edexcel

INTERNATIONAL GCSE (9-1)

Chemistry (4CH1)

First teaching 2017

TOPIC GUIDE:

Chemical bonding

Pearson Edexcel International GCSE in Chemistry (4CH1) and Science (4SD0 and 4SS0)

Pearson

Introduction

Chemical bonding (and related ideas about chemical stability/reactivity) is acknowledged as being a 'tricky to teach' topic, and with good reason. It involves abstract, theoretical ideas that require students to develop and apply increasingly sophisticated ideas in order to make sense of their observations of the macroscopic properties of different substances.

Research has shown that students commonly acquire misconceptions about chemical bonding. Some of these can be persistent and may present significant barriers to students' progression and understanding of more complex ideas in chemistry.

Chemical stability

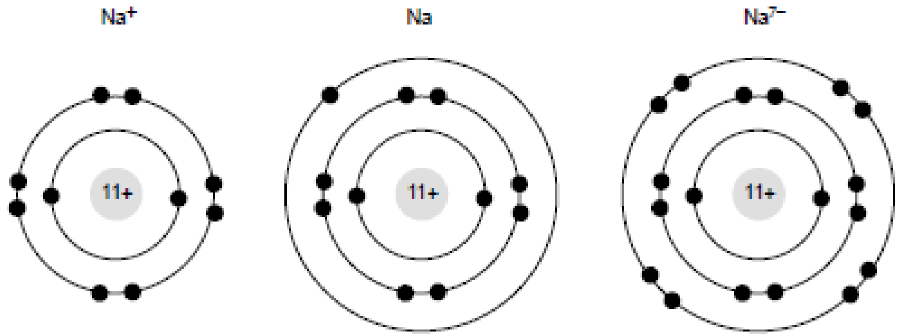
Many students studying at GCSE level, and also some studying at both A level and beyond, appear to develop a firm belief that species with full outer shells or outer shell octet configurations are more energetically stable than species with other electronic configurations. **This is very far from the truth**. The electronic configuration of 2.8 for an atom of neon is not more stable, or less stable, than the electronic configuration of 2.8.1 for a sodium atom.

One study showed that over 97% of the 15-16 year olds tested believed that a Na^{+} ion was more stable than a sodium atom. More worryingly, around 80% of students thought the Na^{7-} ion would also be more stable than the Na atom! Since energy has to be supplied to a sodium atom to either remove one electron, or to add seven electrons, the two ions are both at a higher energy level than the atom, and hence are less stable. It would appear that the octet shell of electrons dominates the thinking of many students.

You can test your own students' views by giving them the following prompt sheet:

Chemical stability (1)

The diagrams below represent three chemical species:-



Sodium ion with electronic configuration of 2.8

Sodium atom with electronic configuration of 2.8.1

Sodium ion with electronic configuration of 2.8.8

1. Tick ✓ one of the four statements:

☐ Na^+ is more stable than Na

☐ Na^+ and Na are equally stable

☐ Na^+ is less stable than Na

☐ I do not know

Why did you think this was the answer?

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2. Tick ✓ one of the four statements:

☐ Na is more stable than Na^{7-}

☐ Na and Na^{7-} are equally stable

☐ Na is less stable than Na^{7-}

☐ I do not know

Why did you think this was the answer?

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3. Tick ✓ one of the four statements:

☐ Na^{7-} is more stable than Na^+

☐ Na^{7-} and Na^+ are equally stable

☐ Na^{7-} is less stable than Na^+

☐ I do not know

Why did you think this was the answer?

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This prompt sheet is just one a several on the topic of chemical stability produced by Keith Taber for the Royal Society of Chemistry. It is available for download via the following link:

http://www.rsc.org/learn-chemistry/resource/res00001102/chemical-stability?cmpid=CMP00002081

Anthropomorphism

In answers to examination questions, students at both GCSE level and A level often suggest that atoms **need** to acquire a full outer shell, or an octet of electrons in the outer shell, and think that this is the 'driver' for chemical reactions. Statements such as 'The sodium atom wants to lose one electron' are very common. This example of anthropomorphism (i.e. giving inanimate objects human characteristics) is not justified. A sodium atom does not want to lose an electron. If it did have desires it would actually want to keep hold of all of its electrons since they are attracted by the protons in the nucleus of the atom.

Instead, sodium atoms only lose electrons when they are placed in an environment where this becomes possible, for example by heating sodium in chlorine gas, or adding sodium to water.

Four statements candidates should avoid making

1. The noble gases are unreactive because their atoms have stable electronic configurations / have a full outer shell / have eight electrons in their outer shell.
2. Chemical bonding involves transferring or sharing electrons in order to achieve the electronic structure of a noble gas.
3. A covalent bond is a shared pair of electrons.
4. Ionic bonds are formed by the transfer of electrons.

The noble gases are unreactive because their atoms do not easily lose or gain electrons. The reason these atoms do not easily lose electrons is that their ionisation energies are relatively large. The noble gases have the highest first ionisation energy of all the elements in their period. The reason the atoms of the noble gases do not easily gain electrons is that incoming electrons are repelled strongly by the electrons already present. This repulsion is so large that the first electron affinities of the noble gases are positive. The first electron affinity of most other elements is negative.

Chemical bonding is the result of electrostatic interactions between positively and negatively charged species. In the case of a covalent bond, it is the electrostatic attraction between the bonding electrons and the two nuclei of the atoms bonded together. In the case of ionic bonding, it is the electrostatic attraction between the positive and negative ions.

A suggested approach to the teaching of chemical bonding

By shifting the focus at an early stage (during their GCSE studies) onto the electrostatic nature of chemical bonding, it is possible to provide a more logical approach and a sound basis for students' progression and facility in dealing with more complex ideas.

Students are familiar with electrostatic phenomena (balloons sticking to walls, hair standing on end) and the 'invisible forces' of attraction and repulsion and can apply these ideas to help them understand chemical bonding.

Although clearly not the whole picture, the electrostatics approach can support more effective progression and help students when they progress to a higher level with related aspects, such as why some bonds are stronger; bond enthalpies; giant lattice structures versus simple molecules; polar covalent bonds and intermolecular interactions.

Formation of ions and ionic bonding

The following is taken from the specification:

Students should:

- **draw dot-and-cross diagrams to show the formation of ionic compounds by electron transfer, limited to combinations of elements from Groups 1, 2, 3 and 5, 6, 7 (*only outer electrons need be shown*)**
- **understand ionic bonding in terms of electrostatic attractions**

For those teachers who are looking for some simple rules to help their students to decide how many electrons will be involved in the formation of ions, the following suggestions may be helpful:

1. The atoms of Group 1 metals lose one electron when they form ions. The atoms of Group 2 metals lose two electrons when they form ions. The atoms of Group 3 metals lose three electrons when they form ions.

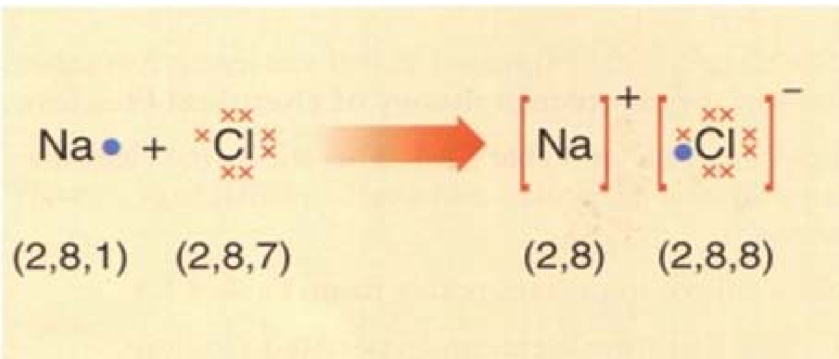
2. The atoms of Group 7 non-metals gain one electron when they form ions. The atoms of Group 6 non-metals gain two electrons when they form ions. The atoms of Group 5 non-metals gain three electrons when they form ions.

The number of electrons lost is equal to the group number of the metal. The number of electrons gained is equal to 8 – the group number.

This information could be presented in the form of a partial Periodic Table:

	Group 1	**Group 2**	**Group 3**	**Group 4**	**Group 5**	**Group 6**	**Group 7**
Period 2	Li^{+}				N^{3-}	O^{2-}	F^{-}
Period 3	Na^{+}	Mg^{2+}	Al^{3+}		P^{3-}	S^{2-}	Cl^{-}
Period 4	K^{+}	Ca^{2+}					Br^{-}
Period 5	Rb^{+}	Sr^{2+}					I^{-}
Period 6	Cs^{+}	Ba^{2+}					

Dot-and-cross diagrams to show the formation of ions from atoms can be shown, for example, as:



Ionic bonding is the electrostatic force of attraction between oppositely charged ions. In the above case it is the electrostatic force of attraction between the positive sodium ions and the negative chloride ions.

Another trap to avoid is to state that ionic bonding is always the results of metals combining with non-metals. Aluminium chloride and beryllium oxide are just two examples of compounds that are covalent.

The following is a link to an animation showing the formation of sodium chloride:

https://www.youtube.com/watch?v=lhC42qxk5kQ

It is one of the few on the internet that does not make spurious references to 'stable octets'.

Formation of simple molecules

The following is taken from the specification:

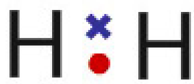
Students should:

- **know that a covalent bond is formed between atoms by the sharing of a pair of electrons**
- **understand covalent bonds in terms of electrostatic attractions**
- **understand how dot-and-cross diagrams can be used to represent covalent bonds in**
  - **i. diatomic molecules, including hydrogen, oxygen, nitrogen, halogens and hydrogen halides**
  - **ii. inorganic molecules, including water, ammonia and carbon dioxide**
  - **iii. organic molecules containing up to two carbon atoms, including methane, ethane, ethene and those containing halogen atoms**

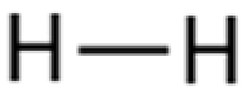
In order to draw a dot-and-cross diagram, it is first necessary to identify how many covalent bonds each atom has formed. Each covalent bond then contains a pair of electrons, one from each atom involved in the bond.

A good starting point might be to look at the diatomic molecules of the elements hydrogen, oxygen and nitrogen to introduce the concept of single, double and triple bonds.

A hydrogen atom, having only one electron, can form only one covalent bond, and this has to be a single bond. The dot-and-cross diagram for the hydrogen molecule is, therefore:



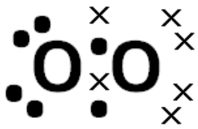
The single bond between the two atoms can be also represented by a single line between the two atoms:



The resulting representation is known as a displayed formula.

The covalent bond is then described as the electrostatic force of attraction between the two nuclei of the hydrogen atoms and the bonding pair of electrons.

An oxygen atom can form two covalent bonds and in the oxygen molecule this is a double bond, containing two sets of paired electrons:



The displayed formula for the oxygen molecule is:

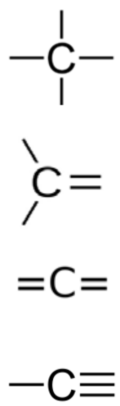
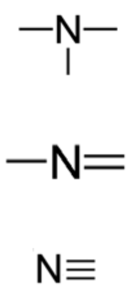
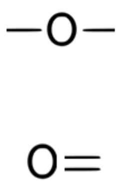
$$O{=}O$$

A nitrogen atom can form three covalent bonds and in the nitrogen molecule this is a triple bond, containing three sets of paired electrons:

The displayed formula for the nitrogen molecule is:

$$N\equiv N$$

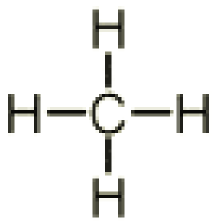
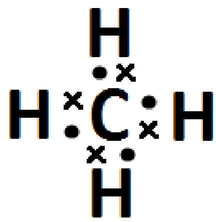
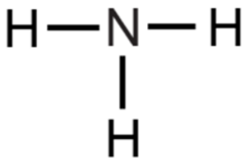
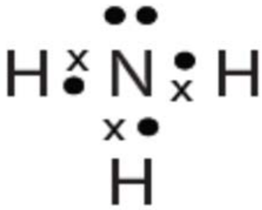
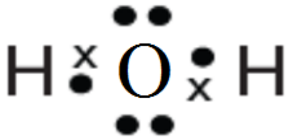
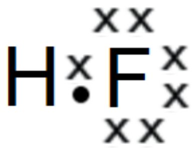
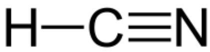
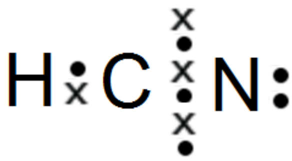
The next move could be to show the number and types bonds that can be formed by the atoms in Period 2 of the Periodic Table, limited to molecules that will be encountered at GCSE level (but not, of course, at AS/A level):

Atom	C	N	O	F
Number of bonds formed	4	3	2	1
Type(s) of bonds formed	4 single bonds 2 single bonds + 1 double bond 2 double bonds 1 single + 1 triple	3 single bonds 1 single bond + 1 double bond 1 triple bond	2 single bonds 1 double bond	1 single bond
Representation as displayed formulae				—F

Now it is possible to introduce specific molecules and match the above types of bonding to the atoms they contain to produce a displayed formula. The dot-and-cross diagram can then be deduced from the displayed formula.

It is important to recognise that, although the non-bonding electrons (lone pairs) need not be shown in a displayed formula, they **must** be shown in a dot-and-cross diagram.

Examples:

Molecule	Displayed formula	Dot-and-cross diagram
CH_4		
NH_3		
H_2O		
HF		
CO_2		
HCN		

You will have noticed that the number of electrons in the outer shell of each of the atoms of the elements C, N, O and F is eight. This is sometimes referred as the 'octet rule'. It is very important that students appreciate that this octet rule applies only to these four elements in Period 2. Both beryllium and boron form simple molecules and they do not obey this rule, as we shall see later.

It is also equally important that students do not gain the impression that these atoms bond with other atoms in order to acquire an octet of electrons in the outer shell of their atoms. This is **not** the reason for bonding taking place, and if students do linger under the misconception then it may, as has been mentioned earlier, hinder their progress and understanding of more complex concepts studied at a higher level.

The ideas learned from the bonding of hydrogen and Period 2 elements can now be extrapolated to other elements.

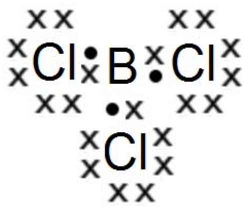
The halogens chlorine, bromine and iodine are capable of forming more than one bond, since the maximum number of electrons that can be accommodated in the outer shell of their atoms is greater than eight. However, at GCSE level the molecules met will be confined to those in which they form only one bond, a single covalent bond.

Phosphorus in Group 5 can, of course, form more than three covalent bonds, but again it is probably wise with the majority of students to confine molecules of phosphorus to those in which it forms only three single bonds. Similarly for sulfur, forming only two single covalent bonds and silicon forming only four single covalent bonds. Hence, students should now be able to draw dot-and-cross diagrams for molecules such as:

PH_3 SCl_2 $SiCl_4$ and $PbBr_4$

You will note that the latter is a compound of a metal and a non-metal. It is important that students are not led (please excuse the pun) to believe that covalent bonds can only be formed between two non-metals.

Returning to beryllium and boron from Period 2, there is no reason why students at GCSE level should not be able to construct a dot-and-cross for molecules such as $BeCl_2$ and BCl_3. It is a simple matter to determine the number of electrons in the outer shell of each atom involved and hence determine the number of covalent bonds that can be formed. In the case of the beryllium atom it is two, and for boron it is three. Hence the dot-and-cross diagrams will be:



Finally, teachers may like to extend their brightest students by introducing them to molecules such as PCl_5 and SF_6 - or even to nitric acid or sulfuric acid - even though this is beyond the scope of what is expected at GCSE level. If students are told that the maximum number of electrons that can be accommodated in the third shell is 18, then they will have no problem in producing diagrams that have 10 electrons in the outer shell of a phosphorous or nitrogen atom and 12 electrons in the outer shell of a sulfur atom.

Metallic bonding

Once again, it is important for students to appreciate that the bonding involved is electrostatic in nature.

This time it is the electrostatic force of attraction between the **nuclei** of the **atoms** and the **delocalised electrons**.

The first three minutes of this video/animation is useful in illustrating metallic bonding. It avoids the over-simplified model of metal cations attracted to delocalised electrons.

https://www.youtube.com/watch?v=eVv3TpaQ2-A

Another link worth viewing is:

https://www.pbslearningmedia.org/resource/lsps07.sci.phys.matter.chembonds/chemical-bonds/#.WgS_bsK7KUk