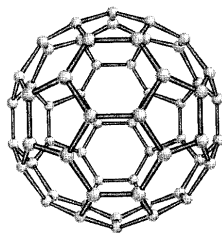
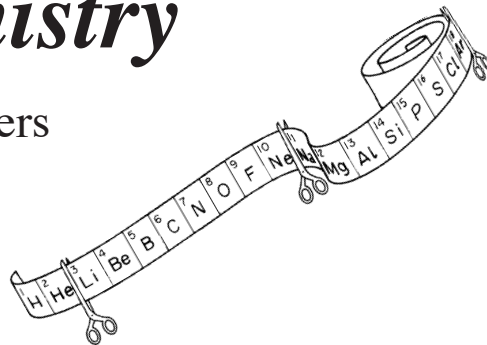


# Higher Chemistry

## Unit 1: Energy Matters



### Section 1:



# Elements & Bonding

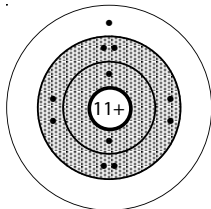
Student:

Lesson	Activities	Done	Checked
1.1 <i>Metallic Elements</i>	1. Metallic Lattice		
	2. Physical Properties		
	3. Boiling Points		
	Check Test	Score: /	
	Home Practice	Score: /	
1.2 <i>Monatomic Elements</i>	1. Very Weak Attractions		
	2. Explanation for Van Der Waal's Forces		
	3. Noble Gas Series		
	Check Test	Score: /	
	Home Practice	Score: /	
1.3 <i>Molecular Elements</i>	1. Diatomic Elements		
	2. Larger Covalent Molecules		
	3. Molecular Carbon (Fullerenes)		
	Check Test	Score: /	
	Home Practice	Score: /	
1.4 <i>Covalent Network</i>	1. Diamond Structure		
	2. Networks and Melting Points		
	3. Graphite Structure		
	Check Test	Score: /	
	Home Practice	Score: /	
1.5 <i>Atomic Size</i>	1. Atomic Radius		
	2. Atomic Size and the Periodic Table		
	3. Atomic Size and Density		
	Check Test	Score: /	
	Home Practice	Score: /	
1.6 <i>Ionisation Energies</i>	1. First Ionisation Energy		
	2. Across a Period and Down a Group		
	3. Second Ionisation Energy		
	Check Test	Score: /	
	Home Practice	Score: /	
1.7 <i>Periodic Pattern</i>	1. Periodic Law		
	2. Mendeleev's Table		
	3. Bonding and Structure		
	Check Test	Score: /	
	Home Practice	Score: /	
<i>Consolidation Work</i>	Consolidation A	Score: /	
	Consolidation B	Score: /	
	Consolidation C	Score: /	
	Consolidation D	Score: /	
	<b>End-of-Section Assessment</b>	Score:	
		%	Grade:

## 1.1 Metallic Elements

This first topic looks at **bonding** in metals and how this helps explain some of the typical **properties** of a metal.

### Metallic Bond



The **inner electrons** of a metal atom are tightly held by the **nucleus**. They are not able to escape from the atom. They are said to be **localised**.

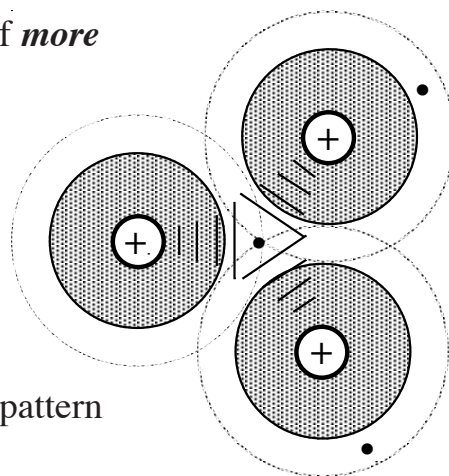
The **outermost electrons** are also strongly attracted by the **nucleus**, but they are **free to move** from atom to atom. They are said to be **delocalised**.

Each delocalised electron can be **attracted** to the **nucleus** of **more than one** neighbouring atom.

The **metallic bond** is the electrostatic **attraction** between the **negatively** charged outermost electrons and the **positively** charged nuclei of neighbouring atoms.

The **metallic bond** is usually **strong** and **acts in all directions** around each atom.

In a solid metal, the atoms are **arranged** neatly in a simple pattern known as a **lattice**



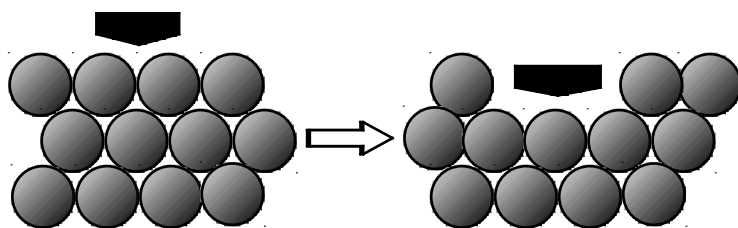
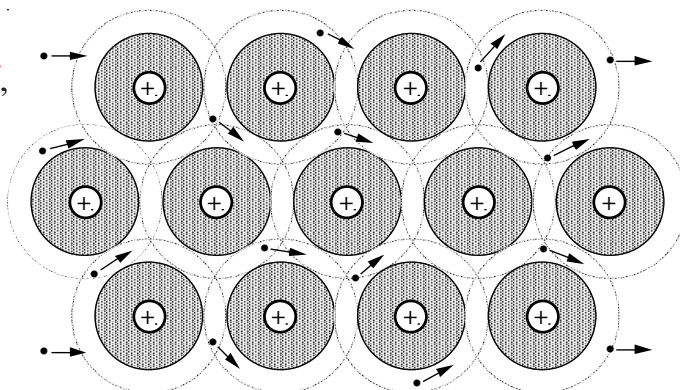
### Physical Properties

*The Physical properties of a substance are those properties which do not involve chemical reactions.*

Metals make **good conductors**.

Since the outermost electrons are **delocalised**, it is easy to make them all move in the **same** direction.

Once a **voltage** is applied across the metal, the **electrons** will move away from the **negative** electrode towards the **positive** electrode.

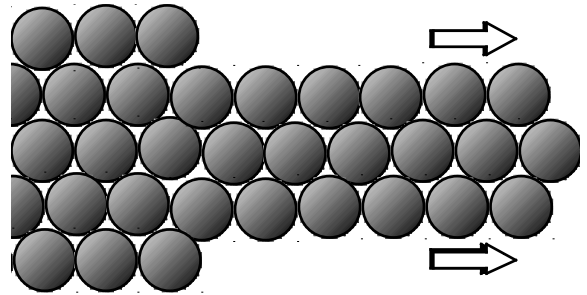


*The shapes of metals can be altered by beating them (horseshoes) or by rolling into sheets (tin foil). This property is called **malleability**.*

**Metallic** bonding is very 'fluid'. This makes it easy for metal **atoms** to be slid around each other and yet remain strongly **bonded** together. Metals are described as **malleable**.

Metals can be drawn into thin wires by stretching. This property is called **ductility**.

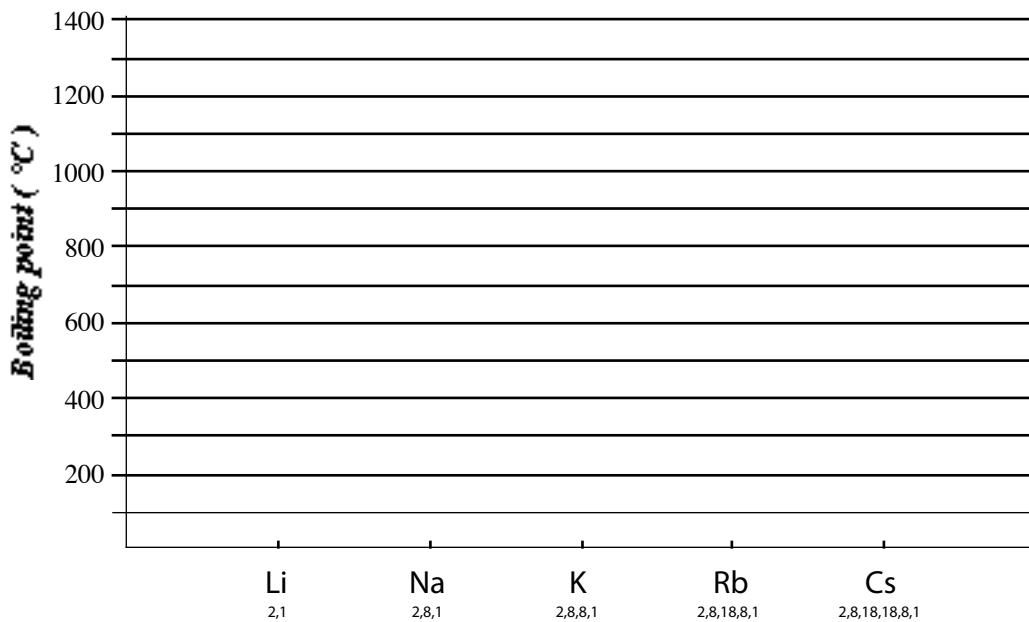
Again, it is because it is easy for metal atoms to **slide** around each other while staying strongly **bonded** to each other. Metals are described as **ductile**.



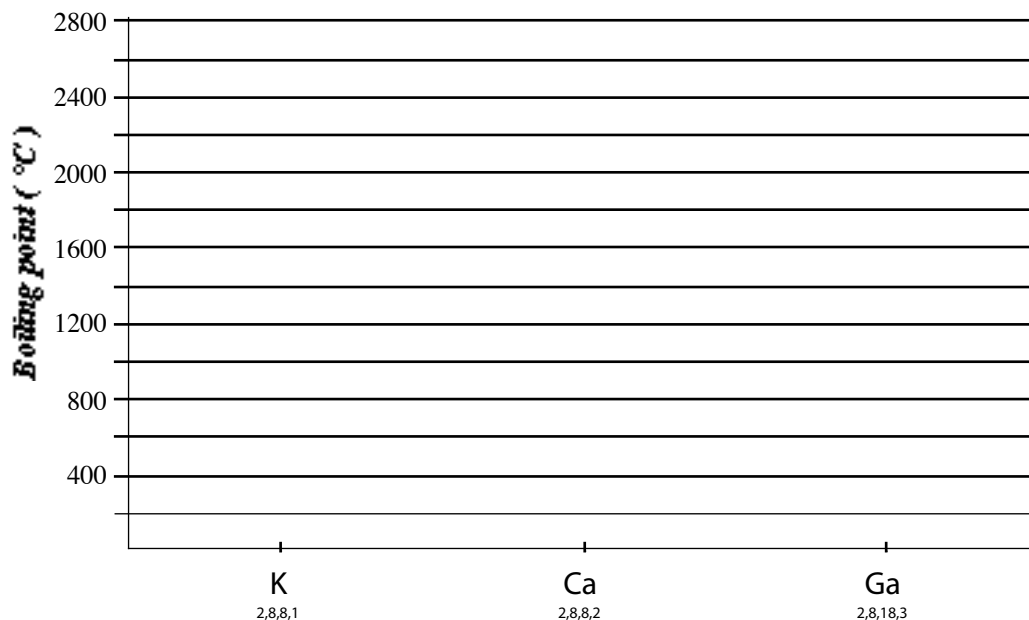
**Boiling Points**

To boil a metal, and allow its atoms to move around freely as a gas, the metallic bond must be overcome.

The Boiling Points of metals are a guide to the relative strengths of the metallic bonds holding the atoms together.



As you go **down** Group 1, the outermost **electron** is further and further away from the **nucleus**. The strength of **attraction** between electrons and nuclei will, therefore, **decrease**. This makes it easier for the metal atoms to move further apart as shown by the **decrease** in Boiling points.



As you go **across** the fourth period, the outermost **electron** are always to be found in the **fourth** shell, at a similar **distance** from the **nucleus**. The **number of electrons** is, however, **increasing**.

Since the Boiling points are **increasing**, the **bond strength** must **increase** as the **number** of outermost electrons in a metal **increases**.

*The strength of the Metallic bond decreases as the outermost electrons become further from the nucleus*

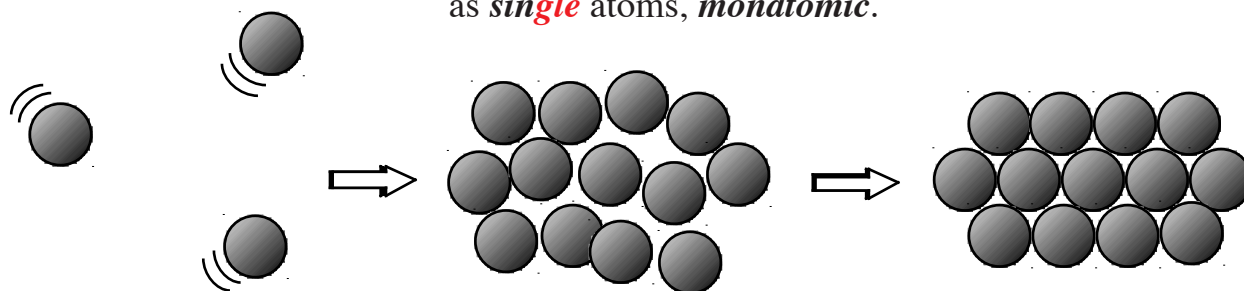
*increases as the number of outermost electrons increase*

## 1.2 Monatomic Elements

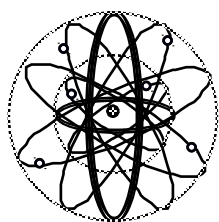
*This topic is about the monatomic elements and the very weak type of bonding force that exists between atoms of these elements.*

### Van Der Waals Forces

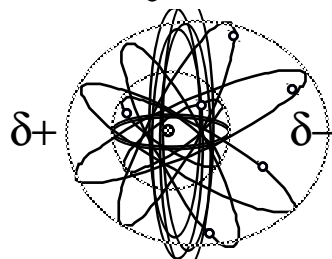
The **Noble** gases are the only elements normally found as **single** atoms, **monatomic**.



At **extremely low temperatures** elements like Neon become **liquid** or even **solid**, like every other substance. At these low temperatures, there must be some **weak forces** holding these atoms together. Very little **energy** is needed to overcome these weak forces, as the low Melting and Boiling Points for the Noble gases demonstrate. *These very weak bonding forces are called Van der Waals Forces.*

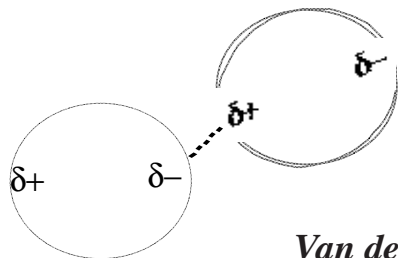


Normally, the **electrons** are **spread evenly** within the set of electron clouds (**orbitals**) that make up an electron **shell**. The centre of **negative** charge and the centre of **positive** charge will coincide at the centre of the atom.



*Uneven distribution of the **electrons** causes the electron cloud to 'wobble' and the centre of **negative** charge no longer coincides with the **positive** charge. One side of the atom is temporarily slightly negative ( $\delta^-$ ), while the other side becomes temporarily positive ( $\delta^+$ ). We call this a temporary **dipole**.*

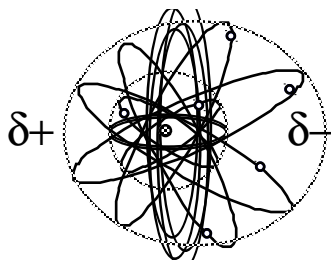
The **dipole** is only **temporary** as electron repulsions quickly act to ensure a more **even distribution**. However, the brief **wobble** on one atom can affect the **electron** clouds on any nearby atoms.



This causes, or *induces*, a temporary *dipole* in the nearby atom. Short lived *attractions* will then exist between the slightly *negative* ( $\delta^-$ ) side of one atom and the slightly *positive* ( $\delta^+$ ) side of the other atom.

*Van der Waals forces* are the shortlived attractions that exist between a slight charge on one atom (*temporary dipole*) and the opposite charge set up on a nearby atom (*induced dipole*).

### Noble Gas Series

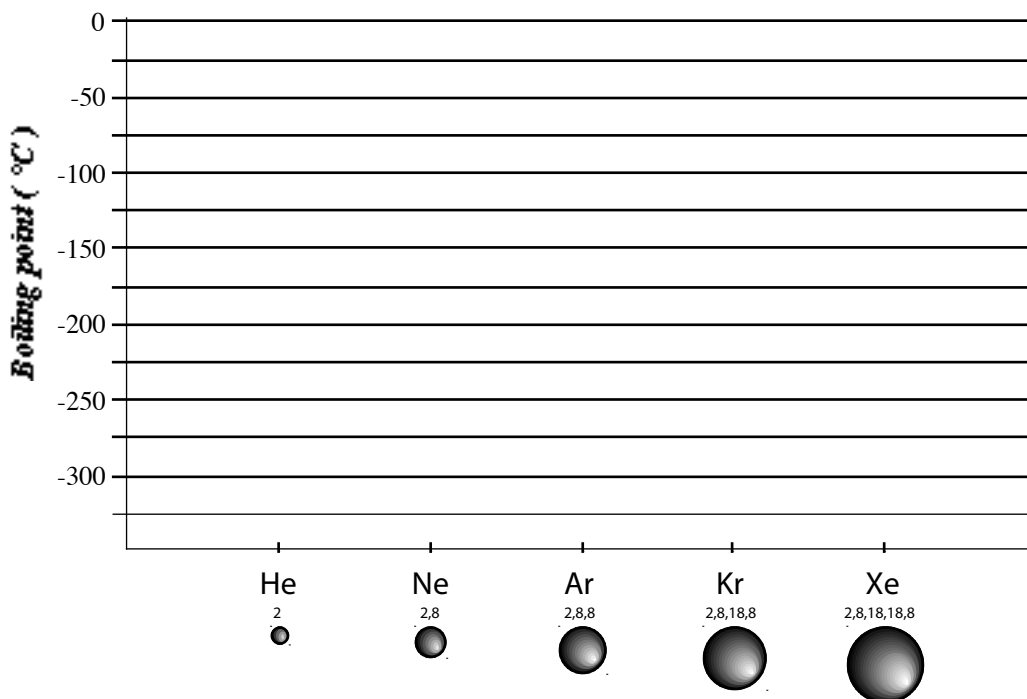


Every atom has a *nucleus* and clouds of *electrons* around it. So all atoms have *Van der Waals forces* acting between them (as long as they are close enough to each other).

They are such *small* forces that they can usually be ignored, especially if other, much *stronger*, bonding forces are present, e.g. *metallic* bonds.

However, for the noble gases, *Van der Waals forces* are the *only* type of bonding force present.

So the melting points and boiling points of the noble gases depend entirely on the strength of the *Van der Waals forces* between their atoms.



The *larger* an atom is, the *further* out its outermost *electrons*, the more likely it is that the electron clouds will be able to ‘wobble’ and set up a *temporary dipole*.

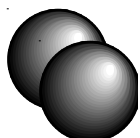
The *larger* the dipole set up, the *stronger* the *Van der Waals forces* that will exist. As the *attractions* between atoms increase, the *energy* needed to move the atoms further apart will *increase* as shown by the *increase* in Boiling points.

*Van der Waals forces* are very small, but they are larger between bigger atoms.

# 1.3 Molecular Elements

This topic is about the elements which exist normally as discrete molecules.

## Diatomic elements

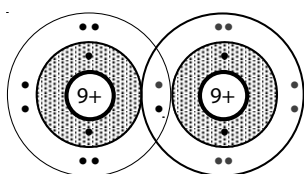


Molecular elements are elements made up of individual molecules i.e. distinct groups of atoms joined together by **covalent bonds**.



The simplest molecular elements are **diatomic** i.e. each molecule is made of just **two atoms**.

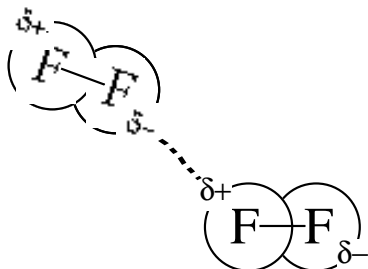
Substances which are molecular have **more than one type** of bonding present.



Within each molecule the atoms are **strongly** held together by a **covalent** bond.

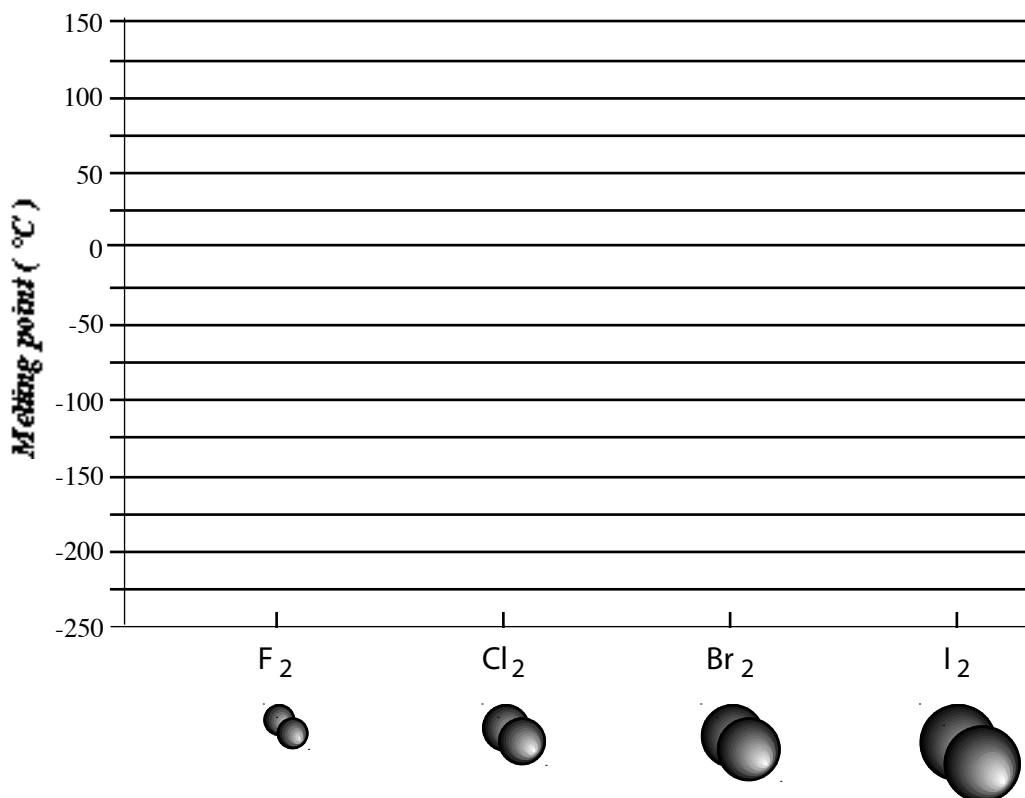
A **covalent bond** is the electrostatic attraction between a **shared pair of electrons** and the nuclei of the two sharing atoms.

Usually, the **covalent bond** is represented by a line.



Between the **molecules** of fluorine there are the same **weak** forces that exist between atoms - **Van der Waals forces**.

Though very weak, these **intermolecular forces** can still determine **physical** properties such as **melting** and **boiling** points.

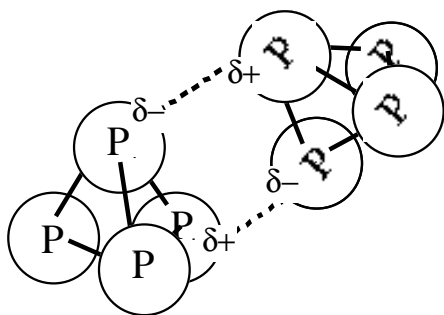
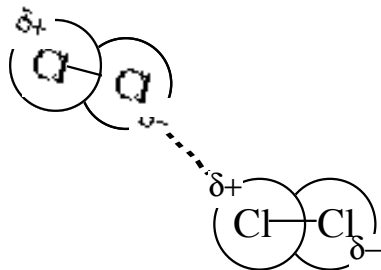


The **larger** a molecule is, the **further** out the outermost **electrons**, the more likely it is that the electron clouds will be able to 'wobble' and set up a **temporary dipole**.

The **larger** the dipole set up, the **stronger** the **Van der Waals** forces that will exist.

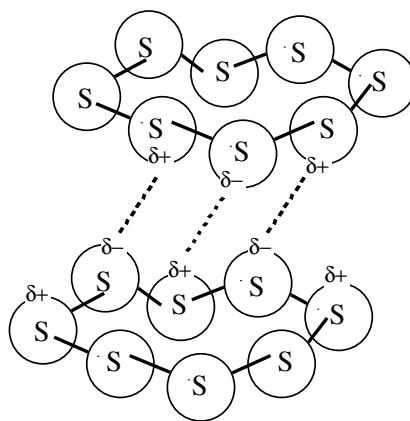
As the **attractions** between molecules increase, the **energy** needed to move the molecules further apart will **increase** as shown by the **increase** in Melting points.

### Larger Covalent Molecules



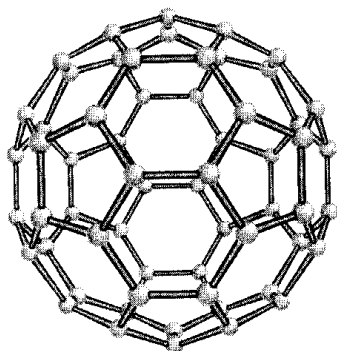
Phosphorus molecules have formula =  $P_4$   
and a melting point =  $44.1$  °C.

Sulphur molecules have formula =  $S_8$   
and a melting point =  $112.8$  °C.



Chlorine, phosphorus and sulphur atoms are of a **similar size** so are likely to produce **similar dipoles**, and **similar Van der Waals attractions**. However, with **more atoms** in a molecule, a **larger** number of Van der Waals attractions are possible.

In all these examples, it is important to remember that there is no need to break the very strong **covalent** bonds **within** these molecules.



Larger molecules like **fullerene** will have even **higher** melting points, though, they are still relatively low compared to other substances which have **stronger** bonds than Van der Waals forces to be broken.

The strength of Van der Waals forces

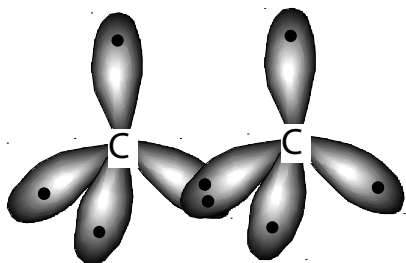
increases as the size of the atom increases

increases as the number of atoms in the molecule increases

## 1.4 Covalent Network Elements

This topic is about elements where the atoms are joined together by strong covalent bonds to form vast networks of atoms.

### Diamond Structure



**Carbon** atoms have 4 *unpaired electrons* which allows them to form a total of **4 bonds**.

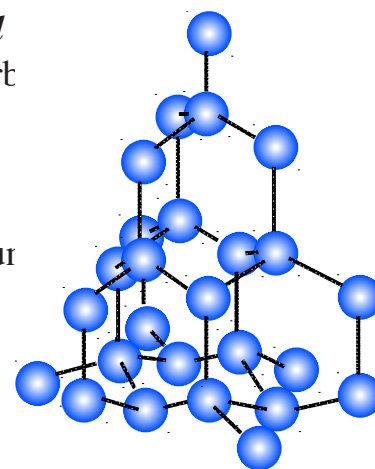
Each **carbon** can bond with another 3 carbons each, and so on until a huge **network** of **covalently** bonded atoms has been formed.

The arrangement of carbon atoms can be described as **tetrahedral** because of the position of the 4 carbon atoms around a central carbon atom.

Rings of 6 carbon atoms, **hexagons**, can also be seen.

**Diamonds** are not malleable - the carbon atoms can not slide around each other without breaking **strong covalent** bonds.

Diamonds are strong and hard since many **strong covalent** bonds would need to be overcome if a diamond is to be cut or broken.

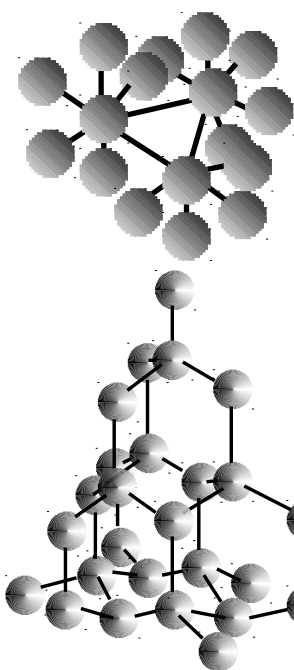


### Other Networks

The only other **covalent networks** in the first 20 elements are **boron** and **silicon** (which has a diamond structure like carbon).

Not surprisingly they are all **high melting point solids**.

A lot of **strong covalent** bonds have to be broken to melt these elements.



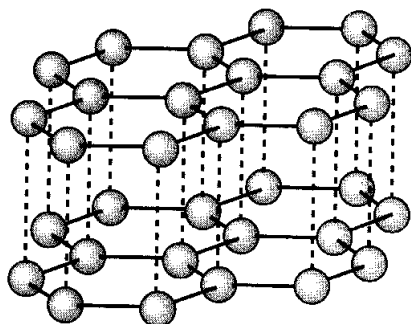
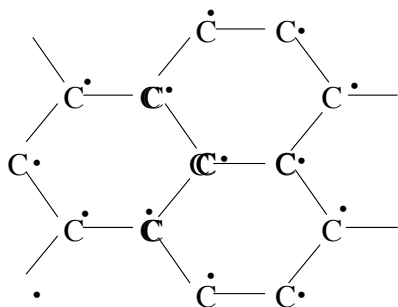
Melting point of **boron** = °C

Melting point of **silicon** = °C

Melting point of **carbon** = °C

Boron and silicon are also very **hard** substances, though not as hard as diamond, of course.

## Graphite Structure



Unusually, **carbon** has two **network** structures - **diamond** and **graphite**. Their **properties** are very different, and so must their **structures** be.

In **graphite**, each carbon atom is only bonded to 3 other carbon atoms by **strong covalent** bonds.

This forms flat **sheets** of 6 membered rings. The fourth outermost electrons are free to move (**delocalised**) between all the carbon atoms in a sheet or layer.

Between the layers are the much weaker **Van der forces** that exist between all atoms.

The distance between the layers is **greater** than the distance between carbon atoms within the same layer

We can use the **structure** of graphite to explain

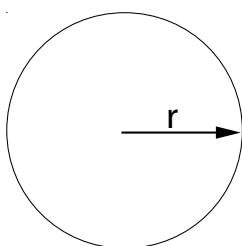
**Conductivity** **graphite** is able to **conduct** along layers because it has some **delocalised** electrons. (Conductivity between layers is not possible).

**Lubrication** powdered graphite makes an excellent solid lubricant because the **weak Van der Waals forces** allow the layers to **slide** over each other.

## 1.5 Atomic Size

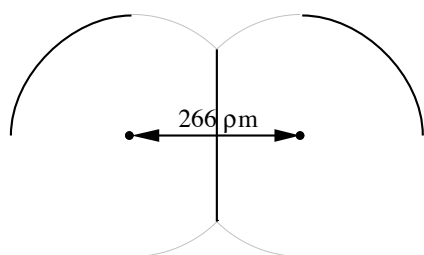
This lesson looks more closely at the size of atoms and the way they vary within the Periodic Table.

### Atomic Radius



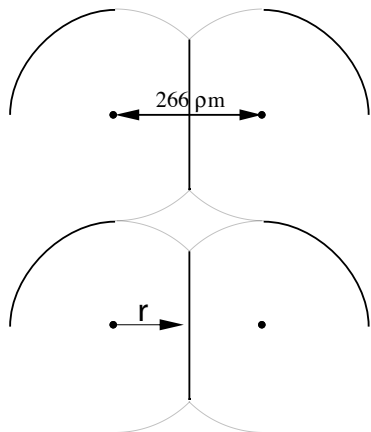
We are used to thinking of atoms as **spheres**, and of drawing them as **circles**, so the size of an atom could be defined by its **radius**; the distance from the **centre** of the atom to its **edge**.

Unfortunately, outside the **nucleus** is mainly **empty space** and there is no obvious edge.



What can be detected (by, for example, X-rays) are the **nuclei** of atoms. This allows the **distance between the nuclei** of neighbouring atoms to be measured, and this can be used to 'calculate' the **atomic radius** of individual atoms.

The radius can be calculated by assuming it is **half the distance between the nuclei of two atoms**. However, there are two ways that this can be done.



### Covalent Radius

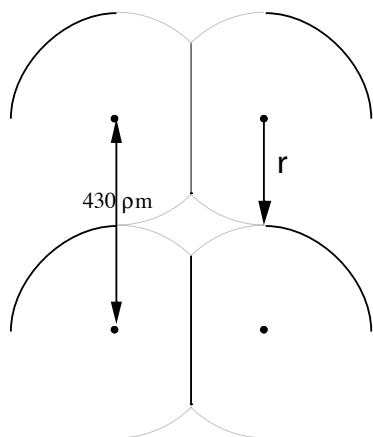
When two atoms form a **covalent bond** they **overlap**.

Half the distance between the two nuclei would give a value slightly **smaller** than the true size of the atom.

In the case of iodine, I<sub>2</sub>,

$$\text{covalent radius} = \frac{266}{2} = 133 \text{ pm}$$

(where pm = picometre = 1 x 10<sup>-12</sup> m)



### Van der Waals Radius

Between some atoms and many **molecules**, there exist weak Van der Waals forces which do not involve any overlap.

Half the distance between the two nuclei of atoms in neighbouring molecules would give a value **closer** to the true size of the atom.

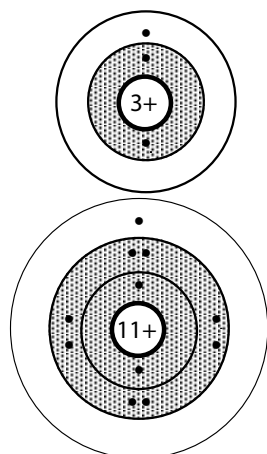
In the case of iodine, I<sub>2</sub>,

$$\text{van der waals radius} = \frac{430}{2} = 215 \text{ pm}$$

(where pm = picometre = 1 x 10<sup>-12</sup> m)

Unfortunately, there are very few atoms whose size could be determined using elements, in the solid state, held together by just Van der Waals forces. Therefore, **Covalent Radii** are the values most often quoted in Data Books and these are the values you will use most often.

## Atomic Size & the Periodic Table



As you go **down** a **Group** (column) there is an **increase** in the size of the atoms, (covalent radii).

Li    2,1                      covalent radius =    pm

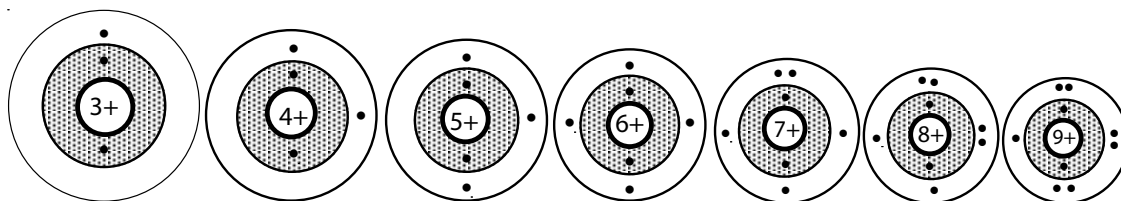
Na    2,8,1                    covalent radius =    pm

K     2,8,8,1                 covalent radius =    pm

This is not surprising, as each atom has an **extra energy level**, (**shell**) further out from the **nucleus**.

More surprisingly is that:-

as you go **across** a **Period** (row) there is a **decrease** in the size of atoms, (covalent radii).



The **number** of energy levels is the **same**, but the increasing size of the **nuclear charge** causes increased **attraction** between the outermost **electrons** and the **nucleus**. This causes the outer shell to be **pulled closer** and closer to the nucleus, and the size of the atoms decrease.

### Atomic Size & Density

*Density is mass per volume.*

So, mass of atom ↑ Density ↑  
and, size of atom ↑ Density ↓

*As you go **down** a Group (column) there is an **increase** in the density*

The **mass** of the atoms will be **increasing**, and the **size** of the atoms will be **increasing**, but the **increase** in mass must be the larger effect.

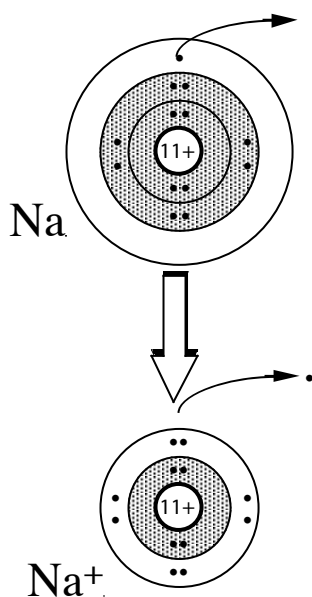
*As you go **across** a Period (row) there is an **increase** in the density*

The **mass** of the atoms will be **increasing**, and the **size** of the atoms will be **decreasing**, so both factors will lead to an **increase** in density. (Changes in structure (**networks**→**molecules**) and state (**solid**→**gases**) explain the low **densities** of many of the elements on the **right** of the Periodic Table.)

## 1.6 Ionisation Energies

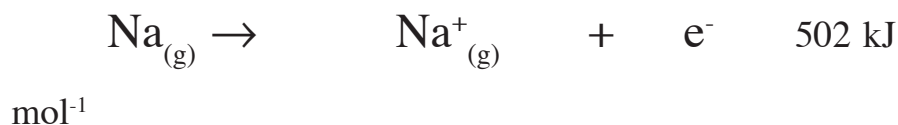
*This lesson is about **ionisation energy** and how it is related to an element's position in the periodic table.*

### First Ionisation Energy



*The **first ionisation energy** is the **energy required** to remove **one electron** from every atom in a mole of **free atoms**.*

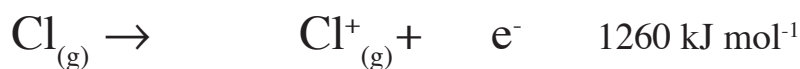
*It is measured in  $\text{kJ mol}^{-1}$ , i.e. **kilojoules per mole***



The sodium atoms have to be **free**, i.e. **not bonded** to any other atom. This can only be done by **heating** the sodium until it is a **gas**.

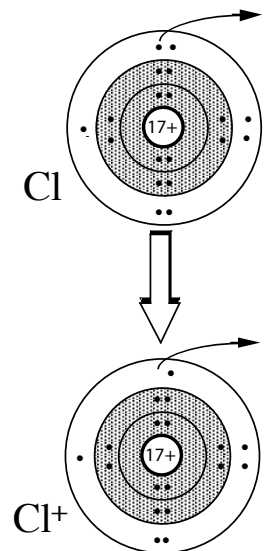
The **electron** lost appears as a **product**, on the **right** hand side of the **equation**.

All atoms, including non-metal atoms, can be forced to lose an **electron**. This results in the formation of a **positive ion**, whereas in normal reactions, non-metals will usually **gain electrons** and form **negative ions**.

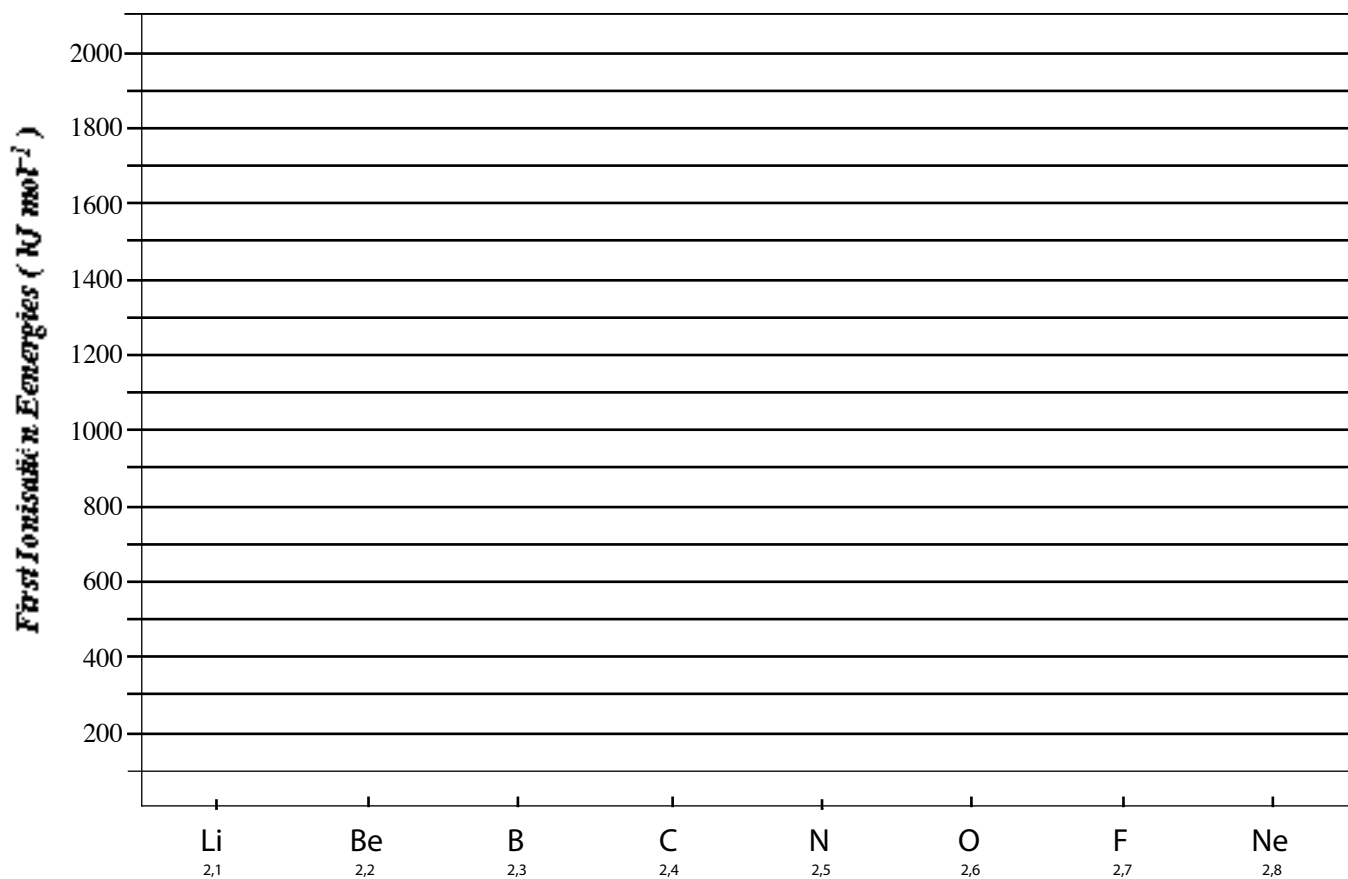


The chlorine atoms have to be **free**, i.e. **not bonded** to any other atom. This can only be done by **breaking** the chlorine **molecules** apart.

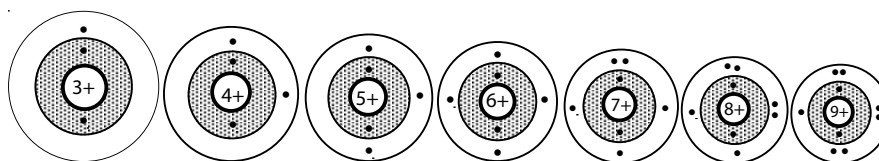
The **electron** lost appears as a **product**, on the **right** hand side of the **equation**.



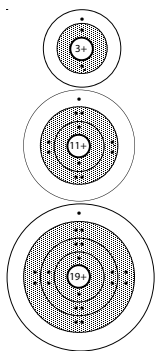
### Across a Period & Down a Group



As you go across a **period** the general trend is for the first ionisation energy to **increase**. (The fluctuations from the general trend require a more advanced explanation that goes beyond this level of chemistry)



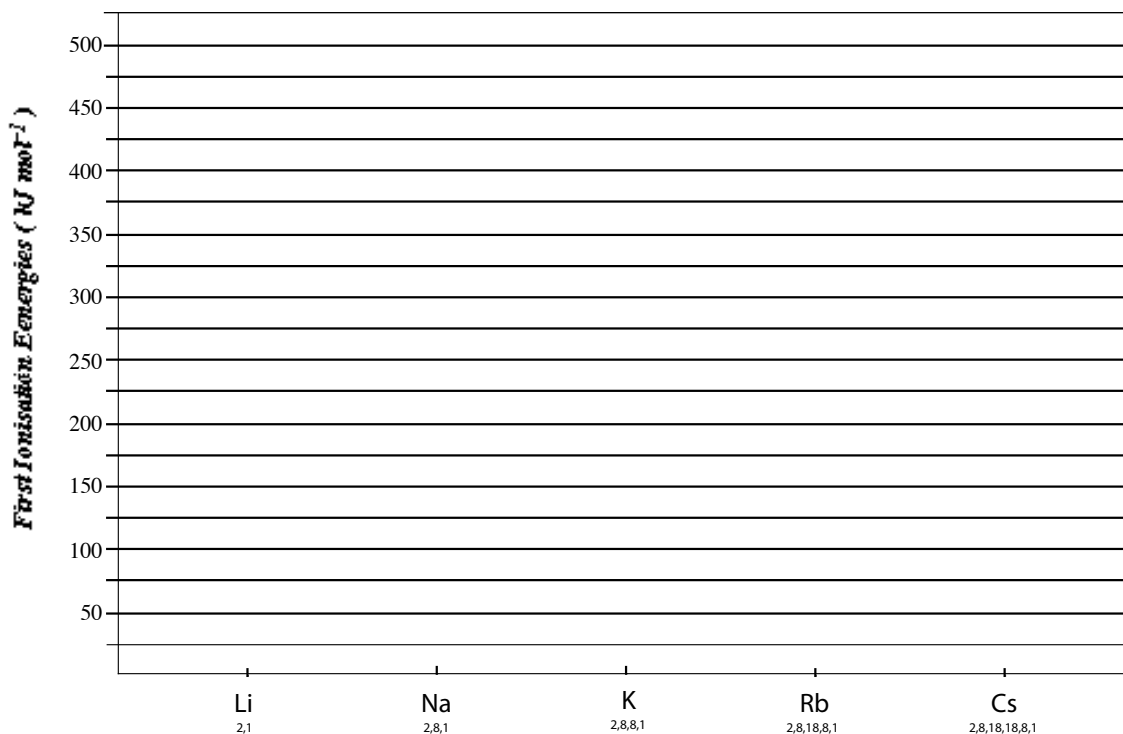
As mentioned earlier, (p11), the **number** of energy levels is the **same**, but the increasing size of the **nuclear charge** causes increased **attraction** between the outermost **electrons** and the **nucleus**. This makes it harder to **remove** an electron from the **outer** shell.



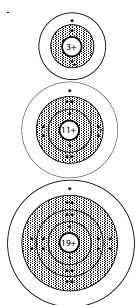
On going **down a group**, a new energy level (shell) is added each time. Each outer layer is further from the nucleus. This could make it **easier** to remove an electron.

The outer shell is screened by more and more shells closer to the nucleus. This could make it **easier** to remove an electron.

On the other hand, the nuclear charge is also increasing. This could make it **harder** to remove an electron.



As you go down a **group** the general trend is for the first ionisation energy to de-



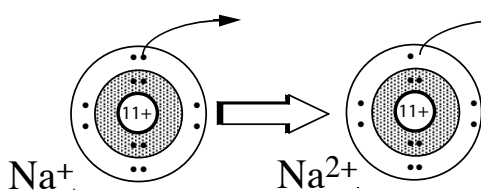
On going **down a group**, an **extra shell** is added each time. Each shell is **further** from the **nucleus**. This makes it **easier** to remove an **electron**.

The outer shell is **screened** by more and more shells **closer** to the **nucleus**. This makes it **easier** to remove an **electron**.

Even though the nuclear charge is increasing, the other two factors mean that it still gets **easier to remove an electron as you go down a group**.

### Second Ionisation Energy

The **second ionisation energy** is the energy required to remove a **second electron** from every 'atom' in a **mole of free 'atoms'**.

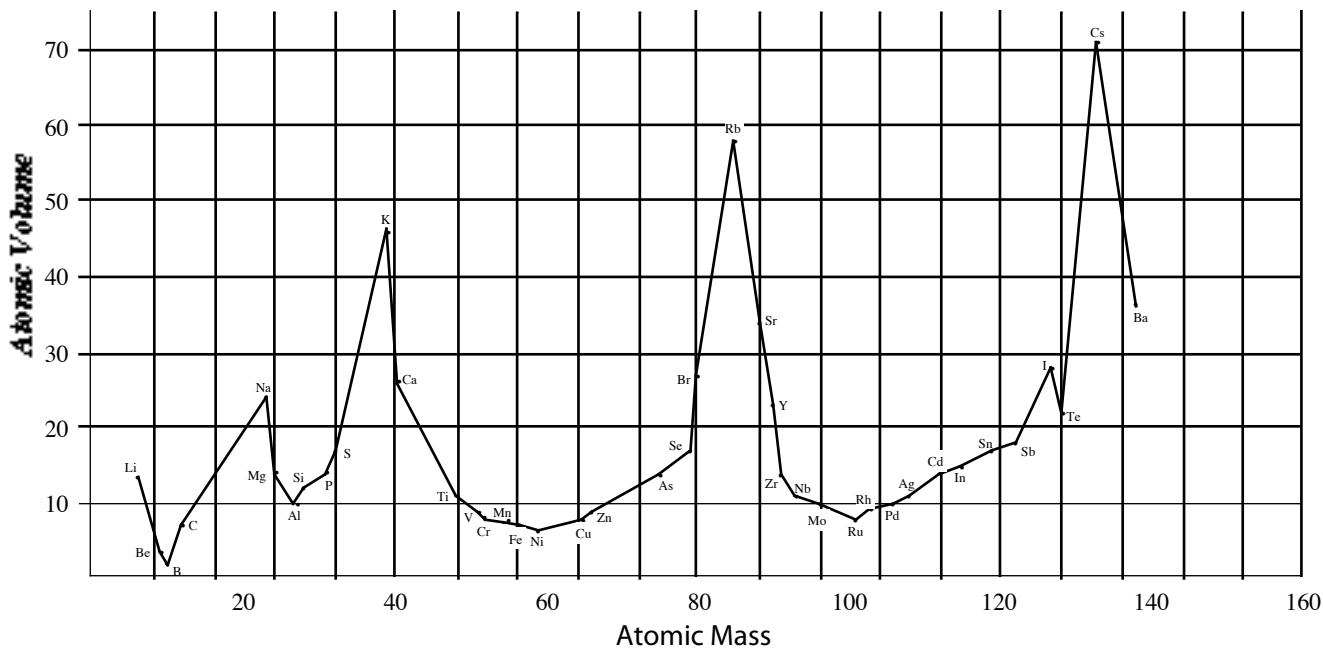
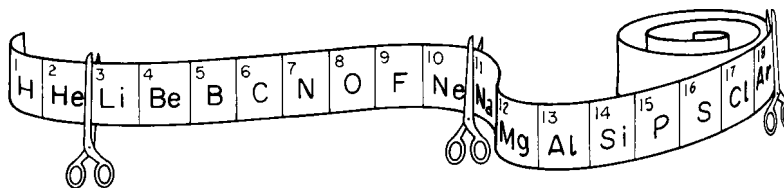


Each ionisation will require **more energy** than the previous one; a **dramatic increase** occurs whenever an electron has to be removed from the **next shell** closer to the nucleus.

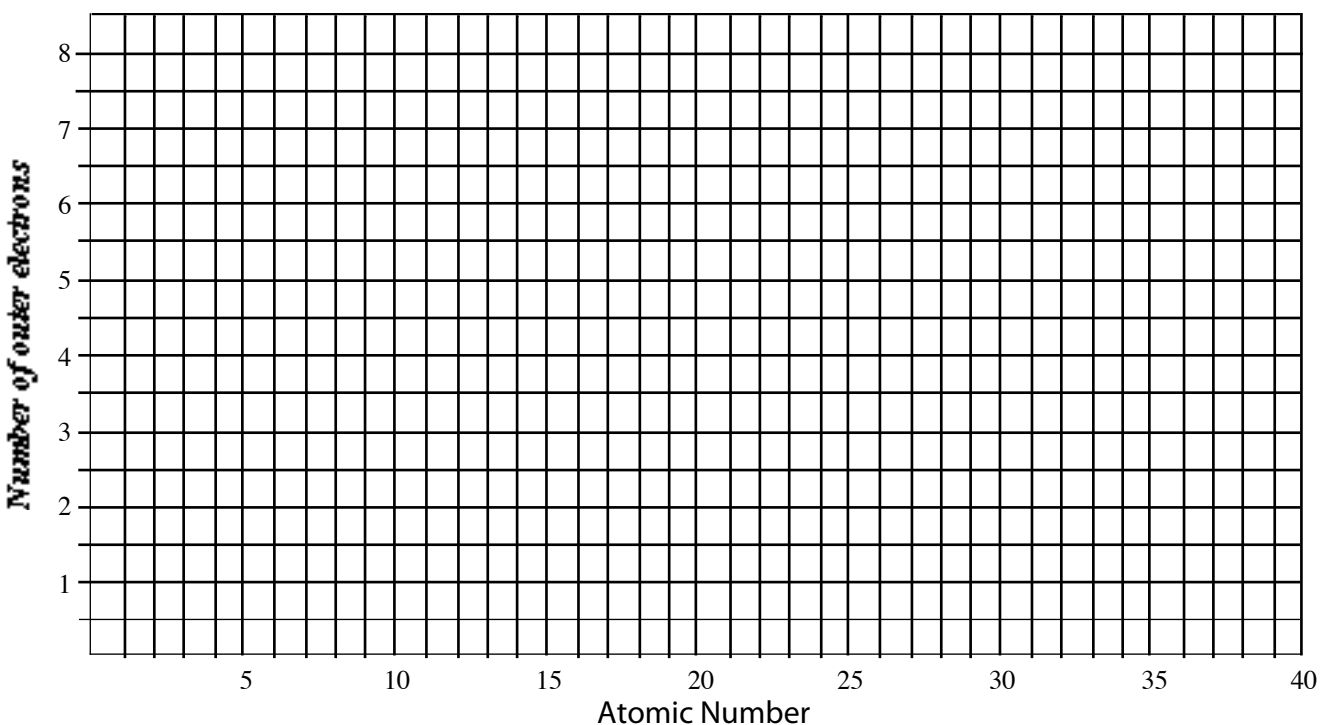
# 1.7 Periodic Pattern

This lesson is about the *periodic pattern* of the elements and how that is represented in the *Periodic Table*.

## Periodic Law



*Periodicity* is when a pattern occurs at *regular* intervals. As early as 1869, a German by the name of Meyer demonstrated this using *atomic volume*. A better demonstration is a plot of *outer electrons* against *atomic number*. Both graphs clearly show the *alkali metals* as the start of each new *period*.



## Mendeleev's Table



	GROUP							
	1	2	3	4	5	6	7	8
Period 1	H							
Period 2	Li	Be	B	C	N	O	F	
Period 3	Na	Mg	Al	Si	P	S	Cl	
Period 4	K Cu	Ca Zn	? ?	? Ti	As V	Se Cr	Br Mn	FeCoNi
Period 5	Rb Ag	Sr Cd	In Y	Sn Zr	Sb Nb	Te Mo	I ?	RuRhPd

In exactly the same year as Meyer, the Russian Mendeleev, came up with the same idea of a periodic relationship within the elements. He published earlier than Meyer and took the ideas much further. Namely;

1. **he changed the order of some elements**- in Meyer's graph you may have noticed that Iodine comes before Tellurium. This is because, on average, **I** atoms are lighter than **Te** atoms (lower RAM) . Mendeleev argued that the properties of the elements meant that **Te** should come before **I**,
2. **he left gaps in his table** - this was to keep elements with similar properties grouped together. Mendeleev argued that there were missing elements still to be discovered,
3. **he predicted the properties of missing elements** - this, more than anything, persuaded other scientists that Mendeleev was on the right track.

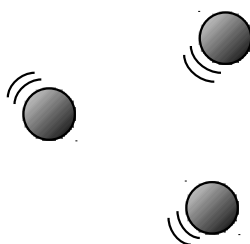
## Bonding &amp; Structure

This section will help you to identify the category of **bonding & structure** for the first 20 elements

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

- **Monatomic gases**

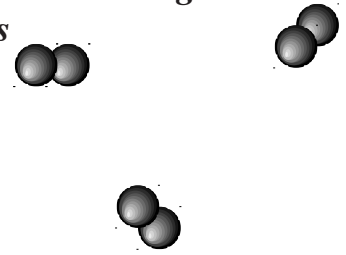
The only bonding present is **Van der Waal's forces** between atoms.



H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

- **Discrete covalent molecular diatomic gases**

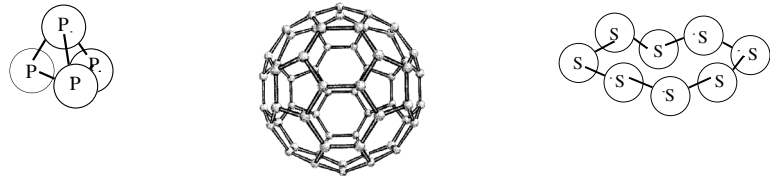
*Strong intramolecular covalent bonding with weak Van der Waal's forces between individual molecules.*



H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

- **Discrete covalent molecular solids**

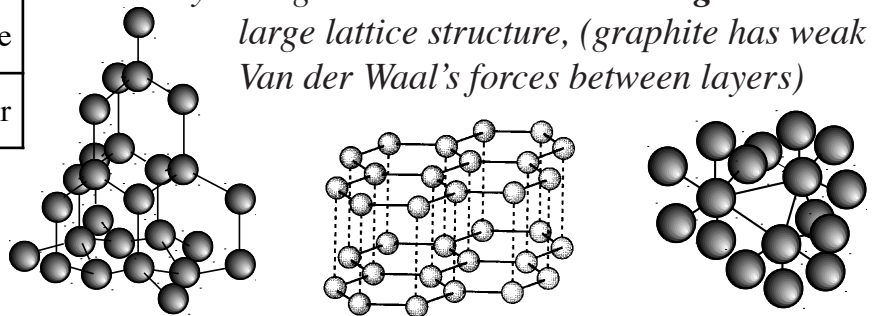
*Strong intramolecular covalent bonding with weak Van der Waal's forces between individual molecules.*



H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

- **Covalent Network solids**

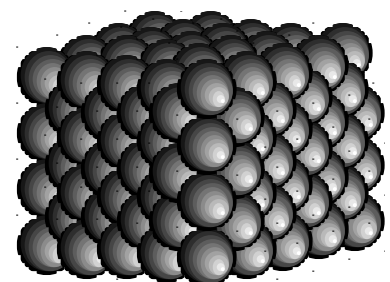
*Very strong internal covalent bonding within large lattice structure, (graphite has weak Van der Waal's forces between layers)*



H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

- **Metallic Lattices**

*Fairly strong metallic bonds within large lattice structure*

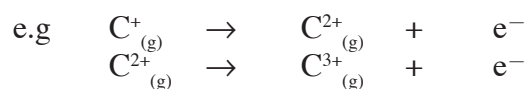


## UNIT 1. ENERGY MATTERS

### Section 1: Elements & Bonding

- |  |  |
|--|--|
| 1. <b>Metallic bonding</b> is the electrostatic attraction of a lattice of positively charged 'ions' and loosely held outer electrons.   | Each metal forms an 'ion' in the sense that they will have 'lost' their outer electron(s)  |
| 2. The properties of metals can be explained in terms of the type (and strength) of bonding and structure.   | Malleability and ductility.<br>Smaller atoms → stronger bonds<br>More outer electrons → stronger bonds   |
| 3. The noble gases are all <b>monatomic</b> .  | Literally "one" (single) atoms   |
| 4. A monatomic structure consists of discrete atoms held together by Van der Waal's forces   | Very weak forces → very low Bpts/Mpts  |
| 5. <b>Van der Waal's forces</b> are attractions which can operate between all atoms and molecules  |  |
| 6. Van der Waal's forces are much weaker than all other types of bonding.  | Only important when there are no other (stronger) forces.  |
| 7. Van der Waal's forces are the electrostatic attractions between <b>temporary dipoles</b> and <b>induced dipoles</b> caused by movement of electrons within atoms and molecules. | Dipole = positive end / negative end<br>Caused when the distribution of electrons becomes unbalanced.<br>Induced = caused by a dipole on a neighbouring atom |
| 8. The strength of Van der Waal's forces is related to the size of the atoms and molecules.  | Larger the atom/molecule → stronger forces   |
| 9. A <b>covalent molecular</b> structure consists of discrete molecules held together by intermolecular forces.  | Discrete = separate/ individual<br>Intermolecular = <b>between</b> molecules<br>(strong covalent bonds hold the atoms <b>within</b> the molecule together)   |
| 10. A <b>covalent network</b> structure consists of a giant lattice of covalently bonded atoms.  |  |
| 11. The <b>atomic size</b> decreases across a period and increases down a group  | The covalent radius of an element is half the distance between the nuclei of two of its bonded atoms   |
| 12. There are variations in the <b>densities</b> of the elements across a period and down a group  | Changes in structure e.g. solid → gas mean that trends don't extend right across a whole period or down a complete group.                                    |
| 13. The <b>first ionisation energy</b> is the energy required to remove one mole of electrons from one mole of gaseous atoms   |  |

14. The **second** and subsequent energies refer to the energies required to remove further moles of electrons.



15. The trends in the first ionisation energy across periods and down groups can be explained in terms of the **atomic size, nuclear charge and screening** effect due to inner shell electrons.

→ **Period:** atomic size ↓ nuclear charge ↑  
 (screening same) so **ionisation** ↑  
 ↓ **Group:** atomic size ↑ (nuclear charge ↑)  
 screening ↑ so **ionisation** ↓

16. The modern Periodic Table is based on the work of Mendeleev who arranged the known elements in order of increasing atomic mass along with similar chemical properties, leaving gaps for yet to be discovered elements.

Changed order when he felt masses were ‘wrong’  
 Left gaps  
 Made predictions.

17. The first 20 elements in the Periodic Table can be categorised according to bonding and structure:

**metallic** ( Li, Be, Na, Mg, Al, K, Ca )

**covalent molecular** ( H<sub>2</sub>, N<sub>2</sub>, O<sub>2</sub>, F<sub>2</sub>, Cl<sub>2</sub>, P<sub>4</sub>, S<sub>8</sub>, and C (fullerenes) )

**covalent network** ( B, C (diamond, graphite), Si )

**monatomic** ( He, Ne, Ar )

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						