Chapter 3

Chemical bonding and structure

Worksheet: Showing bonding with dot-and-cross diagrams

Helpsheet 1: Rules for dot-and-cross diagrams (bonding diagrams)

Helpsheet 2: Pure ionic bonding, pure covalent bonding and in-between

Practical: Physical properties of simple molecular, giant ionic, giant covalent and giant metallic substances

End-of-chapter test

Marking scheme: Worksheet

Marking scheme: End-of-chapter test
Worksheet

Showing bonding with dot-and-cross diagrams [5.1.3b,e]

Helpsheet 1 can be referred to if you find it useful.

1. Draw dot-and-cross diagrams to describe how ionic bonds are formed in the following compounds:
   a. sodium chloride
   b. magnesium oxide [4]

2. Draw dot-and-cross diagrams to describe the covalent bonding in the following elements:
   a. hydrogen
   b. oxygen
   c. chlorine [3]

3. Draw dot-and-cross diagrams to describe the covalent bonding in the following compounds:
   a. hydrogen chloride
   b. water
   c. ammonia
   d. methane
   e. carbon dioxide
   f. ethene
   g. the ammonium ion [7]

4. Describe the difference between:
   a. the bond between the carbon atoms in ethene and the C–H bonds in ethene
   b. one of the N–H bonds in the ammonium ion and the other three N–H bonds. [4]

Questions 1, 2 and 3, above, all involve substances specified in the syllabus. You should therefore ensure you can do them successfully before moving on to the questions below.

5. Draw dot-and-cross diagrams to describe the bonding in the following substances. Your first decision must be whether the bonding is ionic or covalent.
   a. hydrogen sulphide
   b. nitrogen
   c. sodium oxide
   d. fluorine
   e. magnesium chloride
   f. aluminium oxide
   g. ethyne (formula C₂H₂)
   h. potassium fluoride
   i. calcium fluoride
   j. phosphine (PH₃) [15]

6. Identify the two substances in Question 5 that have a triple covalent bond in their molecules. [2]

Total: / 35 Score: %
Helpsheet 1

Rules for dot-and-cross diagrams (bonding diagrams) [5.1.3e]

1. Draw all of the atoms as they are before bonding. Show all the outer electrons, but only show the outer electrons.

2. Every electron you have just shown must be in your final diagram. Don’t lose any of them and don’t introduce any new ones.

3. If the elements involved are one metal and one non-metal the bonding will be ionic.
   a. This means each metal atom loses all its outer electrons and becomes a positive ion, with a 1+, 2+ or 3+ charge.
   b. At the same time each non-metal atom gains enough electrons to fill all the gaps in its outer shell and becomes a negative ion, with a 1–, 2– or 3– charge.

4. If the elements involved are all non-metals the bonding will be covalent.
   a. This means each non-metal atom fills all gaps in its outer shell by sharing electrons.
   b. If two atoms share two electrons between them this is called a single covalent bond.
   c. If two atoms share four electrons between them this is called a double covalent bond.
   d. If two atoms share six electrons between them this is called a triple covalent bond.
Helpsheet 2

Pure ionic bonding, pure covalent bonding and in-between [5.1.3h]

1 If two identical non-metal atoms are covalently bonded (e.g. two chlorine atoms in Cl₂) the bonding electrons are shared equally. This is pure covalent bonding.

2 If two non-identical non-metal atoms are covalently bonded (e.g. hydrogen and chlorine in H–Cl) the bonding electrons are likely to be shared unequally. This is because one atom (Cl in this case) is more electronegative than the other.
   a The more electronegative atom gets the ‘lion’s share’ of the bonding electrons.
   b The less electronegative atom is said to become ‘electron deficient’.
   c The bond formed is said to be polar.

3 If a metal atom and a non-metal atom are ionically bonded the electrons involved should be completely transferred to the non-metal atom. This is a description of pure ionic bonding.

4 However, if the metal ion is small or highly charged (e.g. Al³⁺) it attracts back the electrons it has lost to the non-metal atom. The negative ion is polarised.
   a Bigger negative ions are easier to polarise than small ones.
   b Example 1: in KF the F⁻ ion is not significantly polarised as K⁺ is a large ion and F⁻ is a small one.
   c Example 2: in LiBr the Br⁻ ion is significantly polarised as Li⁺ is a small ion and Br⁻ is a large one.
Practical

Physical properties of simple molecular, giant ionic, giant covalent and giant metallic substances [5.1.3p]

Safety
Always refer to the departmental risk assessment before carrying out any practical work. See the Part 1 Notes on practical work, found in the Additional resources section, for additional guidance and Hazcard references. Wear eye protection at all times.

Apparatus per student

- Wax
- Silicon dioxide
- Sodium chloride
- 4 small pieces of zinc sheet
- 3 spatulas
- Crucible
- Bunsen burner, heatproof mat and tongs
- 12 dry Pyrex test-tubes
- 8 stoppers to fit test-tubes
- Test-tube rack

- Power pack
- 12V bulb
- 3 leads
- 2 crocodile clips
- 2 graphite rods
- Cyclohexane
- 100 cm$^3$ beaker
- Bottle of distilled water
- Eye protection

Procedure

1. The four materials you are going to investigate are wax, silicon dioxide, sodium chloride and zinc sheet. Decide which of these has a simple molecular structure, which is giant ionic, which is giant metallic and which is giant covalent.

2. Place a small sample of each in a dry Pyrex test-tube and heat strongly. Make a simple qualitative statement about the melting point of each substance.

3. Place a small sample of each substance in turn in a dry crucible and test its electrical conductivity. Make a simple qualitative statement about the electrical conductivity of each substance.

4. Place a small sample of each substance in turn in a dry Pyrex test-tube, add water, stopper the tube and shake it. Make a simple qualitative statement about the solubility in water of each substance.

5. Place a small sample of each in a dry Pyrex test-tube, add cyclohexane, stopper the tube and shake it. Make a simple qualitative statement about the solubility in cyclohexane of each substance.

6. Write up the practical. Include in your write-up a results table with the structure of each of the four materials as well as their names.
End-of-chapter test

1. For each of the following substances:
   i. name the type of structure and the type of bonding,
   ii. state whether or not it would conduct electricity when solid,
   iii. state whether or not it would conduct electricity when molten,
   iv. estimate its melting and boiling points as either high or low,
   v. if it is a compound illustrate how the bonding arises with a dot and-cross diagram.
   
   a. hydrogen chloride  
   b. diamond  
   c. sodium chloride  
   d. magnesium oxide  
   e. ethene  
   f. copper  
   g. graphite  
   h. carbon dioxide

2. Explain the meaning of the following terms using a suitable example for each and a diagram if you wish:
   a. ionic bonding  
   b. covalent bond  
   c. dative covalent bond  
   d. lattice structure  
   e. metallic bonding  
   f. electronegativity  
   g. covalent bond polarity  
   h. van der Waals’ forces

3. a. i. Describe the shapes of BF₃, H₂O, CH₄ and NH₃ using words and a diagram for each. Include relevant bond angles.
   ii. Explain why BF₃ and NH₃ have different shapes despite their apparently similar formulae.
   
   b. What would you expect to be the shapes of silane, SiH₄, and phosphine, PH₃?

4. Explain with the aid of a diagram how hydrogen bonding arises between two water molecules.

5. Ammonia, NH₃, hydrogen chloride, HCl, and argon, Ar, are all gases at room temperature but they can all be liquefied if they are cooled sufficiently. Their boiling points are: NH₃, −33 °C; HCl, −88 °C; Ar, −186 °C
   a. Which of these three substances has the greatest intermolecular forces, and which has the least?
   b. Name the most important type of intermolecular force possessed by each substance.

6. Substances that possess ‘pure’ ionic bonding or ‘pure’ covalent bonding are rare. Although most substances can be classified as ionic or covalent, the actual bonding normally involves some of the characteristics of the other type.
   a. Explain what the statement above means. As examples you should use magnesium chloride, which is classified as ionic, and water, which is covalent.
   b. Explain why the element chlorine can be thought of as an example of ‘pure’ covalent bonding.

7. Water possesses a number of strange properties. For example, solid water floats on liquid water, and the melting and boiling points of water seem very high when compared with substances of similar molecular size.
   a. Why does ice have a lower density than water?
   b. Why does water possess such unusually high melting and boiling points for a simple molecular substance with small molecular size?

Total: / 80 Score: %
1 a Before bonding

![Na] ![Cl]

After bonding

![Na⁺] ![Cl⁻]

[1]

1 b Before bonding

![Mg] ![O]

After bonding

![Mg²⁺] ![O²⁻]

[1]

2 a

![H] ![H] ![H] ![H]

[1]

2 b

![Cl] ![Cl]

[1]

3 a

![H] ![Cl]

[1]

3 b

![H] ![H] ![H] ![H]

[1]

3 c

![H] ![N] ![H] ![H]

[1]

3 d

![H] ![C] ![H]

[1]

4 a The bond between the carbon atoms in ethene consists of four electrons, not two as in the C–H bonds [1], it is called a double bond [1]

4 b Both electrons in the ‘different’ bond have come from the nitrogen atom [1], it is called a dative or coordinate bond [1] (although, in practice, all four bonds are indistinguishable)
6  Nitrogen [1] and ethyne [1]
### Marking scheme

#### End-of-chapter test

1

<table>
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<th></th>
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<th>ii &amp; iii</th>
<th>iv</th>
<th>v</th>
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<td>HCl</td>
<td>simple covalent</td>
<td>no, no</td>
<td>low, low</td>
</tr>
<tr>
<td>b</td>
<td>diamond</td>
<td>giant covalent</td>
<td>no, no</td>
<td>high, high</td>
</tr>
<tr>
<td>c</td>
<td>NaCl</td>
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<tr>
<td>d</td>
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<td>no, yes</td>
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<tr>
<td>h</td>
<td>CO₂</td>
<td>simple covalent</td>
<td>no, no</td>
<td>low, low</td>
</tr>
</tbody>
</table>

[1] per answer  
(total = [29])

2  

a  The electrostatic attraction between oppositely charged ions [1];  
iions are formed by the transfer of electrons from one atom to another [1];  
b  The sharing of a pair of electrons between two atoms [1];  
onelectron comes from each atom [1];  
c  As a covalent bond, but both electrons from same atom [1];  
d  A regular repeating pattern [1];  
e  Positive ions [1] embedded in a sea of delocalised electrons [1];  
f  Tendency of an atom to attract the electrons in a covalent bond [1];  
g  Unequal sharing of electrons in covalent bond [1];  
h  An instantaneous dipole in one molecule [1] causes an induced dipole in neighbouring molecule [1]
3 a i BF₃: trigonal planar [1].

\[
\begin{array}{c}
F \\
B \\
F
\end{array}
\] [1], 120° [1];

H₂O: non-linear [1].

\[
\begin{array}{c}
H \\
O \\
H
\end{array}
\] [1], 104.5° [1];

CH₄: tetrahedral [1].

\[
\begin{array}{c}
H \hspace{1cm} \hspace{1cm} \hspace{1cm} \hspace{1cm} H \\
C \\
H
\end{array}
\] [1], 109.5° [1];

NH₃: trigonal pyramidal [1].

\[
\begin{array}{c}
H \\
N \hspace{1cm} \hspace{1cm} \hspace{1cm} \hspace{1cm} H
\end{array}
\] [1], 107° [1];

ii BF₃ has no lone pair [1]. NH₃ has a lone pair [1];

b SiH₄: tetrahedral [1], PH₃: trigonal pyramidal [1]

4 Water molecules drawn non-linear [1], dipole shown [1], O is minus end of dipole, H is plus end [1], H bond marked as dotted line from H of one molecule to O of another [1], lone pair shown on O as start/end of H bond [1], straight line from O on first molecule, through H bond to H atom and covalent bond to O atom on second molecule: i.e. \( \text{O} \sim \text{H} - \text{O} \) is linear [1]

5 a Ammonia greatest [1], argon least [1];

b Ammonia has H bonds [1], HCl has permanent dipole/permanent dipole forces [1], argon has van der Waals’ forces [1]

6 a Sometimes the positive ion will attract some of the charge from the negative ion [1]. This is called polarisation [1]. Use of example [1]. The more electronegative element attracts the two electrons in the covalent bond [1]. This is called polarisation or the creation of a dipole [1]. Use of example [1];

b Both the covalently bonded atoms are the same, therefore they have same electronegativity [1]. There will be totally equal sharing of bonding electrons, so the bond is not polarised [1]

7 a The molecules on average further apart in solid [1], due to rigid network of hydrogen bonds [1];

b Hydrogen bonding [1], average of two per molecule [1]