Energy Matters

Elements and Bonding

Metallic, Monatomic, Molecular
and Covalent Network elements.
Atomic size, Ionisation energies and
Periodic patterns.

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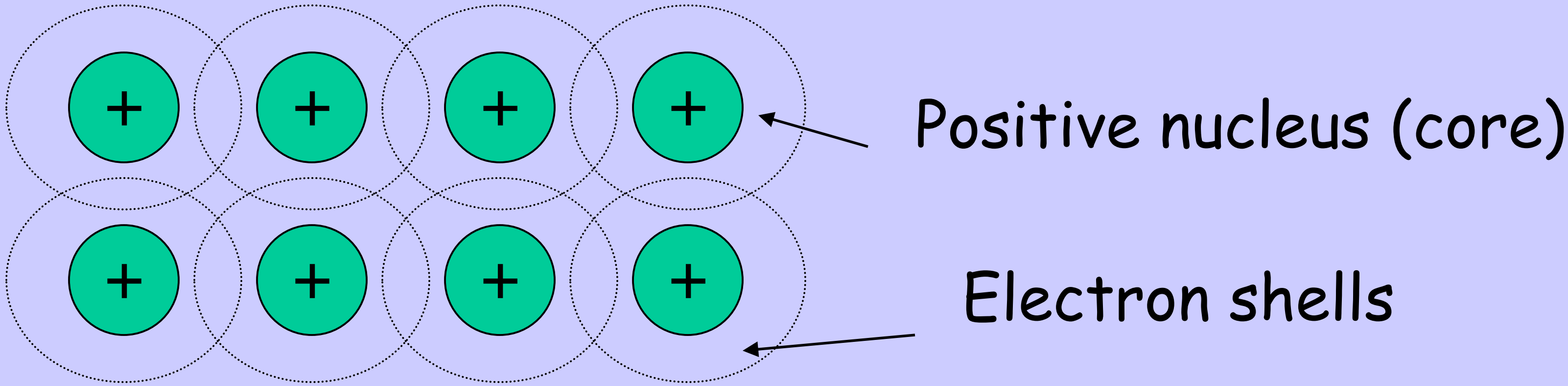
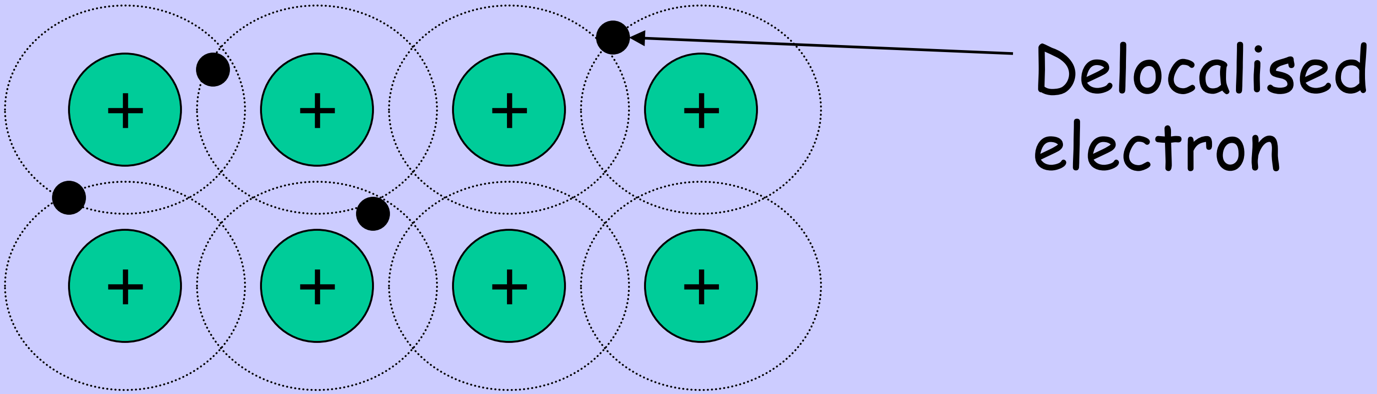
Metallic elements



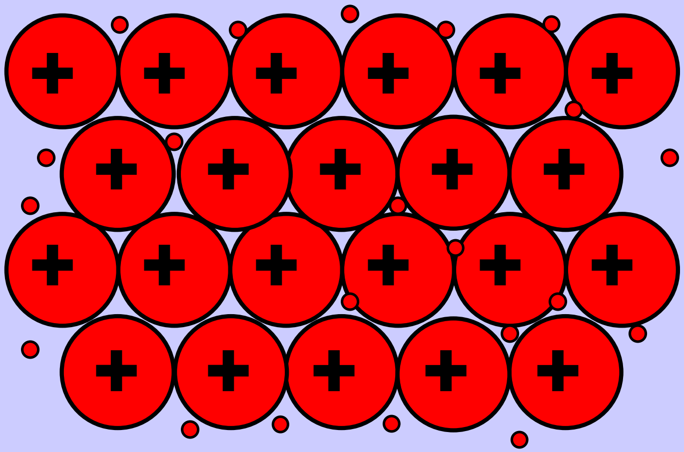
Strong electrostatic forces exist between the positive nuclei (ions) and the delocalised outer shell electrons. These electrostatic attractions are known as metallic bonds.

Metallic elements



The **outer shell** in metals is not full and so metal electrons can move randomly between these partially filled outer shells.
This creates what is sometimes called a 'sea' or 'cloud' of electrons.

Metallic Bonding



The positive metal ions are held together by this electron "Glue"

The outer electrons are delocalised and free to move throughout the lattice.

The greater the number of electrons in the outer shell the <u>stronger</u> the metallic bond.

So the melting point of Al>Mg>Na

Physical properties of metals

A. Metals are malleable and ductile

Metal atoms can ‘slip’ past each other because the metallic bond is not fixed and it acts in all directions.

B. Conduction of electricity and thermal energy.

Physical properties of metals

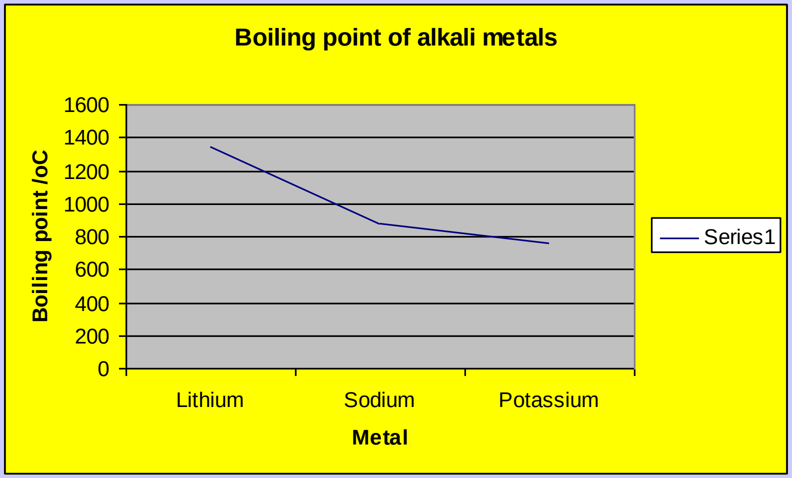
C. Change of state

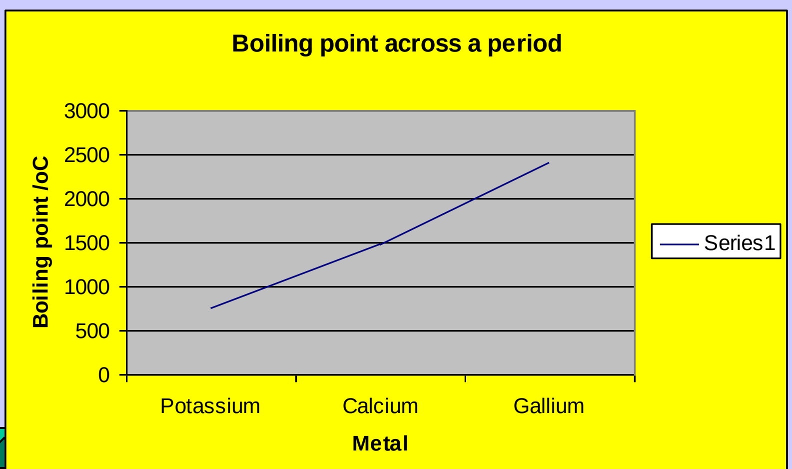
M.p.'s are relatively low compared to the b.p.'s.
When a metal is molten the metallic bond is still present.

B.p.'s are much higher as you need to break the metallic bonds throughout the metal lattice.

Metal b.p.'s are dependant on
(i) How many electrons are in the outer shell
(ii) How many electron shells there are.

Metal boiling point trends





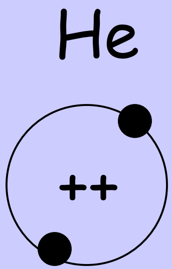
Down a group

The atomic size increases so the outer shell is further away from the positive core.

Across a period

The atomic size decreases as the positive core is increasing in charge, this has the effect of pulling the outer closer to the nucleus.

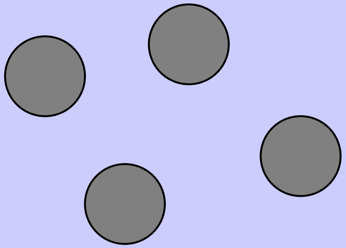
Monatomic elements



Noble gases have full outer electron shells

They do not need to combine with other atoms.

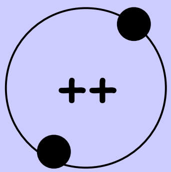
They are said to be monatomic.



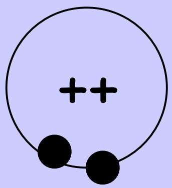
Group 0 are all gases and exist as individual atoms.

However, the monatomic gases do form weak inter-atomic bonds at very low temperatures.

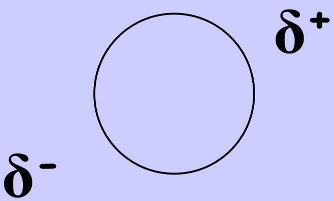
Monatomic elements



Sometimes the electrons can end up on one side of the atom, i.e. the electron cloud can wobble

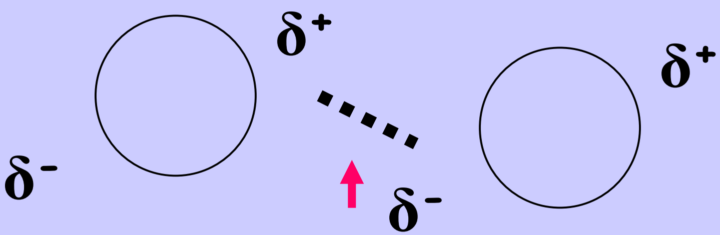


This means that one side of the atom is more negative than the other side. i.e. 2 'electric poles' are formed, otherwise called a dipole.



These charges are given the symbol **δ** 'delta'

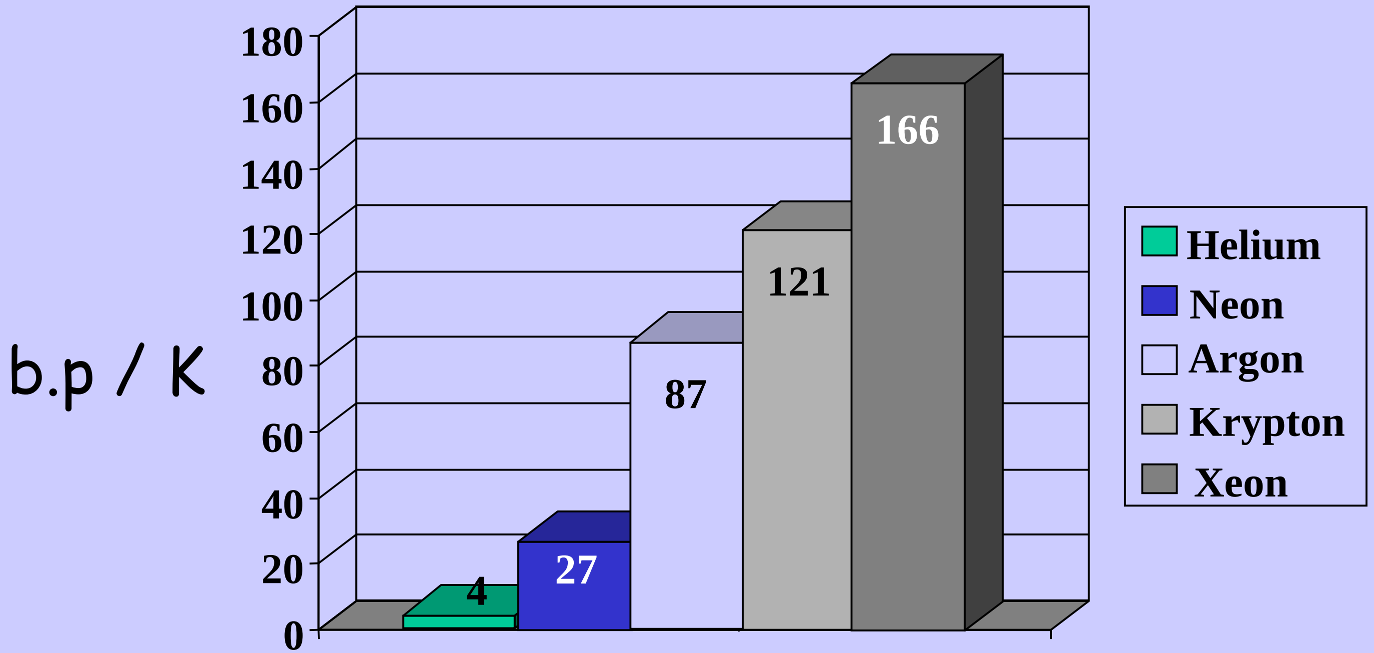
A temporary dipole is therefore formed.



A dipole can induce other atoms to form dipoles, resulting in dipole –dipole attraction.

Van der Waals forces

Noble gases b.p.'s

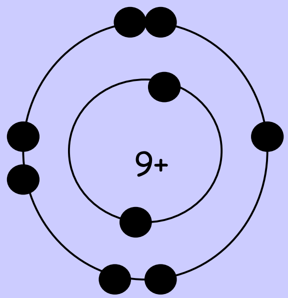


B.p.'s increase as the size of the atom increases

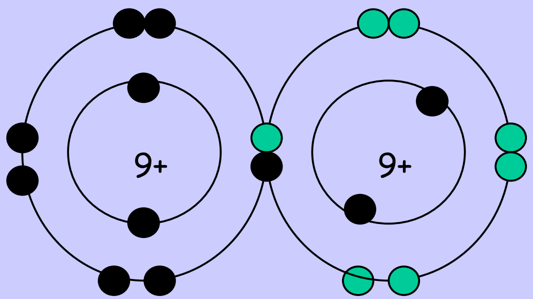
This happens because the Van der Waals' forces increases with increasing size of atoms.

Molecular Elements

Fluorine atom

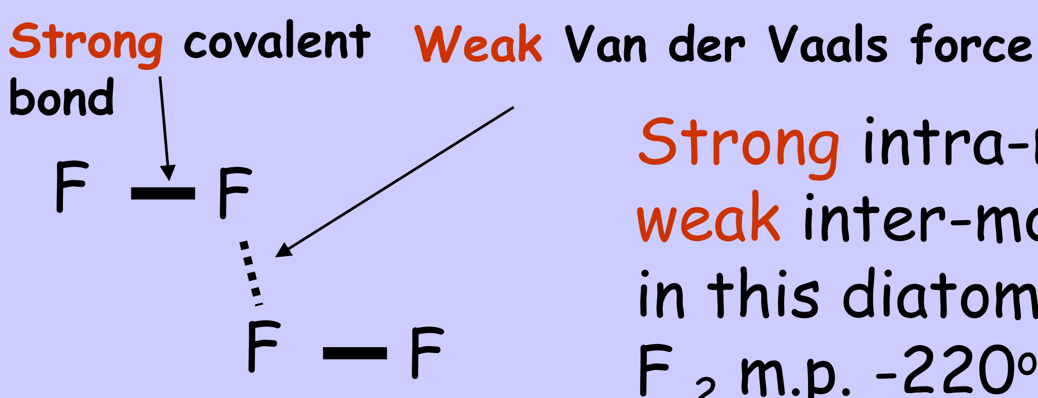


Fluorine molecule F_2



diatomic

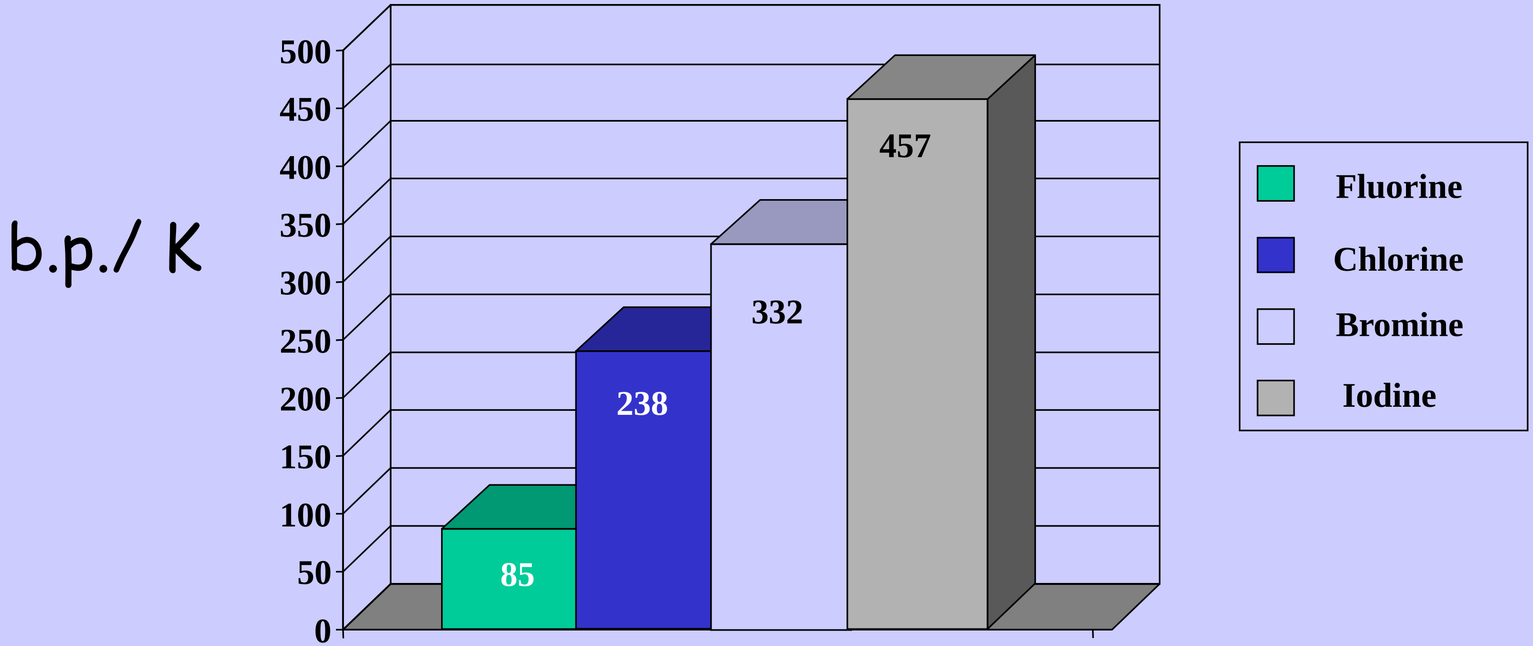
A covalent bond is formed when a pair of electrons are shared. The atoms in a covalent bond are held together by electrostatic forces between positively charged nuclei and negatively charged electrons.



Strong intra-molecular bonding and weak inter-molecular bonding exist in this diatomic molecule.

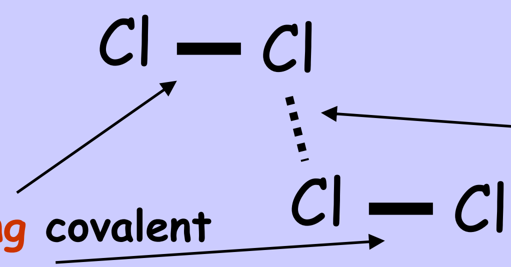
F_2 m.p. -220° C or 53 K

Halogens b.p.'s



As the size of the halogen atom increases, so does the size of the van der waals' forces between the halogen molecule.

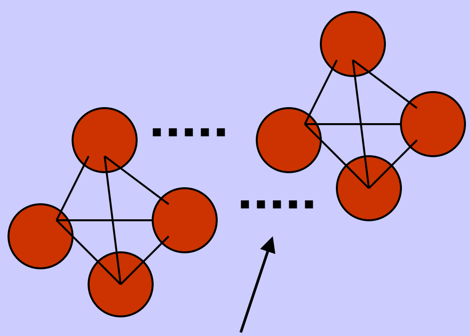
Chlorine, Phosphorus and Sulphur discrete molecules



Chlorine Cl_2 m.p. -101°C

Weak Van der Waals' forces

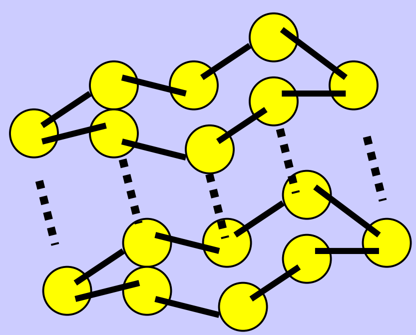
Strong covalent bond



Phosphorus P_4 m.p. 44°C

Sulphur S_8 m.p. 113°C

Weak Van der Waals forces



Higher m.p. because there are stronger Van de Waals' forces between larger molecules.

Fullerenes, molecules of carbon

C_{60} C_{70} C_{240}

Buckminster fullerene (Bucky Balls) were discovered in the 1980's. In 1999 it cost £25 per gram!

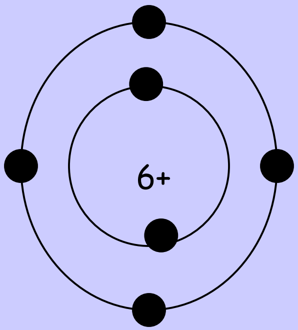
Uses!!!

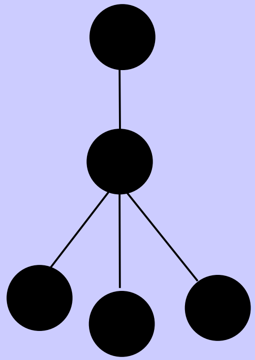
- TV, flat screens
- Bulbs, in paints that could light up a room
- Nanotubes tubes can withstand 1 million x atmospheric pressure

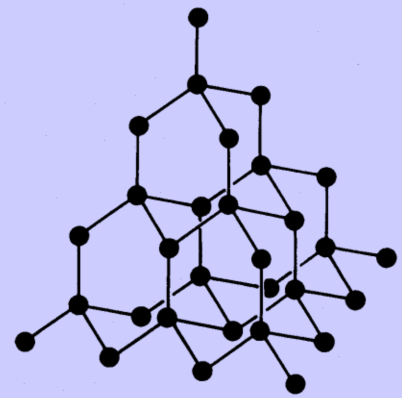
Due to the large molecules , fullerenes have stronger Van der Waals' forces between their molecules, compared to elements made from smaller molecules.

Covalent Network Elements

Carbon







Diamond has a covalent network element

Each of the outer electrons in a **carbon** atom can form a covalent bond with another carbon atom.

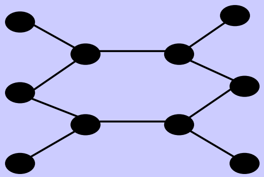
In the first 20 elements, only Boron, Carbon and Silicon have covalent network structures.

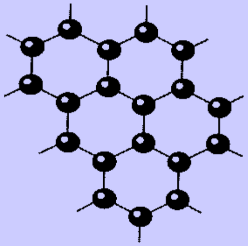
m.p.'s B $2300^{\circ}C$, C > $3642^{\circ}C$ and Si $1410^{\circ}C$

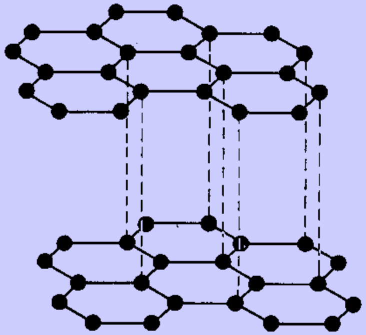
These are high because many covalent bonds have to be broken.

Graphite

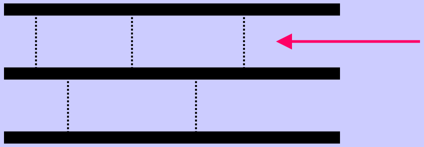
Carbon bonded to only 3 other Carbons







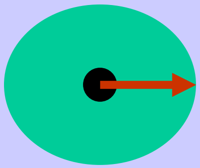
So the spare electrons are delocalised and so free to move. Graphite is a conductor



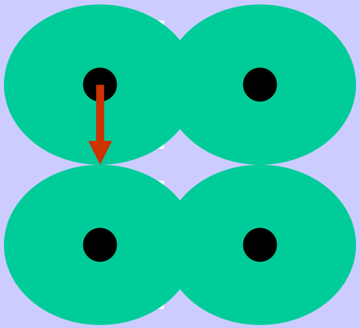
Van der Waals forces between the layers allows layers to slide over each other.

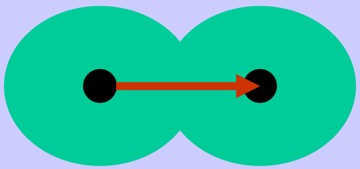
Graphite can be used as a lubricant

Atomic Size

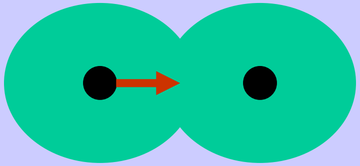


There is no definite edge to an atom.





However, bond lengths can be worked out.



Covalent radius, ½ the distance between nuclei.

To find the bond length, add 2 covalent radii together.

Covalent radius

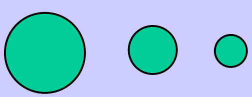
pm = picometre X 10^{-12} m

The covalent radii of the elements in any period decrease with increasing atomic number.

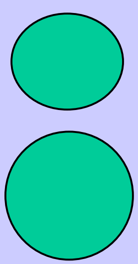
Na 154 pm, Mg 145 pm, Al 130 pm, Si 117 pm, P 110 pm, S 102 pm

The covalent radii of the elements in any group increase with increasing atomic number.

Li 134 pm, Na 154 pm, K 196 pm, Rb 216 pm



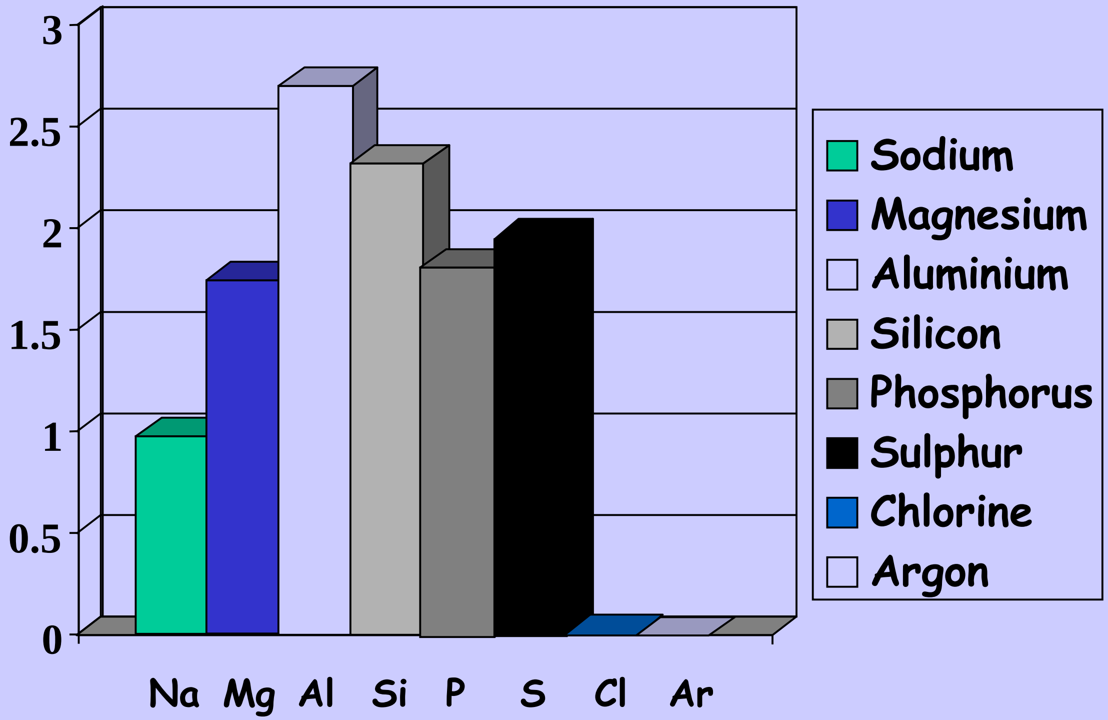
Going across a period, the attraction between the outer shell and the positive nucleus increases.



Going down a group, the attraction between the outer shell and the positive nucleus decreases due to more occupied electron shells 'shielding' the nuclear charge.

Density change across a period

Density g/cm^3



Na to **Al** the atom size decreases leading to greater packing in metal lattice.

Si is a covalent network, tightly packed atoms in covalent lattice.

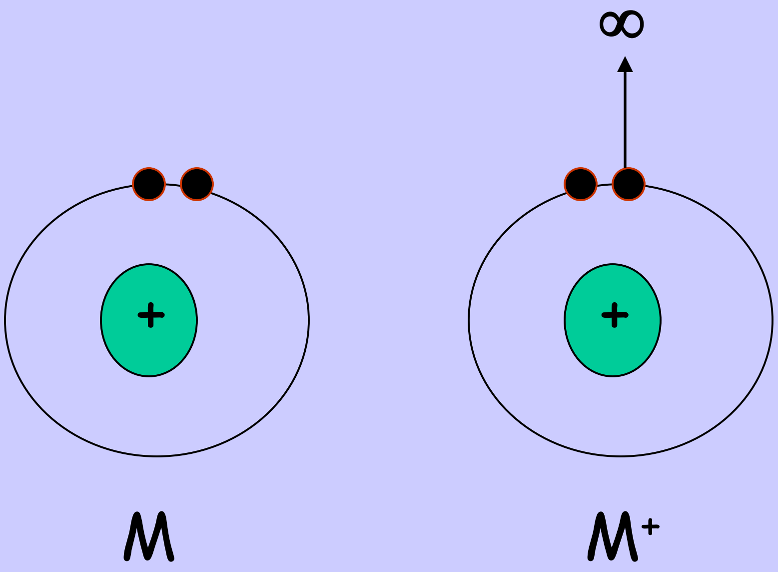
P and **S** are covalent molecular solids with quite densely packed molecules.

Cl and is a covalent molecular gas at room temperature.

Ar and is a monomolecular gas at room temperature.

Ionisation energies

This is defined as *"the amount of energy required to remove one mole of electrons from one mole of gaseous atoms or ions."*



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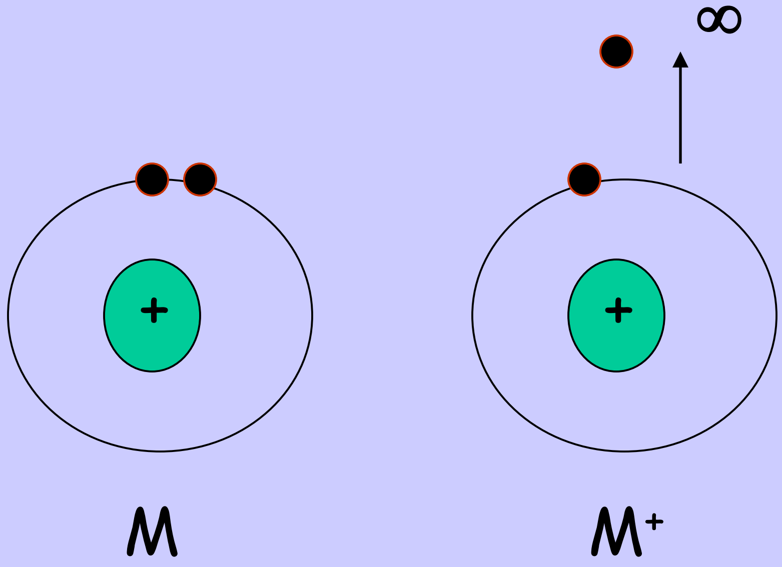

$M_{(g)} \rightarrow M_{(g)}^{+} + e$ 1st ionisation

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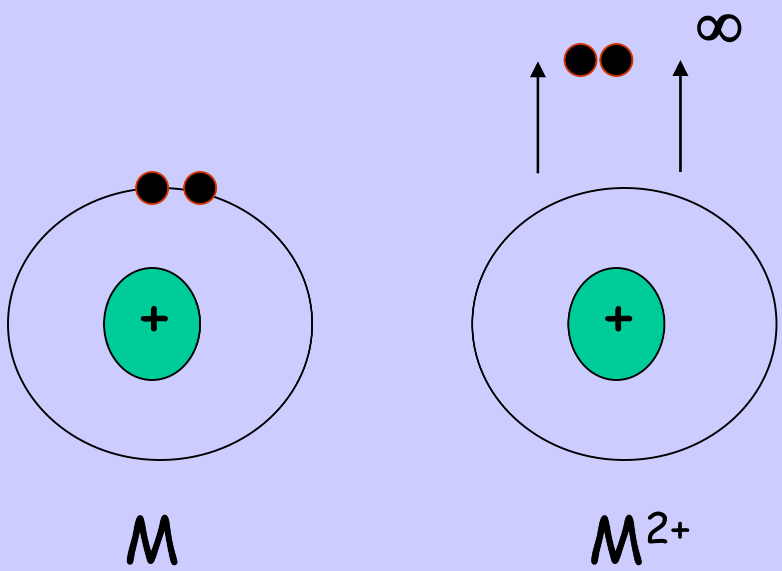


$M_{(g)} \rightarrow M_{(g)}^{+} + e$ 1st ionisation

$M_{(g)}^{+} \rightarrow M_{(g)}^{2+} + e$ 2nd ionisation

Ionisation energies

There are **two** main trends

- **Ionisation energies** tend to **decrease** as we descend a group in the Periodic Table.
- **Ionisation energies** tend to **increase** gradually as we traverse a period in the Periodic Table.

As the distance **increases**, the attraction of the positive nucleus for the negative electron will **decrease** and consequently the ionisation energy will **decrease**. This explains the fall in ionisation energy as we **do down a group**.

As the nuclear charge **increases**, its attraction for the outermost electron/s **increases** and consequently ionisation energy **increases**. This explains the general increase in ionisation energy that occurs as we **move across a period**.

Inner electrons can "**screen**" the outer electrons from the full attractive force of the nucleus.

Ionisation energies kJ mol^{-1}

Li 526	Be 905	B 807	C 1090	N 1410	O 1320	F 1690	Ne 2090
Na 502	Mg 744	Al 584	Si 792	P 1020	S 1010	Cl 1260	Ar 1530
K 425	Ca 596	Ga 577	Ge 762	As 953	Se 941	Br 1150	Kr 1350
Rb 409	Sr 556	In 556	Sn 715	Sb 816	Te 870	I 1020	Xe 1170

Li 2^{nd} ionisation energy is 7310 kJ mol^{-1}

Periodic Pattern

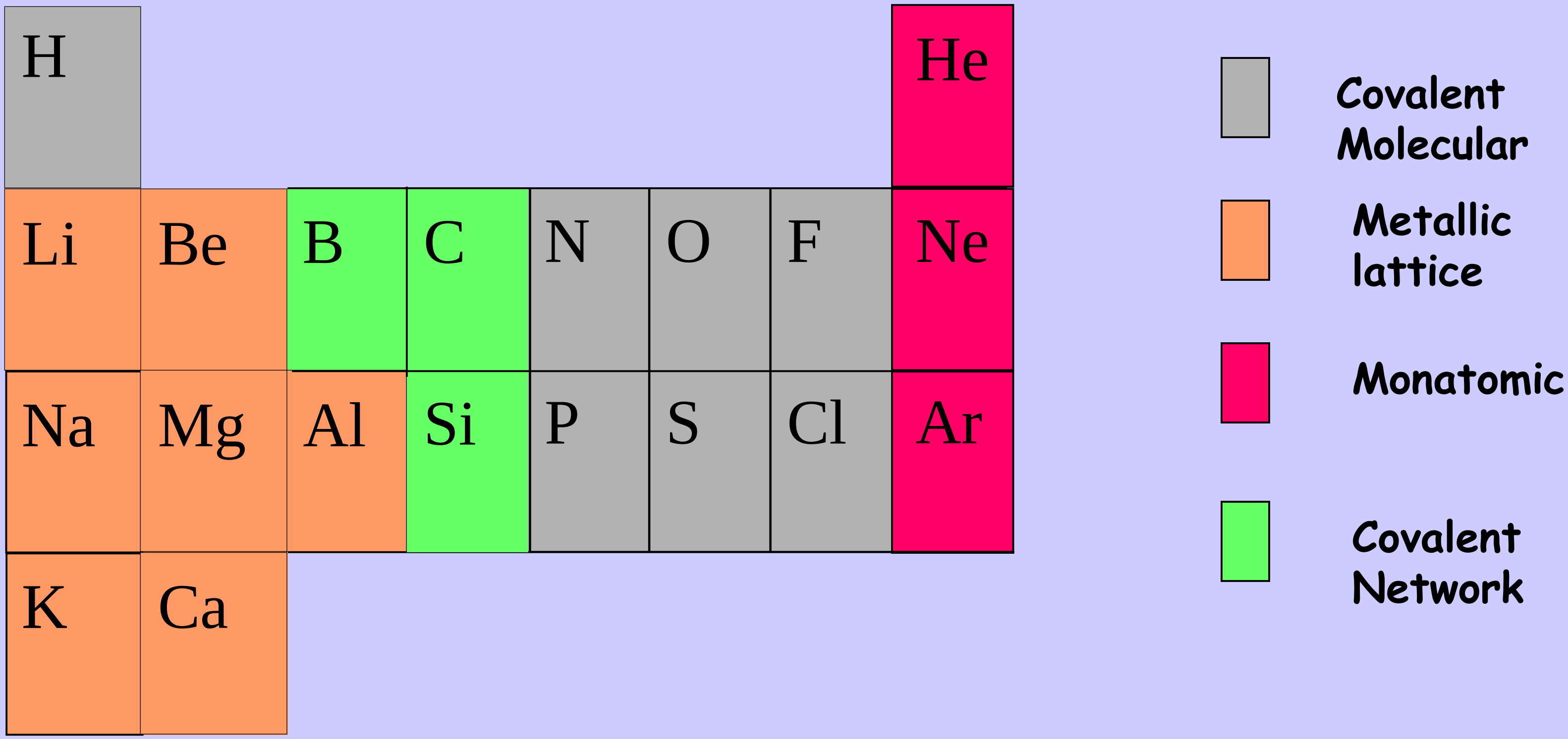
The modern Periodic Table is based on the work of Dimtri Mendeleev in 1869

Period	I	II	III	IV	V	VI	VII	VII
1	H							
2	Li	Be	B	C	N	O	F	
3	Na K	Mg Ca	Al *	Si Ti	P V	S Cr	Cl Mn	Fe Co Ni
4	Cu Rb	Zn Sr	* Y	* Zr	As Nb	Se Mo	Br *	Ru Rh Pd
5	Ag	Cd	In	Sn	Sb	Te	I	

Periodicity means reoccurring at regular intervals

History of the periodic table

Bonding patterns of the 1st 20 elements



C , in the form of fullerenes, is covalent molecular

Bond Strengths

Bond Type	Strength (kJ mol $^{-1}$)
Metallic	80 to 600
Ionic	100 to 500
Covalent	100 to 500
Hydrogen	40
Dipole-Dipole	30
Van der Waals	1 to 20