

## The consequences of Bonding and Structure **viz. Periodicity/Melting Points/Boiling Points/Thermal Conductivity/ Electrical Conductivity**

*[When we talk about Melting Points/Boiling Points/Thermal Conductivity/Electrical Conductivity/etc in this Chapter, I shall assume that you have re-read Chapters 19-24 inclusive and that you are thus familiar with all the concepts regarding the nature and strength of different types of bonds.]*

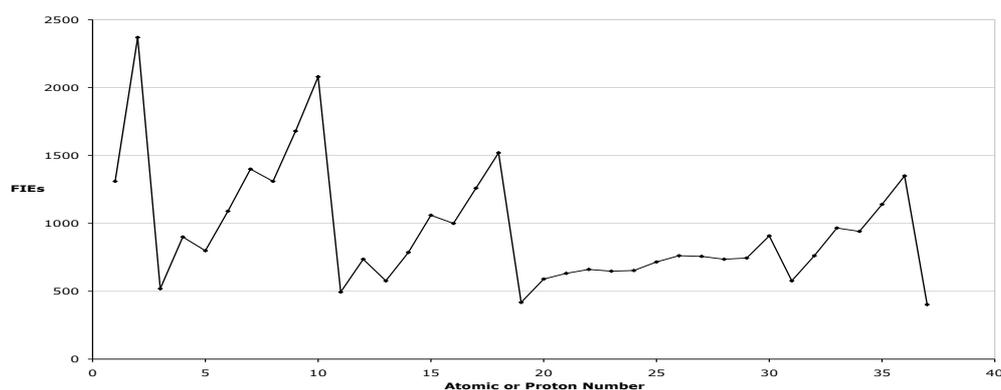
### GROUPS

- **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!**

All the Group I elements have	<b>1 electron</b> in their outermost shells
All the Group II elements have	<b>2 electrons</b> in their outermost shells
All the Group III elements have	<b>3 electrons</b> in their outermost shells
All the Group IV elements have	<b>4 electrons</b> in their outermost shells
All the Group V elements have	<b>5 electrons</b> in their outermost shells
All the Group VI elements have	<b>6 electrons</b> in their outermost shells
All the Group VII elements have	<b>7 electrons</b> in their outermost shells
All the Group VIII/Group "0" elements have	<b>8 electrons</b> in their outermost shells.

### PERIODS

- Today the English word "period" means an interval of time, and there is an implication of 'recurrence' over that period of time. Many years ago, the word "period" was used to indicate a regularly repeating pattern, and this word was therefore used to indicate that certain cycles/certain patterns were repeated regularly in the groupings of elements in the *Periodic* Table.
- Could you now please note carefully that although all the elements in a Group behave in a similar manner (because a vast number of chemical reactions are influenced and are even *determined* solely by the number of electrons present in the outermost shell of an atom); nevertheless, **the elements in any one particular Period do not (I repeat do NOT) all behave in a similar manner**, and instead **it is the pattern of behaviour exhibited across the elements in a Period that is repeated again and again throughout all the Periods** (i.e. it is the pattern of changes in structure/the changes in melting points/the changes in conductivity/etc across one Period that are repeated across other Periods).



- Do you know why there is a fall in FIE between Mg (12) and Al (13), and between Phosphorous (15) and Sulphur (16) – if you need reminding, there are always four factors to consider viz. distance/number of protons/shielding and mutual repulsion.
- 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.

## METALS vs NON-METALS

- The most important grouping in the Periodic Table is that of 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** electrostatic 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 such as one of the three van der Waals' forces of attraction.*]
- The most important characteristic of a metal is that **when a compound is formed between a metal and a non-metal, then the atoms of the metal give up electron(s)<sup>1</sup> to the atoms of the non-metal element**. A reasonable description of a piece of metal would be **“a giant lattice structure of positively charged metal ions of the metal set in a sea of delocalised electrons with an ion and its delocalised electrons together constituting an atom of the metal, and where the whole structure is held together by metallic bonding”<sup>2</sup>**.

## 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.

*What sort of bonding:* Electrovalent / ionic.

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<sup>1</sup> If you have not encountered this usage [i.e. “electron(s)”] before, then it just means “one or more electrons”.

<sup>2</sup> “Set” means situated. It does not mean “set” in any one specific location.

## 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**

**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 **inter-molecular** forces of attraction. They are

#### **a) Hydrogen bonds**

**b) van der Waals'** forces of attraction viz.

- b1) Permanent dipole to Permanent dipole
- b2) Permanent dipole to Induced dipole, and
- b3) Instantaneous dipole to Induced dipole forces of attraction.

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

## THE CHARACTERISTICS/PROPERTIES OF METALS

- Metals can be elements (or compounds/or alloys of metal elements), and on the whole metal elements
  - 1 have relatively high melting and boiling points
  - 2 are good conductors of electricity and heat
  - 3 are malleable, ductile and lustrous (i.e. they can be made to shine)
  - 4 are (*other than for Mercury*)<sup>3</sup> usually highly dense solids at room temperature, and
  - 5 metals in the same Group
    - have similar reactions with water
    - have similar Oxidisation Numbers
    - have similarly coloured compounds.
  - 6 Metals are significantly less electronegative than non-metal elements in the same Period, and
  - 7 the **Oxides of metals** are "basic" (e.g. under the Bronsted-Lowry definition they are **proton acceptors** with proton donating substances being classified as "acidic") whilst the **Oxides of non-metal elements** are **acidic** (i.e. they are proton donors).
- Let us now talk about the Periodic trends in differing elements with regard to
  - A) Melting Points and Boiling Points
  - B) Electrical Conductivity, and
  - C) Thermal Conductivity.

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<sup>3</sup> Caesium and Gallium are liquid at just above RTP.

## 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), and **Boiling** refers to the process whereby a substance, when it *boils*, goes from the liquid state (l) to the gaseous state (g)<sup>4</sup>, 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).
- OK, since we know that there are extremely strong bonds in pieces of metal/and in ionic substances/and in **intra**-molecular covalent bonds, then 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.
- In contrast, **inter**-molecular bonds are very weak, cf. page 3, therefore simple molecular structures have very low MPs and BPs, and become gases at very low temperatures. They are therefore 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** bonds<sup>5</sup>.

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

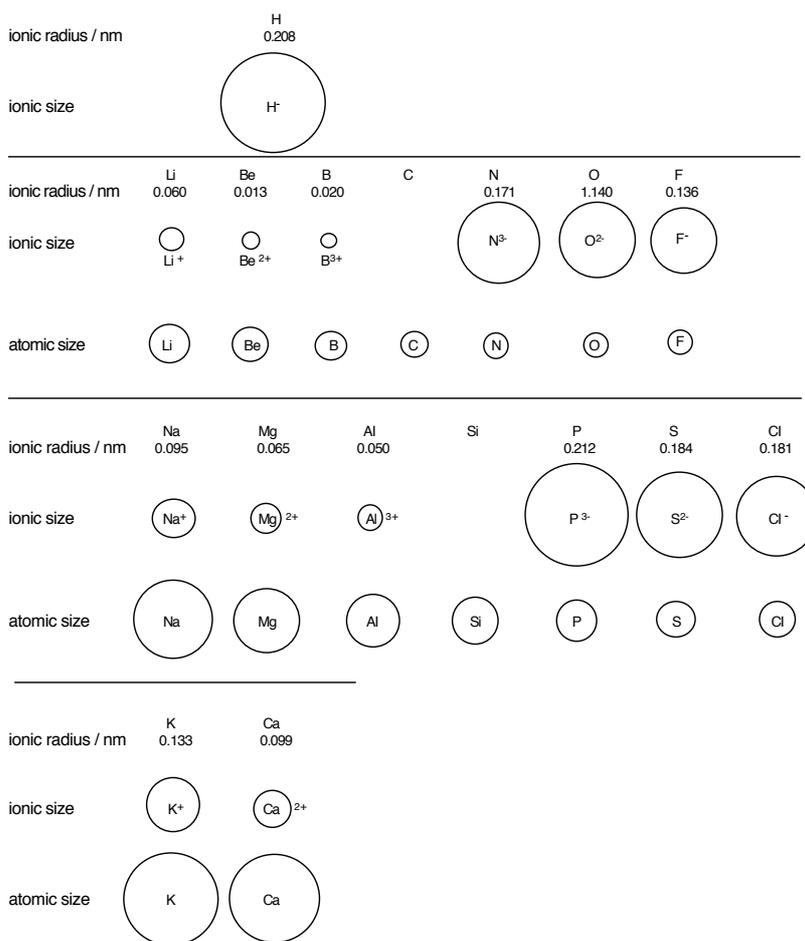
- 1) **More bonds** : The more electrons that there are in the pool of delocalised electrons the more bonds there will be, and the more bonds that there are then the greater will be the amount of energy required to break those bonds and to melt and subsequently to boil the piece of metal. It is then 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 that follows 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.

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<sup>4</sup> 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.

<sup>5</sup> 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!**

## Chemistry in Context



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

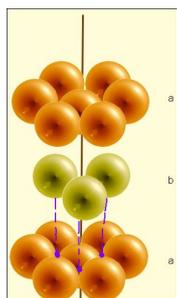
## INCREASING ATOMIC RADIUS

← INCREASING ATOMIC RADIUS																					
1 <b>H</b> Hydrogen 1.00794																	2 <b>He</b> Helium 4.003				
3 <b>Li</b> Lithium 6.941	4 <b>Be</b> Beryllium 9.012182															5 <b>B</b> Boron 10.811	6 <b>C</b> Carbon 12.01107	7 <b>N</b> Nitrogen 14.00674	8 <b>O</b> Oxygen 15.9994	9 <b>F</b> Fluorine 18.998432	10 <b>Ne</b> Neon 20.1797
11 <b>Na</b> Sodium 22.989770	12 <b>Mg</b> Magnesium 24.3050															13 <b>Al</b> Aluminum 26.981538	14 <b>Si</b> Silicon 28.0855	15 <b>P</b> Phosphorus 30.973761	16 <b>S</b> Sulfur 32.066	17 <b>Cl</b> Chlorine 35.4527	18 <b>Ar</b> Argon 39.948
19 <b>K</b> Potassium 39.0983	20 <b>Ca</b> Calcium 40.078	21 <b>Sc</b> Scandium 44.955910	22 <b>Ti</b> Titanium 47.867	23 <b>V</b> Vanadium 50.9415	24 <b>Cr</b> Chromium 51.9961	25 <b>Mn</b> Manganese 54.938049	26 <b>Fe</b> Iron 55.845	27 <b>Co</b> Cobalt 58.933200	28 <b>Ni</b> Nickel 58.6934	29 <b>Cu</b> Copper 63.546	30 <b>Zn</b> Zinc 65.39	31 <b>Ga</b> Gallium 69.723	32 <b>Ge</b> Germanium 72.61	33 <b>As</b> Arsenic 74.92160	34 <b>Se</b> Selenium 78.96	35 <b>Br</b> Bromine 79.904	36 <b>Kr</b> Krypton 83.80				
37 <b>Rb</b> Rubidium 85.4678	38 <b>Sr</b> Strontium 87.62	39 <b>Y</b> Yttrium 88.90585	40 <b>Zr</b> Zirconium 91.224	41 <b>Nb</b> Niobium 92.90638	42 <b>Mo</b> Molybdenum 95.94	43 <b>Tc</b> Technetium (98)	44 <b>Ru</b> Ruthenium 101.07	45 <b>Rh</b> Rhodium 102.90550	46 <b>Pd</b> Palladium 106.42	47 <b>Ag</b> Silver 107.8682	48 <b>Cd</b> Cadmium 112.411	49 <b>In</b> Indium 114.818	50 <b>Sn</b> Tin 118.710	51 <b>Sb</b> Antimony 121.760	52 <b>Te</b> Tellurium 127.60	53 <b>I</b> Iodine 126.90447	54 <b>Xe</b> Xenon 131.29				
55 <b>Cs</b> Cesium 132.90545	56 <b>Ba</b> Barium 137.327	57 <b>La</b> Lanthanum 138.9055	72 <b>Hf</b> Hafnium 178.49	73 <b>Ta</b> Tantalum 180.9479	74 <b>W</b> Tungsten 183.84	75 <b>Re</b> Rhenium 186.207	76 <b>Os</b> Osmium 190.23	77 <b>Ir</b> Iridium 192.222	78 <b>Pt</b> Platinum 195.078	79 <b>Au</b> Gold 196.96655	80 <b>Hg</b> Mercury 200.59	81 <b>Tl</b> Thallium 204.3833	82 <b>Pb</b> Lead 207.2	83 <b>Bi</b> Bismuth 208.98038	84 <b>Po</b> Polonium (209)	85 <b>At</b> Astatine (210)	86 <b>Rn</b> Radon (222)				
87 <b>Fr</b> Francium (223)	88 <b>Ra</b> Radium (226)	89 <b>Ac</b> Actinium (227)	104 <b>Rf</b> Rutherfordium (261)	105 <b>Db</b> Dubnium (262)	106 <b>Sg</b> Seaborgium (263)	107 <b>Bh</b> Bohrium (262)	108 <b>Hs</b> Hassium (265)	109 <b>Mt</b> Meitnerium (266)	110	111 (269)	112 (272)	113 (273)	114								

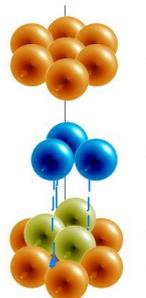
Source : University of California, Davis

3) **Shorter bonds with regard to differing Periods** : Equally, atomic radii increase going downwards in a Group in the Periodic Table<sup>6</sup> 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.**<sup>7</sup>
- Actually, 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. I do not know of any two-dimensional teaching source that teaches Packing well, but all that you need to know at ‘A’ Level is that there are different types of packing and that the more tightly packed metallic ions are, then the closer they will be to each other.



Cubic Close (abcabc) Packing



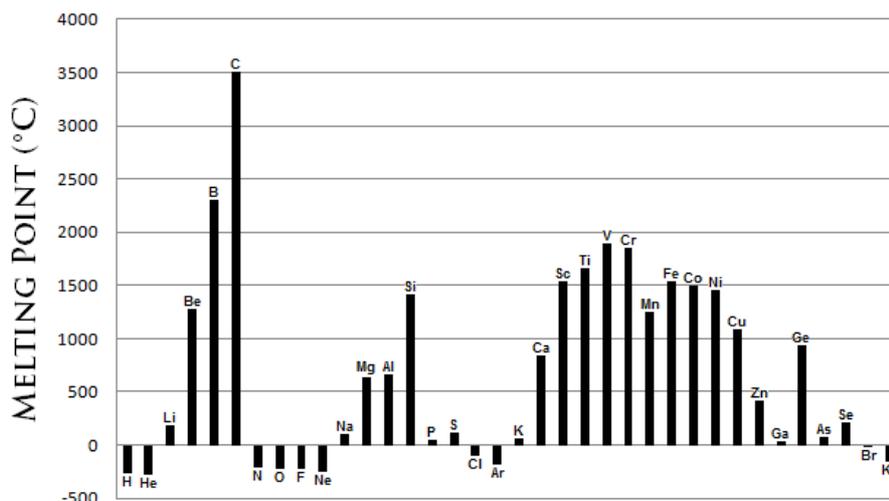
Hexagonal Close (ababab) Packing

Source : UCD

- Jim Clark makes the following observation “Beryllium and Magnesium are both hexagonal close-packed; Calcium and Strontium are face-centred cubic; Barium is body-centred cubic (like Group I metals)”.

## 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.



Source : University of California, Davis

<sup>6</sup> As the Principal Quantum Energy Level goes from  $n=1$  to  $n=2$  to  $n=3$  and so on.

<sup>7</sup> 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.

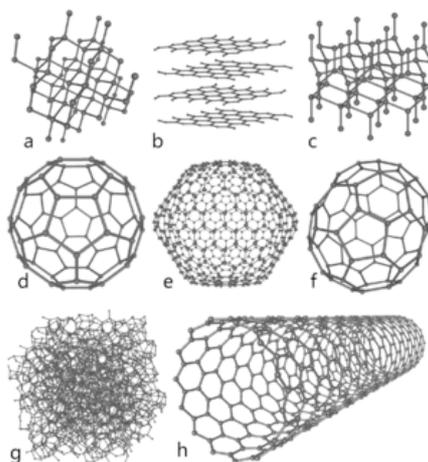
- The preceding table shows the Melting Points of the first 36 elements, and from it you will see that Metal elements have high Melting Points but that, in contrast, those Non-Metal elements that have **simple molecular structures** have low Boiling Points – and, indeed, at 0°C/273K most of the Non-Metal elements are gases.<sup>8</sup> 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.**

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 <sup>-1</sup>
Hydrogen “bonds”	about 10-50 kJ mol <sup>-1</sup>
Permanent dipole to Permanent dipole	about 5 kJ mol <sup>-1</sup>
Permanent dipole to Induced dipole	about 3 kJ mol <sup>-1</sup>
Instantaneous dipole to Induced dipole	about 1 kJ mol <sup>-1</sup>

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.

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

- As we have seen, 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.
- Allotropes are differing structural forms of an element, and the diagram below shows eight different allotropes of Carbon viz. (a) diamond (b) graphite (c) lonsdaleite (d) buckminsterfullerene (e) C<sub>540</sub> (f) C<sub>70</sub> and (g) amorphous Carbon (i.e. very small particles of graphite) and (h) nanotube Carbon



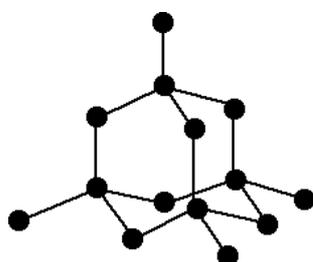
Source : Wikipedia

and at ‘A’ Level you are required to know about only diamond and graphite.

<sup>8</sup> 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 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’.)

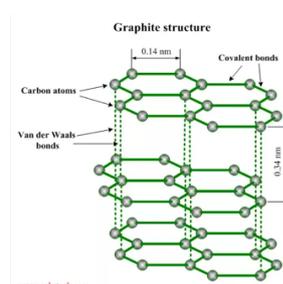
- Every single Carbon atom in diamond is linked by a covalent bond to four other Carbon atoms which in turn are linked to four other Carbon atoms by covalent bonds and they in turn are linked to four other C atoms by covalent bonds ..... and so on. It is hardly surprising therefore that (elemental carbon) diamond has the highest Melting Point of all the elements – and this is entirely to do with the macromolecular structure of Carbon (which gives a much higher internal bond strength than any ionic or metallic substance).
- Group IV elements form more bonds (i.e. 4 rather than 1, or 2, or 3 bonds) than the elements in any other Group in the Periodic Table, and the shorter the distance between the shared pair (or pairs) of electrons and the nuclei of the atoms that are bonded together, then the stronger will those bonds be. Carbon is the topmost element in Group IV in the Periodic Table, therefore the outermost electrons in a Carbon atom will be **closer** to the nucleus of the Carbon atom than will be the case for any other element in Group IV – and therefore Carbon will have the shortest bonds of any element in Group IV. Group IV elements have similar structures to Carbon, but Carbon is the topmost element in Group IV therefore its bonds are shorter than any other Group IV element.

### The structure of Diamond



Source : Jim Clark

### The structure of Graphite



Source : substech.com

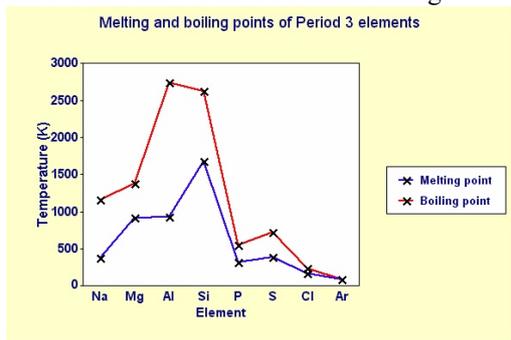
- However, Carbon in the form of Graphite has a slightly different structure and is layered, and each layer is stacked one on top of another (with slightly different stacking patterns between different forms of graphite), and in each layer each C atom is covalently bonded to three other C atoms – and it is the fact that in diamond each C atom is covalently bonded to **four** other C atoms that gives diamond a higher Melting Point than Graphite.<sup>9</sup>
- 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. I do not know of any ‘*a priori*’ way of working out the structure of an element, and as far as I know the only way to know it is to learn the structures off by heart. Hydrogen and Helium both have simple molecular structures, and then the structure of the next few elements is

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

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

<sup>9</sup> Carbon has such a high Melting Point that blocks of Carbon are actually used to line furnaces as refractory linings. I could not believe it when I first started visiting steel plants, but that actually is the case (and **they do not catch fire** because **you need Oxygen for Combustion** and Oxygen is excluded in such furnaces).

- OK, we have now talked enough about how the differences in Melting Points and Boiling Points are related to the Periodic Table, so let us talk a bit about Electrical Conductivity. However, before we do so, you might want to see a graphic representation of what we have been discussing.



Source: creative-chemistry.org

## B) Electrical Conductivity

- Let us now consider an explanation for electrical conductivity. The critical thing with each of those pieces 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, are delocalised electrons that wander around freely in a piece of metal.
- If every piece of metal has delocalised electrons in it, and if *the number of delocalised electrons that wander around freely through the piece of metal is directly related to the number of metal cations in the piece of metal*, then if the piece of metal consisted of an insulated wire of an element and theoretically if we inserted an electron into one end of the wire, then it would not be at all surprising if that action caused an electron to pop out of the other end of the wire – and, in essence, that is how an electrical current flows through a piece of metal i.e. electrons enter one end of the piece of metal and displaced electrons exit the other end of the piece of metal – and it is the fact that the piece of metal has delocalised electrons in it that allows this to happen. Clearly, in practice 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, “ $\delta^+$ ” is meant to represent a metal ion, and “ $e^-$ ” is meant to represent an electron. (Actually, the “ $\delta^+$ ” should in fact have been a “+”, because if a species has lost an electron then there is no partial separation of charge involved, but instead there is a total separation of charge into ions.)



Source: unknown

- Now, 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 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 metallic structures “**have no mobile charged particles in them**”. Even if a substance has charged particles in it but the particles are **not** mobile, then electric current cannot be carried within/by the substance. For example, solid ionic substances do have charged particles in them, but they cannot carry an electric current because the charged particles are **not mobile** and instead they are locked into giant lattice structures! However, when an ionic substance is

melted and has thus become a liquid then it can carry an electric current because the charged particles are now indeed **mobile** – and the same is true when an ionic substance has become dissolved in a solvent, because once more the charged particles in the ionic substance are mobile (they are free to move around in the solvent) and thus they can carry a current of electricity.

- We have now seen that 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. 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 probably safer 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 motion<sup>10</sup> (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.<sup>11</sup> 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/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 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>

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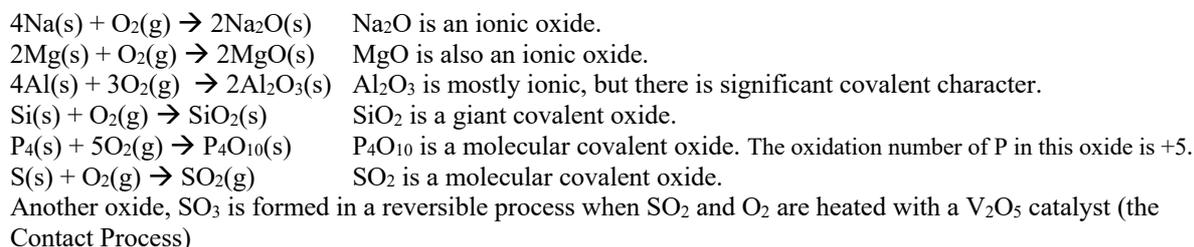
<sup>10</sup> We will talk more about this next year when we talk about infra-red spectroscopy/NMR spectroscopy/and other types of methods of analysis.

<sup>11</sup> 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.



### 2. Physical properties of oxides

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

$\text{Na}_2\text{O}$ ,  $\text{Al}_2\text{O}_3$  and  $\text{MgO}$  are ionic oxides and hence have a high melting point.  $\text{MgO}$  and  $\text{Al}_2\text{O}_3$  have a higher melting point than  $\text{Na}_2\text{O}$  since the charges are higher, resulting in a stronger attraction between the ions.

$\text{SiO}_2$  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.

$\text{P}_4\text{O}_{10}$  and  $\text{SO}_3$  are molecular covalent and so only intermolecular forces exist between the molecules. The melting points are thus much lower.  $\text{P}_4\text{O}_{10}$  is a much bigger molecule than  $\text{SO}_3$  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	$\text{Na}_2\text{O}$	$\text{MgO}$	$\text{Al}_2\text{O}_3$	$\text{SiO}_2$	$\text{P}_4\text{O}_{10}$	$\text{SO}_3$
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  $\text{O}^{2-}$  ion. This is a strongly basic ion which reacts with water to produce hydroxide ions:

$$\text{O}^{2-}\text{(aq)} + \text{H}_2\text{O(l)} \rightarrow 2\text{OH}^-\text{(aq)}$$

**Thus all ionic oxides are BASIC.**

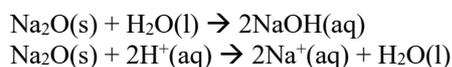
Covalent oxides do not contain ions, but have a strongly positive dipole on the atom which is not oxygen. This attracts the lone pair on water molecules, releasing  $\text{H}^+$  ions:



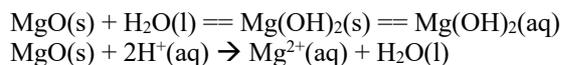
**Thus all covalent oxides are ACIDIC.**

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**.

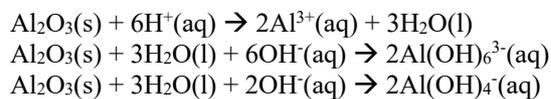
$\text{Na}_2\text{O}$  is a **basic oxide**. It dissolves in water to give an alkaline solution (pH = 14). It also reacts with acids:



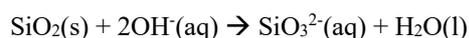
MgO is a **basic oxide**. It is only slightly soluble in water and so the solution is only slightly alkaline (pH = 9). It reacts readily with acids:



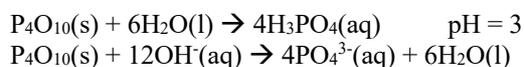
Al<sub>2</sub>O<sub>3</sub> is an **amphoteric oxide**. It is insoluble in water (pH = 7) but dissolves in both acids and alkalis:



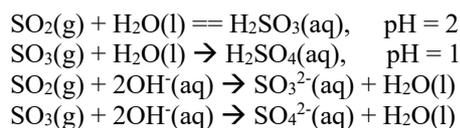
SiO<sub>2</sub> is an **acidic oxide**. It is insoluble in water (pH = 7) but dissolves in hot concentrated alkalis:



P<sub>4</sub>O<sub>10</sub> is an **acidic oxide**. It dissolves in water to give acidic solutions and is also soluble in alkalis:



SO<sub>2</sub> and SO<sub>3</sub> are **acidic oxides**. They dissolve in water to give acidic solutions, and also react with alkalis:



SO<sub>2</sub> is a waste gas in many industrial processes. It is harmful because it dissolves in rain water 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:

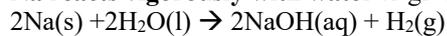
Element	Na	Mg	Al	Si	P	S
Formulae of oxides	Na <sub>2</sub> O	MgO	Al <sub>2</sub> O <sub>3</sub>	SiO <sub>2</sub>	P <sub>4</sub> O <sub>10</sub>	SO <sub>2</sub> SO <sub>3</sub>
Acid-base character of oxide	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic
pH of solution when dissolved in water	12 - 14	8 - 9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

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

### THE REACTION OF PERIOD 3 ELEMENTS WITH WATER

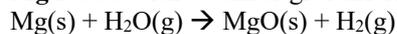
Na, Mg, Al and Si are more electropositive than H and can reduce the water to hydrogen gas:

**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 Cl<sub>2</sub> 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.