

Revision on the Properties of the Elements in Period 3

1st December 2018

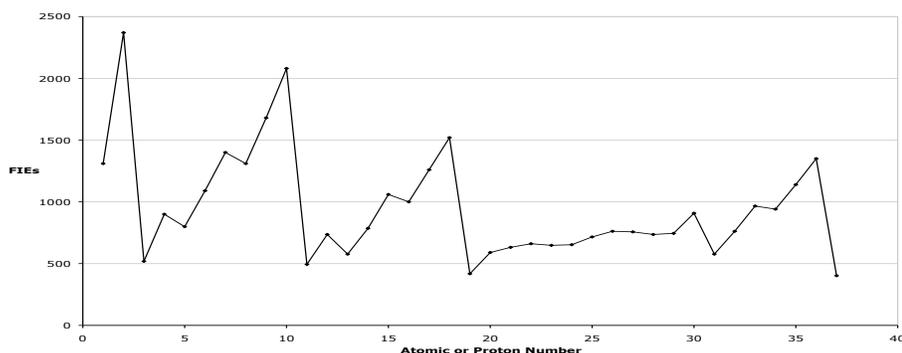
The number of protons in an atom defines the element to which the atom belongs. As you traverse from left to right in the Periodic Table each successive element has one more proton than the previous element.

GROUPS

The reason for the similarities between different elements in a group is due to the fact that all the elements in a Group have exactly the same number of electrons in their outermost shells.

PERIODS

The word “period” is used to indicate a regularly repeating pattern, and this word is therefore used to indicate that certain cycles/certain patterns are repeated regularly in the groupings of elements in the *Periodic* Table.



FIRST IONISATION ENERGIES

Would you be able to tell the examiners why there is a fall in FIE between Mg (12) and Al (13), and between Phosphorous (15) and Sulphur (16) – and if you need reminding, there are always four factors to consider viz. (a) distance (b) number of protons (c) shielding and (d) mutual repulsion.

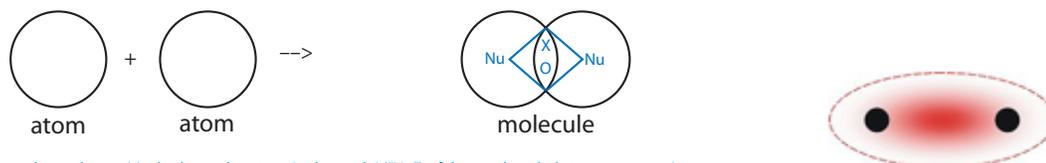
METALS vs. NON-METALS

- **In the formation of an ionic substance**, metal atoms react with non-metal atoms in such a way that the atoms of the metal element **TRANSFER** the electron or electrons in their outermost shells to the atoms of the non-metal element thereby creating *positively* charged metal cations and *negatively* charged non-metal anions. It is the electrostatic forces of attraction (called **ionic** or **electrovalent** bonds) between every ion and every oppositely charged ion (i.e. the **omni-directional electrovalent bonds**) that hold the ionic substance together.

In contrast,

- **in the formation of a molecular substance**, the atoms of two or more non-metal elements react together in such a way that the atoms of the non-metal elements **SHARE** one or more pairs of bonding electrons thus forming molecules, and the *atoms* inside the molecules are held together by highly **directional intra-molecular** forces of attraction (called **covalent** bonds). [*Neighbouring molecules in a molecular substance are held together by much weaker **inter-molecular** forces of attraction such as Hydrogen Bonds or one of the three van der Waals'/London/Debye forces of attraction.*]

The molecular covalent bonds on page 1 are called **intra-molecular** bonds (i.e. the bonds *within* a molecule¹) and they are just one sort of “**covalent bond**”. The one below is a “sigma” (σ) bond, and the other intra-molecular bond that is discussed at ‘A’ Level is the “pi” (π) bond. (I believe that the schema in red on the right was created by Professor Stephen Lower, Simon Fraser University.)



In the diagram above, the positively charged protons in the nuclei (“Nu”) of the two bonded atoms are exerting an electrostatic force of attraction (shown here as “<>”) on the pair of negatively charged bonding electrons “xo”.

When we look at the properties of the Period 3 elements we will need to keep in mind both the structure and the bonding involved (as well as other factors which I will come to). One of the things that the Syllabus requires you to know is the properties (i.e. melting points/boiling points/electrical conductivity/and thermal conductivity) of the Period 3 elements, and about their Oxides. We therefore need to understand the nature of these substances.

IONIC SUBSTANCES

How is the substance formed: The atoms of a metal element **TRANSFER** the electrons in their outermost shell to the atoms of a non-metal element thus creating positively charged metal ions and negatively charged non-metal ions.

What holds the substance together: The electrostatic force of attraction between positively charged **IONS** and negatively charged **IONS** hold ionic substances together.

These forces of attraction are called **electrovalent** or **ionic** bonds, and these bonds exist between **EVERY** ion and **EVERY** oppositely charged ion in the ionic substance viz. they are **non-** or **omni-directional** bonds, and there is a **COLOSSAL** number of electrovalent or ionic bonds inside an ionic substance (which is why so much energy is needed to melt/boil ionic substances).

What sort of bonding is there: Electrovalent / ionic.

MOLECULAR SUBSTANCES

How is the substance formed: The atoms of two or more non-metal elements **SHARE** one or more pairs of bonding electrons.

What holds the substance together

- A) INSIDE each molecule** *VERY STRONG* **Intra-molecular** electrostatic forces of attraction exist between
 (i) positively charged **protons** in the nuclei of the bonded atoms and
 (ii) one or more pairs of negatively charged bonding **electrons**.
 These forces of attraction are called **covalent bonds**, and they are highly **directional** bonds.

- B) BETWEEN the molecules** there are four *VERY WEAK* **inter-molecular** forces of attraction. They are

| | <u>STRENGTH (in kJ mol⁻¹)</u> |
|---|--|
| a) Hydrogen bonds | 10-50 |
| b) van der Waals’ forces of attraction viz. | |
| b1) Permanent dipole to Permanent dipole | 5 |
| b2) Permanent dipole to Induced dipole, and | 3 |
| b3) Momentary or Instantaneous dipole to Induced dipole forces of attraction. | 1 |

¹ In English, the Latin word “**intra**” means “inside” or “within”, and the Latin word “**inter**” means “between”.

What sort of bonding: Covalent bonds **INSIDE** the molecules, and “inter-molecular forces of attraction” **BETWEEN** molecules.

If you do not understand the stuff on the previous page, then you will struggle to understand the properties involved in Period 3 (or in any other Period).

A) Differences in the Periodic Table in Melting Points and Boiling Points

Melting refers to the process whereby a substance, when it *melts*, goes from the solid state (s) to the liquid state (l).

Boiling refers to the process whereby a substance, when it *boils*, goes from the liquid state (l) to the gaseous state (g)²,

and the amount of energy that it takes to convert a substance from the solid state to the liquid state, and from the liquid state to the gaseous state will depend entirely on the strength of the bonds that are within the substance – and we then need to examine the strength of the bonds inside substances composed of differing elements.

The length of a bond determines how strong it is, and the shorter a bond the stronger it is – and in order to analyse the strength of differing bonds, we must first determine the nature of the bond that is being examined i.e. how many bonds there are in the substance/how short or how long they are/and what type they are (i.e. whether they are metallic, electrovalent or covalent bonds).

There are extremely strong bonds in pieces of metal/and in ionic substances/and in **intra**-molecular covalent bonds, therefore it must follow that it will require a large amount of energy in order to break all these bonds – and it is this that causes metals (and ionic substances) to have high melting and boiling points. **Giant or macro-molecular substances held together by molecular bonds (e.g. diamond and SiO₂) have high MPs/BPs, but simple molecular substances held together by vdW forces have low MPs/BPs.**

Inter-molecular bonds are very weak, therefore simple molecular structures have very low MPs and BPs, and become gases at very low temperatures. *Simple molecular structures are gases at RTP.*

A.1) The Melting and Boiling Points of Metals

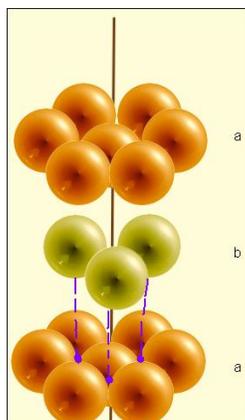
Every metal ion in a piece of metal has an interest in ALL (not just one or two or three or four or five) but ALL of the huge pool of electrons possessed by that piece of metal! Every positively charged metal ion in a piece of metal exerts a force of attraction on every, on EVERY single one of the negatively charged electrons in the huge pool of delocalised electrons in the piece of metal (and vice versa) because these bonds are non-directional or omni-directional bonds³.

² The difference between boiling and *evaporating* is that in the process of “boiling” (i) bonds are being broken and liquid units/molecules are being converted into gas **right throughout the whole structure of the liquid** and *not just at the surface of the liquid*, and (ii) irrespective of how much energy is being poured into the liquid, **the temperature of the whole liquid substance remains constant at the Boiling Point of the liquid until every single unit/molecule of the substance has been converted into gaseous form**, whereas in “evaporation” (a) it is only specific **localised** bonds that are being broken and this happens **only at the surface of the liquid** and thus (b) the temperature of the liquid has absolutely nothing whatsoever to do with the Boiling Point of the liquid, and the energy needed in evaporation is miniscule compared to the energy needed in boiling. In fact, in the evaporation associated with the drying of washing on a clothes line, virtually all the energy used is provided by the kinetic energy possessed by the molecules of the wind that collide with the molecules of water on the surface of the damp clothes.

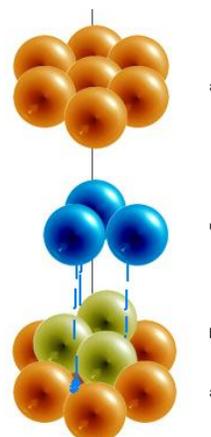
³ The force of attraction diminishes by $[1 \div (\text{distance}^2)]$ and, clearly, the further away that a metal ion and an electron are from each other, the very much less will that the force of attraction be – **but that does not stop them from being attracted to each other!**

The Melting and Boiling points of metals are related to the Group/the Period in which the element resides

- 1) **More bonds** : The more electrons that there are in the pool of delocalised electrons the more bonds there will be, and *the more bonds that there are then the greater will be the amount of energy required to break those bonds and to melt and subsequently to boil the piece of metal*. It is then but a short stride to move from there to seeing that since Group III metals delocalise three electrons per atom then there will be more metallic bonds in a piece of Group III metal than there are in a piece of Group II metal, and that a piece of a Group II metal will have more metallic bonds in it than a Group I metal.
 - 2) **Shorter bonds with regard to any one Period** : In the table showing atomic radii (on the next page) you can see that atomic radii *decrease in going from left to right across the Periodic Table*, (and therefore increase in going from right to left), and if shorter bonds are stronger than longer bonds, then pieces of *metal* made from an element to the right of another metal element in any Period of the table will be stronger than a piece of metal made from a metal element to the left of it in the Periodic Table – and thus the metal on the right will (all other things being equal) have a higher melting point than the metal element on the left.
 - 3) **Shorter bonds with regard to differing Periods** : Equally, *atomic radii increase going downwards in a Group* in the Periodic Table⁴ therefore pieces of metal made from elements lower in the Periodic Table will have longer/weaker bonds and thus have lower Melting and Boiling points.
- You can see therefore why it is that **textbooks say that the amount of energy needed to melt a piece of a metal will depend on the Group and the Period to which the metal belongs.**⁵
 - There is one other very important factor that affects the Melting and Boiling Points of Metals, and that is to do with how metallic ions in a piece of metal are **packed** on top of and around each other – but the subject of Packing is very complicated and needs the actual physical handling of something like billiard balls to be understood properly.



Cubic Close (abab) Packing



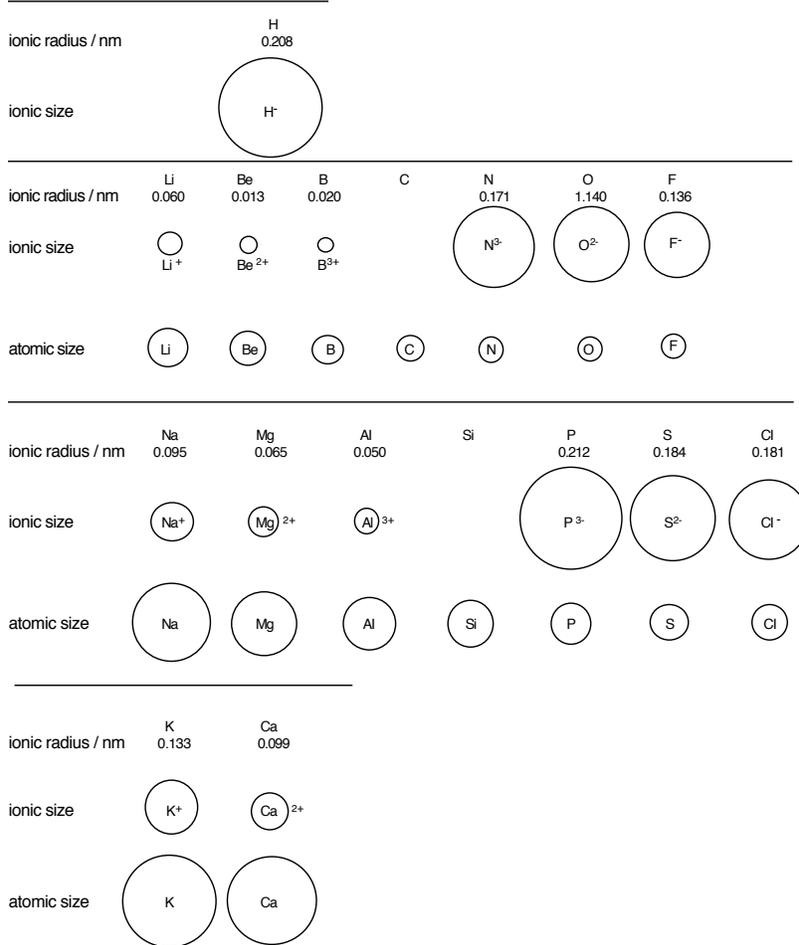
Hexagonal Close (abcabc) Packing

Source : UCD

⁴ As the Principal Quantum Energy Level goes from $n=1$ to $n=2$ to $n=3$ and so on.

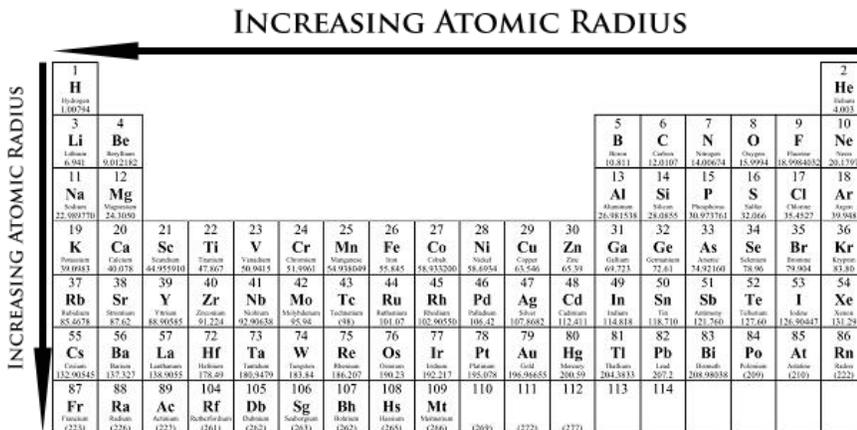
⁵ Actually, it is more complicated than this when it comes to Transition Metals/etc, but I want to concentrate only on the principles that are involved.

Chemistry in Context



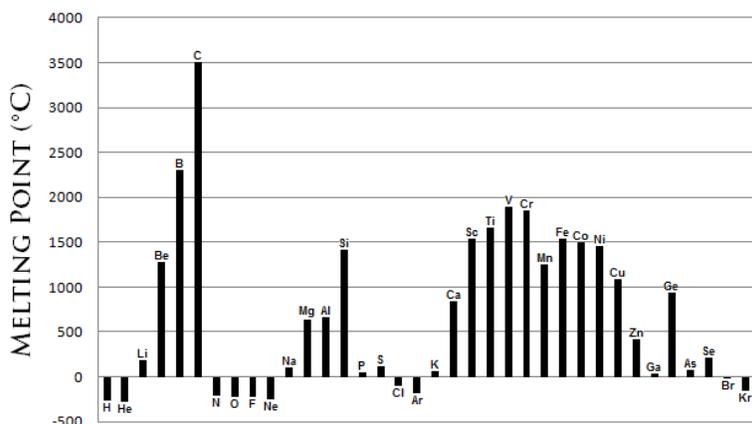
Atomic radii decrease as you go from left to right across a Period. This happens because the increasing number of protons draws the electron cloud closer and closer to the nucleus of the atom.

Radii of the most stable ions for some of the first 20 elements in the periodic table.



Source : University of California, Davis

The result of all those different factors gives the following MPs.



Source : University of California, Davis

A2) The Melting and Boiling Points of Non-metal Elements that have simple molecular structures

The high Melting and Boiling Points of Metals are entirely related to the fact that every piece of metal has strong omni-directional metallic bonding holding the piece of metal together.

The table above shows the Melting Points of the first 36 elements, and from it you will see that Metal elements have high Melting Points but that, in contrast, those Non-Metal elements that have **simple molecular structures** have low Boiling Points – and, indeed, at 0°C/273K most of the Non-Metal elements are gases.⁶ The reason for this is very simple viz. **Non-Metal elements that have simple molecular structures have very weak inter-molecular forces (such as Hydrogen bonds and vdW forces) of attraction holding them together.**

I showed you this on page 2, but I will repeat the table

| Type of bond/force of attraction | Energy required to break the bond/the force of attraction (but can vary enormously) |
|--|---|
| Covalent and electrovalent bonds | Many hundreds and sometimes even thousands of kJ mol ⁻¹ |
| Hydrogen “bonds” | about 10-50 kJ mol ⁻¹ |
| Permanent dipole to Permanent dipole | about 5 kJ mol ⁻¹ |
| Permanent dipole to Induced dipole | about 3 kJ mol ⁻¹ |
| Instantaneous dipole to Induced dipole | about 1 kJ mol ⁻¹ |

and you can see that compared to *covalent* and *electrovalent* bonds, *Hydrogen bonds* are weak, and that *vdW forces* are even weaker still. It is thus no surprise that the Melting and Boiling Points of elements that have simple molecular structures are very low indeed.

⁶ 0°C is a temperature that relates entirely to the Earth and has no importance whatsoever to any other part of the Universe. It is important to human beings because most life forms on earth are totally dependent on Water, but there is no ‘*a priori*’ reason for believing that other life forms must have evolved around this need for Water. For example, if you can obtain a copy of Professor Fred Hoyle’s excellent “The Black Cloud”, then you really ought to read it. If you have had (as I had in my youth) a very restricted intellectual upbringing because of religious factors, then it will broaden your intellectual horizons. (If you have never encountered the Latin phrase ‘*a priori*’ before, then in this context it means ‘beforehand’/‘before you study the relevant facts’.)

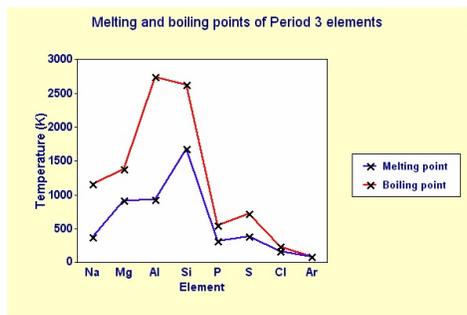
A3) The Melting and Boiling Points of elements that have giant/macromolecular structures

Pieces of metal are held together by very strong omni-directional metallic bonds, and those elements that have simple molecular structures are held together by very weak inter-molecular bonds – but those elements that have giant or macromolecular structures are held together by very strong covalent bonds, and thus they have very high Melting and Boiling Points.

The thing that you then need to know is which Non-Metal elements have giant/macromolecular structures and which have simple molecular structures, and then we will have covered Melting Points and Boiling Points in more than sufficient depth for ‘A’ Level purposes, and you just have to know the next table off by heart.

| Group | I | II | III | IV | | V | VI | VII | VIII |
|------------------|----|----|-----|--------------|-------------|----------------|----------------|-----|------|
| Period 2 | | | | | | | | | |
| Element | Li | Be | B | C (graphite) | C (diamond) | N | O | F | Ne |
| Structure | M | M | GM | GM | GM | SM | SM | SM | SM |
| Period 3 | | | | | | | | | |
| Element | Na | Mg | Al | Si | | P ₄ | S ₈ | Cl | Ar |
| Structure | M | M | M | GM/Md | | SM | SM | SM | SM |

M = Metallic Structure Md = Metalloid GM = Giant Molecular SM = Simple Molecular Structure



Source: creative-chemistry.org

Anything that has a giant molecular structure is held together by strong covalent bonds. If therefore, in the exams, you are asked why the giant molecular substance XYZ (e.g. Diamond) has a *high* MP/BP, then the answer that you give must state that “the substance is held together by nothing but strong covalent bonds”, and if you are asked why the simple molecular substance ABC (e.g. P₄/S₈/Cl₂) has a relatively *low* MP/BP, then the answer you give must be that “the molecules of the substance are held together by nothing but weak vdW forces of attraction”.

When it comes to the oxides of the Period 3 elements, it will be up to you to know that

- metal oxides are **ionic** and they therefore have high MPs/BPs, whereas
- some non-metal oxides such as SiO₂ are **giant molecular** substances (just like Diamond) and form just one huge molecule therefore they have high MPs/BPs, whereas
- others (such as SO₂ and SO₃) are **simple molecular** structures and thus have relatively low MPs/BPs, and in between
- there are **big multiple molecule** structures (e.g. polymers) held together by vdW forces of attraction but the molecules are so large (therefore their electron clouds are very large and their Momentary Dipoles are therefore very large) that the vdW force of attraction is much larger than normal and they therefore have higher MPs/BPs than simple molecular structures.

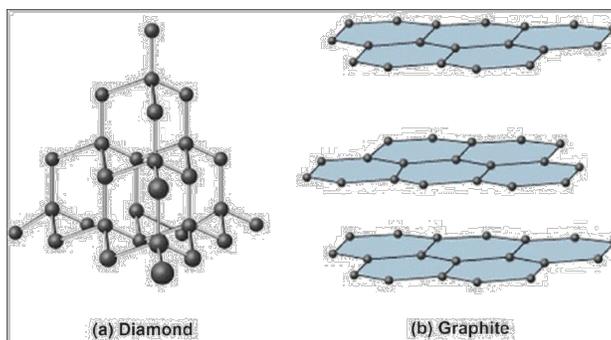
B) Electrical Conductivity

- Let us now consider electrical conductivity. The critical thing with each piece of metal is the **delocalised electrons** which do not stay in any one place and are not associated with any one metal ion but instead, as their name implies, wander around freely in a piece of metal.
- When a current flows through a wire, **many** electrons enter one end of the wire and just as many electrons pop out of the other end where in the diagram below, “ δ^+ ” is meant to represent a metal ion, and “ e^- ” is meant to represent an electron.



Source: unknown

- If it is the presence of delocalised electrons that allows a metal to conduct electricity, then the more delocalised electrons that there are, the greater will be the ability of the metal element to conduct electricity – and (all other things being equal) this means that Group III metals are better conductors of electricity than Group II elements, and they in turn are better conductors than Group I elements.
- When it comes to the elements, it is only **elements with metallic structures** that are normally able to conduct electricity.
- Generally speaking, those with molecular structures (whether simple or giant) are not able to conduct electricity, because electricity has to be carried by something, and molecular and solid ionic structures “**have no mobile charged particles in them**”. The critical thing for the *conduction* of electricity by a substance is that the substance must have **mobile charged** particles, and metallic structures and both molten and dissolved ionic substances have mobile charged particles.
- In contrast to Diamond, Graphite **does** have mobile electrons between each covalently-bonded layer of giant molecular sheets of Carbon atoms. In the diagram below you will see that in every sheet, each C atom is bonded to only three other C atoms. That means there is a fourth delocalised electron connecting each C atom to a C atom in the sheet below – and it is this mobile charged particle that carries the electric current in Graphite. In Diamond, every C atom is bonded to 4 other C atoms and thus there are no “spare” mobile electrons to carry current.



- However, with a very few exceptions, giant **molecular** structures do not have **mobile** particles nor do they have any **charged** particles therefore they cannot conduct electricity, and simple molecular substances can have *mobile* particles (because on earth many of them are gases) but they have no *charged* particles – therefore, on the whole, simple molecular structures do not conduct electricity.
- It is safe to say that **molecular structures make poor conductors of electricity**. However,
 - i) if energy is put into a gas, then electrons can be stripped off the molecules and then the molecules become gaseous ions and this ionised molecular substance will now conduct electricity – otherwise molecular substances do not conduct electricity, and
 - ii) there are molecular substances such as Silicon that can be turned into semi-conductors (but semi-conductors are not part of the ‘A’ Level Syllabus).

C) Thermal Conductivity

- All the particles in a substance are in constant/continuous motion⁷ (except at 0K/-273.15°C) and the amount of motion is dependent on the amount of energy that any particular particle possesses (where the amount of energy can normally be measured by the temperature of the particle) – and if the particles in one part of a substance are moving faster than their neighbouring particles, then (through collisions) they pass on some of their energy to their neighbours, and the energy of those neighbours will increase until eventually all the particles in a substance possess almost exactly the same amount of energy, and the temperature of the body will be almost the same in every part of the body.⁸ In this context, a “particle” can be any species such as an electron/an atom/an ion/a molecule/etc. **Anything that possesses more energy than another particle with whom it collides or with whom it is in contact will pass some of its energy on to the other particle**, and this is why heat is always transferred from a hotter body to a colder body rather than the other way around (i.e. a hotter body will not suck heat out of a colder body).
- As I have stated it above, in contrast to Electrical Conductivity (where for example an absolutely pure piece of diamond will not conduct electricity) every substance will therefore conduct heat – except that in some substances heat will be conducted more easily/more readily than in other substances. For example, a piece of metal will conduct heat more readily than will a piece of wood or a piece of asbestos, and gases will conduct heat more readily than will liquids (where heat is *convected* more than it is *conducted*).
- In terms of the Periodic Table, Metals are extremely good conductors of heat whereas molecular structures are very poor conductors of heat - but all liquids and gases are good conductors of heat.

The Syllabus can ask about Oxides and you might want to read the great Jim Clark on

<https://www.chemguide.co.uk/inorganic/period3/oxidesphys.html>

but it has more detail than you need. You might also like to look at

<https://studylib.net/doc/7268351/physical-properties-of-the-period-3-oxides>

The following is **not** my writing but comes from

<http://www.a-levelchemistry.co.uk/topic-14---reactions-of-period-3-elements-and-their-oxides.html> .

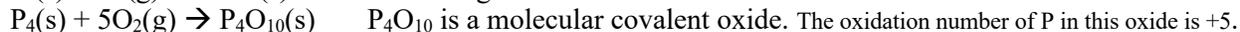
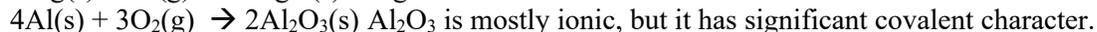
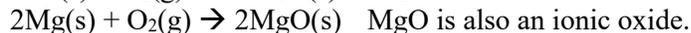
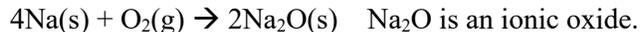
⁷ We will talk more about this next year when we talk about infra-red spectroscopy/NMR spectroscopy/and other types of methods of analysis.

⁸ Energy can be shared in many different ways, but I do not want to make the subject overcomplicated.

THE OXIDES OF PERIOD 3 ELEMENTS

1. Formation of oxides

All the elements in Period 3 except chlorine and argon combine directly with oxygen to form oxides.



Another oxide, SO_3 is formed in a reversible process when SO_2 and O_2 are heated with a V_2O_5 catalyst (the Contact Process).

2. Physical properties of oxides

The physical properties of these oxides depend on the type of bonding.

Na_2O , Al_2O_3 and MgO are ionic oxides and hence have a high melting point. MgO and Al_2O_3 have a higher melting point than Na_2O since the charges are higher, resulting in a stronger attraction between the ions.

SiO_2 has a giant covalent structure and hence a high melting point. There are strong covalent bonds between all the atoms and thus lots of energy is required to break them.

P_4O_{10} and SO_3 are molecular covalent and so only intermolecular forces exist between the molecules. The melting points are thus much lower. P_4O_{10} is a much bigger molecule than SO_3 and so has a much higher melting point, as the van der Waal's forces are stronger.

| Element | Na | Mg | Al | Si | P | S |
|----------------------------|-----------------------|--------------|-------------------------|----------------|---------------------------|--------------------|
| Formulae of oxide | Na_2O | MgO | Al_2O_3 | SiO_2 | P_4O_{10} | SO_3 |
| Structure of oxide | Ionic | Ionic | Mostly ionic | Giant covalent | Molecular covalent | Molecular covalent |
| Melting point of oxide /°C | 1275 | 2852 | 2072 | 1703 | 300 | -10 |

3. Acid-base character of oxides

Ionic oxides contain the O^{2-} ion. This is a strongly basic ion which reacts with water to produce hydroxide ions:



Thus all ionic oxides are BASIC.

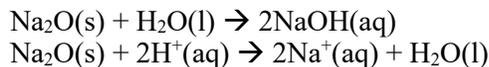
Covalent oxides do not contain ions, but have a strongly positive dipole on the atom which is not oxygen. This attracts the lone pair on water molecules, releasing H^+ ions:



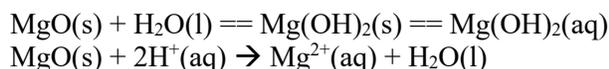
Thus all covalent oxides are ACIDIC.

Intermediate oxides can react in either of the above ways, depending on the conditions. They can thus behave as either acids or bases and are thus **AMPHOTERIC**.

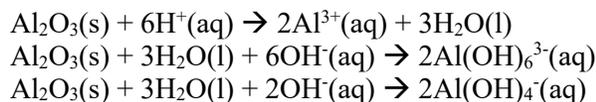
Na₂O is a **basic oxide**. It dissolves in water to give (depending on its concentration) an alkaline solution (with a pH = 14). It therefore reacts with acids:



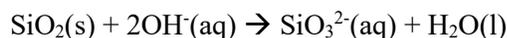
MgO is a **basic oxide**. It is only slightly soluble in water and so the solution is only slightly alkaline (with a pH = 9). It reacts readily with acids:



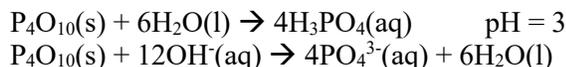
Al₂O₃ is an **amphoteric oxide**. It is insoluble in water (pH = 7) but dissolves in both acids and alkalis:



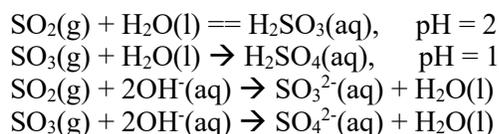
SiO₂ is an **acidic oxide**. It is insoluble in water (pH = 7) but dissolves in hot concentrated alkalis:



P₄O₁₀ is an **acidic oxide**. It dissolves in water to give acidic solutions and is also soluble in alkalis:



SO₂ and SO₃ are **acidic oxides**. They dissolve in water to give acidic solutions, and also react with alkalis:



SO₂ is a waste gas in many industrial processes. It is harmful because it dissolves in rain to give acid rain. It can be removed from waste gases because it dissolves in alkali and so it is passed through an alkaline solution in waste gas outlets to minimise the amount which escapes into the atmosphere.

The acid-base properties of the oxides of Period 3 can be summarised in the following table:

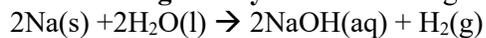
| Element | Na | Mg | Al | Si | P | S |
|--|-------------------|-------|--------------------------------|------------------|--------------------------------|--|
| Formulae of oxides | Na ₂ O | MgO | Al ₂ O ₃ | SiO ₂ | P ₄ O ₁₀ | SO ₂ SO ₃ |
| Acid-base character of oxide | Basic | Basic | Amphoteric | Acidic | Acidic | Acidic |
| pH of solution when dissolved in water | 12 - 14 | 8 - 9 | 7 (insoluble) | 7 (insoluble) | 2 - 4 | 2 - 4 (SO ₂) 1 - 3 (SO ₃) |

The oxides therefore become more acidic on moving from left to right in the periodic table.

THE REACTION OF PERIOD 3 ELEMENTS WITH WATER

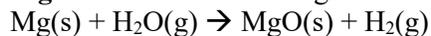
Na, Mg, Al and Si are more electropositive than H and can reduce the water to hydrogen gas:

Na reacts vigorously with water to give the hydroxide and hydrogen:



The resulting solution is strongly alkaline, and will have a pH of 14.

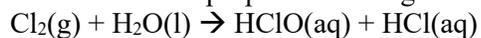
Mg reacts with steam to give the oxide and hydrogen:



The resulting solution is weakly alkaline, since the oxide is slightly basic (pH = 9).

Al and Si also react with steam under certain conditions.

P, S and Cl₂ do not reduce water to hydrogen gas. Phosphorus and Sulphur do not react with water but Chlorine will disproportionate to give an acidic solution:



The resulting solution contains HCl(aq) and is thus acidic (pH = 2).

The reactivity of the elements of period 3 towards water thus decreases from Na to Si, and then increases from P to Cl. The resulting solutions become increasingly acidic.