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Chemistry
Teach Yourself Series
Topic 9: Chemical bonding

A: Level 14, 474 Flinders Street Melbourne VIC 3000
T: 1300 134 518 **W:** tssm.com.au **E:** info@tssm.com.au

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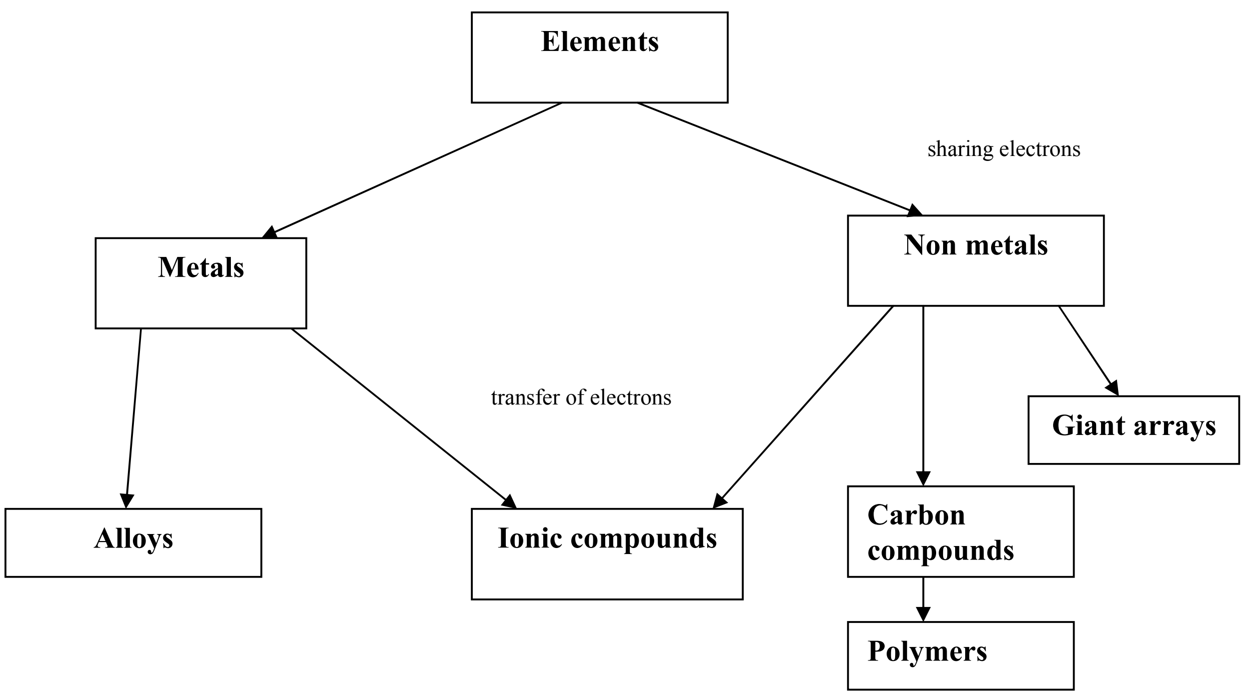
Chemical bonding

CO_2 is a gas that is slightly soluble in water.
SiO_2 has a melting point over 2000 °C. It is brittle and insoluble.
$MgCl_2$ is a solid but it is soluble in water and its solutions conduct electricity.
SO_2 is a gas that is a pollutant. It is soluble in water.

Note: The substances above have similar looking chemical formulas but their properties are extremely different. **A knowledge of chemical bonding can lead to an understanding of each substance and its likely properties**. Once you categorise a substance, you should be able to predict its chemical behaviour. A knowledge of the Periodic Table is the basis of any prediction.

Chemical Bonding

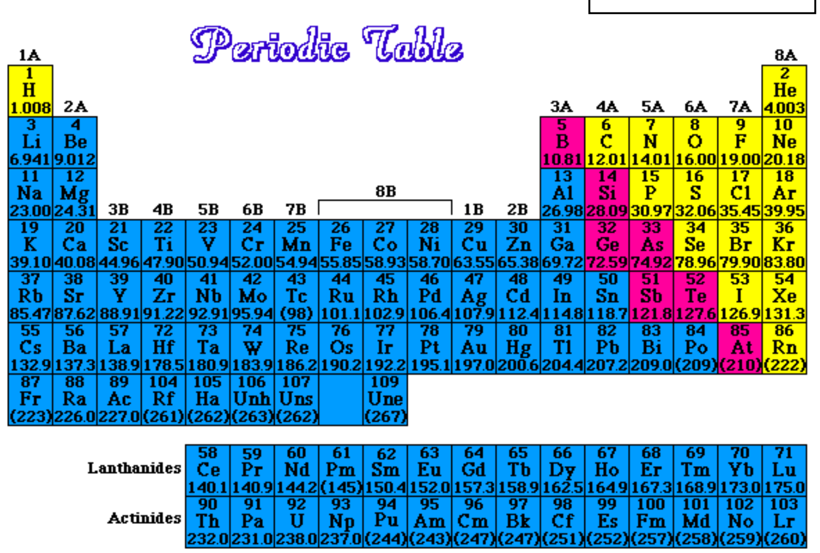
As it appears in Unit 1



Periodic Table

Most elements on the left side and middle of the table are **metals**.

Elements on the top right corner are **non metals**



Metals: elements with 1,2 or 3 electrons in the outer shell
Non metals: elements with 4 or more electrons in the outer shell

Review questions

1. Give one example of each of the following

a. alloy b. metal c. ionic solid d. polymer e. organic substance

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2. Classify each of the following as metal or non metal

titanium bromine nitrogen lead vanadium xenon

Metals

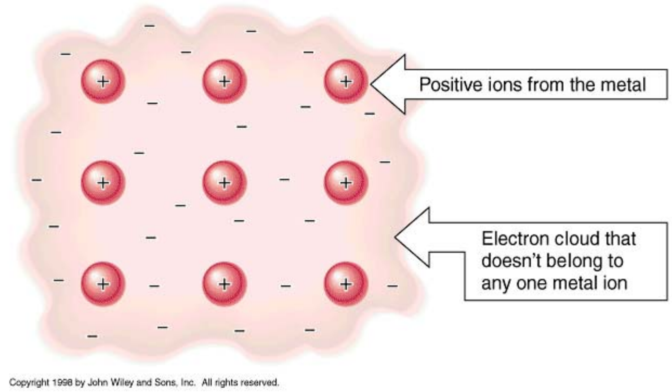
As it appears in Unit 1

Aluminium foil is an example of a **metal**. It conducts, it is shiny and it has been pressed into a thin sheet.

Metals atoms have a low number of electrons in the outer shell. These outer shell **electrons** become **delocalized**, moving throughout the structure. The positive ions are held in a particular arrangement.

Metals are elements with the following **properties**:

- **malleable and ductile**. Mallaeable and ductile refers to the fact that metals can be beaten into shapes. They are not brittle when struck with a hammer. The layers of metal ions can slide past each other and maintain a stable structure.
- good **conductors** of heat and electricity. The delocalized electrons can move towards a positive electrode if the metal is connected to a power source.
- **lustrous**. Light is reflected by the mobile electrons.
- usually high density and **high melting point**. Metallic bonding is usually fairly strong and the metal ions are held very close together.



Examples
Magnesium, Mg: Melting point 650 °C, density 1.7 g mL^{-1}. Electrons $1s^2 2s^2 2p^6 3s^2$
Aluminium, Al: Melting point 659 °C, density 2.7 g mL^{-1}. Electrons $1s^2 2s^2 2p^6 3s^2 3p^1$
Gold, Au: Melting point 1063 °C, density 19.3 g mL^{-1} Electrons **[Xe]** $\mathbf{4f^{14}\ 5d^{10}\ 6s^1}$

Group 1 Metals: Lithium, sodium, potassium, rhubidium etc all have 1 electron in the outer shell. This leads to their reactivity. Their reactivity limits their usefulness.

Li $1s^2\mathbf{2s^1}$
Na $1s^2 2s^2 2p^6\mathbf{3s^1}$ Note: the similar outer shell of each Group 1 metal
K $1s^2 2s^2 2p^6 3s^2 3p^6\mathbf{4s^1}$

Transition metals
Many of the **useful metals** are found in the **Transition Metal** section of the Periodic Table i.e. copper, zinc, iron and nickel. These metals are more **stable and have high melting points**.
The d-orbitals in the atoms are being filled as you move across the transition series

Ti $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
V $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ **Note**: the 3d shell is filling. Each element has $4s^2$ for its outer shell
Fe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

Transition metals have coloured compounds and **several oxidation states** i.e. iron can form the following oxides, each with a different oxidation number. FeO, Fe_2O_3, Fe_3O_4

Alloys

Alloys are mixtures of metals. The **properties** of an alloy are often **far superior** to those of the individual metals used. Bronze is an alloy of copper and tin but it is harder than either of the two metals. It also corrodes less than pure tin. Most of our coins are examples of alloys.

Review questions

3. Explain how the delocalised electrons in metals dictate most of the properties of metals.

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4. Most metals are dense and most metals have high melting points. Why is the word 'most' required?

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5. List three metals that are used for specific purposes i.e. tungsten – light filaments due to the high melting point

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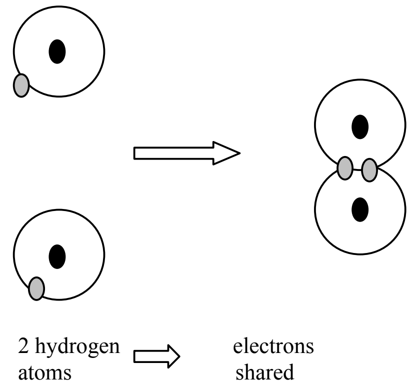
Molecules

As it appears in Unit 1



Non metals combine with each other by **sharing** electrons.
They obtain complete outer shells of electrons this way.

Hydrogen

Each H atom shown has a single outer shell electron.
When the two atoms share the outer shell electron,
a **H_2 molecule** is formed and both atoms now have
a complete outer shell.

Different ways of representing a molecule of hydrogen gas.

H : H — **Electron dot diagram** outer shell electrons only

H_2

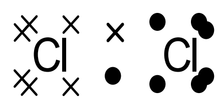
H – H ↑ single covalent bond

H_2 H_2 H_2 H_2 H_2

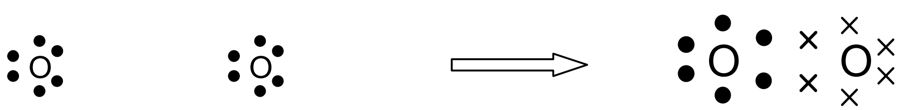
Sample of hydrogen gas: container with many discrete H_2 molecules moving rapidly in it.

Hydrogen gas is made up of **covalently bonded diatomic** molecules.

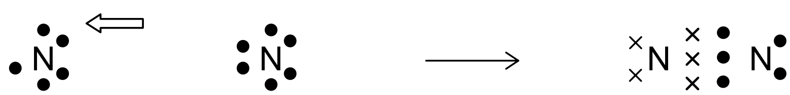
Many other molecular substances exist. They share electrons to complete their outer shells. The chemical formulas of the molecules reflect the outer shell electrons.
Chlorine has 7 electrons in its outer shell, hence it will form Cl_2 molecules with a single covalent bond between each chlorine atom, Cl – Cl



Oxygen has 6 electrons in its outer shell so it shares 4 electrons



O = O O_2 double covalent bond



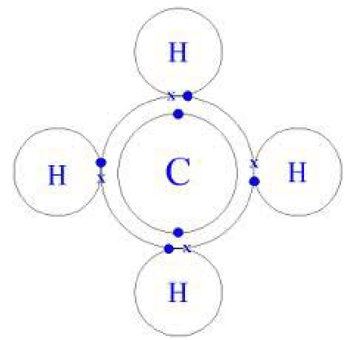
Nitrogen atoms form a molecule that has a **triple covalent bond**, N_2

There is no reason for molecules to be elements. Several different elements can share electrons i.e. H_2O, CO_2

Two oxygen atoms plus one carbon combine to form carbon dioxide. Each element has its outer shell completed by this sharing of electrons.

$O + C + O \longrightarrow O{=}C{=}O$

carbon + 4 hydrogen atoms → methane, CH_4



Review questions

6. Water has a formula H_2O

a. Represent water as an electron dot diagram.

b. Why is its formula H_2O?

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c. Explain why water is a V shaped molecule.

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7. Draw electron dot diagrams of each of the following

propane, C_3H_8 carbon tetrachloride, CCl_4 chlorine, Cl_2 hydrogen iodide, HI

8. Nitrogen, oxygen and fluorine all form diatomic gases. The nature of the covelent bonding in each is different, however. Explain how the bonding differs.

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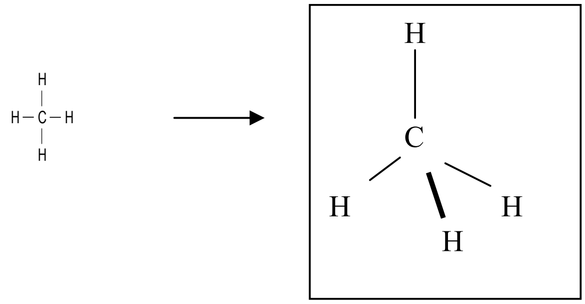
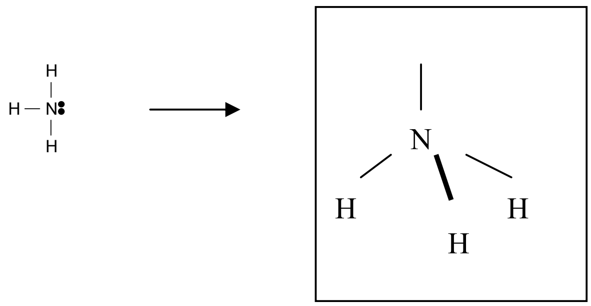
Molecular shape

As it appears in Unit 1

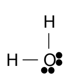
The shape of the molecule is **determined by the arrangement of outer shell electron pairs**. Electron pairs repel each other and strive to get as far away from each other as possible.
Methane, CH_4 has 4 pairs of electrons. This leads to **a tetrahedral** arrangement.



Methane has a formula CH_4. The molecule will form a tetrahedral shape as this allows the electrons to be more than 90° apart.



Ammonia, NH_3 has 4 electron pairs but one of these is non bonding. Its shape is therefore a pyramid



Water, H_2O will form a V-shaped molecule.

Properties of molecular substances

- usually low boiling point
- non conductors of heat and electricity

Forces between molecules

If molecular substances are to form liquids or solids, there must be **forces between one molecule and the next.**

this covalent bond is strong but it does not explain why HCl can bond to another HCl

H — Cl Cl — H

an intermolecular force needs to be present between separate HCls

Cl — H H — Cl

These forces are due to **dipoles**

$$H^{\delta+} \quad Cl^{\delta-} \qquad\qquad H^{\delta+} – Cl^{\delta+}$$

The 2 electrons in the HCl bond **are not shared evenly**. They are more likely to be closer to the chlorineatom. As electrons are negatively charged, this makes the **chlorine slightly negative** and it leaves the **hydrogen slightly positive.** This charge separation is referred to as a **dipole.**

Electronegativity tables are needed to determine whether dipoles exist

Typical electronegativity values

H	P	C	S	N	Cl	O	F
2.1	2.2	2.5	2.6	3.0	3.2	3.5	4.0

The atom with the **higher electronegativity** will have the slight negative charge

Example

Show the dipole that will exist between a. O and H b. N and H

Solution

a. $O^{\delta-} – H^{\delta+}$ (3.5 compared to 2.1) b. $N^{\delta-} – H^{\delta+}$ (3.0 compared to 2.1)

HCl as a solid

$H^{\delta+}$ — $Cl^{\delta-}$ $H^{\delta+}$ — $Cl^{\delta-}$

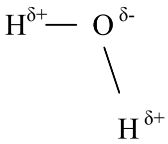
- $Cl^{\delta-}$ $H^{\delta+}$ — $Cl^{\delta-}$ $H^{\delta+}$ — $Cl^{\delta-}$

dipoles have led to attractions between molecules
the stronger the dipoles, the higher the boiling point

When the dipole involves hydrogen, it is often a relatively strong dipole. This is referred to as **hydrogen bonding.**

Polar and non polar molecules
Molecules with dipoles may or may not be **polar**. Sometimes dipoles are equal and opposite in a molecule and they cancel each other out as far as the molecule as a whole is concerned. Such molecules are **non polar.**

$O^{\delta-} == C^{\delta+} == O^{\delta-}$ is **non polar** as the two dipoles are equal and opposite, hence cancel.



is **polar** as the two dipoles are not symmetrical

Polar molecules include HCl, CH_3Cl, CH_3CH_2OH, NH_3
Non polar molecules include O_2, CO_2, CH_4

Dispersion forces
What holds molecules such as helium, oxygen and nitrogen together when they exist as solids? They have no dipoles. The very weak force responsible is referred to as **dispersion forces**. The melting point of helium is in fact -270 °C, just three degrees above absolute zero.

Covalent Lattices (Giant Arrays)
Some substances have covalent bonds but the bonds continue to form **very large structures** instead of discrete molecules.
Diamond, graphite, sand and tungsten carbide are examples.
These substances are hard and brittle because covalent bonds are **strong bonds**.

Review questions

9. Draw electron dot diagrams of the following molecules and use these diagrams to predict the shape of the molecules

a. F_2 **b** H_2S **c**. PH_3 **d**. CCl_4

10. **a**. For the molecules in Q.9 show any dipoles that might exist.

b. Decide if the molecules will be polar or not.

11. Rank each of the following in order of boiling points, giving reasons for your choices

H_2, HCl, H_2S

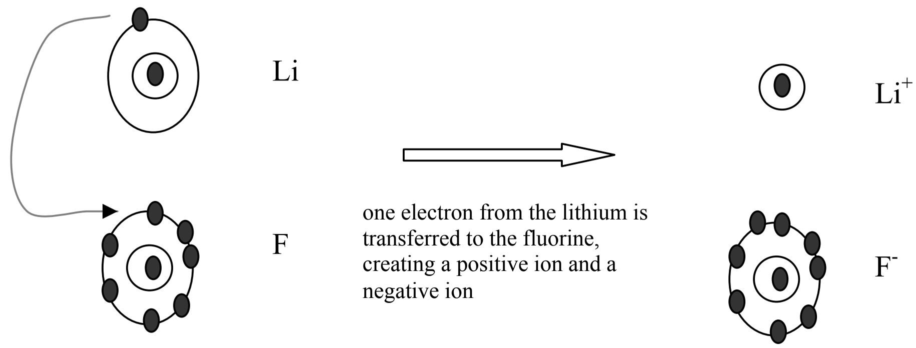
12. Diamond and methane both contain covalent bonds. The melting point of diamond is over 3000 °C higher than that of methane. Explain why this is the case.

13. Use CCl_4 as an example to explain how a molecule can have dipoles yet be non polar.

Ionic Solids

As it appears in Unit 1

Metals and non metals are often fairly reactive. **Since metals like to lose electrons and non metals gain electrons**, it is natural they will react with each other. Atoms that have lost or gained electrons are called **ions** – they have a charge. Metals form **positive ions** and non metals form **negative ions**.



Lithium fluoride is an ionic substance. It contains a network of positive Li^+ ions and negative F^- ions.

Properties Ionic solids

- Conduct electricity as liquids or solutions
- Are hard and brittle
- Have high melting points, as ionic bonds are strong

In table salt, sodium atoms donate electrons to chlorine atoms to form NaCl.
NaCl consists of Na^+ ions and Cl^- ions.
The charge on an ion is determined by its outer shell electrons

$Na \rightarrow Na^+$
$Mg \rightarrow Mg^{2+}$
$Al \rightarrow Al^{3+}$
$F \rightarrow F^-$
$O \rightarrow O^{2-}$

Polyatomic ions also exist i.e. SO_4^{2-} , CO_3^{2-}

Writing formulas
Ionic substances are **neutral**, the positive ions balance the negative ions
The **positive ions are written first**

i.e. aluminium chloride $AlCl_3$ as Al^{3+} and Cl^-
calcium bromide $CaBr_2$ as Ca^{2+} and Br^-
aluminium sulfate $Al_2(SO_4)_3$ as Al^{3+} and SO_4^{2-}

Review questions

14. Write correct chemical formulas for the following

a. calcium iodide | **b**. lithium carbonate
c. magnesium phosphate | **d**. copper iodide

15. Name each of the following compounds

a. $CuCO_3$ | **b.** LiBr
c. Ag_2SO_4 | **d**. Ag_2S
e. $Al_2(CO_3)_3$ | **f**. NH_4I

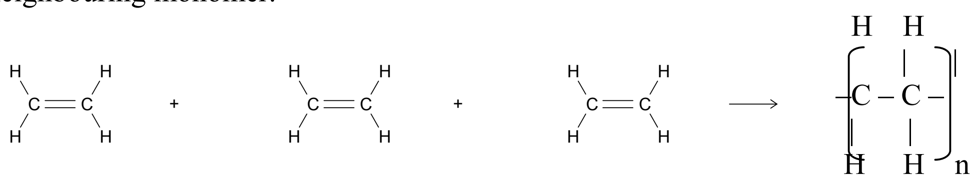
16. Explain which category each of the following substances falls into i.e. water = H_2O = covalent molecular

Mg | MgI_2 | SO_2
NH_3 | $CaSO_4$ | H_2O

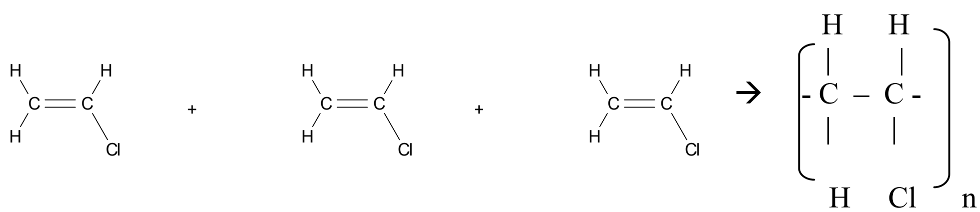
Polymers

As it appears in Units 1 & 3

Polymers are very **long molecules** made from a repeating unit known as a **monomer**.
The most common polymers are addition polymers where the **double bond** in a monomer is reacted with a neighbouring monomer.



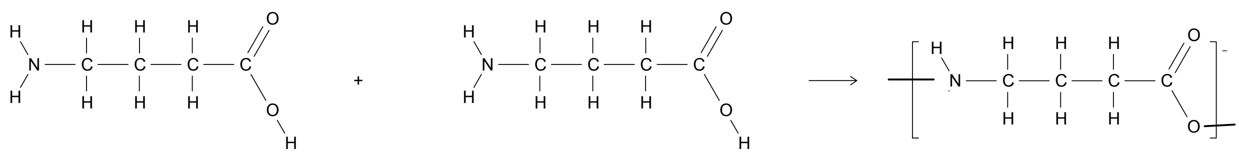
ethene + itself several thousand times → polyethene



chloroethene + itself → polychloroethene or PVC

These are **addition polymers**

Condensation polymers can form when monomers with two functional groups can react with each other.



This is a **condensation reaction** that occurs on both ends of the monomer. **Polyesters and nylons** are similar reactions to this.

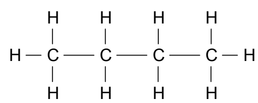
Bonding in Unit 3 exams

As it appears in Unit 3

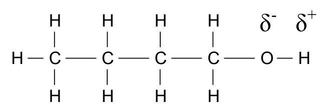
Although bonding itself is not listed as a separate topic in Unit 3, questions often expect an understanding of the impact of bonding on properties of molecules.

Sample questions

1. Butane vs butanol
The structure of butanol is fairly similar to that of butane, yet their properties are very different.
Explain why this is the case.



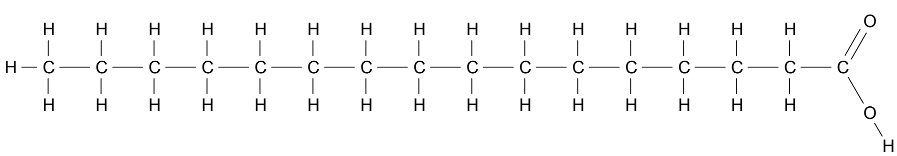
butane



butanol

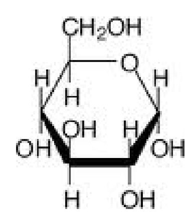
The presence of an oxygen atom in butanol leads to butanol having a dipole and being a polar molecule. The dipole in one molecule of butanol will attract dipoles in neighbouring molecules, leading to the melting point and boiling point of butanol being much higher than that of butane. Butane relies upon dispersion forces. Butanol will be soluble in water due to the polar nature of its molecule.

2. Fatty acids
Fatty acid molecules are relatively non polar. Explain why this is the case.

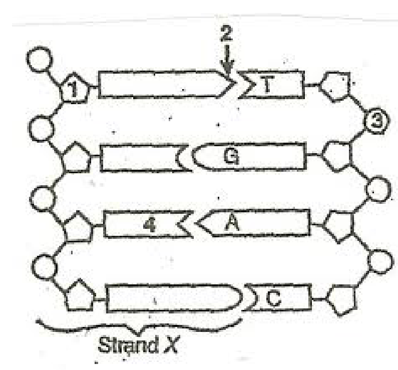


Fatty acids are very long molecules. The alkane part of the chain is non polar so most of the molecule is non polar. The length of the non polar section dominates the chemical behaviour of the molecule.

3. Glucose
Glucose is a relatively soluble molecule. Explain why.
A glucose molecule has 5 hydroxy groups, –O – H, as part of its structure. The hydroxyl groups causeglucose to be a quite polar molecule that will dissolve fairly easily in polar solvents like water.



4. DNA
DNA is a biopolymer that exists as two separate strands. Explain what type of bonding is present in a DNA polymer.
DNA contains many covalent bonds. The sugar molecules and base molecules are constructed with covalent bonds and are joined by covalent bonds. The separate strands of DNA are held together by hydrogen bonding between the base molecules. The nitrogen and oxygem atoms in the bases lead to significant dipoles.



5. Aspirin
Aspirin is often sold as a sodium salt.
Write an equation for the formation of the sodium salt and explain why the sodium salt might be preferred.

Aspirin is a weak acid, represented here for simplicity as AsH.

AsH(s) + NaOH(aq) → AsNa(aq) + $H_2O(l)$ where AsNa is Na^+ ions with As^- ions.

The sodium salt of aspirin is an ionic compound and not a covalent compound. As such, its solubility is greater. It is useful for a pharmaceutical to be soluble in water.

Solutions to Review Questions

1. Sample answers a. alloy - brass b. metal - magnesium c. ionic solid - sodium chloride d. polymer - PVC e. organic substance - glucose

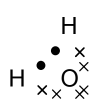
2. Metals - titanium lead vanadium xenon
Non metals - bromine nitrogen xenon

3. Metals conduct because the free electrons move towards the charged plates. Metals are malleable because the layers of ions can slide over each other.

4. There are always exceptions: aluminium and magnesium are light in weight. Mercury is a liquid at room temperatures.

5. Platinum as a catalyst, gold for electrical condictive wiring, titanium for light weight strength in airplane wings

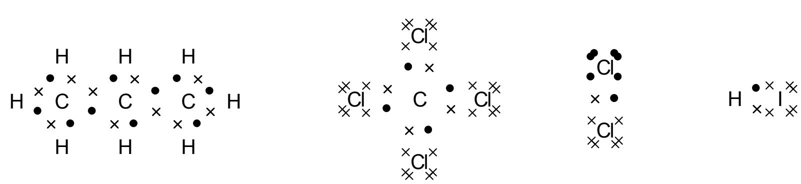
6. **a**. Represent water as an electron dot diagram.



b. hydrogen needs to share one lectron but oxygen needs two, hence H_2O.

c. There are 4 pairs of electrons making the electrons tetrahedral in arrangement. Two of these electron pairs are non bonding, therefore the shape is V-shape.

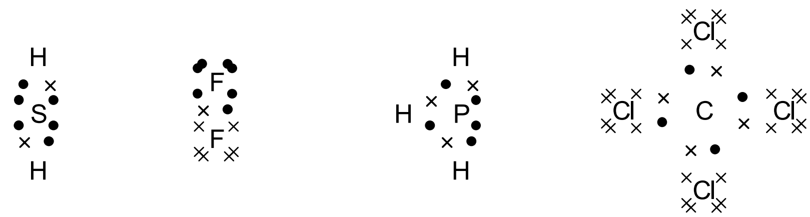
7.



8. The bonding goes from single, to double and then to triple

F — F O = O N ≡ N

9. **a**. F_2 **b** H_2S **c**. PH_3 **d**. CCl_4



V-shape linear trigonal pyramid tetrahedral

10. **a**. H_2S will have dipoles, the S being negative and the H positive.
F_2 no dipoles
PH_3 dipoles between P and H, the P being the negative atom
CCl_4 has dipoles, with the Cl as the negative

b. H_2S and PH_3 will be the polar molecules

11. H_2 has only dispersion forces. H_2S and HCl both have dipoles but the dipoles in HCl will be stronger than in H_2S.

12. Diamond is a giant array but methane is molecular. The giant array is entirely covalent bonds hence the melting point is strong. Methane has weak bonding between one molecule and another molecule.

13. There is a dipole between each carbon and chlorine atom but the 4 dipoles are in a symmetrical arrangement where their impact on the molecule as a whole cancels out.

14. **a**. CaI_2 **b**. Li_2CO_3
c. $Mg_3(PO_4)_2$ **d**. CuI_2

15. **a**. copper carbonate **b.** lithium bromide
c. silver sulfate **d**. silver sulfide
e. aluminium carbonate **f**. ammonium iodide

16. Mg - metal MgI_2 - ionic SO_2 - covalent molecular
NH_3 - covalent molecular $CaSO_4$ - ionic H_2O - covalent molecular