Unit 2
ATOMIC STRUCTURE & BONDING **10 minutes**
Atomic Structure

Test Yourself

1 The patterns shown in the atomic _____ of gases caused Niels Bohr to develop a new _____ for the atom. He suggested that the _____ of an atom contained protons and neutrons, with _____ rotating around it in fixed _____. The number of protons in the nucleus is called the _____ _____, with the symbol __. The mass number of an element is the sum of the number of _____ plus _____ in its _____.

2 When a chemical reaction occurs between two elements _____ are transferred from one to the _____. Mendeleev then classified elements which underwent the same kind of reactions into _____, such as the Alkali Metals.
From the patterns shown in the _____ Table it was realised that electrons exist in particular energy _____, each with sub-shells.

3 Each energy level can only hold a maximum number of _____ before it is _____. Each sub-level orbital, such as s, p, d has a _____ number of lobes, each holding one _____ of electrons (e.g px, py, pz lobes). The p sub-level has 3 parts to it ('lobes'), whereas the d has ____ lobes and the f sub-level has ____ lobes.

4 The electron configuration of an atom shows how many _____ there are in each level and what _____ of orbital it is. For instance $1s^2 2s^2 2p^5$ represents the element called _____.

5 The Periodic Table shows every _____ in order of filling of _____. The period number shows the principal _____ level and the columns show the _____ number. Elements with completely filled shells are called the _____ _____.

6 Fill in the symbols for the elements with given atomic numbers below:

Z value	Z = 3	Z = 8	Z = 12	Z = 18	Z = 21	Z = 25	Z = 30
Element symbol							

7 The electrons contained in the outer shell only are called the _____ electrons. These are responsible for all of its chemical _____. Reactions between elements occur by _____ transfer so that each element can achieve a _____ shell, which is the lowest _____ value attainable. Metals, which are mostly on the _____-hand side of the Periodic Table, tend to _____ their electrons when bonding, whereas the _____ - metals have more than 4 valence electrons and tend to _____ electrons when bonding.

8 Metal to non-metal bonds are called _____ bonds as the metal _____ its valence electrons to become a + ion. In accepting electrons the non-metal will become a __ ion and therefore be _____ to the metal ion and form a strong _____.

9 Bonds formed between 2 non-metals are called _____ bonds. Here electrons are _____ so that each atom achieves the _____ gas structure. There can also be the case

where one atom is responsible for _____ both of the electrons in a bonding pair. This kind of _____is called a _____ covalent bond.

Answers:

1. Spectra; model; nucleus; electrons; orbitals; atomic number; z; protons; neutrons; nucleus.
2. Electrons, other, groups; Periodic; levels.
3. Electrons; full; particular; pair; 5; 7.
4. Electrons; type; fluorine.
5. Element; shells; energy; group; noble gases.
6. 3 = Li; 8 = O; 12 = Mg; 18 = Ar; 25 = Mn; 30 = Zn.
7. Valence; reactions; electron; full; energy; left; lose; non; gain.
8. Ionic; loses; -; attracted; bond.
9. Covalent; shared; noble; donating; bond; co-ordinate.

Unit 2
ATOMIC STRUCTURE & BONDING **20 minutes**
Atomic Structure

Key Points

1 In the early 1900s Niels Bohr derived a new model for the atom from evidence gained from **atomic spectra**. Bohr's model has a central nucleus containing protons and neutrons with **electrons rotating in orbitals**, or shells.

- The atomic number (z) of an element = number of protons in the nucleus and the mass number (A) = sum of protons + neutrons in the nucleus.

2 All chemical reactions simply involve **transfer of electrons** from one chemical species to another (the neutrons and protons play no part). Similarities existing between elements are due to their similar **electron configurations**.

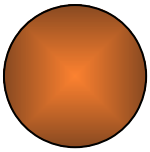
- Mendeleev had previously placed elements with similar properties into families, or **Groups** e.g. alkali metals, halogens, etc.
- This classification into groups led to an extension of the Bohr model to include ideas of **energy levels** and **sub-shells** which explains the structure of the **periodic table** better**.**

3 A refinement of the Bohr Model now shows that the number of electrons in the atomic orbitals increases as we go through the periodic table. These are arranged in energy levels that can only hold a certain number before they are filled.

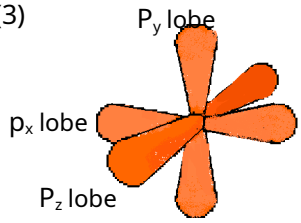
- Each principal energy level holds electrons in sub-levels with differently shaped orbitals.

Orbital shapes

s-orbital (1 lobe)



p-orbitals (3)



d-orbitals (5) Complex shape **f-orbitals** (7) Complex shape

- Each individual sub-level orbital lobe (s, p_x, p_y, etc) can hold one pair of electrons – one as "spin-up" (↑) and one as "spin-down" (↓) before it is full.

- Sub-level type No. of orbitals No. of electrons to fill sub-level

s	1	2	s↑↓

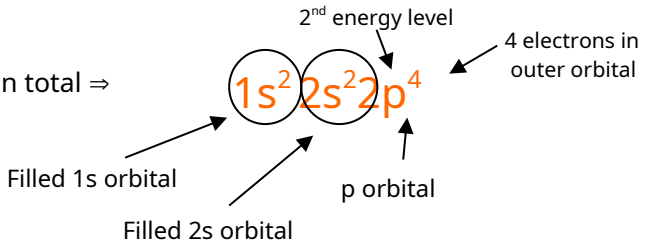
p	3	6	p_x↑↓ p_y↑↓ p_z↑↓
d	5	10	↑↓ ↑↓ ↑↓ ↑↓ ↑↓
f	7	14	↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓

- Each time an energy level is added an extra sub-level is introduced

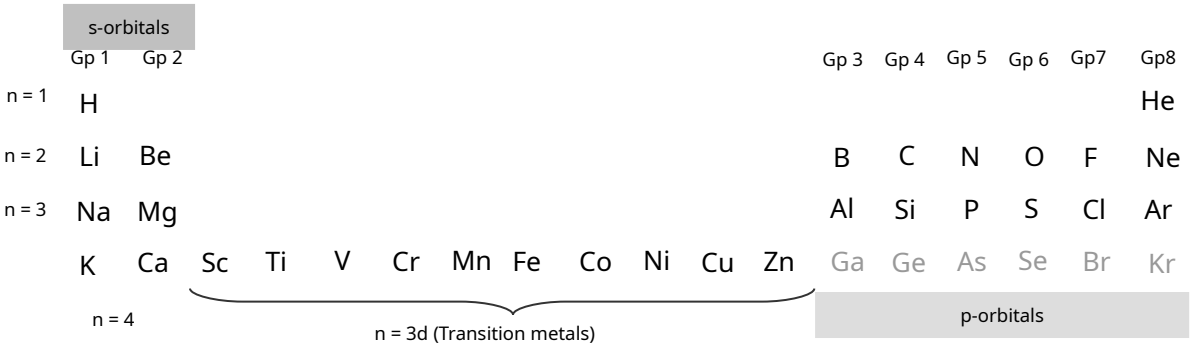
Energy level	Sub-levels (electrons to fill)	Max. No. of electrons to fill
n = 4	s(2), p(6), d(10), f(14)	32
n = 3	s(2), p(6), d(10)	18
n = 2	s(2), p(6)	8
n = 1	s(2)	2

4 The **electron configuration of elements** shows energy level, orbital type and No. of electrons held in that orbital.
e.g. The oxygen electron configuration is shown below:

Oxygen has 8 electrons in total ⇒ $1s^2 2s^2 2p^4$



5 The structure of the periodic table of elements corresponds to the filling of these sub-shells from lowest to highest energy.
The **energy level** corresponds to the **period.**
The sub-shell **electron content** corresponds to the **group**.



Note: The 4s orbital has as lower energy value than the 3d orbitals and is filled before them (**4s before 3d)**

(e.g Calcium has a configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ and scandium has a configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$)
Note: You will need to remember the top half of this table up to element No. 20 (Ca) and it would also be an advantage to remember the symbols for the transition metals.

6 **Remembering** the first 20 elements by sentences.

Harry **H**e **Li**kes **Be**er **B**ut **C**old, **N**ot **O**ver **F**rothy **Ne**llie's **Na**nny's **MG** **Al**ways **Si**gnals **P**romptly – **S**ome **Cl**ots **Ar**e **K**angaroo **Ca**tchers

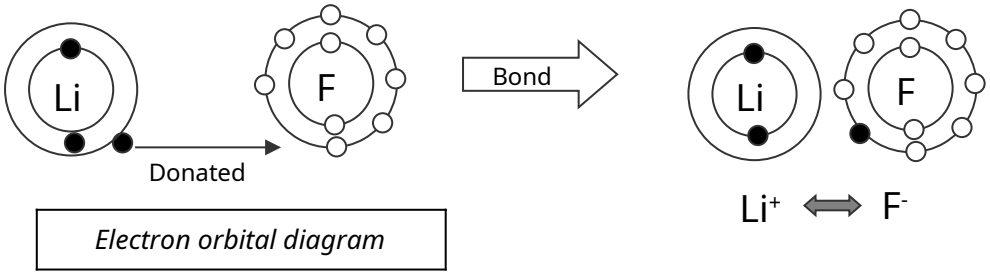
- **Transition metals** mnemonic
  Scandinavian **T**raction **V**ehicles **Cr**eate **M**ore **Fe**ar and **Co**ngestion at **Ni**ght than **Cu**rteous (sp) **Z**ulus

7 **Group number** shows the number of **valence electrons** (outer electrons free to be used in bonding reactions). Hence all elements in a particular group undergo **similar reactions** and form compound with similar bonds and formulae.

- Bonds between elements are formed to lower the total energy by achieving full electron shells of 2 (n = 1), 8 (n = 2), 8 (n = 3), etc. Elements with < 4 valence electrons tend to **donate** them in a bond (e.g. Li, Be, B) and those with > 4 electrons tend to **accept** them in a bond (e.g. N, O, F). They can then achieve a **noble gas configuration**. Noble gases undergo **no reactions** because their electron configuration is very stable (lowest energy attainable).

8 Metal to non-metal bonding uses **ionic bonds** where the metal donates electrons to the non-metal.

e.g 1 Lithium fluoride



Electron orbital diagram

e.g. 2 Sodium sulfide

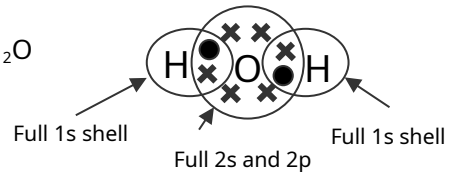
$2Na^+ \iff S^{2-}$

Electron dot diagram
(just considering valence electrons)

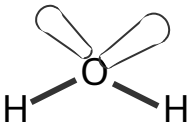
$2Na^+ \iff S^{2-}$

9 **Covalent bonds** occur between 2 non-metals that both need to accept electrons. The compromise is the **sharing** of electrons in a **covalent** bond to each achieve a noble gas configuration.

e.g. H_2O



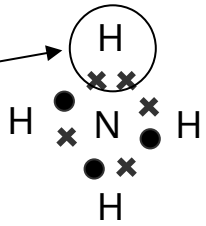
⇒



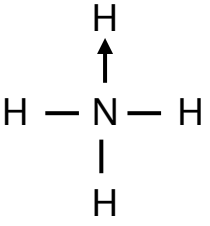
- A covalent **co-ordinate** bond is formed when one element donates **both** electrons in a covalent bond.

e.g. 1 The NH_4^+ ion (1 electron missing)

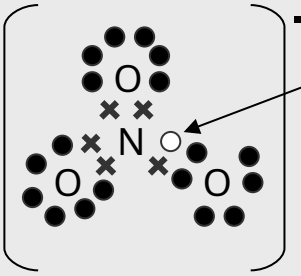
Shell is full. Both of these electrons come from the N atom



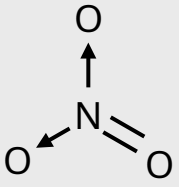
Another representation ⇒



e.g. 2 The NO_3^- ion (1 electron extra)



Another representation ⇒



Checklist – can you:

1. State Bohr's Orbital model for the atom?
2. Define the Atomic Number and Mass Number for an element?
3. Explain the arrangement of the Periodic Table into periods and groups?

4. Explain how the electrons of an atom are arranged into energy levels and orbitals?
5. Recall the shape of s and p orbitals?
6. Recall the number of orbital lobes and total number of electrons held in the s, p, d and f orbitals?
7. Write the electron configuration for any element from its position in the Periodic Table ($1s^2 2s^2 2p^6$, etc)?
8. Recall the order of filling of shells and orbitals for elements?
9. Recall the names and symbols for the first 20 elements?
10. Recall the names and symbols for the Transition Metals?
11. Explain how the valence electrons in an element influence its reactions and bonding in compounds?
12. Recall the donating/ receiving patterns of electron movement for the reactions of metals and non-metals?
13. Explain and give graphical examples for the formation of ionic bonds?
14. Explain and give graphical examples for the formation of covalent bonds?
15. Explain and give graphical examples for the formation of co-ordinate covalent bonds?

Unit 2
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Atomic Structure

Key Exam Questions

1. Fill in the entire table below with any missing numbers relating to the 5 atoms shown in the top row.

Symbol $^{12}_{12}Mg$ $^{85}_{37}Mg^{2+}$ $^{17}_{}Cl^{2-}$ $^{226}_{}$ Protons Neutrons 18 25 19 Electrons 17 23 Atomic No. Mass No. 48 Net charge 0

Hint 1 [12 marks]

2. In the table above identify:
a) A halogen
b) A Transition metal
c) A non-metals with a valency of -2
d) A metal that is not from the Transition series. [4 marks]

3. a) Use the Periodic Table to explain how
(i) Evidence shows that the 1s shell only needs 2 electrons for completion. [2marks]
(ii) We know that it takes 10 electrons to fill a d-orbital. [2 marks]
b) Give the electron configurations for
(i) Si [1 mark]
(ii) Ti [1 mark]
c) Which element is represented by the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$? [1 mark]

4. a) Explain why magnesium has a valency of +2 [2 marks]
b) Draw an electron-dot diagram of the compound magnesium sulfide [2 marks]
c) Explain why non-metals like sulfur always form negative ions when bonding with metals *Hint 2* [2 marks]
d) Draw an electron-dot diagram for the phosphide ion (P^{3-}) [2 marks]

5. a) Use an electron-dot diagram to explain why oxygen only exists as the molecule O_2. [3 marks]

b) What is meant by the term "non-bonding" pair of electrons? Give an example, Illustrated with a diagram. [3 marks]

c) Using an electron-dot diagram, show the bonding arrangement for the compound Cl_2O *Hint 3* [2 marks]

d) Illustrate the idea of a co-ordinate covalent bond, using the sulfate ion as an example. [3 marks]

Hint 1 : How does an atom become charged?
Hint 2 : Think about stable configurations
Hint 3 : Remember- each atom must have 8

Unit 2
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Atomic Structure

Key Exam Questions - Answers

1. a)

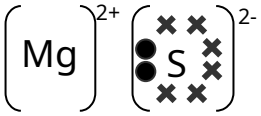
	25		26			35			48
Symbol		$_{12}Mg$		$_{12}Mg^{2+}$			$_{17}Cl$	$_{23}V$	S^{2-}
Protons	12		12		17		23		16
Neutrons	13		14		18		25		19
Electrons	12		10		17		23		18
Atomic No.	12		12		17		23		16
Mass No.	25		26		35		48		35
Net charge	0		+2		0		0		2-

[½ mark each]

2. a) Cl [1 mark]
   b) V [1 mark]
   c) S [1 mark]
   d) Mg [1 mark]

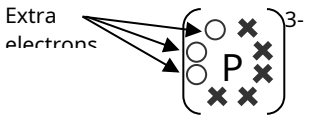
3. a) (i) Helium has the same properties as neon which has a full shell and is at the end of the next period [1 mark]. Helium shows full shell characteristics after only 2 electrons and so it too must have a full shell [1 mark].
   (ii) The gap between Ca (Gp 2) and Ga (Gp 3) is filled with the transition metals [1 mark]. This gap shows that another shell must be starting to be filled
   b) (i) Si is $1s^2 2s^2 2p^6 3s^2 3p^2$ [1 mark].
   (ii) Ti is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$ [1 mark].
   c) Potassium [1 mark].

4. a) Magnesium has 2 valence electrons. To achieve a noble gas structure it can either lose 2 or gain 6 electrons. [1 mark]. Losing 2 requires less energy and so it becomes Mg^{2+}.[1 mark].
   b)

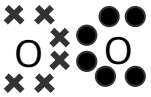


c) Non-metals are in groups 4 – 6 (right hand side of the table) which all find it easier to **gain** electrons in becoming iso-electronic with the noble gases [1 mark]. By gaining electrons atoms become negatively charged ions [1 mark].

d)



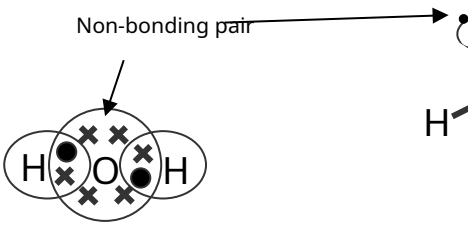
5. a) An oxygen atom needs another 2 electrons to form the stable octet[1 mark]. If it forms a double covalent bond with another oxygen atom each can achieve this. [1 mark].



[Diagram 1 mark].

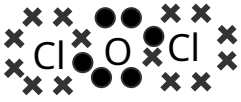
b) After forming a bond some elements have a pair, or two pairs of electrons contributing to the stable 8 that are not used for bonding. These are non-bonding pairs. [1 mark].

Example: Water



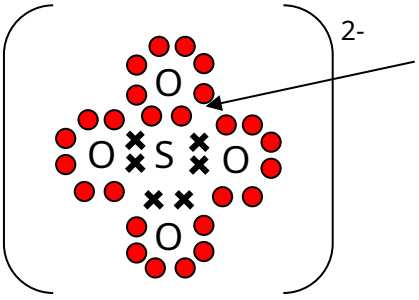
[Diagrams 2 marks].

c)



[Diagram 2 marks].

d)



[Diagram 3 marks].