in the ions formed when potassium and sulfur react to form potassium sulfide.

Give the charge on each of the ions.

(3)

Element	Arrangement of electrons in atom	Arrangement of electrons in ion	Charge on ion
K	2.8.8.1		
S	2.8.6		

(b) (i) Explain why potassium sulfide conducts electricity when molten.

(1)

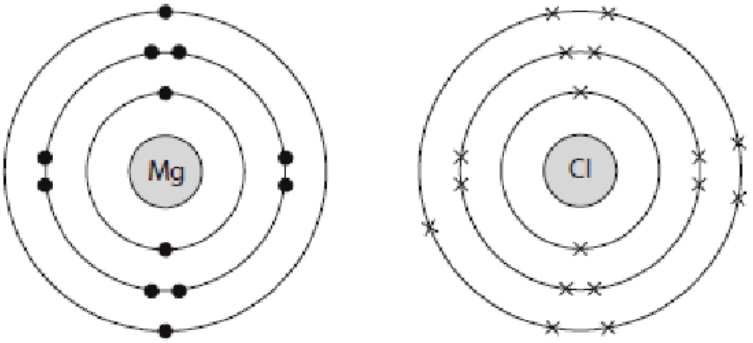
(ii) Explain why potassium sulfide has a high melting point.

(3)

(Total for Question 6 = 7 marks)

(b) Magnesium chloride ($MgCl_2$) is an ionic compound.

The diagram shows the electronic configurations of atoms of magnesium and chlorine.



This lesson involves ...

*Practice
*Perseverance
*Resilience

(i) Describe how magnesium atoms and chlorine atoms form magnesium ions and chloride ions.

(3)

(ii) Draw a diagram to represent the electronic configurations of each of the ions in magnesium chloride.

Show the charge on each ion.

(3)

(c) Calcium reacts with chlorine to form the ionic compound calcium chloride ($CaCl_2$).

(i) Describe, in terms of electrons, how an atom of calcium reacts with two chlorine atoms to form calcium chloride.

You may use a diagram in your answer.

(3)

(ii) Write the formula of a calcium ion.

(1)

5 Potassium chloride, KCl, is very similar to sodium chloride, NaCl. They have the same

(iii) Explain why magnesium chloride has a high melting point.

(3)

This lesson involves …

*Practice
*Perseverance
*Resilience

PES/

(c) Sodium is manufactured by the electrolysis of molten sodium chloride, NaCl(l).

Sodium is produced at the negative electrode and chlorine is produced at the positive electrode.

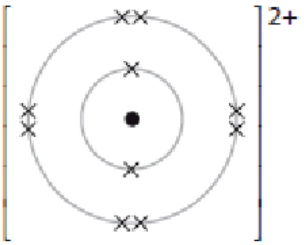
(i) Why does the sodium chloride have to be molten before it will conduct electr

This lesson involves …

*Automaticity
*Speed & Accuracy

(c) Magnesium ions and oxide ions are formed when magnesium reacts with oxyge

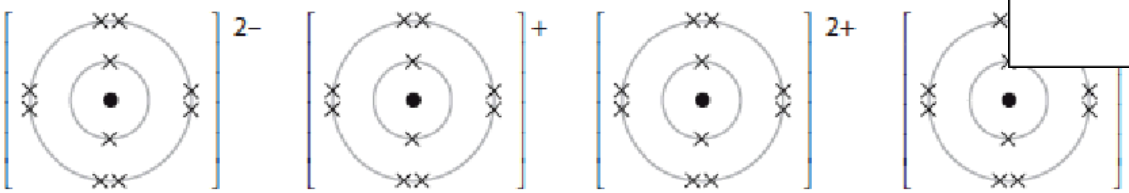
The diagram shows the electronic configuration and charge of a magnesium ion.



This lesson involves …

*Practice
*Perseverance
*Resilience

Put a cross in a box to indicate the diagram that shows the electronic configurati and charge of an oxide ion.



A ☐ B ☐ C ☐ D ☐

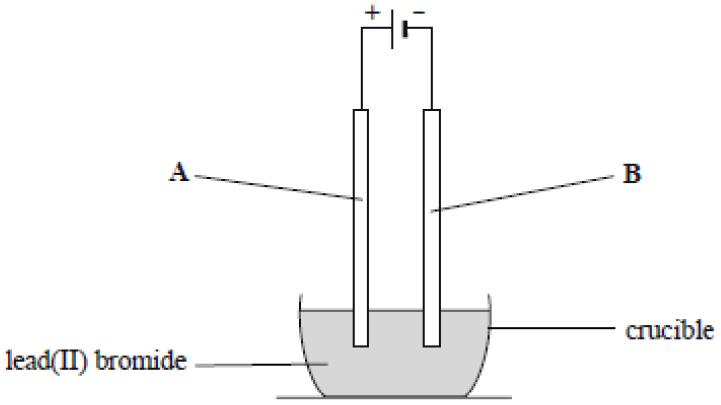
(d) A major use of magnesium oxide is as a refractory material, which is a material that can withstand very high temperatures.

Explain, in terms of its structure and bonding, why magnesium oxide has a very high melting point.

(4)

(b) Bromine is formed by the electrolysis of molten lead(II) bromide.

The diagram shows the apparatus used.



(i) Solid lead(II) bromide contains ions.

Why does solid lead(II) bromide not conduct electricity?

(1)

9.08 Metallic Bonding

I have reviewed the syllabus statements for this topic	
I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	

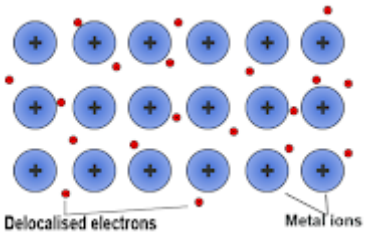
Prep Grade	
Test Grade	

Target	Pupil Signature

9.08 Metallic Bonding

Syllabus

- Metals consist of giant structures of atoms arranged in a regular pattern. The electrons in the outer shell of metal atoms are delocalised and so are free to move through the whole structure.
- The sharing of delocalised electrons gives rise to strong metallic bonds. The bonding in metals may be represented in the following form:



- Metals have giant structures of atoms with strong metallic bonding. This means that most metals have high melting and boiling points.
- Metals are good conductors of electricity because the delocalised electrons in the metal carry electrical charge through the metal. Metals are good conductors of thermal energy because energy is transferred by the delocalised electrons.

- In pure metals, atoms are arranged in layers, which allows metals to be bent and shaped. Pure metals are too soft for many uses and so are mixed with other metals to make alloys which are harder.

- Students should be able to explain why alloys are harder than pure metals in terms of distortion of the layers of atoms in the structure of a pure metal.

Metallic Bonding

The electronic structures of metal elements have something in common.

QUESTION: What are the electronic structures of the following metals:

$_{11}$Na_____ $_{20}$Ca_____ $_{3}$Li_____ $_{13}$Al _____ $_{12}$Mg _____

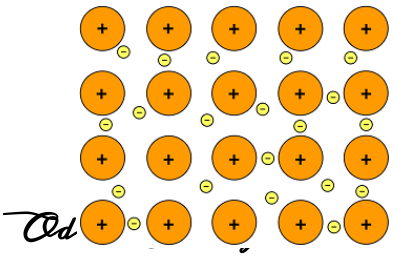
Metal atoms have extra electrons they need to _____ to achieve a full outer shell.

So metal ***atoms*** tend to form ***ions*** with _____ive charges.

Unfortunately, ions with the same charges repel rather than attracting.

So, how do metals stick together?

Look at this block of Sodium atoms.



In reality, they have all become 1+ Sodium ions.

Q1. What are the negative particles between the + ions?

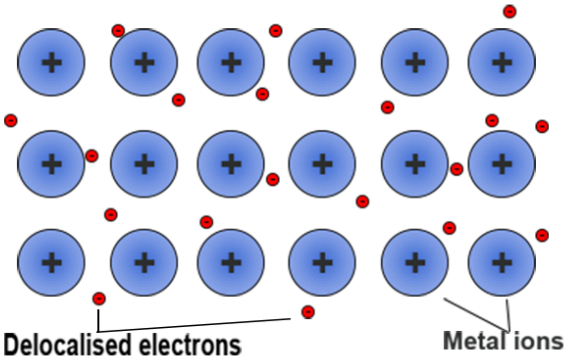
Q2. What causes attractive forces in this block? What causes repulsive forces in this block?

Q3. What do you notice about the distance from a +ion to a negative particle compared to the distance from one + ion to the next + ion?

Q4. Which type of force (attraction or repulsion) will be strongest?

Q5. Why do metals tend to have high melting points?

Metallic Properties



The electrons between the ions are said to be “delocalised” or “free”.

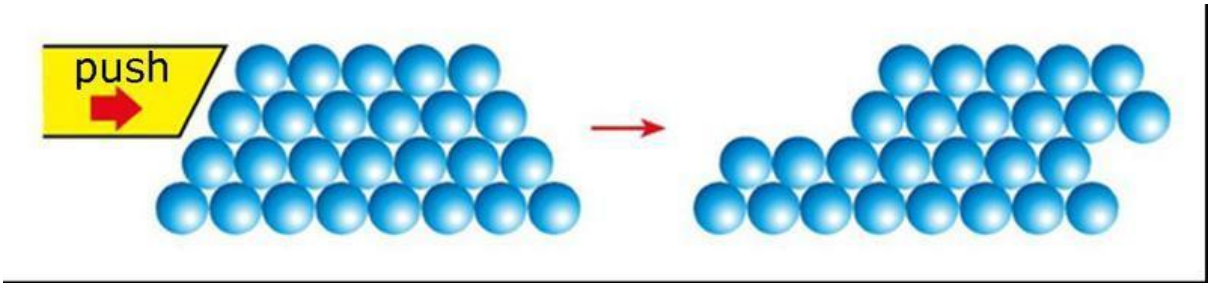
Q1. What do you think a “delocalised” electron means?

This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Q2. Which typical property/properties of metals is explained by delocalised electrons?

Q3. Metal ions tend to line up in rows. Look at the diagram below, how could you use it in an exam to explain why metals are malleable?



Alloys

An alloy is a mixture of a metal element with another element (usually another metal).

1. You should know that Brass is an alloy of C_____ and Z_____

2. Bronze is an alloy of C_____ and T_____.

3. Steel is an alloy of I_____ and C_____.

4. Adding Chromium to Steel makes Stainless steel – what is different about Stainless steel compared to normal steel?


Metal element structure


Alloy structure – harder than original metal

5. We saw that pure metals are soft because the layers of ions slide over each other. Why would adding ions of other elements that are different sizes make alloys harder and stronger?

6. Adding a lot of Carbon to steel makes it very hard. But it will also make it prone to shattering – it makes it b_____.

Extended Response Questions

Question **Ionic and metallic bonding**

Magnesium is a metal and magnesium chloride is an ionic solid.

Compare the bonding and structure in magnesium and magnesium chloride.
Include diagrams to illustrate the bonding in each substance.

[6 marks]

Mark scheme

Answers
Level 2: Scientifically relevant features are identified; the way(s) in which they are similar/different is made clear and (where appropriate) the magnitude of the similarity/difference is noted.
Level 1: Relevant features are identified and differences noted.
No relevant content

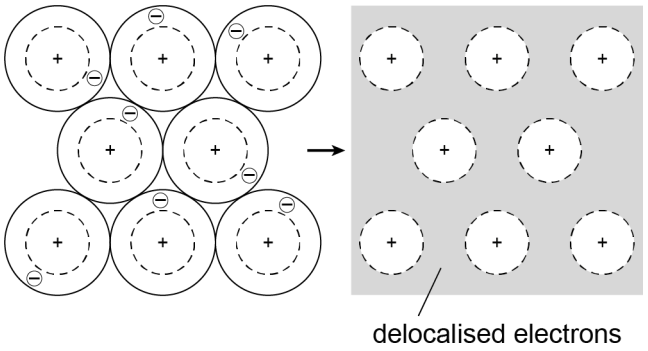
Indicative content

Similarities

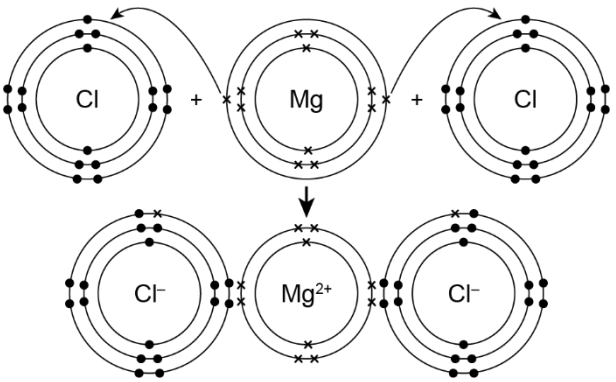
- Both have strong bonding.
- Both have giant structures.

Differences

- The electrons in the outer shell of magnesium atoms are delocalised.
- The delocalised electrons are free to move through the whole structure.
- Diagram of metallic bonding:



- Whereas in magnesium chloride the magnesium ions each lose two electrons from their outer shell
- Each chlorine atom gains one electron.
- Mg^{2+} and $2Cl^-$ ions are formed.
- Diagram of ionic bonding:



- Metallic bonding is the strong attraction between the positive ions and delocalised electrons
- Whereas ionic bonding is the strong attraction between positive ions and negative ions.
- Magnesium is a giant structure of atoms.
- The atoms are arranged in layers.
- Magnesium chloride is a giant structure of positive and negative ions.
- Theses ions are arranged in a lattice.

(c) The metallic structure of molybdenum gives it some typical properties.

(i) Describe the metallic structure of molybdenum.

(2)

This lesson involves …

*Practice
*Perseverance
*Resilience

This lesson involves …

*Practice
*Perseverance
*Resilience

(ii) Explain why molybdenum is a good conductor of electricity.

(2)

(iii) Explain why molybdenum is malleable.

(2)

6 Sodium (Na) and sodium chloride (NaCl) both have lattice structures.

Their melting points are shown in the table.

	Melting point in °C	Type of lattice structure
sodium	98	giant metallic
sodium chloride	801	

(a) Complete the table by stating the type of lattice structure in sodium chloride.

(1)

(b) Explain why sodium and sodium chloride have different melting points.

In your answer you should refer to

- the types of particle
- the types of forces between the particles in each substance

(5)

6 Copper and iron are metals. They have the typical properties of metals, including a high melting point and good electrical conductivity.

(a) (i) Draw a labelled diagram to show the arrangement of the particles involved in the bonding within a metal.

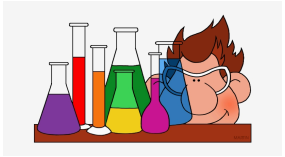
(3)

(ii) Explain why metals have good electrical conductivity.

(2)

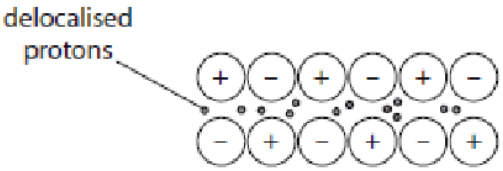
(iii) Suggest why some metals have high melting points.

(2)



10 This question is about magnesium and its compounds.

(a) A student draws this labelled diagram to show the particles in magnesium metal.



He makes two mistakes.

State the two corrections he should make to his labelled diagram.

(2)

1 .....

.....

2 .....

.....

(b) Explain why magnesium metal is malleable and a good conductor of electricity.

(4)

.....

.....

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9.08 Covalent Bonding

I have reviewed the syllabus statements for this topic	

I have completed the questions in this section	
I have read the relevant sections of the College Website	
I have made some revision material (mind-map, key-words & definitions etc)	
Prep Grade	
Test Grade	

Target	Pupil Signature

9.09 Covalent Bonding

Syllabus

- When atoms ***share pairs of electrons***, they form ***covalent bonds***.
- Bonds between atoms are ***strong***. Covalently bonded substances may consist of ***small molecules***.
- Some covalently bonded substances have very large molecules, such as polymers.
- Students should be able to

1. ***recognise*** common substances that consist of small molecules from their chemical formula.
2. draw ***dot cross diagrams*** for the molecules of ***Hydrogen, Chlorine, Oxygen, Nitrogen, Hydrogen Chloride, Water, Ammonia*** and ***Methane***

3. describe the ***limitations*** of using dot and cross, ball and stick, two and three-dimensional diagrams to represent molecules or giant structures
4. ***deduce the molecular formula*** of a substance from a given model or diagram in these forms showing the atoms and bonds in the molecule.

- Substances with small molecules are gases/ liquids with ***low melting points*** & ***boiling points***.
- These substances have only ***weak forces*** between the molecules (***intermolecular forces***).
- These intermolecular forces that break, ***not*** the covalent bonds, when the substance melts or boils.
- The intermolecular forces ***increase with the size of the molecules***, so ***larger molecules*** have ***higher*** melting and boiling points.
- These substances ***don't conduct*** because the molecules don't have overall electric charge.
- Students should be able to use the idea that intermolecular forces are weak compared with covalent bonds to explain the bulk properties of molecular substances.

- Substances that consist of ***giant covalent structures*** are ***solids*** with very ***high melting point***s. All of the atoms in these structures are linked to other atoms by strong covalent bonds. These bonds must be overcome to melt or boil these substances. ***Diamond*** and ***Graphite*** (forms of Carbon) and ***Silicon Dioxide*** (silica) are examples of giant covalent structures.
- Students should recognise giant covalent structures from diagrams of their bonding & structure.
- In ***Diamond***, each Carbon atom forms ***4 covalent bonds*** with other C atoms in a giant covalent structure, so diamond is ***very hard***, has a ***very high melting point*** and ***does not conduct electricity***.
- Students should be able to explain the properties of diamond in terms of its structure and bonding.

- In ***Graphite***, each Carbon atom forms ***three covalent bonds*** with three other Carbon atoms, forming ***layers*** of ***hexagonal rings*** which have ***no covalent bonds between the layers***.
- In Graphite, ***one electron*** from each Carbon atom is ***delocalised***.
- Students should be able to explain the properties of graphite in terms of its structure and bonding.
- Students should know that graphite is similar to metals in that it has delocalised electrons.

- ***Graphene*** is a single layer of Graphite and has properties that make it useful in ***electronics*** and ***composites***.

- Students should be able to explain the properties of Graphene in terms of its structure & bonding.
- ***Fullerenes*** are molecules of Carbon atoms with ***hollow shapes***.
- The structure of ***Fullerenes are*** based on ***hexagonal rings of Carbon atoms*** but they may also contain rings with ***five*** or ***seven*** Carbon atoms.
- The first ***Fullerene*** to be discovered was ***Buckminsterfullerene (C_{60})*** which has ***spherical*** shape.
- Carbon ***nanotubes*** are ***cylindrical*** fullerenes with very high length to diameter ratios.
- Their properties make them useful for ***nanotechnology***, ***electronics*** and ***materials***.
- Students should be able to:

1. recognise graphene and fullerenes from diagrams and descriptions of their bonding and structure
2. give examples of the uses of fullerenes, including Carbon nanotubes.

Covalent Bonding

In earlier years you learned that when elements combine they make a compound that is very different to the elements that form it.

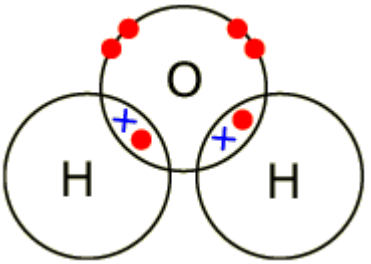
Hydrogen + Oxygen □ Water

Oxygen is a gas that makes fires burn.
Hydrogen is a gas that explodes on contact with flames.
Water is a liquid that is used to put out fires.

So how do these atoms stick to each other?

Covalent Bonding

In water, the outer (**valence**) shell of the Oxygen atom overlaps with the Hydrogen atoms' shells.

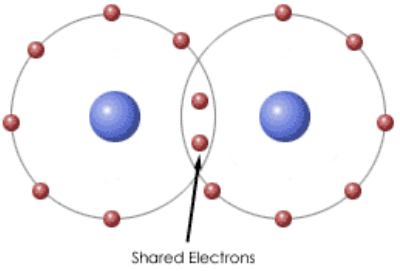


This means that both Hydrogen atoms now have **full outer shells**- their shells now have 2 electrons each.

The Oxygen's **valence shell** now has 8 electrons and is also full.

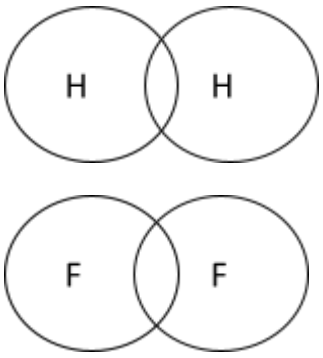
But this doesn't explain why the atoms stick together when you would expect the paired electrons to **repel** each other.

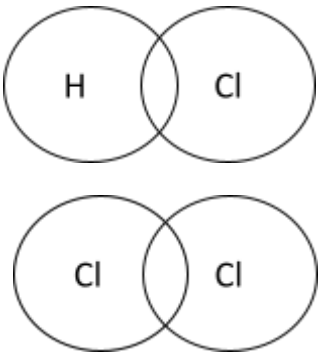
But each atom has a **positive** nucleus that is **attracted** to the **bond pair**. And the **bond pair** is attracted to the nucleus. Effectively, the two nuclei are pulled towards each other. A **molecule** has been formed – 2 or more atoms bonded together.

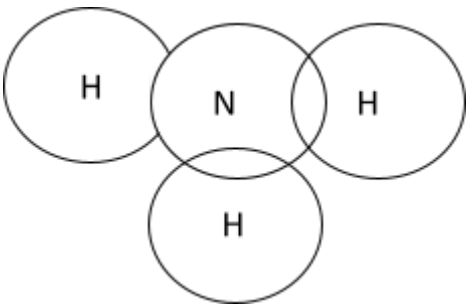


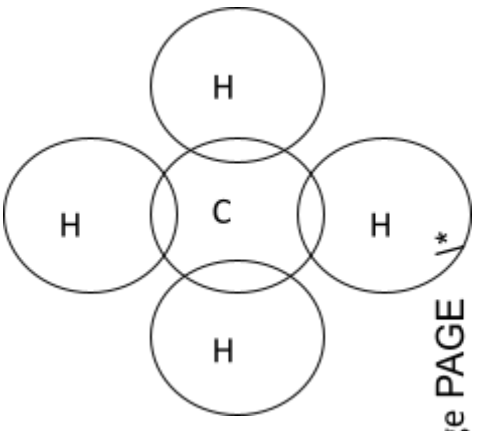
❖ Task

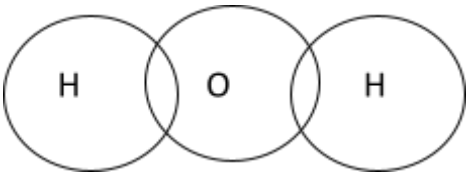
Look up the **Atomic Numbers** of the elements in each molecule and work out their electronic structure. Use **dots and crosses** to show how the elements share the unpaired electrons in their outer shells to form **covalent bonds**:











This lesson involves …

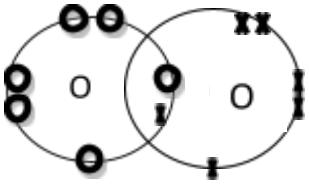
*Logical Thinking
*Precision
*Multi-step problem-solving

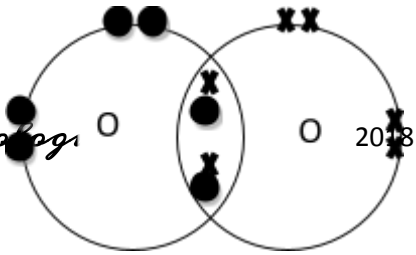
Double and Treble Bonds.

Oxygen exists as **diatomic molecules –** pairs of atoms – O_2

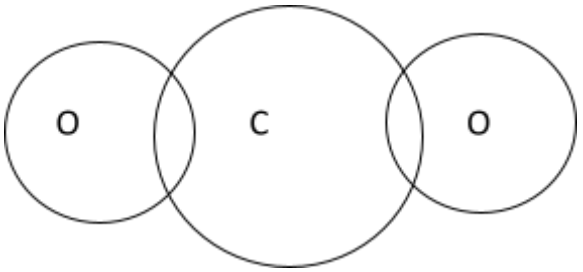
If we pair these molecules to form a single covalent bond then both atoms still has an **incomplete valence shell**.

But it is possible to have an overlap in which there are remaining single electrons so that there are two bond pairs – a double bond

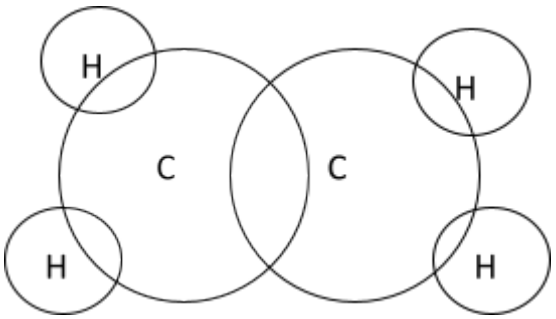




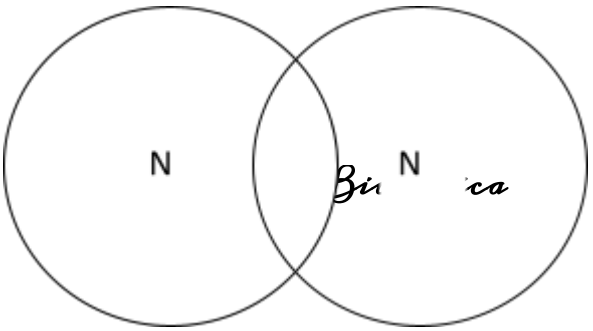
- Task: The following molecules will require at least one double or treble bond to be complete. Use dots and crosses to represent the electrons in all their covalent bonds.



Not on exams!



Not on exams!



Often on exams!

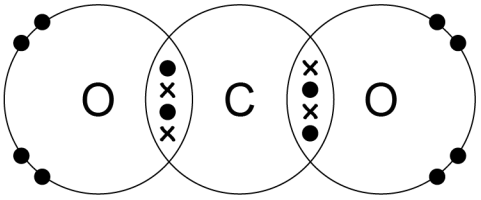
This lesson involves …

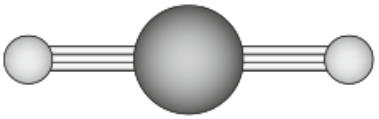
*Logical Thinking
*Precision
*Multi-step problem-solving

Question

Models

Some models representing a molecule of carbon dioxide are shown below

Type of model	Model
Structural formula	O=C=O
Dot and cross	

Ball and stick	

Explain what is meant by covalent bonding in terms of electrostatic attractions, shells, & electrons, and how these hold the atoms together in the molecule.

[6 marks]

Mark scheme

Answers	Mark
Level 3: Most relevant points are made and Science is used well to explain the nature of the attraction well so that a clear, well-ordered explanation of how atoms covalently bond is given.	5–6
Level 2: More relevant points are made and Science is used to partially explain the nature of the attraction	3–4
Level 1: Relevant points are made. They are not logically linked.	1–2
No relevant content	0
Indicative content ● A covalent bond is a shared pair of electrons. ● Some atoms may share 2 or 3 pairs of electrons in double and triple bonds ● Atoms share electrons by overlapping their outer shells ● Atoms share electrons in order to have a full outer shell. ● There is an attraction between the shared pair of electrons and the nuclei/nucleuses of the two atoms… ● …because the nuclei are positive and the electrons are negative ● Covalent bonds are electrostatic forces (of attraction). ● It takes a lot of energy to overcome these strong forces ● The atoms remain bonded unless the forces are broken	

Structure

The structure of a substance has a huge bearing on its properties.

Some covalent substances are solids with melting points over 1000°C.

Others are gases that don't condense until well below -100°C.

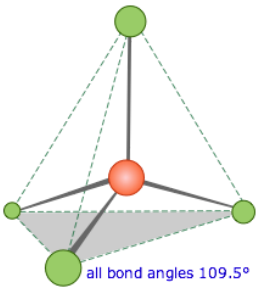
Why?

GIANT STRUCTURES.

Many covalent substances make individual **molecules** containing between 2 and dozens of atoms.

Giant structures can contain hundreds of thousands of billions of atoms! The largest ever **diamond** contained about 5,300,000,000,000,000,000,000,000 carbon atoms in one giant crystal. There are larger ones (and smaller ones) still in the ground so, although the arrangement of atoms is the same in each diamond, the number of Carbon atoms is not.

Diamond Structure



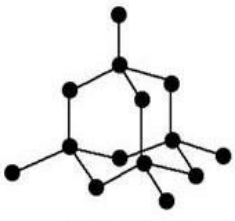
Each Carbon atom in a diamond is covalently bonded to **four** others.

They are arranged in a pattern called a **tetrahedron**.

❖ **Task**

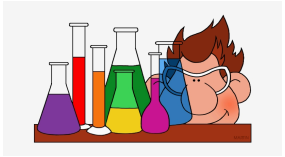
You are likely to be asked to draw a small diamond structure in exams.

Check Page 28 of your text-book if unsure how, or copy the diagram here at least twice to practise.


Structure of Diamond

Q1. Explain why Diamond has such a high **melting point** (3550°C).

Q2. Explain why Diamond is so useful for industrial **drills** and **saws**.



This lesson involves …

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

Q3. Why do you think Diamond does not conduct electricity?

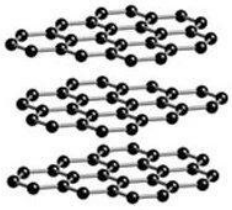
Q4. Why do you think Diamond doesn't **dissolve**?

Q5. Explain what the difference between the meanings of "**hard**" and "**strong**".

Giant Covalent Structures #2

<u>Graphite Structure</u>.

Graphite is another ***allotrope*** of Carbon but common enough and different enough from Diamond to be used in pencils as "***lead***".



Unlike the Carbon atoms in Diamond the Carbon atoms in Graphite are only **covalently bonded** to ***three*** other Carbon atoms.

This results in **flat sheets** of atoms that are as difficult to break up as the Carbons in Diamond.

But between the layers there are only **weak attractions**, allowing the layers to slide over each other. These attractions are caused by **delocalised electrons** (as in metal structures).

Q1. Explain why Graphite can conduct electricity when Diamond cannot.

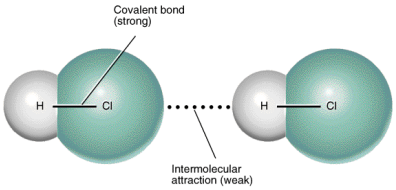
Q2. Explain why, despite the weak attractions between layers, Graphite still has a **high melting point** (over 3000°C)

Q3. Explain how the structure of Graphite makes it suitable for using in a **pencil**.

Q4. Both Diamond and Graphite are pure Carbon but Graphite is less dense. Why?

Q5. Draw the structure of Graphite.

ar Su



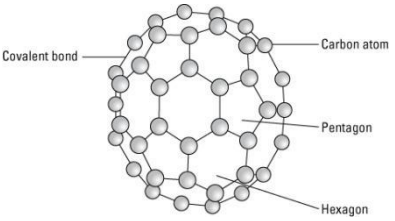
The bonds within a covalent substance are strong.

But the forces attracting one molecule to its neighbours are much weaker.

This means it doesn't take much energy to melt or boil them.

Another allotrope of Carbon is the Buckminsterfullerene. It has a total of 60 Carbon atoms C_{60}.

You can see why it might be called a Buckyball.



One C_{60} molecule attracts another only weakly because there are no covalent bonds between the molecules - only inside them.

So Buckminsterfullerene is an allotrope of carbon that would easily turn into a gas – unlike Graphite or Diamond.

Although each C_{60} molecule has delocalised electrons they can't move from one molecule to another – so, like diamond, it is not a conductor of electricity.

Big molecules like this tend to have larger boiling points because they have stronger intermolecular forces.

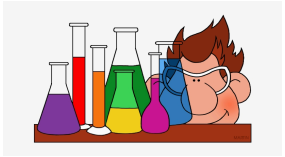
This lesson involves ...

*Generalisation
*Connection-finding
*Big Picture Thinking
*Abstraction
*Imagination

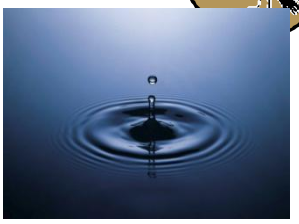
Q1. Why don't simple molecular substances conduct electricity?

Q2. The intermolecular forces in water are unusually **strong**. Why do most other simple molecular substances not dissolve in water?

Q3. **Organic solvents** like paint thinner or dry-cleaning fluid have much weaker intermolecular forces. Why do so many simple molecular substances dissolve in Organic solvents?



Q4. Most simple molecular substances are gases. Why is water a liquid at room temperature and at temperatures as high as 100°C?



Q5. Sand is mostly Silicon Oxide which is a covalently bonded substance. It doesn't melt until 1700 °C. What does this tell you about the structure of Silicon Oxide?

Q6

Table 1 *Hydrocarbon Boiling Points*

Hydrocarbon	Formula	Boiling Point (°C)
Butane	C_4H_{10}	-0.5
Decane	$C_{10}H_{22}$	174.0
Ethane	C_2H_6	-88.6
Heptane	C_7H_{16}	98.4
Hexane	C_6H_{14}	68.7
Methane	CH_4	-161.7
Nonane	C_9H_{20}	150.8
Octane	C_8H_{18}	125.7
Pentane	C_5H_{12}	36.1
Propane	C_3H_8	-42.1

Look at the formulas and boiling points of these molecules. What is the trend? Suggest an explanation.

Extended Response Questions

Methane and diamond both contain carbon and have covalent bonding.

Some of their properties are shown below

Table 3

Property	Methane	Diamond
Melting point in °C	–182	3730
Conductor of electricity	no	no

Explain these two properties in terms of the structure and bonding in methane and diamond.

[6 marks]

Mark scheme

Answers	Mark
Level 3: Relevant points (reasons/causes) are identified, given in detail and logically linked to form a clear account.	5–6
Level 2: Relevant points (reasons/causes) are identified, and there are attempts at logical linking. The resulting account is not fully clear.	3–4
Level 1: Points are identified and stated simply, but their relevance is not clear and there is no attempt at logical linking.	1–2
No relevant content	0
Indicative content ● A covalent bond is formed when two atoms share a pair of electrons. ● Methane has a low melting point because it has a simple molecular structure. ● It consists of CH_4 molecules ● With strong covalent bonds between the C and H atoms. ● There are weak intermolecular forces / weak forces between the molecules ● So not much energy is needed to overcome the forces ● To separate the molecules. ● Diamond has a high melting point because it has a giant covalent structure. ● The carbon atoms are held together in a giant lattice with each atom joined to four others ● With strong covalent bonds between all the atoms. ● A lot of energy is needed to break the covalent bonds ● To separate the atoms. ● Methane and diamond do not conduct electricity as there are no charge carriers / ions / delocalised electrons ● To carry the current. Credit may be given for diagrams showing the structure and bonding in methane and/or diamond.	

(e) A molecule of silicon tetrafluoride (SiF_4) contains covalent bonds.

Draw a dot and cross diagram to show the outer electrons in this molecule.

(2)

F

F Si F

F

This lesson involves …

*Practice
*Perseverance
*Resilience

(f) The table shows the boiling points of some compounds containing silicon. All of these compounds contain covalent bonds.

Compound	Boiling point in °C
SiF_4	–86
$SiCl_4$	58
SiO_2	2950

This lesson involves ...

*Practice
*Perseverance
*Resilience

SiF_4 and $SiCl_4$ have simple molecular structures.

SiO_2 has a giant covalent structure.

(i) Explain why the boiling point of $SiCl_4$ is greater than the boiling point of SiF_4

(2)

This lesson involves ...

*Practice
*Perseverance
*Resilience

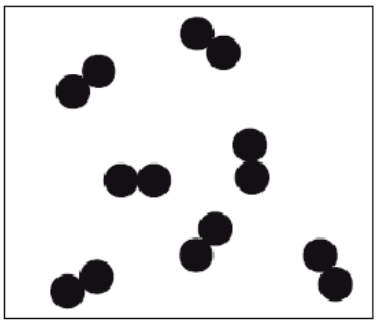
(ii) Explain why the boiling point of SiO_2 is very much greater than the boiling point of $SiCl_4$

(2)

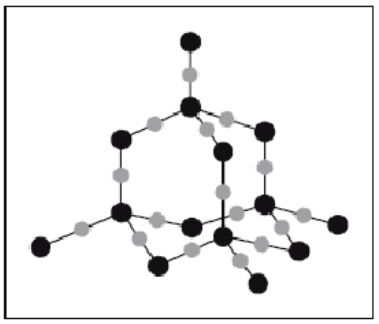
(b) The main substance that acts as an insulator in this method of transmission of electricity is air, which is mostly nitrogen.

The power lines are supported by solid insulators. Most solid insulators are manufactured using silica.

The diagram shows the structures of nitrogen and silica.



Nitrogen



Silica

Explain, in terms of bonding and structure, why nitrogen is a gas at room temperature but silica is a solid with a high melting point.

(5)

Nitrogen .....

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Silica .....

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.....

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5 The table shows some properties of four substances A, B, C and D.

Substance	Melting point in °C	Boiling point in °C	Conducts electricity when solid?	Conducts electricity when molten?
A	−101	−35	no	no
B	1063	2970	yes	yes
C	801	1413	no	yes
D	3550	4830	no	no

(a) Use the information in the table to identify the substance that

(i) is a metal

(1)

☐ **A** ☐ **B** ☐ **C** ☐ **D**

(ii) could be diamond

(1)

☐ **A** ☐ **B** ☐ **C** ☐ **D**

(iii) is a gas at 20 °C

(1)

☐ **A** ☐ **B** ☐ **C** ☐ **D**

(iv) contains oppositely charged ions

(1)

☐ **A** ☐ **B** ☐ **C** ☐ **D**

(b) Some of the substances in the table are compounds.

What is meant by the term **compound**?

(2)

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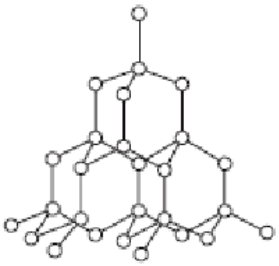
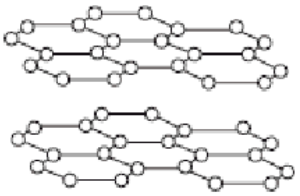
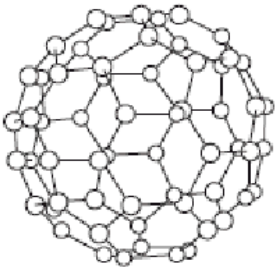
13 The diagram shows three different forms of carbon.



diamond structure



graphite structure



fullerene molecule

(a) Name the type of bond that exists between the carbon atoms in all three structures.

(1)

(b) (i) Explain why diamond has a very high melting point.

(4)

(ii) Fullerene has a simple molecular structure.

Explain why it has a low melting point.

(2)

(b) Hydrazine (N_2H_4) is a useful compound that can be manufactured from ammonia.

(i) Hydrogen peroxide can be used to convert ammonia to hydrazine.

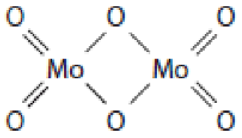
Balance the equation for this reaction.

(1)

..... NH_3 + ... H_2O_2 → ..... N_2H_4 + ... H_2O

(iii) The melting point of molybdenum oxide suggests that it has ionic bonding. However, it is often represented as a molecular structure.

Deduce the molecular formula of molybdenum oxide as shown in this structure.

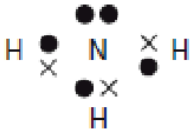


(1)

This lesson involves ...

*Automaticity
*Speed & Accuracy

1 The diagram represents a particle of ammonia.



(a) This particle of ammonia is

(1)

- ☒ **A** an atom
- ☒ **B** an ion
- ☒ **C** a lattice
- ☒ **D** a molecule

(b) Which type of bonding is present in this particle of ammonia?

(1)

- ☒ **A** covalent
- ☒ **B** hydrogen
- ☒ **C** ionic
- ☒ **D** metallic

(c) What is the formula of ammonia?

(1)

(Total for Question 1 = 3 marks)

This lesson involves ...

*Automaticity
*Speed & Accuracy

*Resilience

9 This question is about bonding, structures and properties.

(a) The box gives four types of structure.

giant covalent	giant ionic	giant metallic	simple molecular

The table shows some properties of four substances, A, B, C and D.

Complete the table by giving the correct type of structure for each substance.

You may use each structure once, more than once or not at all.

(4)

Substance	Electrical conductivity		Melting point	Type of structure
	of the solid	of the liquid		
A	poor	poor	low	
B	poor	poor	high	
C	good	good	high	
D	poor	good	high	

(c) A molecule of carbon dioxide contains double covalent bonds.

Complete the diagram, using dots and crosses, to show the arrangement of the outer electrons in a molecule of carbon dioxide.

O C O

(2)

(d) Indium is a metal in Group 3 of the Periodic Table.

(i) Describe the structure and bonding in indium.

(3)

(ii) Explain why indium is malleable.

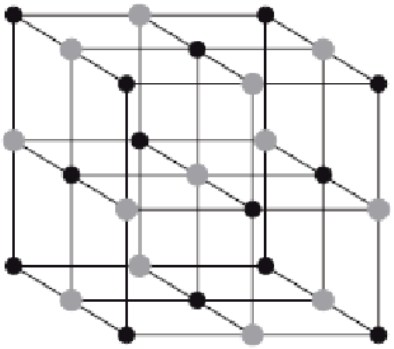
(2)

14 Sodium chloride (NaCl) and silicon dioxide (SiO_2) both have giant lattice structures.

Sodium chloride is an ionic compound.

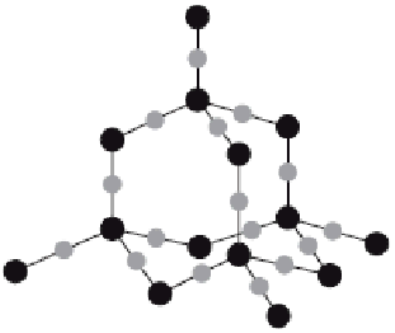
Silicon dioxide is a covalent compound.

Structure of sodium chloride



● Na^+ ● Cl^-

Structure of silicon dioxide



● Si ● O

The table shows some properties of each compound.

Sodium chloride	Silicon dioxide
melting point = 801 °C	melting point = 1610 °C
soluble in water	insoluble in water
conducts electricity when molten	does not conduct electricity when molten

(a) (i) Explain why silicon dioxide has a high melting point.

(2)

(ii) Suggest why the melting point of silicon dioxide is higher than the melting point of sodium chloride.

(1)

(b) State why sodium chloride conducts electricity when molten.

(1)

(c) Carbon dioxide is described as a simple molecular substance.

State why carbon dioxide (CO_2) is a gas at room temperature.

(1)

(b) Bromine reacts with phosphorus to form the covalent compound phosphorus tribromide.

Draw a dot and cross diagram to show the outer electrons in a molecule of phosphorus tribromide.

(2)

Br P Br

Br

This lesson involves …

*Practice
*Perseverance
*Resilience

(b) The electronic configuration of a sodium atom is 2.8.1
Sodium oxide, Na_2O, is an ionic compound formed when sodium reacts with oxygen.

PE (i) Describe, in terms of electrons, what happens when sodium oxide is formed in this reaction.

(3)

(c) Explain why water has a much lower melting point than sodium oxide.

(2)

(b) Hydrogen bromide (HBr) and sodium bromide (NaBr) are compounds of bromine.

(i) Draw a dot and cross diagram to represent a hydrogen bromide molecule.

Show only the outer electrons in each atom.

(2)

(ii) Explain how the atoms are held together in a hydrogen bromide molecule.

(2)

.....

.....

.....

(iii) Explain why sodium bromide has a higher melting point than hydrogen bromide.

(3)

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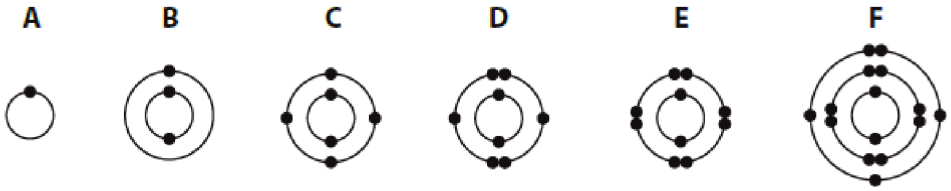
4 The diagram shows the electronic configurations of six different atoms.



(b) Atoms of A and D combine to form a compound containing covalent bonds.

(i) Complete the sentence to describe a covalent bond.

(2)

A covalent bond is the electrostatic attraction between a pair of .....

and the ..... of two atoms.

(ii) Suggest, with reference to electronic configurations, the most likely formula of the compound formed between atoms of A and D.

(1)

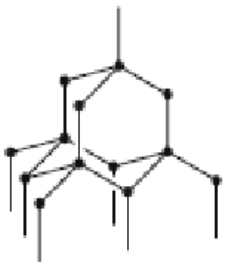
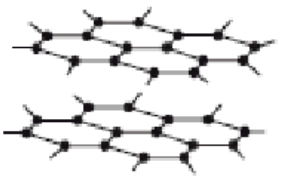
.....

This lesson involves …

*Practice
*Perseverance
*Resilience

6 Diamond and graphite are two naturally-occurring forms of carbon.

The diagrams below show the arrangement of the carbon atoms in diamond and in graphite. The black dots (•) represent carbon atoms.

Diamond	Graphite
	

This lesson involves …

*Automaticity
*Speed & Accuracy

(a) Name the type of structure in diamond and explain, in terms of its bonding, why diamond has a high melting point.

(4)

(b) Explain, in terms of its structure, why graphite can act as a lubricant.

(2)

(c) The structure of graphite has one feature in common with that of metals. This feature allows graphite to conduct electricity.

Suggest what this feature is and why it allows graphite to conduct electricity.

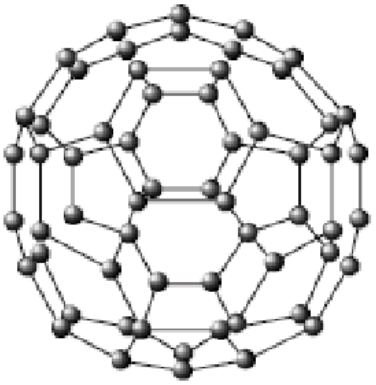
(2)

(d) In 1985, a new form of carbon was discovered. It was called buckminsterfullerene after the architect Buckminster Fuller, who designed buildings with complex geometric shapes.

Buckminsterfullerene (C_{60}) has a simple molecular structure containing 60 carbon atoms per molecule. It looks a little bit like a football.

This lesson involves …

*Practice
*Perseverance
*Resilience



Suggest why buckminsterfullerene has a much lower melting point than diamond.

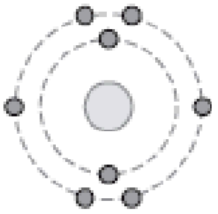
(2)

(Total for Question 6 = 10 marks)

6 The diagram shows how the electrons are arranged in an atom of oxygen.

This lesson involves ...

*Automaticity
*Speed & Accuracy



Oxygen atoms form both covalent and ionic bonds.

(a) Water is formed when two atoms of hydrogen combine with one atom of oxygen.

(i) Draw a dot and cross diagram of a molecule of water. You need only show the electrons in the outer shells.

(2)

(ii) Explain how the covalent bonds in the water molecule hold the hydrogen and oxygen atoms together.

(2)

9 When water is added to a mixture of sand and cement, a reaction takes place between silicon dioxide in the sand and calcium oxide in the cement. The reaction produces a salt called calcium silicate.

The equation for the reaction is:

$$SiO_2 + CaO \rightarrow CaSiO_3$$