### National 5 Chemistry

#### Unit 1:

# Chemical Changes & Structure

### Topic 3

**Bonding & Bonding Structures**

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<td>Consolidation C</td>
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<td>Score: /</td>
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</tr>
</tbody>
</table>

**End-of-Topic Assessment**

Score: %

Grade:
3.1 Bonding

Noble gases like neon are unusual because they do not seem to need to form bonds. They do not react. Their valencies are zero.

The noble gases also have a full outer shell and this seems to be a very stable electron arrangement.

**Ionic Bonding**

Sodium is element 11 and very close to neon in the Periodic Table.

Sodium atoms cannot change the number of protons in their nuclei, but they can lose one electron to have the same stable electron arrangement as a neon atom.

The sodium ion formed has the same nucleus as the sodium atom, but has the same electron arrangement as neon; the nearest noble gas.

Chlorine is element 17 and very close to argon in the Periodic Table.

Chlorine atoms also cannot change the number of protons in their nuclei, but they can gain 1 electron to have the same stable electron arrangement as an argon atom.

The chloride ion formed has the same nucleus as the chlorine atom, but has the same electron arrangement as argon; the nearest noble gas.

The sodium can only lose an electron because the chlorine is willing to gain the electron.

An electron is transferred.

The ionic bond is the attraction between the ions of opposite charge.
The Ionic Bond is the mutual attraction between positive and negative ions.

Though 1 sodium atom will give 1 electron to 1 chlorine atom and the formula for sodium chloride will be NaCl, each Na\(^+\) ion will attract several Cl\(^-\) ions and vice versa, hence a network structure.
**Covalent Bonding**

In *covalent compounds*, both the *elements* involved are usually *non-metal*.

*Non-metal* atoms prefer to *gain electrons*.

For *both* atoms to *gain extra electrons*, the atoms have to *share electrons*.

![Diagrams of covalent bonding examples](image)

Atoms *overlap* shells in order to *share electrons*.

Both atoms achieve a stable electron arrangement (a *full outer shell*).

![Electron orbitals](image)

Each shared pair of electrons is a *covalent bond*. When necessary atoms can share more than one pair and form *double* or even *triple* covalent bonds.

**Examples:**

- **O—C—O**
- **H—C≡N**
A molecule of methane has four hydrogen atoms joined to one carbon atom.

‘Dots and crosses’ can be used to stand for the outermost electrons in both types of atoms.

All the electrons are now paired up.

In cov compounds atoms join together by sha electrons.

Only the ou electrons are involved.

Sha allows odd electrons from different atoms to pair up.

Cov compounds usually only involves n -metal atoms.

The sha pair of electrons hold the atoms together.

Molecular Shapes

The shapes of some simple molecules need to be known and understood. Their shapes are all based on the need for the 4 orbitals, found in the outer shell of many atoms, to remain as far apart as possible to minimise repulsions. This 3-dimensional arrangement is tetrahedral.

Methane, CH₄

Carbon atoms have single electrons in all 4 orbitals available to form bonding pairs.

The 4 orbitals will be arranged tetrahedrally and with 4 hydrogen atoms overlapping with each orbital, the molecular shape is also described as tetrahedral.
Ammonia, \( \text{NH}_3 \)

Nitrogen atoms have \textit{single electrons} in 3 of the 4 \textit{orbitals} available to form \textit{bonding pairs}. The 4th orbital is full - these form a \textit{lone pair} or \textit{non-bonding pair}.

The \textit{4 orbitals} will be \textit{arranged tetrahedrally} but with \textit{only 3 hydrogen} atoms overlapping with 3 \textit{orbitals}, the \textit{molecular shape} will be different and is described as \textit{trigonal pyramidal}.

Water, \( \text{H}_2\text{O} \)

Oxygen atoms have \textit{single electrons} in \textit{2 of the 4 orbitals} available to form \textit{bonding pairs}. The other 2 orbitals are full - these form \textit{lone pairs} or \textit{non-bonding pairs}.

The \textit{4 orbitals} will be \textit{arranged tetrahedrally} and with \textit{only 2 hydrogen} atoms overlapping with 2 \textit{orbitals}, the \textit{molecular shape} will be different and is described as \textit{bent} or \textit{v-shaped}.
Metallic Bonding

Here, all the atoms want to lose electrons but none are prepared to gain electrons. At first sight there is no way that this can happen.

However, the next best thing is to temporarily ‘lose’ the outer electron(s) by allowing them to drift freely between all the separate metal atoms.

This results in temporary metal ‘ions’ forming which immediately attract an electron back to reform the atom.

The outer electrons end up ‘belonging’ to more than one atom;

the metal atoms are bonded together when they attract the same electron at the same time.

A metallic structure can be described as a regular arrangement (network or lattice) of positive metal ions held together by a ‘sea’ of constantly moving negative electrons. The outer electrons are said to be ‘delocalised’.

Since all metal atoms bond in this way, introducing atoms of a different metal does not disturb the structure at all. This is why mixtures of metals (alloys) are usually very stable and would be very difficult to separate again.
Q1. Int2
Which of the following pairs of elements combine to form an ionic compound?

A  Lead and fluorine
B  Sulphur and oxygen
C  Carbon and nitrogen
D  Phosphorus and chlorine

Q2. SG
Identify the covalent compound

A  zinc chloride
B  magnesium sulphate
C  lead carbonate
D  hydrogen sulphide

Q3. Int2
Metallic bonds are due to

A  pairs of electrons being shared equally between atoms
B  pairs of electrons being shared unequally between atoms
C  the attraction of oppositely charged ions for each other
D  the attraction of positively charged ions for delocalised electrons.

Q4. Int2
Atoms of an element form ions with a single positive charge and an electron arrangement of 2,8.

The element is

A  fluorine
B  lithium
C  sodium
D  neon

Q5. Int2
The table shows information about an ion.

<table>
<thead>
<tr>
<th>Particle</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>protons</td>
<td>19</td>
</tr>
<tr>
<td>neutrons</td>
<td>20</td>
</tr>
<tr>
<td>electrons</td>
<td>18</td>
</tr>
</tbody>
</table>

The charge on the ion is

A  1+
B  1-
C  2+
D  2-

Q6. SC
The shapes and names of some molecules are shown below.

The shapes are tetrahedral, pyramidal, bent, and linear.

Phosphine is a compound of phosphorus and hydrogen. The formula is PH₃. The shape of a molecule of phosphine is likely to be

A  tetrahedral
B  pyramidal
C  bent
D  linear

Q7. Int2
Carbon forms many compounds with other elements such as hydrogen.

a) Draw a diagram to show how the outer electrons are arranged in a molecule of methane, CH₄.

b) Draw a diagram to show the shape of a molecule of methane, CH₄.

c) Identify the two elements which react together to form a molecule with the same shape as a methane molecule.

A  H
B  N
C  Si
D  Al
E  Mg
F  O
3.2 Bonding Structures

**Single Atoms**

At *room temperature*, about 20 °C, there are only really 5 substances that we consider as having *single atom* structure - sometimes referred to as *monatomic*.

These are the *Noble Gases*; He, Ne, Ar, Kr and Xe

They have extremely weak attractions between the atoms so have very low densities - used in *balloons* and would *distort your voice box*. However, as the atoms get bigger, the attractions increase so, by the time we reach Xenon, it is dense enough to 'pour' and allow an *aluminium* foil boat to float.

If electrically 'excited', these atoms release light with characteristic *colours* making them suitable for *strip lights* and *advertising signs*.

**Covalent Molecular**

This is undoubtedly the *largest* and most *diverse* grouping.

Though they can have extremely weak attractions *between* the molecules and form very low density *gases* - they can also have strong enough attractions to form *liquids* and *solids* as well.

Whilst most are *compounds*, there are a reasonable number of *elements* with a *covalent molecular* structure. *For example*;

<table>
<thead>
<tr>
<th>Covalent Molecular Elements</th>
<th>Covalent Molecular Compounds</th>
</tr>
</thead>
</table>
By contrast, this is a much smaller group and you are only likely to meet 6 examples.

**Diamond**

In this form each *carbon* atom uses *all 4* of its *electrons* to form *4 bonds* to *4 different carbon* atoms.

The *pattern* can be described in two ways: it is called *tetrahedral* because the *4 carbon* atoms lie at the *corners* of a *pyramid* or *tetrahedron*, but you could also call it *hexagonal* as there are *rings* of *6* carbons.

**Graphite**

In this form each *carbon* atom uses *only 4* of its *electrons* to form *3 bonds* to *4 different carbon* atoms.

This produces *flat sheets* of *carbon* atoms joined in *rings* of *6, hexagons*. The *fourth* electrons are *free to move* within the *sheet* and produce *very weak attractions* between the sheets.

As well as *carbon*, there are two other *elements* that have a *covalent network* structure -

**Boron** and **Silicon**

There are two main *compounds* that have a *covalent network* structure -

**Silicon Dioxide** and **Silicon Carbide**
Ionic Networks are quite straightforward, if we stick to the normal 'rule of thumb';

**metal / non-metal** ➞ **Ionic Compound** ➞ **Ionic Network**

### Ionic Networks

**Examples include**:

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cu$^{2+}$ SO$_4^{2-}$ (s)</td>
<td>copper (II) sulphate</td>
</tr>
<tr>
<td>(K$^+$)$_2$ CO$_3^{2-}$ (s)</td>
<td>potassium carbonate</td>
</tr>
<tr>
<td>Ni$^{2+}$ (I$^-$)$_2$ (s)</td>
<td>nickel (II) iodide</td>
</tr>
<tr>
<td>Na$^+$ Cl$^-$ (s)</td>
<td>sodium chloride</td>
</tr>
<tr>
<td>(K$^+$)$_2$ O$_2^-$ (s)</td>
<td>potassium oxide</td>
</tr>
<tr>
<td>Mg$^{2+}$ ( I$^-$)$_2$ (s)</td>
<td>magnesium iodide</td>
</tr>
<tr>
<td>NH$_4^+$NO$_3^-$ (s)</td>
<td>ammonium nitrate</td>
</tr>
<tr>
<td>H$^+$Cl$^-$ (aq)</td>
<td>hydrochloric acid</td>
</tr>
</tbody>
</table>

However, there are **ionic networks** which do not contain any metal ions and **ionic compounds** that only exist in solutions so **don't form networks**.

Subtle differences in the **arrangements of ions** with different **Ionic Networks** need not concern us.

Our 'rule of thumb' works best with **metals** well over to the **left** (in Periodic Table) and with **non-metals** well over to the **right**.

**In the end, it is the properties that will confirm structure, not a 'rule of thumb'**.
Metallic Networks are also quite straightforward, with little difference between metal elements and mixtures of metals called alloys.

Though they have many properties in common there are enough differences to ensure that metals have a wide variety of uses.

<table>
<thead>
<tr>
<th>Name</th>
<th>Elements</th>
<th>Properties</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mercury</td>
<td>Hg</td>
<td>liquid (low Mpt)</td>
<td>thermometers</td>
</tr>
<tr>
<td></td>
<td></td>
<td>very dense</td>
<td>barometers</td>
</tr>
<tr>
<td></td>
<td></td>
<td>poisonous</td>
<td>smoothing felt (hats)</td>
</tr>
<tr>
<td>Gold</td>
<td>Au</td>
<td>melts (low MPt)</td>
<td>coins</td>
</tr>
<tr>
<td></td>
<td></td>
<td>easily shaped (soft)</td>
<td>jewellery etc</td>
</tr>
<tr>
<td></td>
<td></td>
<td>corrosion resistant</td>
<td></td>
</tr>
<tr>
<td>Tungsten</td>
<td>W</td>
<td>hard</td>
<td>filaments in light bulbs</td>
</tr>
<tr>
<td></td>
<td></td>
<td>heavy</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>very high BPt</td>
<td></td>
</tr>
<tr>
<td>Steel</td>
<td>Fe , Cr ,</td>
<td>corrosion resistant</td>
<td>various types of</td>
</tr>
<tr>
<td>(alloy)</td>
<td>Ni , Mo</td>
<td>unreactive</td>
<td>stainless steel</td>
</tr>
<tr>
<td></td>
<td></td>
<td>strong</td>
<td></td>
</tr>
<tr>
<td>Alkali</td>
<td>K , Na , Li</td>
<td>very reactive</td>
<td>having fun at school!</td>
</tr>
<tr>
<td>Metals</td>
<td></td>
<td>low MPt</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>low density</td>
<td></td>
</tr>
<tr>
<td>Nitinol</td>
<td>Ni , Sn</td>
<td>shape memory</td>
<td>arterial stents</td>
</tr>
<tr>
<td>(alloy)</td>
<td></td>
<td>superelasticity</td>
<td>artificial tendons</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>self-adjusting clothing</td>
</tr>
</tbody>
</table>
SUMMARY

**Single Atoms**
*(monatomic elements)*

Very rare - only the *Noble Gases* exist as single atoms

He Ne Ar Kr Xe

**Covalent Molecular**
*(non-metal elements & compounds)*

Small - \( \text{H}_2 \text{O, N}_2 \text{ etc} \)  
\( \text{HCl, H}_2 \text{O, NH}_3, \text{CH}_4 \)

Medium - \( \text{C}_6 \text{H}_12 \text{O}_6, \text{C}_6 \text{H}_{14} \text{, C}_{60} \)

Large - starch, polythene etc

**Covalent Network**
*(non-metal elements & compounds)*

Only a few examples

the *element carbon* (diamond)

the *element carbon* (graphite)

the *compound SiO\(_2\*)

**Ionic Network**
*(metal/non-metal compound)*

Almost all metal/non-metal compounds

Na\(^+\) Cl\(^-\)  
Mg\(^{2+}\) O\(^2-\)  
Cu\(^{2+}\) \((\text{NO}_3\_)_2\)

**Metallic Network**
*(metal elements & alloys)*

All metals elements & metal alloys

Cu Na Mg Fe brass
Q1. Int 2

An element, X, has the following properties.
- It is a gas.
- It is not made up of molecules.
- It does not react with other elements.

Element, X, is likely to be in group

A 0  
B 1  
C 2  
D 7

Q2. Int 2

Which of the following diagrams could be used to represent the structure of sodium chloride?

A  
B  
C  
D

Q3. Int 2

A section of a covalent network is shown below

Write the formula for this covalent network compound ______________

Q4. Int 2

a) To which family of metals does copper belong?
   
   (You may wish to use page 8 of the data booklet to help you)

b) Copper can be used to make other metals such as brass and bronze.

What term is used to describe metals such as brass and bronze?

Q5. Int 2

Which of the following diagrams could be used to represent the structure of sodium chloride?

A  
B  
C  
D

Q6. Int 2

The element carbon can exist in the form of diamond. The structure of diamond is shown in the diagram.

a) Name the type of bonding and structure present in diamond.
3.3 Melting Points

To be of any value our Bonding Theory must be able to explain the properties observed for a substance.

**Melting Point** tells us the *temperature* at which a substance changes state from

\[
solid \quad \iff \quad liquid
\]

**Boiling Point** the *temperature* at which it changes from

\[
liquid \quad \iff \quad gas.
\]

Both these changes in state will probably involve a change in *structure* but may also involve the breaking of *bonds*.

### Single Atoms

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass</th>
<th>Mpt (°C)</th>
<th>Bpt (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ne</td>
<td></td>
<td>-189</td>
<td>-186</td>
</tr>
<tr>
<td>Kr</td>
<td></td>
<td>-112</td>
<td>-107</td>
</tr>
</tbody>
</table>

**Conclusion:** The *noble gases* have extremely low melting and boiling points :- are *all* gases at room temperature.

**Explanation:** Since there are *no* formal bonds *between* atoms at all, it requires very little *energy* to move them *faster* and *further* apart. Only very weak attractions.

### Covalent Molecular

**Gas** :- molecules moving extremely fast, total freedom to travel anywhere

**Liquid** :- molecules moving faster, free to change position within body of liquid

**Solid** :- molecules vibrate slowly in fixed position
### Elements

<table>
<thead>
<tr>
<th>Name</th>
<th>Bromine</th>
<th>Hydrogen</th>
<th>Iodine</th>
<th>Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Formula</td>
<td>Br₂</td>
<td>H₂</td>
<td>I₂</td>
<td>O₂</td>
</tr>
<tr>
<td>Molecule</td>
<td><img src="image" alt="Bromine Molecule" /></td>
<td><img src="image" alt="Hydrogen Molecule" /></td>
<td><img src="image" alt="Iodine Molecule" /></td>
<td><img src="image" alt="Oxygen Molecule" /></td>
</tr>
<tr>
<td>Melting Pt (°C)</td>
<td>-7</td>
<td>0</td>
<td>150</td>
<td>200</td>
</tr>
<tr>
<td>Boiling Pt (°C)</td>
<td>59</td>
<td>100</td>
<td>-150</td>
<td>-250</td>
</tr>
<tr>
<td>State at room T</td>
<td>liquid</td>
<td>room T</td>
<td>room T</td>
<td>room T</td>
</tr>
</tbody>
</table>

### Compounds

<table>
<thead>
<tr>
<th>Name</th>
<th>Ammonia</th>
<th>Ethanol</th>
<th>Methane</th>
<th>Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Formula</td>
<td>NH₃</td>
<td>C₂H₅OH</td>
<td>C₂H₆</td>
<td>H₂O</td>
</tr>
<tr>
<td>Molecule</td>
<td><img src="image" alt="Ammonia Molecule" /></td>
<td><img src="image" alt="Ethanol Molecule" /></td>
<td><img src="image" alt="Methane Molecule" /></td>
<td><img src="image" alt="Water Molecule" /></td>
</tr>
<tr>
<td>Melting Pt (°C)</td>
<td>-78</td>
<td>-33</td>
<td>-70</td>
<td>-100</td>
</tr>
<tr>
<td>Boiling Pt (°C)</td>
<td>-33</td>
<td>-100</td>
<td>-150</td>
<td>-200</td>
</tr>
<tr>
<td>State at room T</td>
<td>gas</td>
<td>gas</td>
<td>room T</td>
<td>room T</td>
</tr>
</tbody>
</table>

**Conclusion:** Covalent molecules, depending on mass, are gases or liquids mostly. Any solids are usually easily melted (Low Mpt)

**Explanation:** Since there is no need to actually break the strong bonds within the molecule and forces between molecules are usually weak, it requires very little energy to move them faster and further apart.
**Conclusion:** All *networks*, except the metal *mercury*, are *solids* at room temperature. Most have *very high* Mpt and Bpts.

**Explanation:** Since there is a need to *break* the *strong bonds* between the particles making up the *network*, it requires *large* amounts of *energy* to move them *faster* and *further* apart.
3.4 Conductivity

Metallic Network

<table>
<thead>
<tr>
<th>Name</th>
<th>Element or Alloy</th>
<th>State</th>
<th>Appearance</th>
<th>Conducts?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper</td>
<td>element</td>
<td>solid</td>
<td>brown shiny</td>
<td>yes</td>
</tr>
<tr>
<td>Duraluminum</td>
<td>alloy of Al &amp; Cu</td>
<td>solid</td>
<td>silvery shiny</td>
<td>yes</td>
</tr>
<tr>
<td>Mercury</td>
<td>element</td>
<td>liquid</td>
<td>silvery shiny</td>
<td>yes</td>
</tr>
<tr>
<td>Solder</td>
<td>alloy of Pb &amp; Sn</td>
<td>solid</td>
<td>silvery shiny</td>
<td>yes</td>
</tr>
<tr>
<td>Magnesium</td>
<td>element</td>
<td>solid</td>
<td>silvery shiny</td>
<td>yes</td>
</tr>
<tr>
<td>Zinc</td>
<td>element</td>
<td>solid</td>
<td>grey shiny</td>
<td>yes</td>
</tr>
</tbody>
</table>

**Conclusion:** All metals are good *conductors* when *solid* and when *liquid*.

**Explanation:** As previously seen, the metallic bond has *delocalised* electrons making metals ideal as *conductors*;

When a *voltage* (push) is applied across the metal, by a *battery* or power supply, all the *delocalised electrons* move in the same direction; attracted towards the *positive* end of the battery.
An electric current is simply a flow of negatively charged electrons, and metals complete the circuit by allowing a current to flow easily through them. The metal is completely unchanged and when the voltage is switched off the electrons will revert to drifting freely in all directions throughout the metal.

In fact, all parts of an electric circuit; bulb, leads, crocodile clips etc, will be made of metal and are just a reservoir of delocalised electrons which the battery forces to move in a particular direction.

The battery is like an ‘electron pump’, that pulls electrons in at the positive terminal and pumps them out at the negative terminal.

### Covalent Network & Molecular

<table>
<thead>
<tr>
<th>Name</th>
<th>Network or Molecule</th>
<th>Element or Compound</th>
<th>State</th>
<th>Appearance</th>
<th>Conducts?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Silicon</td>
<td>network</td>
<td>element</td>
<td>solid</td>
<td>shiny dark crystalline</td>
<td>no</td>
</tr>
<tr>
<td>Sulphur</td>
<td>molecular</td>
<td>element</td>
<td>solid</td>
<td>yellow powdery</td>
<td>no</td>
</tr>
<tr>
<td>Wax</td>
<td>molecular</td>
<td>compound</td>
<td>solid</td>
<td>shiny translucent</td>
<td>no</td>
</tr>
<tr>
<td>Graphite</td>
<td>network</td>
<td>element</td>
<td>solid</td>
<td>shiny black</td>
<td>yes</td>
</tr>
<tr>
<td>Water</td>
<td>molecular</td>
<td>compound</td>
<td>liquid</td>
<td>transparent colourless</td>
<td>no</td>
</tr>
<tr>
<td>Ethanol</td>
<td>molecular</td>
<td>compound</td>
<td>liquid</td>
<td>transparent colourless</td>
<td>no</td>
</tr>
<tr>
<td>Carbon Dioxide</td>
<td>molecular</td>
<td>compound</td>
<td>gas</td>
<td>transparent colourless</td>
<td>no</td>
</tr>
</tbody>
</table>

**Conclusion:** All Covalent Networks and Covalent Molecules are non-conductors when solid, liquid or in solution.
**Explanation:** Once a covalent bond forms, the shared electrons are fixed in place and will not be available to move between electrodes to produce an electric current.

**Exception:** In the graphite form of carbon, only 3 of the 4 electrons are being used for the covalent bonds. The 4th electron is free to move - it is delocalised - and graphite is similar to a metal in that it can conduct electricity.

### Ionic Network

<table>
<thead>
<tr>
<th>Name</th>
<th>State</th>
<th>Formula</th>
<th>Conducts?</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium chloride</td>
<td>solid</td>
<td>Na(^+)Cl(^-)(_{(s)})</td>
<td>no</td>
</tr>
<tr>
<td>lead bromide</td>
<td>liquid</td>
<td>Pb(^{2+})(Br(^-))(<em>2)(</em>{(l)})</td>
<td>yes</td>
</tr>
<tr>
<td>copper (II) sulphate</td>
<td>solid</td>
<td>Cu(^{2+})(Cl(^-))(<em>2)(</em>{(s)})</td>
<td>no</td>
</tr>
<tr>
<td>copper (II) sulphate</td>
<td>solution</td>
<td>Cu(^{2+})(Cl(^-))(<em>2)(</em>{(aq)})</td>
<td>yes</td>
</tr>
<tr>
<td>potassium iodide</td>
<td>solution</td>
<td>K(^+)I(^-)(_{(aq)})</td>
<td>yes</td>
</tr>
</tbody>
</table>

**Conclusion:** Ionic networks cannot act as conductors when solid, but are good conductors if melted to form liquids or if dissolved in water to make a solution.
Explanation: Once formed, ions hold onto their electrons very strongly indeed, so there are never any delocalised electrons free to move through an ionic solid.

However, when liquid or in solution, the individual ions are free to move and, being charged, will move towards the electrode of opposite charge.

It appears that moving ions are able, in some way, to complete the circuit.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Formula</th>
<th>Reaction at Cathode (negative electrode)</th>
<th>Reaction at Anode (positive electrode)</th>
</tr>
</thead>
<tbody>
<tr>
<td>copper (II) chloride</td>
<td>Cu²⁺(Cl⁻)₂(aq)</td>
<td>brown solid forms (copper atoms)</td>
<td>bubbles of gas 'bleachy' smell (chlorine molecules)</td>
</tr>
<tr>
<td>zinc iodide</td>
<td>Zn²⁺(I⁻)₂(aq)</td>
<td>grey solid forms (zinc atoms)</td>
<td>brown solution forms (iodine molecules)</td>
</tr>
</tbody>
</table>

Ionic compounds are the only compounds that can conduct electricity but they are chemically changed by the process.

Metal elements form positively charged ions by losing electrons and are always attracted towards the negative electrode.

Non-metal elements form negatively charged ions by gaining electrons and are always attracted towards the positive electrode.

Electrons flow through metal wires to the Cathode (negative electrode). These electrons are effectively removed by metal ions which are converted back into atoms in the process.

Meanwhile electrons appear at the Anode (positive electrode) as non-metal ions lose electrons and are converted back into atoms. These electrons flow back to the battery through the wire. It is as if electrons flow round the circuit as normal but, in reality, only ions move through the solution - not electrons.
**Electrolysis** - is the name given to the process of *splitting apart* an ionic compound using electricity.

**Electrolytes** - are chemicals that can be used to complete an electrical circuit - effectively ionic solutions.

Electrolytes (ionic solutions) play a number of important roles in maintaining various processes within the body.

During sport, for example, electrolytes are lost through sweating and must be regularly topped up to prevent a deterioration in performance - particularly 'bad decision making' late in games. Sportspeople now try and 'top up' electrolytes during games.

Many ionic compounds have characteristic colours due to the presence of particular ions:

<table>
<thead>
<tr>
<th>Colourless ions</th>
<th>Coloured ion name</th>
<th>Coloured ion formula</th>
<th>Colour observed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na⁺, K⁺, Ca²⁺, NH₄⁺, Mg²⁺, Zn²⁺</td>
<td>copper (II)</td>
<td>Cu²⁺</td>
<td>blue</td>
</tr>
<tr>
<td></td>
<td>nickel (II)</td>
<td>Ni²⁺</td>
<td>green</td>
</tr>
<tr>
<td>Cl⁻, SO₄²⁻, CO₃²⁻, NO₃⁻, OH⁻, I⁻</td>
<td>dichromate</td>
<td>Cr₂O₇²⁻</td>
<td>orange</td>
</tr>
<tr>
<td></td>
<td>chromate</td>
<td>CrO₄²⁻</td>
<td>yellow</td>
</tr>
<tr>
<td></td>
<td>iron (II)</td>
<td>Fe²⁺</td>
<td>pale bluey-green</td>
</tr>
<tr>
<td></td>
<td>iron (III)</td>
<td>Fe³⁺</td>
<td>browny red (rust)</td>
</tr>
<tr>
<td></td>
<td>manganate</td>
<td>MnO₄⁻</td>
<td>purple</td>
</tr>
<tr>
<td></td>
<td>cobalt (II)</td>
<td>Co²⁺</td>
<td>pink</td>
</tr>
</tbody>
</table>

Compounds containing *Transition Metal ions* are usually coloured.

Electrolytes:
- Conduct energy
- Regulate fluid balance
- Transport nutrients
- Support proper muscle function
- Support mental function
- Help convert calories into energy
- Regulate pH
- & much, much more
**Single Atoms** (monatomic elements)

Very rare - only the Noble Gases exist as single atoms

- He, Ne, Ar, Kr, Xe

**Covalent Molecular** (non-metal elements & compounds)

Small - H₂, O₂, N₂, etc
HCl, H₂O, NH₃, CH₄

Medium - C₆H₆, C₆H₁₂, O₆

Large - starch, polythene, etc

**Covalent Network** (non-metal elements & compounds)

- Only 3 examples
- the element carbon (diamond)
- the element carbon (graphite)
- the compound SiO₂

**Ionic Network** (metal/non-metal compounds)

- All metal/non-metal compounds
- Na⁺ Cl⁻
- Mg²⁺ O₂⁻
- Cu³⁺ (NO₃⁻)₂

**Metallic Network** (metal elements & alloys)

- All metal/non-metal compounds
- Cu, Na, Mg, Fe, brass

**Conductivity**

- Solid - NO
- Liquid - NO
- Solution - NO

**Melting / Boiling**

- MPt - extremely low
- BPt - extremely low
- All Gases

- MPt - very low
- BPt - very low
- All Liquids, Solids

- MPt - extremely high
- BPt - extremely high
- All Solids

**Conductivity**

- Solid - NO
- Liquid - NO
- Solution - NO

* except for graphite

**Conductivity**

- Solid - NO
- Molten - Yes
- Solution - Yes *

* if soluble (Data Book)

**Conductivity**

- Solid - Yes
- Liquid - Yes
- Solution - insoluble

**Melting / Boiling**

- MPt - high to very high
- BPt - high to very high
- All Solids

* All Solids (except Hg)
Q1. Int2
Solid ionic compounds do not conduct electricity because
A the ions are not free to move
B the electrons are not free to move
C solid substances never conduct electricity
D there are no charged particles in ionic compounds

Q2. SC
Several conductivity experiments were carried out using the apparatus below.

Low voltage
d.c. supply

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Substance X</th>
<th>Substance Y</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>glucose solution</td>
<td>sodium chloride solution</td>
</tr>
<tr>
<td>B</td>
<td>copper nitrate solution</td>
<td>solid potassium nitrate</td>
</tr>
<tr>
<td>C</td>
<td>molten tin</td>
<td>liquid mercury</td>
</tr>
<tr>
<td>D</td>
<td>potassium sulphate solution</td>
<td>liquid hexane</td>
</tr>
<tr>
<td>E</td>
<td>lithium chloride solution</td>
<td>molten nickel bromide</td>
</tr>
</tbody>
</table>

Identify the two experiments in which the bulb would light.

Q3. Int2
During the electrolysis of molten copper (II) bromide
A copper atoms lose electrons to form copper ions
B bromine molecules gain electrons to form bromide ions
C bromide ions gain electrons to form bromine molecules
D copper ions gain electrons to form copper atoms

Q4. SC
The table contains information about some substances.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Melting point/°C</th>
<th>Boiling point/°C</th>
<th>Conducts as a solid</th>
<th>Conducts as a liquid</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>-7</td>
<td>59</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>B</td>
<td>1492</td>
<td>2897</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>C</td>
<td>1407</td>
<td>2357</td>
<td>no</td>
<td>yes</td>
</tr>
<tr>
<td>D</td>
<td>606</td>
<td>1305</td>
<td>no</td>
<td>yes</td>
</tr>
<tr>
<td>E</td>
<td>-39</td>
<td>357</td>
<td>yes</td>
<td>no</td>
</tr>
<tr>
<td>F</td>
<td>-78</td>
<td>-33</td>
<td>no</td>
<td>yes</td>
</tr>
</tbody>
</table>

a) Identify the substance which is a gas at 0 °C.

b) Identify the two substances which exist as molecules.

Q5. Int2
Glass is made from the chemical silica, SiO₂, which is covalently bonded and has a melting point of 1700 °C
Carbon dioxide, CO₂, is also covalently bonded but has a melting point of -78 °C.

a) What does the melting point of silica suggest about its structure?

b) What does the melting point of carbon dioxide suggest about its structure?

Q6. Int2
The properties of a substance depend on its type of bonding and structure.

Here are four types of bonding and structure.

<table>
<thead>
<tr>
<th>Discrete covalent molecular</th>
<th>Covalent network</th>
<th>Ionic lattice</th>
<th>Metallic lattice</th>
</tr>
</thead>
</table>
| a) Which type of bonding structure is missing?

b) Complete the table to match up each type of bonding and structure with its properties.

<table>
<thead>
<tr>
<th>Bonding and structure type</th>
<th>Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>do not conduct electricity and have high melting points</td>
<td></td>
</tr>
<tr>
<td>have high melting points and conduct electricity when liquid but not when solid</td>
<td></td>
</tr>
<tr>
<td>conduct electricity when solid and have a wide range of melting points</td>
<td></td>
</tr>
<tr>
<td>do not conduct electricity and have low melting points</td>
<td></td>
</tr>
</tbody>
</table>
3.5 Polarity & Solubility

Unequal Sharing

In a molecule like F₂, both atoms are exactly the same. They have equal attraction for the bonding pair of electrons.

The electrons are equally shared. This is a pure covalent bond.

A fluorine atom has a stronger attraction for electrons than a hydrogen atom, the bonding pair is pulled closer to the fluorine.

The fluorine becomes slightly negative (δ−), while the hydrogen becomes slightly positive (δ+). This is a polar covalent bond.

Water Molecules

A water molecule is a good example of a polar molecule.

It is polar, because the oxygen atom (8 protons) can attract electrons more strongly than the hydrogen atoms (1 proton), making the O—H bonds polar covalent.

Importantly, the shape of the water molecule means that one side has a slight negative charge while the other side is slightly positive. This makes water a polar molecule.

The water molecules will flip round so that their positive side is closer to the negatively charged balloon. This will make the attractions even stronger, causing the stream of water to deflect towards the balloon.
There is a 'rule of thumb' in *chemistry* that states that *like dissolves like*. In other words, chemicals (*solutes*) will dissolve in liquids (*solvents*) that are very similar to themselves in terms of the kind of *attractions* that exist between them.

### 3 Solvents:
- **Hexane** (pure covalent)
- **Ethanol** (polar covalent)
- **Water** (strongly polar covalent)

<table>
<thead>
<tr>
<th>Solute</th>
<th>Hexane (pure)</th>
<th>Ethanol (polar)</th>
<th>Water (very polar)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Wax (pure covalent solid)</td>
<td>soluble</td>
<td>insoluble</td>
<td>insoluble</td>
</tr>
<tr>
<td>Glucose (polar covalent solid)</td>
<td>insoluble</td>
<td>soluble</td>
<td>soluble</td>
</tr>
<tr>
<td>Bromoethane (polar covalent liquid)</td>
<td>soluble</td>
<td>insoluble</td>
<td>insoluble</td>
</tr>
<tr>
<td>White Spirit (pure covalent liquid)</td>
<td>soluble</td>
<td>insoluble</td>
<td>insoluble</td>
</tr>
<tr>
<td>Copper (II) sulphate (ionic solid)</td>
<td>insoluble</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
</tbody>
</table>

Predictably, pure covalent solutes tend to only dissolve in pure covalent solvents.

Polar covalent solutes can dissolve in any of the solvents - it will depend on how strongly polar they are.

The attractions set up by water molecules can be strong enough to overcome the ionic attractions in certain ionic compounds causing them to break up and dissolve.
Individually, the polar water attractions are not as strong as the ionic attractions ...

... but several water molecules will surround each ion and can succeed in pulling it away from the Ionic Network causing it to dissolve.

'Theoretical' Chemistry can provide us with rules which allow us to predict the properties of a substance if we know it's bonding and structure.

Bonding & Structure $\Rightarrow$ Properties

However, we are a Practical subject for good reason. In reality, it is the properties of a substance that often provide us with the information needed to predict it's bonding and structure

Properties $\Rightarrow$ Bonding & Structure

For example, if a substance

- **dissolves in water** $\Rightarrow$ we deduce strongly polar covalent or ionic

- **dissolves in ethanol** $\Rightarrow$ we deduce polar covalent

- **dissolves in bromoethane** $\Rightarrow$ we deduce weakly polar covalent or pure covalent

- **dissolves in hexane** $\Rightarrow$ we deduce pure covalent or very weakly polar covalent

Whenever, possible we would also want to measure Melting Points & Boiling Points as well as Conductivity etc.
Q1. KHS  
The table gives information about the attraction some atoms have for bonded electrons.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Attraction for electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>least</td>
</tr>
<tr>
<td>I</td>
<td></td>
</tr>
<tr>
<td>Br</td>
<td>greatest</td>
</tr>
<tr>
<td>Cl</td>
<td></td>
</tr>
<tr>
<td>F</td>
<td></td>
</tr>
</tbody>
</table>

Which of the following bonds is the least polar?

A  C — F  
B  C — Cl 
C  C — Br 
D  C — I 

Q2. KHS  
The table contains information about the attractions of some atoms for bonded electrons.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Relative attraction for bonded electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.2</td>
</tr>
<tr>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>N</td>
<td>3.0</td>
</tr>
<tr>
<td>O</td>
<td>3.5</td>
</tr>
</tbody>
</table>

Ammonia and water are two covalent molecules.

a) In both these molecules the electrons are not shared equally. What name is given to these types of bonds?

b) In both these molecules there are electrons not used for bonding. What name is given to these electrons?

c) Both these molecules have partial charges. Using the symbols δ+ and δ−, mark on the molecules above the positions of these charges on each.

Q3. Int2  
Synthetic nappies contain hydrogel polymers which attract and absorb water molecules.

The diagram below shows how water molecules are attracted to the hydrogel.

a) What type of bonding is present in water molecules?

b) What attracts the water molecules to the hydrogel.

Q4. Int2  
Some of the bonds in an amino acid molecule are polar covalent.

The table contains information about the attraction of some atoms for bonded electrons.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Relative attraction for bonded electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.2</td>
</tr>
<tr>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>N</td>
<td>3.0</td>
</tr>
<tr>
<td>O</td>
<td>3.5</td>
</tr>
</tbody>
</table>

The most polar bond in the amino acid molecule will be

A  C — H  
B  N — H  
C  O — H  
D  C — O 

Q5. KHS  
With the help of your data book, decide which of the following ionic compounds would dissolve in water.

<table>
<thead>
<tr>
<th>Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium iodide</td>
</tr>
<tr>
<td>lead(II) iodide</td>
</tr>
<tr>
<td>barium sulphate</td>
</tr>
<tr>
<td>aluminium hydroxide</td>
</tr>
</tbody>
</table>
**Knowledge Met in this Topic**

**Bonding**
- Only the noble gases exist as *single atoms* not permanently bonded to other atoms.
- In all other substances, atoms are held together by *bonds*.
- All bonds rely on the attraction between *positive* and *negative* charge.
- Bonding usually only involves *unpaired electrons* in the *outer shell*.
- All bonding involves *orbitals* in the *outer shell* coming close enough to *overlap*.
- Compounds of *metals* and *non-metals* usually result in *electrons being transferred* - *ionic bonding*.
- Substances containing only *non-metals* usually result in *electrons being shared* - held together by *covalent bonds*.

** Ionic Bonding**
- *Metal* atoms *lose electrons* to form more stable *positively* charged ions (*cations*),
  - e.g. Na\(^+\), Mg\(^{2+}\), Al\(^{3+}\), Sn\(^{4+}\).
- *Non-metal* atoms *gain electrons* to form more stable *negatively* charged ions (*anions*),
  - e.g. Cl\(^-\), O\(^2-\), P\(^3-\).
- *Non-metal* atoms often form molecules which *gain electrons* to form more stable *negatively* charged ions (*anions*),
  - e.g. NO\(_3^\)\(^-\), CO\(_3^{2-}\), PO\(_4^{3-}\).
- Very rarely, *non-metal* atoms form molecules which *lose electrons* to form more stable *positively* charged ions (*cations*),
  - e.g. NH\(_4^+\).
- An *ionic bond* is the force of *attraction* between *oppositely charged ions*.
- An ion can form attractions with *many* (6 - 8) oppositely charged ions.

** Covalent Bonding**
- When atoms bond covalently, they *share electrons* in such a way as to obtain the same stable electron arrangement as the *nearest noble gas*.
- A covalent bond is the result of *two positive nuclei* attracting the same *shared pair of electrons* in overlapping orbitals.
- Sometimes electrons are *not* shared equally - resulting in *polar covalent bonds*. 
Metallic Bonding

- When metal atoms bond they overlap orbitals and *share/lose electrons* to become more stable
- This results in *delocalised electrons* *constantly moving* between the orbitals of metal atoms
- A metallic bond is the result of *many positive nuclei* attracting the same *delocalised electrons* as they move between atoms
- Metallic bonding can also be described as *'a sea of electrons drifting amongst temporary positive ions'*

Single Atoms*Compounds*

- Only the *Noble Gases* exist as single atoms at room temperature
- *Attractions between* single atoms are *extremely weak* resulting in *very low Melting & Boiling Points*
- Other substances can be broken down into single atoms (*atomised*) but only at extremely high temperatures.
- Substances made up of single atoms (*monatomic*) *cannot conduct electricity*

Covalent Molecules*Compounds*

- Molecules have a fixed number of atoms bonded together by shared electrons
- Molecules can be all sizes:
  - **small** - *diatomic* - H₂, HCl, CO
  - *triatomic* - H₂O, SCl₂, CO₂ etc
  - **medium** - glucose - C₆H₁₂O₆
  - *fat* - C₅₇H₁₁₀O₉
  - **large** - starch
  - protein
- *Attractions between* molecules are *usually weak* resulting in *low Melting & Boiling Points* but attractions increase with molecular size
- *Attractions between* molecules with *polar covalent bonds* can be *stronger*
- Substances made up of molecules *cannot conduct electricity*
- Some metals from the middle of the Periodic Table can form covalent molecules
  - eg BeCl₂, AlCl₃
- as shown by their *lower than expected Melting & Boiling Points* and states at room temperature
Molecular Shapes

- Most central atoms in molecules have **4 pairs of electrons** surrounding them.
- To minimise repulsions, the electron pairs will arrange themselves *tetrahedrally*
- The shape of a molecule will depend on how many of the electron pairs are being used to bond to other atoms.

  eg  4 bonds - CH$_4$ - tetrahedral shape  
  3 bonds - NH$_3$ - pyramid shape 
  2 bonds - OH$_2$ - bent shape 
  1 bond - FH - linear shape

Ionic Networks

- Substances made up of molecules **cannot conduct electricity**
- A *network* (sometimes called a *lattice*) is a very regular arrangement
- All ionic compounds are solids at room temperature
- Ionic compounds do **not conduct** electricity when **solid** because the ions are not free to move.
- Ionic compounds do **conduct** electricity when **molten** or **dissolved** because the ions are free to move.
- When ionic compounds conduct, *chemical changes* take place at the electrodes.
- *Metals* are produced at the **negative** electrode, *non-metals* at the **positive**.

Covalent Networks

- Covalent *elements*, such as *silicon* and *carbon*, exist as giant *networks* of atoms.
- Covalent *compounds*, such as *silica* (SiO$_2$) and *carborundum* (SiC), exist as giant *networks* of atoms.
- All covalent networks are **solids** at room temperature
- Covalent networks, except *graphite*, do not conduct electricity in any state as they have *no delocalised electrons* and *no charged particles* are present.
- *Graphite* has some *delocalised electrons* which allows *graphite* to conduct electricity.

Metallic Networks

- Metallic *elements*, such as *silver* and *copper*, exist as giant *networks* of atoms.
- Metallic *alloys*, such as *bronze* (Cu & Sn) and *brass* (Cu & Zn), exist as giant *networks* of atoms.
• All metallic networks, with the exception of mercury, are solids at room temperature
• Metallic networks all conduct electricity due to the presence of delocalised electrons.

Coloured Ions Compounds
• Whilst most ionic compounds are colourless there are some which are coloured
• Most Transition Metals produce compounds with characteristic colours
  eg compounds containing Cu$^{2+}$ are always blue in colour
• Some Transition Metals have several ions each with characteristic colours
  eg Co$^{2+}$ ions are pink in colour
  Co$^{3+}$ ions are green in colour
  Fe$^{2+}$ ions are pale blue/green in colour
  Fe$^{3+}$ ions are rust in colour

Solubility Compounds
• Covalent Molecules tend to dissolve in pure covalent solvents
• Polar Covalent Molecules tend to dissolve in polar covalent solvents such as water.
• Some Ionic Compounds can also dissolve in water
**CONSOLIDATION QUESTIONS**

**Q1.**  
A nitrogen molecule is held together by three covalent bonds.  

Circle the correct words to complete the sentence.  

*In a covalent bond the atoms are held together by the*  

\{ electrons \neutrons \protons \} and the  

\{ electrons \neutrons \protons \} .

**Q2.**  
Metals can be extracted from metal compounds by heat alone, heating with carbon or by electrolysis.  

a) What is meant by the term electrolysis?  

b) A solution of copper (II) chloride was electrolysed.  

\[ \text{d.c. supply} \]

\[ \text{carbon electrode} \]

\[ \text{brown solid} \]

\[ \text{bubbles of gas} \]

\[ \text{copper(II) chloride solution} \]

i) Complete the table by adding the charge for each electrode  

<table>
<thead>
<tr>
<th>Observation at electrode</th>
<th>Observation at electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>bubbles of gas</td>
<td>brown solid formed</td>
</tr>
</tbody>
</table>

ii) How could the gas be identified?  

**Q3.**  
Metallic bonding is a force of attraction between  

A positive ions and delocalised electrons  

B negative ions and delocalised electrons  

C negative ions and positive ions  

D a shared pair of electrons and two nuclei.

**Q4.**  
Identify the covalent compound  

A zinc chloride  

B magnesium sulphate  

C lead carbonate  

D hydrogen sulphide

**Q5.**  
Which of the following diagrams could be used to represent the structure of a metal?
CONSOLIDATION QUESTIONS

Q1.
Identify the covalent compound
A  zinc chloride
B  magnesium sulphate
C  lead carbonate
D  hydrogen sulphide

Q2.
Which line in the table shows the properties of an ionic compound?

<table>
<thead>
<tr>
<th></th>
<th>Melting point (°C)</th>
<th>Boiling point (°C)</th>
<th>Conducts electricity?</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Solid</td>
<td>Liquid</td>
<td></td>
</tr>
<tr>
<td>A</td>
<td>181</td>
<td>1347</td>
<td>yes</td>
</tr>
<tr>
<td>B</td>
<td>-95</td>
<td>69</td>
<td>no</td>
</tr>
<tr>
<td>C</td>
<td>686</td>
<td>1330</td>
<td>no</td>
</tr>
<tr>
<td>D</td>
<td>1700</td>
<td>2230</td>
<td>no</td>
</tr>
</tbody>
</table>

Q3.
Chlorofluorocarbons (CFCs) are a family of compounds which are highly effective as refrigerants and aerosol propellants. However, they are now known to damage the ozone layer.

One example of a CFC molecule is shown.

\[ \text{C} \quad \text{Cl} \quad \text{C} \quad \text{F} \quad \text{F} \]

\[ \text{δ}^+ \quad \text{C} \quad \text{δ}^- \]

a) What term is used to describe the shape of this molecule?

b) What type of bonding is found in this molecule?

c) What does the symbol \( \delta^+ \) mean?

d) Which atom in this molecule has the strongest attraction for electrons?

Q4.
Which of the following elements has similar properties to argon?
A  Fluorine
B  Krypton
C  Potassium
D  Zinc

Q5.
A student set up the following experiment to investigate the colour of ions in nickel(II) chromate solution.

![Diagram of experiment]

The results are shown.

Green colour moves towards electrode A
Yellow colour moves towards electrode B

a) What is meant by d.c. supply?

b) Why must a d.c. supply be used?

c) What is meant by the term electrolyte?

d) State the colour of the nickel (II) ions?
CONSOLIDATION QUESTIONS

Q1. Identify the covalent compound

- A zinc chloride
- B magnesium sulphate
- C lead carbonate
- D hydrogen sulphide

Q2. Information on some two-element molecules is shown in the table.

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Shape of molecule</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen fluoride</td>
<td>HF</td>
<td>![Diagram of HF molecule]</td>
</tr>
<tr>
<td>water</td>
<td>H₂O</td>
<td>![Diagram of H₂O molecule]</td>
</tr>
<tr>
<td>ammonia</td>
<td>NH₃</td>
<td></td>
</tr>
</tbody>
</table>

a) Complete the table to show the shape of a molecule of ammonia.

b) The hydrogen fluoride molecule can be represented as:

![Diagram of HF molecule]

Showing all outer electrons, draw a similar diagram to represent a molecule of water, H₂O.

Q3. The table shows the colours of some ionic compounds in solution.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium chloride</td>
<td>colourless</td>
</tr>
<tr>
<td>potassium chromate</td>
<td>yellow</td>
</tr>
<tr>
<td>copper chromate</td>
<td>green</td>
</tr>
<tr>
<td>copper sulphate</td>
<td>blue</td>
</tr>
</tbody>
</table>

The colour of the chromate ion is

- A colourless
- B yellow
- C green
- D blue

Q4. Tin and its compounds have many uses.

a) Why do metals such as tin conduct electricity?

b) Tin (IV) chloride is a liquid at room temperature and is made up of discrete molecules.

What type of bonding does this suggest is present in tin (IV) chloride?

c) What is the most likely shape of a tin (IV) chloride molecule?

Q5. Which of the following substances is made up of molecules containing polar covalent bonds?

- A Calcium oxide
- B Chlorine
- C Sodium bromide
- D Water
CONSOLIDATION QUESTIONS

Q1 Many ionic compounds are coloured.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>nickel(II) nitrate</td>
<td>green</td>
</tr>
<tr>
<td>nickel(II) sulphate</td>
<td>green</td>
</tr>
<tr>
<td>potassium permanganate</td>
<td>purple</td>
</tr>
<tr>
<td>potassium sulphate</td>
<td>colourless</td>
</tr>
</tbody>
</table>

a) Give the symbol for a potassium ion.

___________________________________________________________________

b) Using the information in the table, state the colour of the potassium ion.

___________________________________________________________________

c) A student set up the following experiment to investigate the colour of the ions in copper(II) chromate.

<table>
<thead>
<tr>
<th>Observation</th>
</tr>
</thead>
<tbody>
<tr>
<td>yellow colour moves to the positive electrode</td>
</tr>
<tr>
<td>blue colour moves to the negative electrode</td>
</tr>
</tbody>
</table>

i) Lithium nitrate solution is used as the electrolyte. What is the purpose of an electrolyte?

___________________________________________________________________

ii) Suggest why lithium phosphate can not be used as the electrolyte in this experiment. You may wish to use the data booklet to help you.

___________________________________________________________________

d) State the colour of the chromate ion.

___________________________________________________________________