AN INTRODUCTION TO
BONDING

A guide for A level students

2008
SPECIFICATIONS

KNOCKHARDY PUBLISHING

BONDING

INTRODUCTION

This *Powerpoint* show is one of several produced to help students understand selected topics at AS and A2 level Chemistry. It is based on the requirements of the AQA and OCR specifications but is suitable for other examination boards.

Individual students may use the material at home for revision purposes or it may be used for classroom teaching if an interactive white board is available.

Accompanying notes on this, and the full range of AS and A2 topics, are available from the KNOCKHARDY SCIENCE WEBSITE at...

`www.knockhardy.org.uk/sci.htm`

Navigation is achieved by...

either clicking on the grey arrows at the foot of each page

or using the left and right arrow keys on the keyboard

BONDING

CONTENTS

STRUCTURE AND BONDING

The physical properties of a substance depend on its structure and type of bonding present. Bonding determines the type of structure.

Basic theory

- the noble gases (He, Ne, Ar, Kr, Xe and Rn) are in Group VIII
- they are all relatively, or totally, inert
- their electronic structure appears to confer stability
- they have just filled their ‘outer shell’ of electrons
- atoms without the electronic structure of a noble gas try to get one
- various ways are available
- the method depends on an element’s position in the periodic table

STRUCTURE AND BONDING

The physical properties of a substance depend on its structure and type of bonding present. Bonding determines the type of structure.

TYPES OF BOND

CHEMICAL ionic (or electrovalent)
strong bonds covalent
dative covalent (or co-ordinate)
metallic

PHYSICAL van der Waals‘ forces - weakest
weak bonds dipole-dipole interaction
hydrogen bonds - strongest

IONIC

BONDING

THE IONIC BOND

Ionic bonds tend to be formed between elements whose atoms need to “lose” electrons to gain the nearest noble gas electronic configuration (n.g.e.c.) and those which need to gain electrons. The electrons are transferred from one atom to the other.

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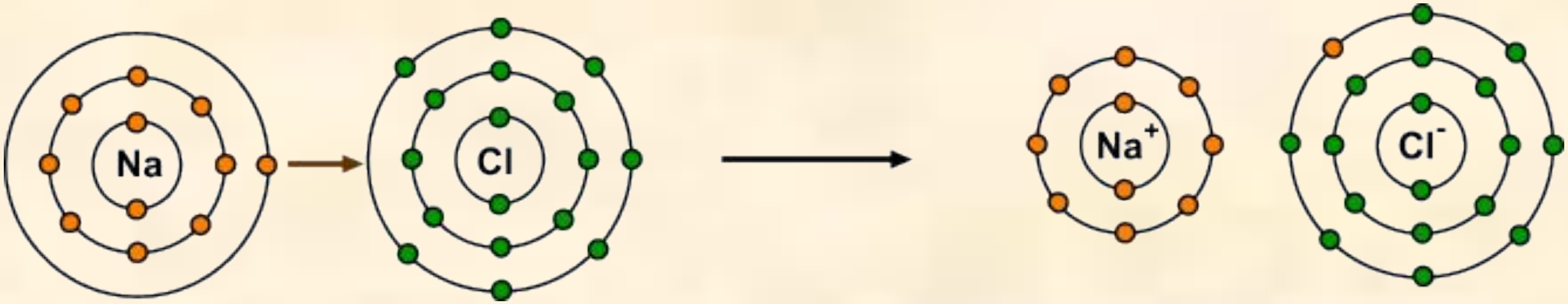
Sodium Chloride



Na ——> Na+ + e^- and Cl + e^- ——> Cl^-

$1s^2\ 2s^2\ 2p^6\ 3s^1$ $1s^2\ 2s^2\ 2p^6$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

or 2,8,1 2,8 2,8,7 2,8,8

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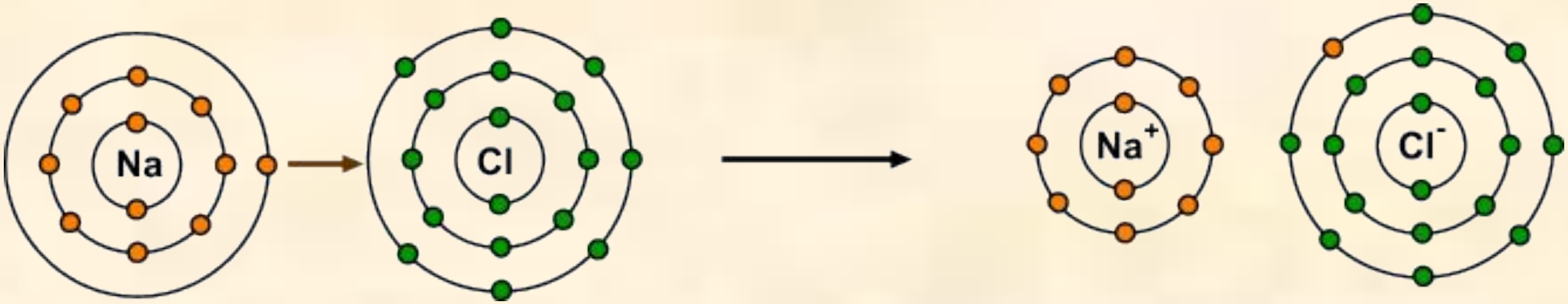

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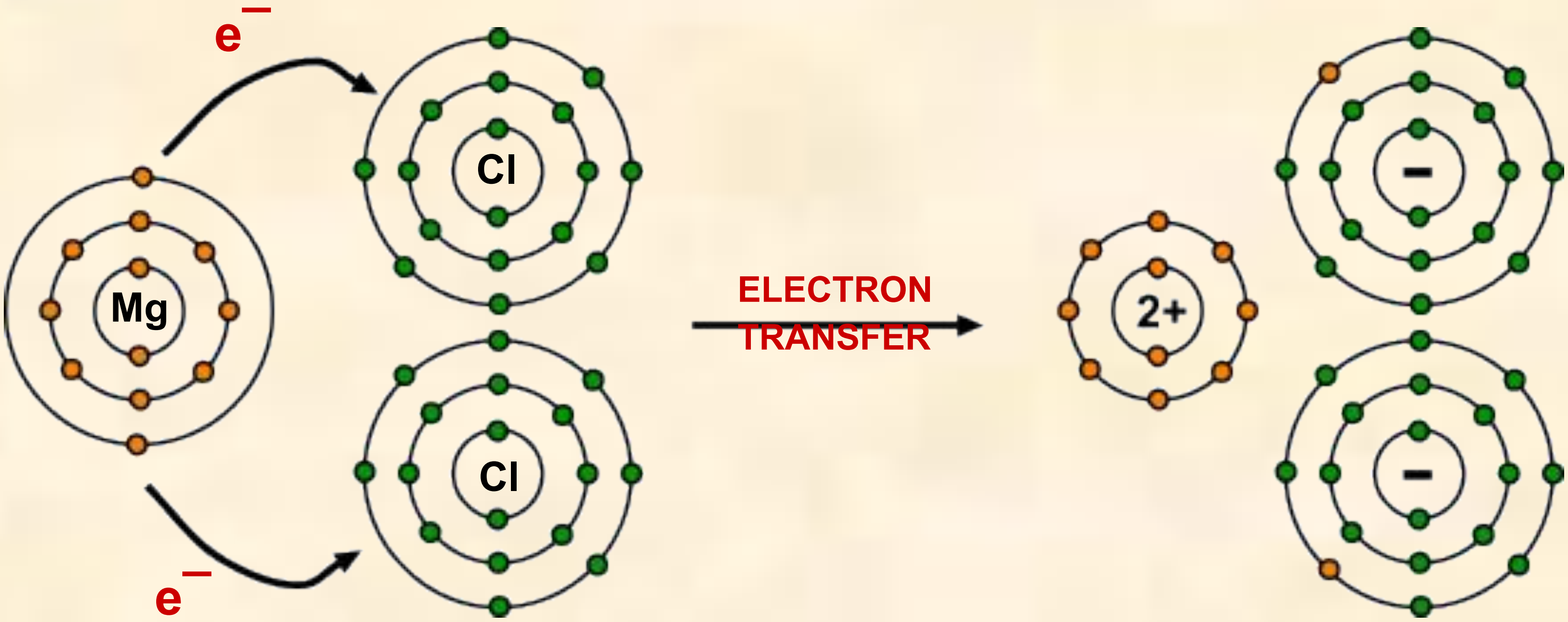
or 2,8,1 2,8 2,8,7 2,8,8

An electron is transferred from the 3s orbital of sodium to the 3p orbital of chlorine; both species end up with the electronic configuration of the nearest noble gas the resulting ions are held together in a crystal lattice by electrostatic attraction.

THE IONIC BOND

FORMATION OF MAGNESIUM CHLORIDE



$Mg \longrightarrow Mg^{2+} + 2e^-$ and $2Cl + 2e^- \longrightarrow 2\,Cl^-$

THE FORMATION OF IONS

Positive ions

- also known as cations; they are smaller than the original atom.
- formed when electrons are removed from atoms.
- the energy associated with the process is known as the ionisation energy

1st IONISATION ENERGY (1st I.E.)

The energy required to remove one mole of electrons (to infinity) from the one mole of gaseous atoms to form one mole of gaseous positive ions.

e.g. $Na(g) \longrightarrow Na^+(g) + e^-$ or $Mg(g) \longrightarrow Mg^+(g) + e^-$

Other points

Successive IE's get larger as the proton:electron ratio increases.

Large jumps in value occur when electrons are removed from shells nearer the nucleus because there is less shielding and more energy is required to overcome the attraction. If the I.E. values are very high, covalent bonding will be favoured (e.g. beryllium).

THE FORMATION OF IONS

Negative ions

- known as anions
- are larger than the original atom due to electron repulsion in outer shell
- formed when electrons are added to atoms
- energy is released as the nucleus pulls in an electron
- this energy is the electron affinity.

ELECTRON AFFINITY

The energy change when one mole of gaseous atoms acquires one mole of electrons (from infinity) to form one mole of gaseous negative ion

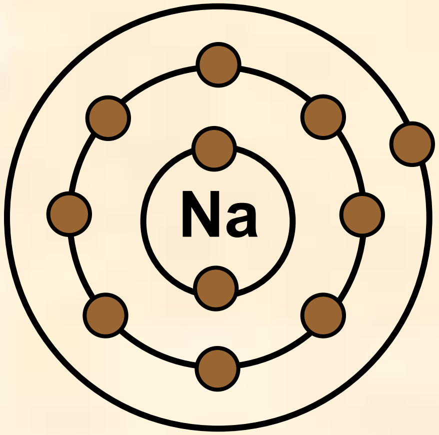
e.g. $Cl(g) + e^- \longrightarrow Cl^-(g)$ and $O(g) + e^- \longrightarrow O^-(g)$

The greater the effective nuclear charge (E.N.C.) the easier an electron is pulled in.

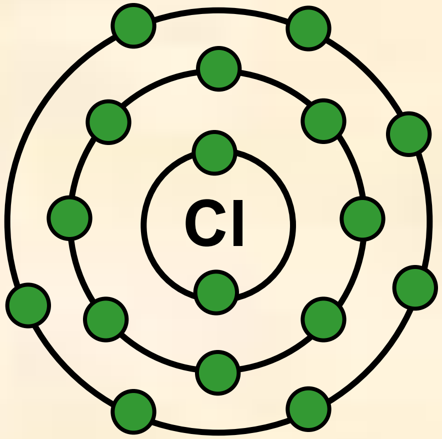
IONIC BONDING

Animations

SODIUM CHLORIDE

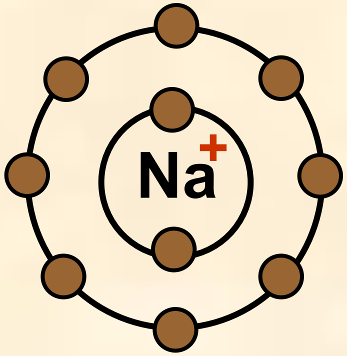


SODIUM ATOM
2,8,1

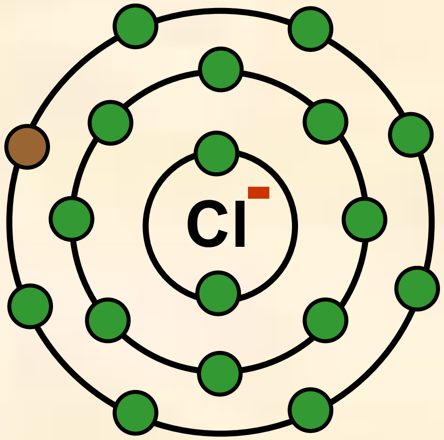


CHLORINE ATOM
2,8,7

SODIUM CHLORIDE



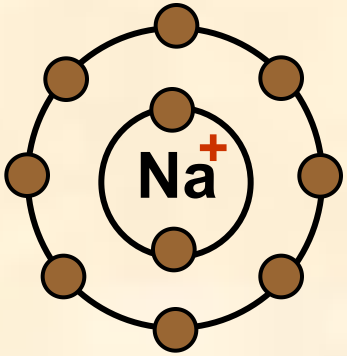
SODIUM ION
2,8



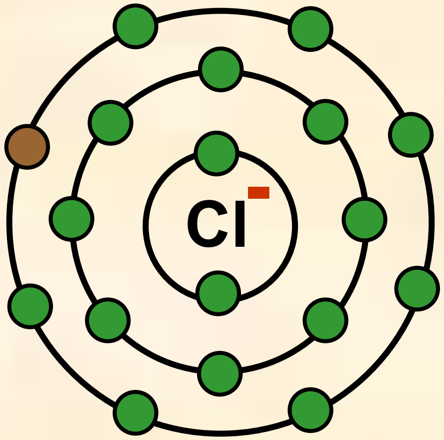
CHLORIDE ION
2,8,8

both species now have ‘full’ outer shells; ie they have the electronic configuration of a noble gas

SODIUM CHLORIDE



SODIUM ION
2,8



CHLORIDE ION
2,8,8

$$Na \longrightarrow Na^+ + e^-$$

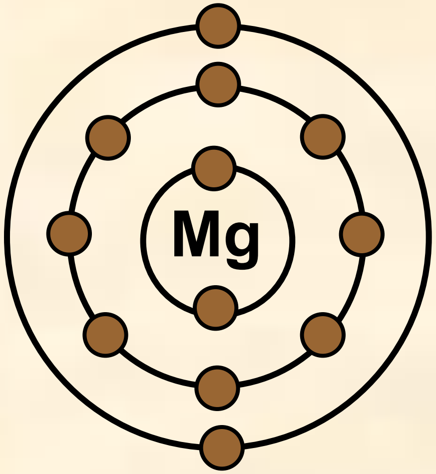
2,8,1 2,8

ELECTRON TRANSFERRED

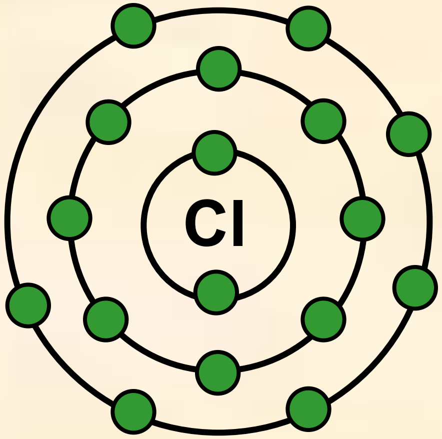
$$Cl + e^- \qquad Cl^-$$

2,8,7 2,8,8

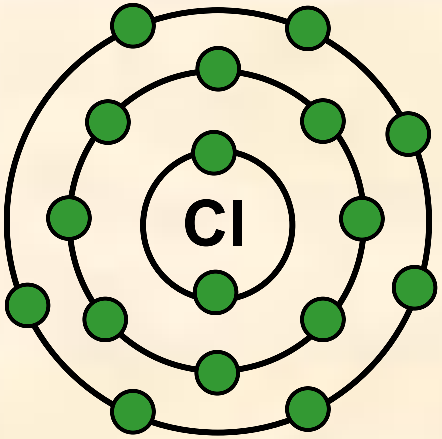
MAGNESIUM CHLORIDE



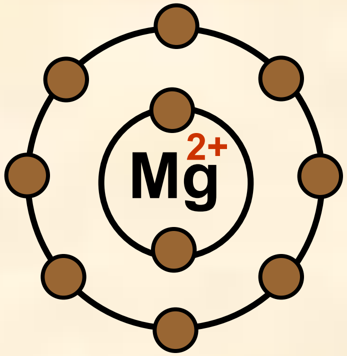
MAGNESIUM ATOM
2,8,2



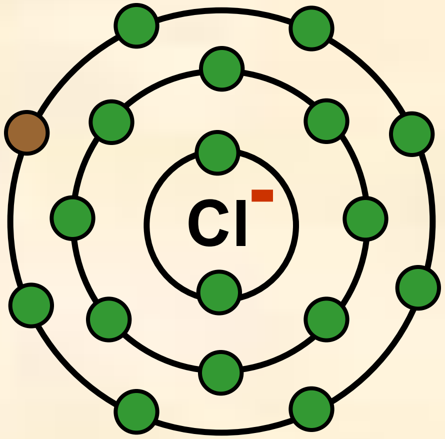
CHLORINE ATOMS
2,8,7



MAGNESIUM CHLORIDE



MAGNESIUM ION
2,8



CHLORIDE IONS
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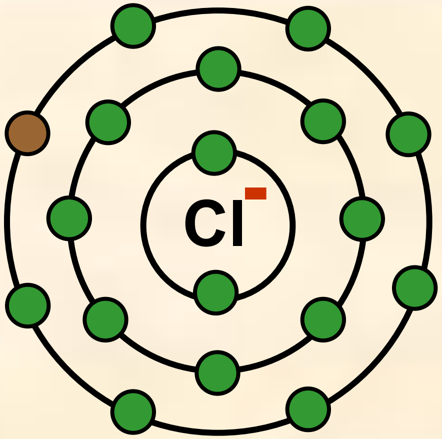
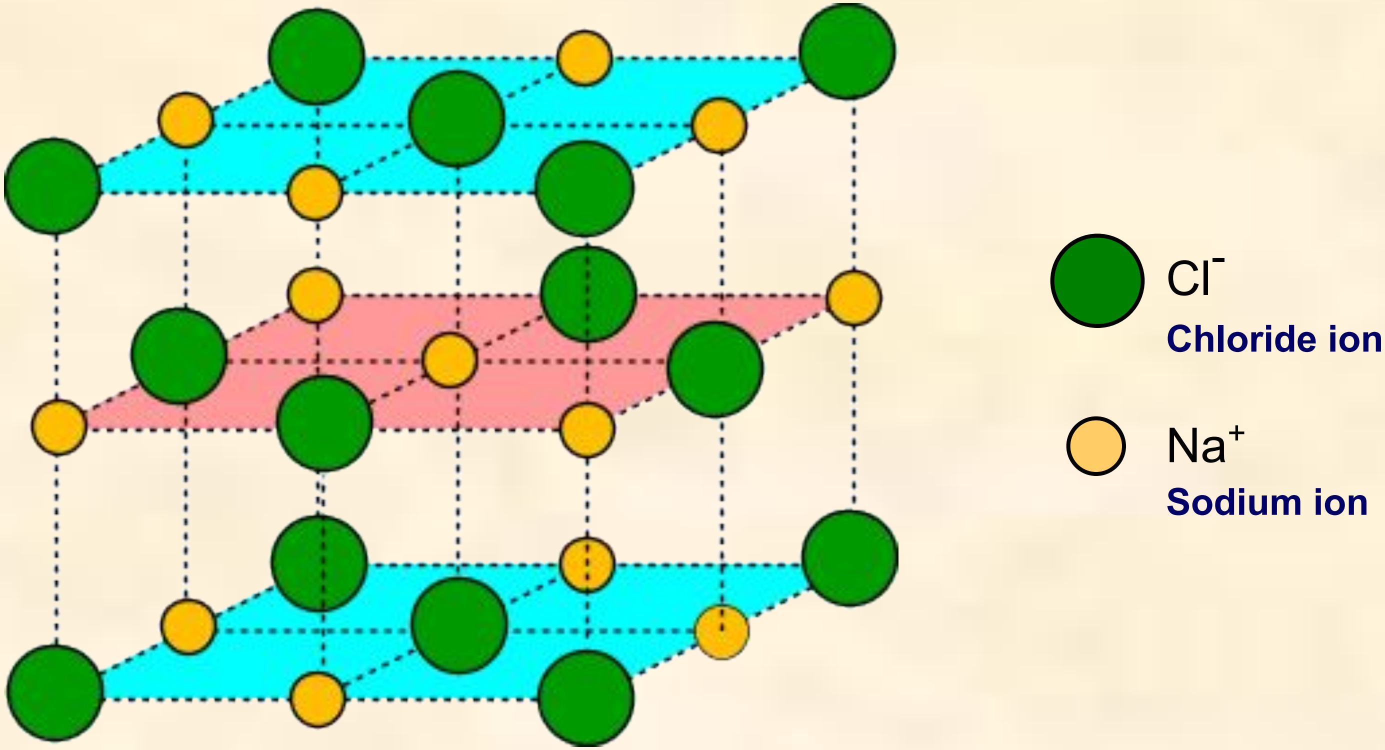


GIANT IONIC CRYSTAL LATTICE

Oppositely charged ions held in a regular 3-dimensional lattice by electrostatic attraction

The arrangement of ions in a crystal lattice depends on the relative sizes of the ions



The Na^+ ion is small enough relative to a Cl^- ion to fit in the spaces so that both ions occur in every plane.

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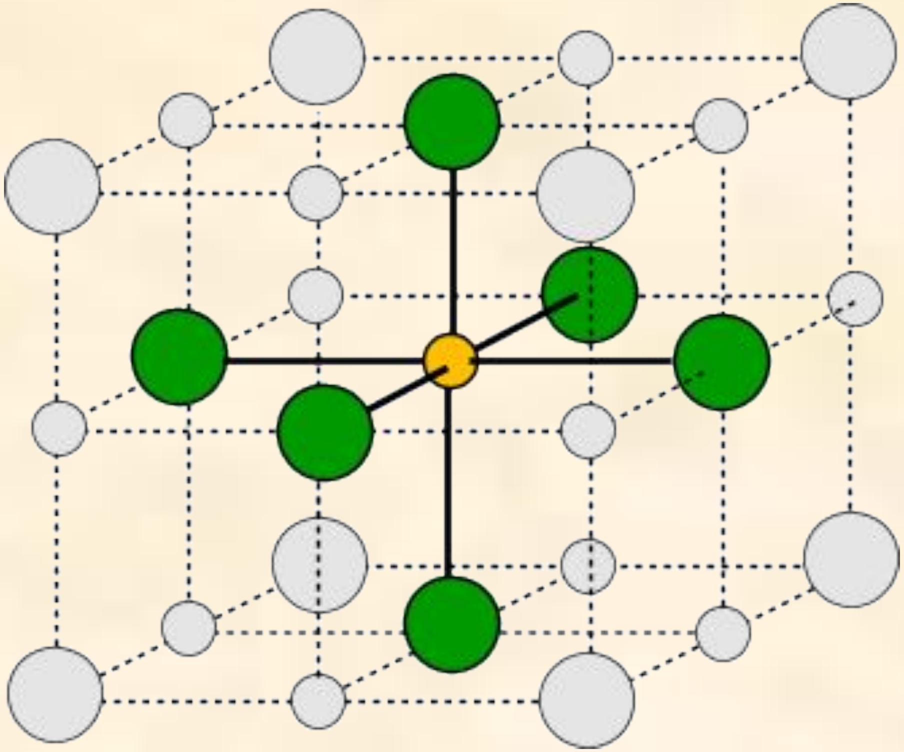

Each Na^+ is surrounded by 6 Cl^- (co-ordination number = 6)

and each Cl^- is surrounded by 6 Na^+ (co-ordination number = 6).

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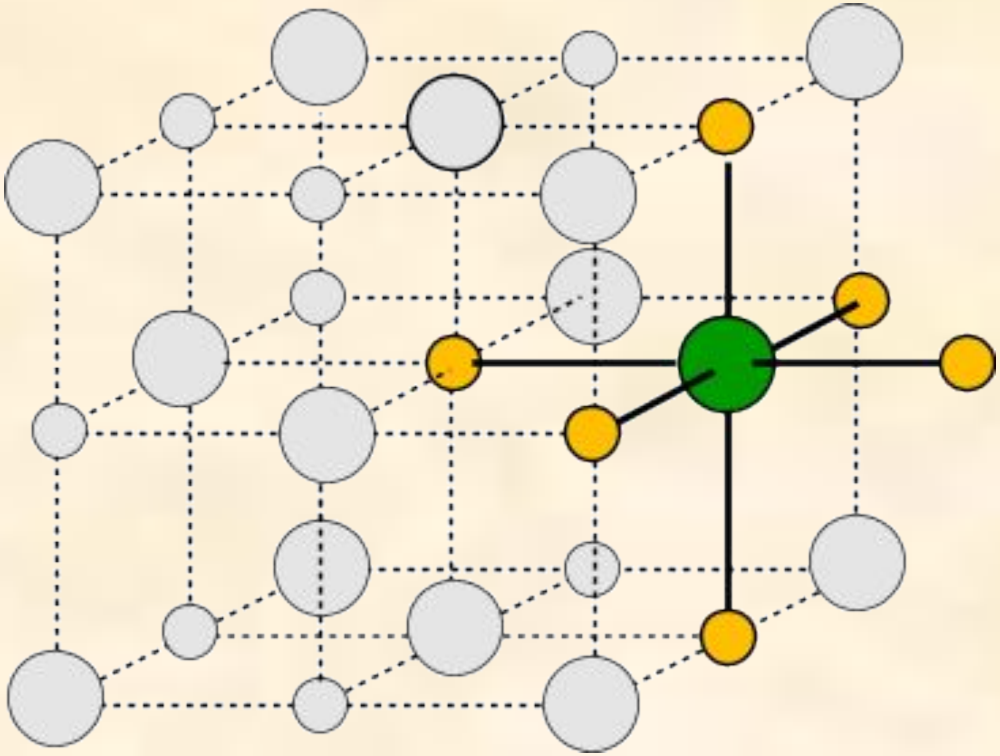

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Physical properties of ionic compounds

Melting point

very high A large amount of energy must be put in to overcome the strong electrostatic attractions and separate the ions.

Strength

Very brittle Any dislocation leads to the layers moving and similar ions being adjacent. The repulsion splits the crystal.

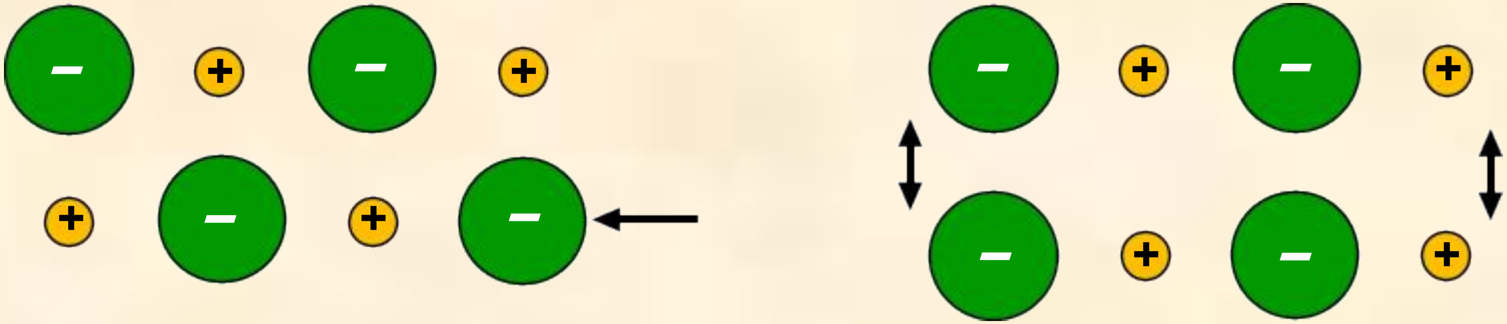
Electrical don't conduct when solid - ions held strongly in the lattice
conduct when molten or in aqueous solution - the ions become mobile and conduction takes place.

Solubility Insoluble in non-polar solvents but soluble in water
Water is a polar solvent and stabilises the separated ions.

Much energy is needed to overcome the electrostatic attraction and separate the ions stability attained by being surrounded by polar water molecules compensates for this

IONIC BONDING

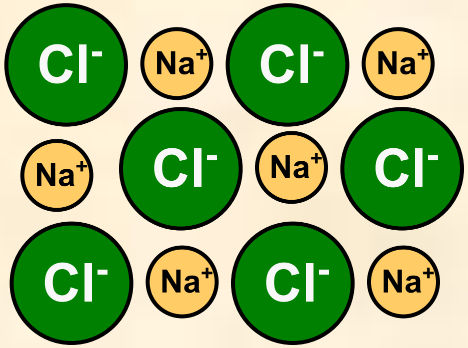
BRITTLE IONIC LATTICES



IF YOU MOVE A LAYER OF IONS, YOU GET IONS OF THE SAME CHARGE NEXT TO EACH OTHER. THE LAYERS REPEL EACH OTHER AND THE CRYSTAL BREAKS UP.

IONIC COMPOUNDS - ELECTRICAL PROPERTIES

SOLID IONIC COMPOUNDS DO NOT CONDUCT ELECTRICITY

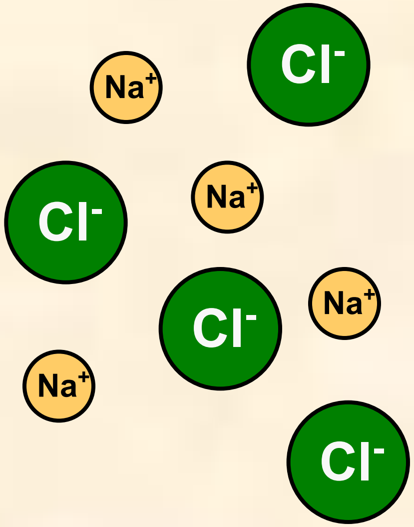


IONS ARE HELD STRONGLY TOGETHER

+ IONS CAN'T MOVE TO THE CATHODE

- IONS CAN'T MOVE TO THE ANODE

MOLTEN IONIC COMPOUNDS DO CONDUCT ELECTRICITY



IONS HAVE MORE FREEDOM IN A LIQUID SO CAN MOVE TO THE ELECTRODES

SOLUTIONS OF IONIC COMPOUNDS IN WATER DO CONDUCT ELECTRICITY

DISSOLVING AN IONIC COMPOUND IN WATER BREAKS UP THE STRUCTURE SO IONS ARE FREE TO MOVE TO THE ELECTRODES

COVALENT

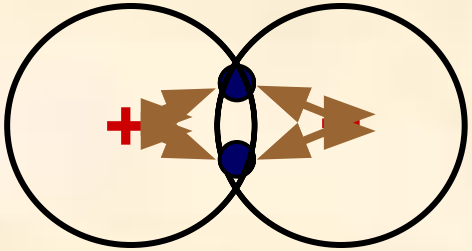
BONDING

COVALENT BONDING

Definition consists of a shared pair of electrons with one electron being supplied by each atom either side of the bond.

compare this with dative covalent bonding

atoms are held together
because their nuclei which
have an overall positive charge
are attracted to the shared electrons



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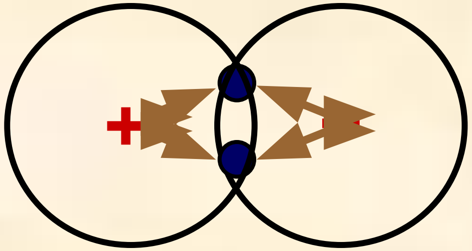

atoms are held together
because their nuclei which
have an overall positive charge
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Formation between atoms of the same element N_2, O_2, diamond, graphite

between atoms of different elements on the RHS of the table; CO_2, SO_2

when one of the elements is in the middle of the table; CCl_4, $SiCl_4$

with head-of-the-group elements with high ionisation energies; $BeCl_2$

COVALENT BONDING

- atoms share electrons to get the nearest noble gas electronic configuration

- some don’t achieve an “octet” as they haven’t got enough electrons

  eg Al in $AlCl_3$

- others share only some - if they share all they will exceed their “octet”

  eg NH_3 and H_2O

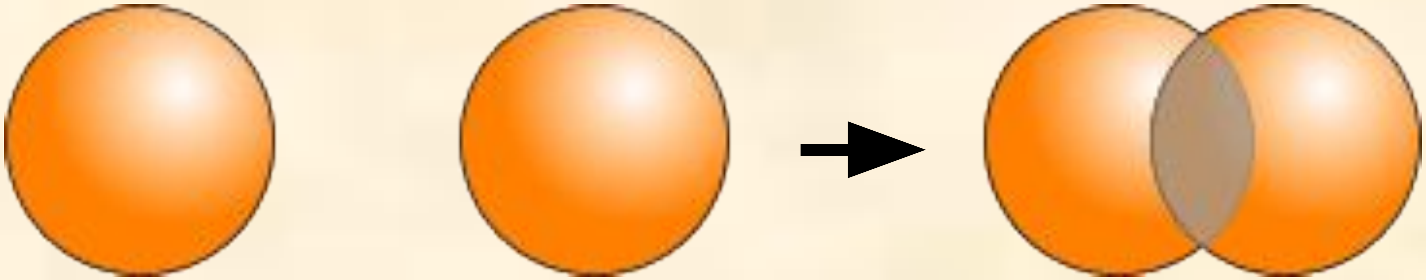
- atoms of elements in the 3rd period onwards can exceed their “octet” if they wish as they are not restricted to eight electrons in their “outer shell”

  eg PCl_5 and SF_6

SIMPLE MOLECULES

Orbital theory

Covalent bonds are formed when orbitals, each containing one electron, overlap. This forms a region in space where an electron pair can be found; new molecular orbitals are formed.



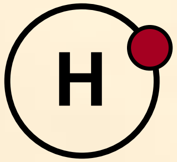
orbital containing 1 electron

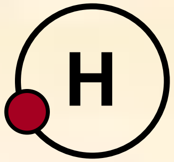
orbital containing 1 electron

overlap of orbitals provides a region in space which can contain a pair of electrons

The greater the overlap the stronger the bond.

HYDROGEN





Hydrogen atom needs
one electron to
complete its outer shell

Another hydrogen atom
also needs one electron to
complete its outer shell

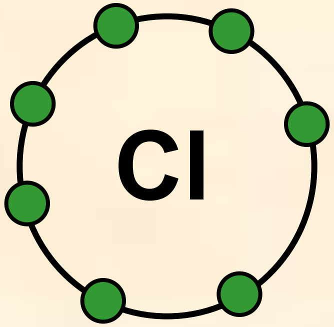
atoms share a pair of electrons to
form a single covalent bond

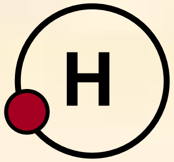
A hydrogen **MOLECULE** is formed

WAYS TO REPRESENT THE MOLECULE

PRESSING THE SPACE BAR WILL ACTIVATE EACH STEP OF THE ANIMATION

HYDROGEN CHLORIDE



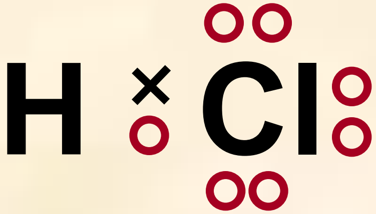


Chlorine atom
needs one electron
to complete its
outer shell

atoms share a pair of
electrons to form a
single covalent bond

Hydrogen atom also
needs one electron
to complete its outer
shell

WAYS TO REPRESENT THE MOLECULE



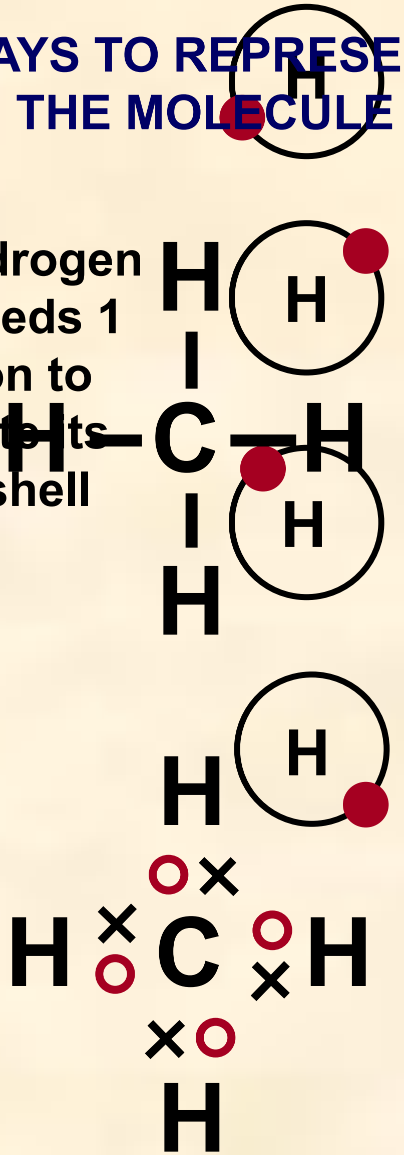
H—Cl

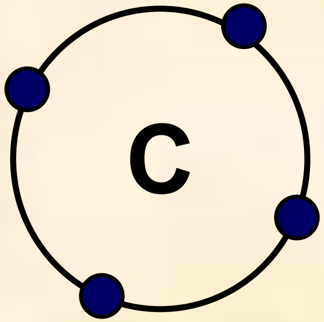
PRESSING THE SPACE BAR WILL ACTIVATE EACH STEP OF THE ANIMATION

METHANE

WAYS TO REPRESENT THE MOLECULE

Each hydrogen atom needs 1 electron to complete its outer shell





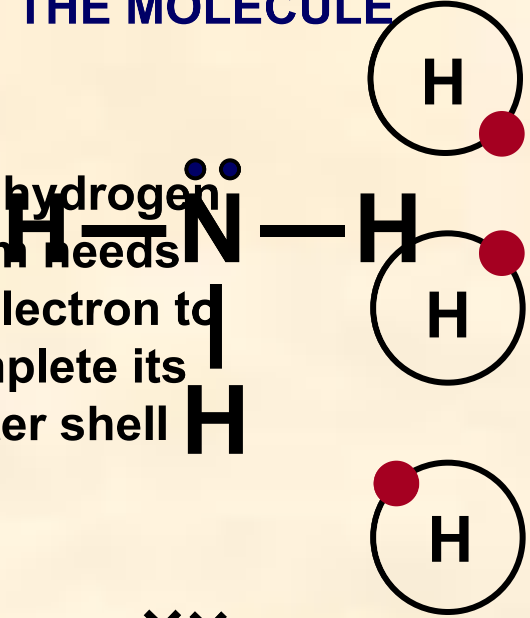
A carbon atom needs 4 electrons to complete its outer shell

Carbon shares all 4 of its electrons to form 4 single covalent bonds

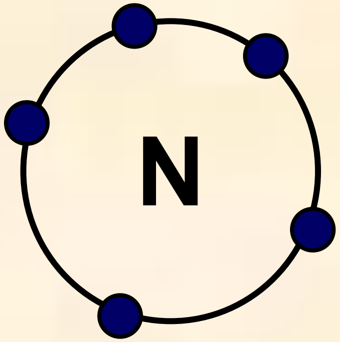
PRESSING THE SPACE BAR WILL ACTIVATE EACH STEP OF THE ANIMATION

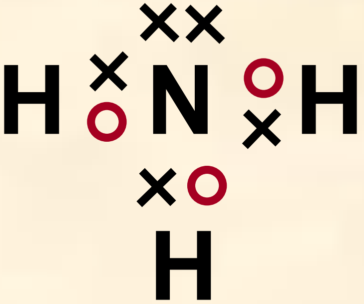
AMMONIA

WAYS TO REPRESENT THE MOLECULE



Each hydrogen atom needs one electron to complete its outer shell





Nitrogen atom needs 3 electrons to complete its outer shell

Nitrogen can only share 3 of its 5 electrons otherwise it will exceed the maximum of 8

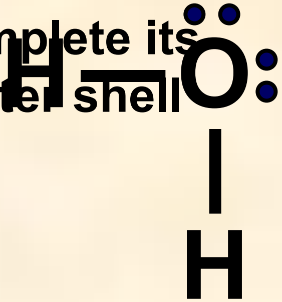
A LONE PAIR REMAINS

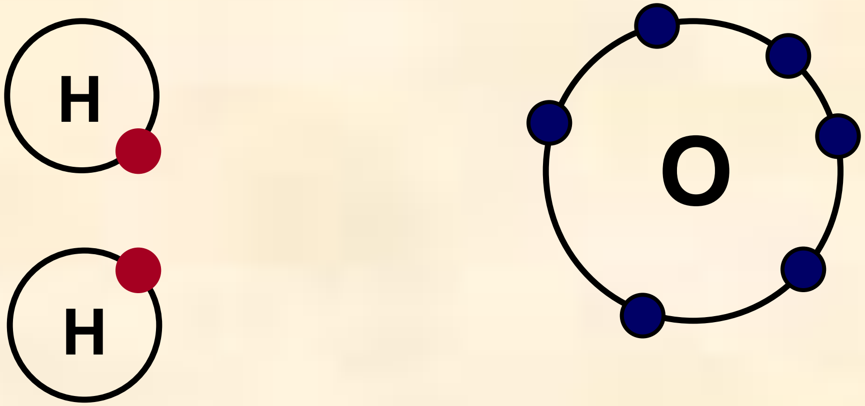
PRESSING THE SPACE BAR WILL ACTIVATE EACH STEP OF THE ANIMATION

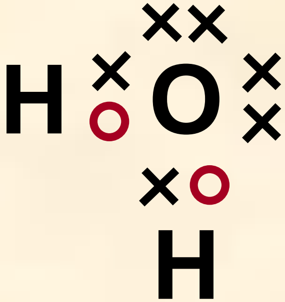
WATER

WAYS TO REPRESENT THE MOLECULE

Each hydrogen atom needs one electron to complete its outer shell







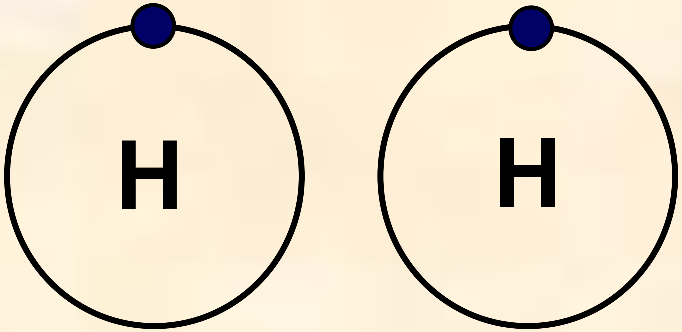
Oxygen atom needs 2 electrons to complete its outer shell

Oxygen can only share 2 of its 6 electrons otherwise it will exceed the maximum of 8

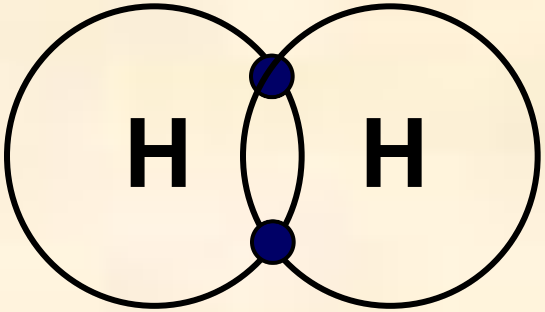
2 LONE PAIRS REMAIN

PRESSING THE SPACE BAR WILL ACTIVATE EACH STEP OF THE ANIMATION

HYDROGEN

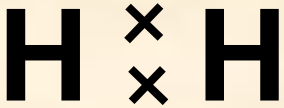


both atoms need one electron
to complete their outer shell



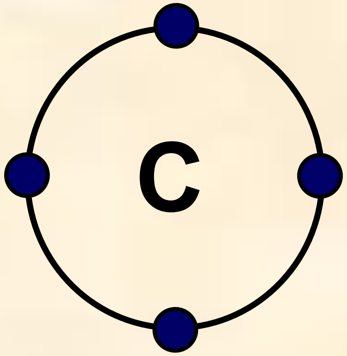
atoms share a pair of electrons
to form a single covalent bond

**DOT AND
CROSS
DIAGRAM**

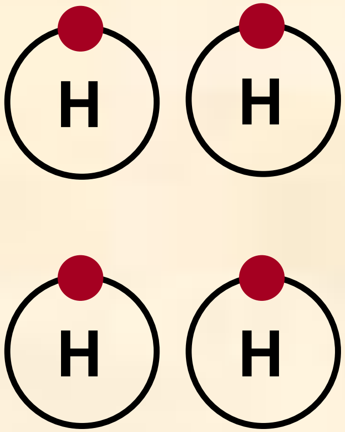


H—H

METHANE



atom needs four
electrons to complete
its outer shell



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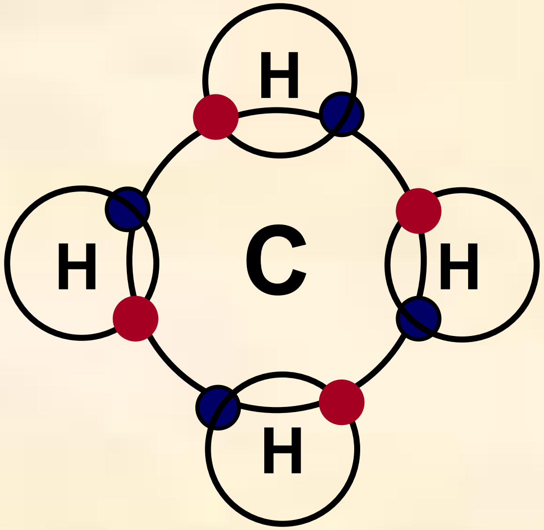
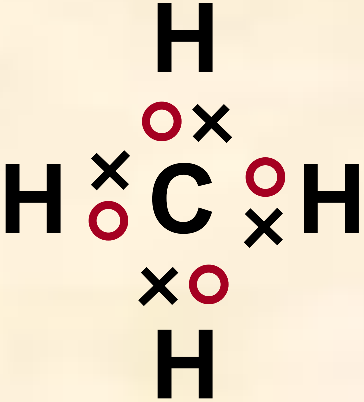
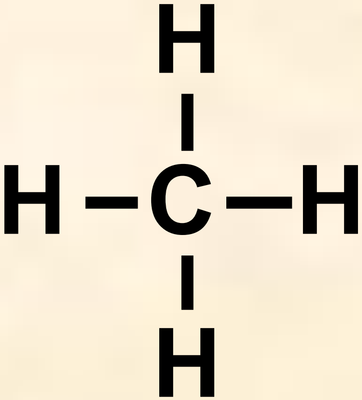


Carbon shares all 4 of its
electrons to form 4 single
covalent bonds

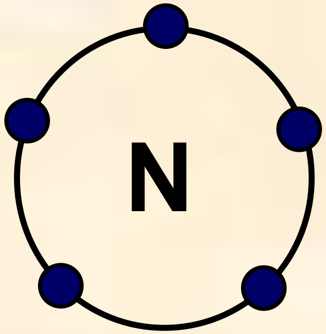
DOT AND
CROSS
DIAGRAM



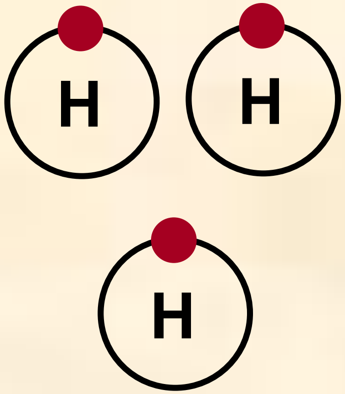


$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

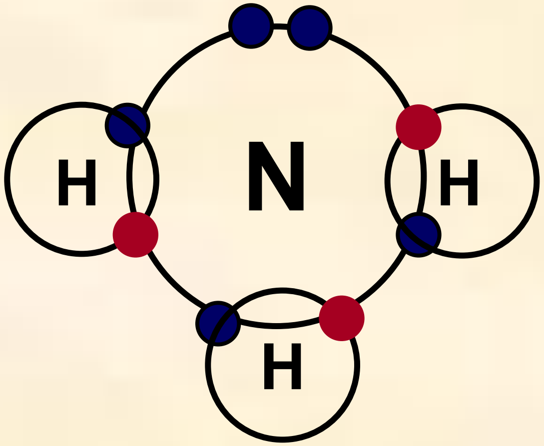
AMMONIA



atom needs three electrons to complete its outer shell

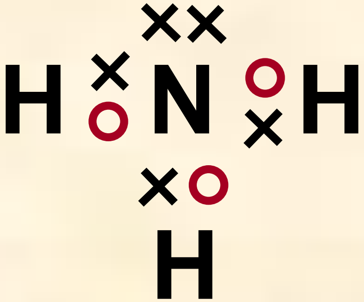


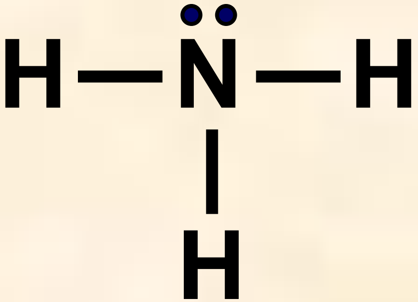
each atom needs one electron to complete its outer shell



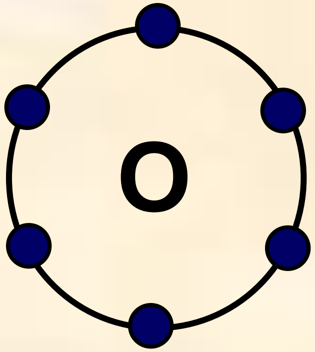
Nitrogen can only share 3 of its 5 electrons otherwise it will exceed the maximum of 8

A LONE PAIR REMAINS

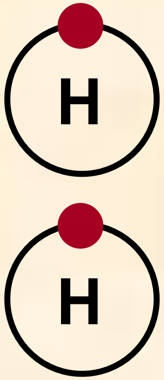




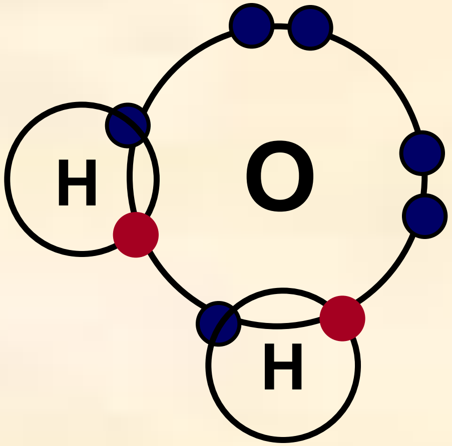
WATER



atom needs two
electrons to complete
its outer shell

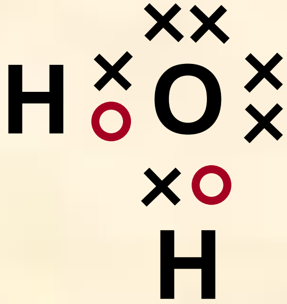


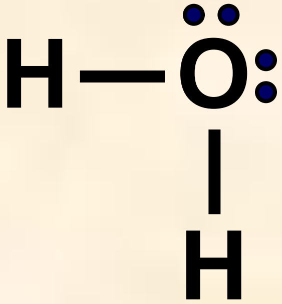
each atom needs one
electron to complete
its outer shell



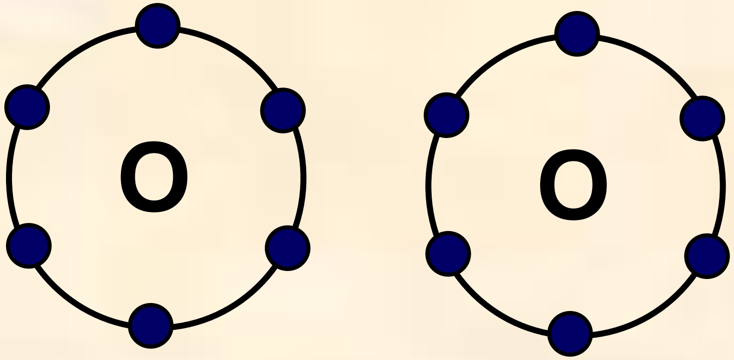
Oxygen can only share 2 of its
6 electrons otherwise it will
exceed the maximum of 8

TWO LONE PAIRS REMAIN

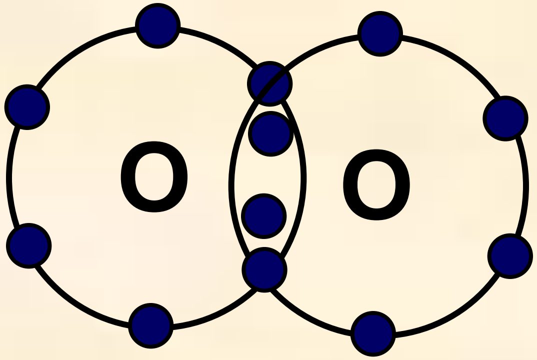




OXYGEN



each atom needs two electrons
to complete its outer shell



each oxygen shares 2 of its
electrons to form a

DOUBLE COVALENT BOND

O=O

SIMPLE COVALENT MOLECULES

Bonding Atoms are joined together within the molecule by covalent bonds.

Electrical Don’t conduct electricity as they have no mobile ions or electrons

Solubility Tend to be more soluble in organic solvents than in water; some are hydrolysed

Boiling point Low - intermolecular forces (van der Waals’ forces) are weak; they increase as molecules get a larger surface area

e.g. CH_4 -161°C C_2H_6 - 88°C C_3H_8 -42°C

as the intermolecular forces are weak, little energy is required to to separate molecules from each other so boiling points are low

some boiling points are higher than expected for a given mass because you can get additional forces of attraction

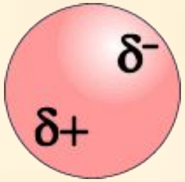
VAN DER WAALS' FORCES

INSTANTANEOUS DIPOLE-INDUCED DIPOLE FORCES

Although the bonding within molecules is strong, that between molecules is weak. Molecules and monatomic noble gases are subject to weak attractive forces.

Instantaneous dipole-induced dipole forces

Because electrons move quickly in orbitals, their position is constantly changing; at any given instant they could be anywhere in an atom. The possibility will exist that one side will have more electrons than the other. This will give rise to a dipole...



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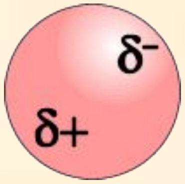
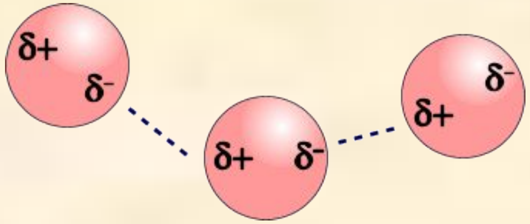


The dipole on one atom induces dipoles on nearby atoms

Atoms are now attracted to each other by a weak forces

The greater the number of electrons, the stronger the attraction and the greater the energy needed to separate the particles.



VAN DER WAALS’ FORCES

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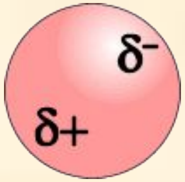
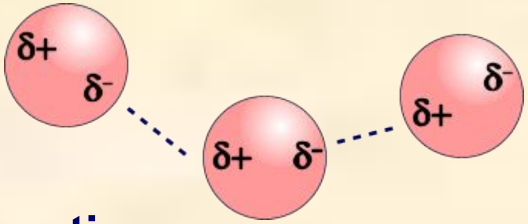


The dipole on one atom induces dipoles on others

Atoms are now attracted to each other by a weak forces



The greater the number of electrons, the stronger the attraction and the greater the energy needed to separate the particles.

NOBLE GASES	Electrons	B pt.
He	2	-269°C
Ne	10	-246°C
Ar	18	-186°C
Kr	36	-152°C

ALKANES	Electrons	B pt.
CH_4	10	-161°C
C_2H_6	18	- 88°C
C_3H_8	26	- 42°C

ELECTRONEGATIVITY

‘The ability of an atom to attract the electron pair in a covalent bond to itself’

Non-polar bond similar atoms have the same electronegativity
they will both pull on the electrons to the same extent
the electrons will be equally shared

Polar bond different atoms have different electronegativities
one will pull the electron pair closer to its end
it will be slightly more negative than average, d-
the other will be slightly less negative, or more positive, d+
a dipole is formed and the bond is said to be polar
greater electronegativity difference = greater polarity

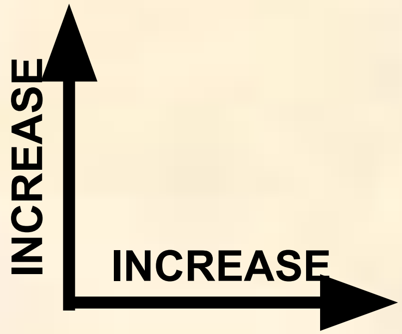
Pauling Scale a scale for measuring electronegativity

ELECTRONEGATIVITY

‘The ability of an atom to attract the electron pair in a covalent bond to itself’

Pauling Scale a scale for measuring electronegativity

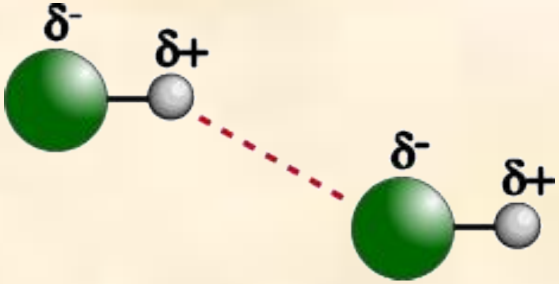
values increase across periods
values decrease down groups
fluorine has the highest value



H						
2.1						
Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.2	1.5	1.8	2.1	2.5	3.0
K				Br		
0.8				2.8		

DIPOLE-DIPOLE INTERACTION

Occurrence occurs between molecules containing polar bonds
acts in addition to the basic van der Waals' forces
the extra attraction between dipoles means that
more energy must be put in to separate molecules
get higher boiling points than expected for a given mass



Boiling points of hydrides

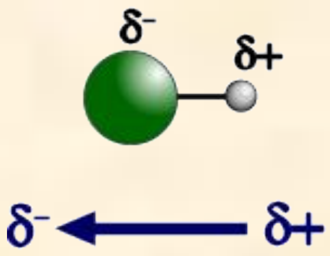
	Mr	°C
CH_4	16	-161
SiH_4	32	-117
GeH_4	77	-90
SnH_4	123	-50
NH_3	17	-33
PH_3	34	-90
AsH_3	78	-55
SbH_3	125	-17

	Mr	°C
H_2O	18	+100
H_2S	34	-61
H_2Se	81	-40
H_2Te	130	-2
HF	20	+20
HCl	36.5	-85
HBr	81	-69
HI	128	-35

POLAR MOLECULES

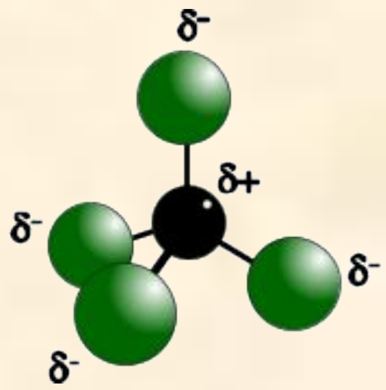
Occurrence not all molecules containing polar bonds are polar overall
if bond dipoles ‘cancel each other’ the molecule isn’t polar
if there is a ‘net dipole’ the molecule will be polar

HYDROGEN CHLORIDE



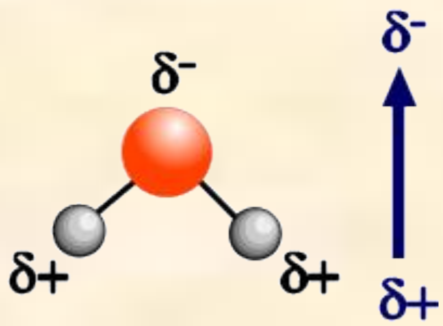
NET DIPOLE - POLAR

TETRACHLOROMETHANE



NON-POLAR

WATER



NET DIPOLE - POLAR

POLAR MOLECULES

Evidence place a liquid in a burette
allow it to run out
place a charged rod alongside the stream of liquid
polar molecules are attracted by electrostatic attraction
non-polar molecules will be unaffected

NET DIPOLE - POLAR

NON-POLAR

BOILING POINTS OF HYDRIDES

GROUP IV

	Mr	°C
CH_4	16	-161
SiH_4	32	-117
GeH_4	77	-90
SnH_4	123	-50

GROUP V

NH_3	17	-33
PH_3	34	-90
AsH_3	78	-55
SbH_3	125	-17

GROUP VI

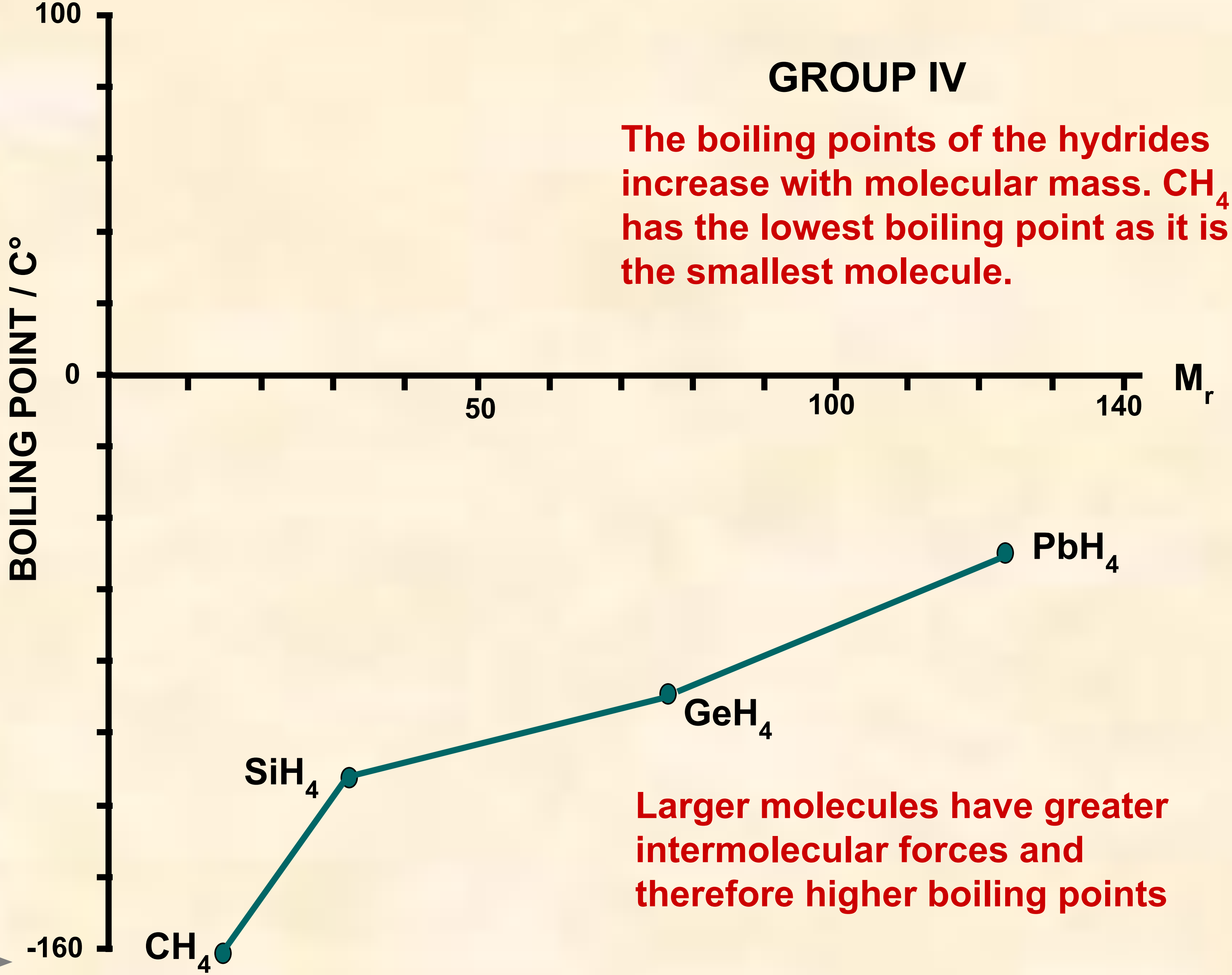
	Mr	°C
H_2O	18	+100
H_2S	34	-61
H_2Se	81	-40
H_2Te	130	-2

GROUP VII

HF	20	+20
HCl	36.5	-85
HBr	81	-69
HI	128	-35

The values of certain hydrides are not typical of the trend you would expect

BOILING POINTS OF HYDRIDES

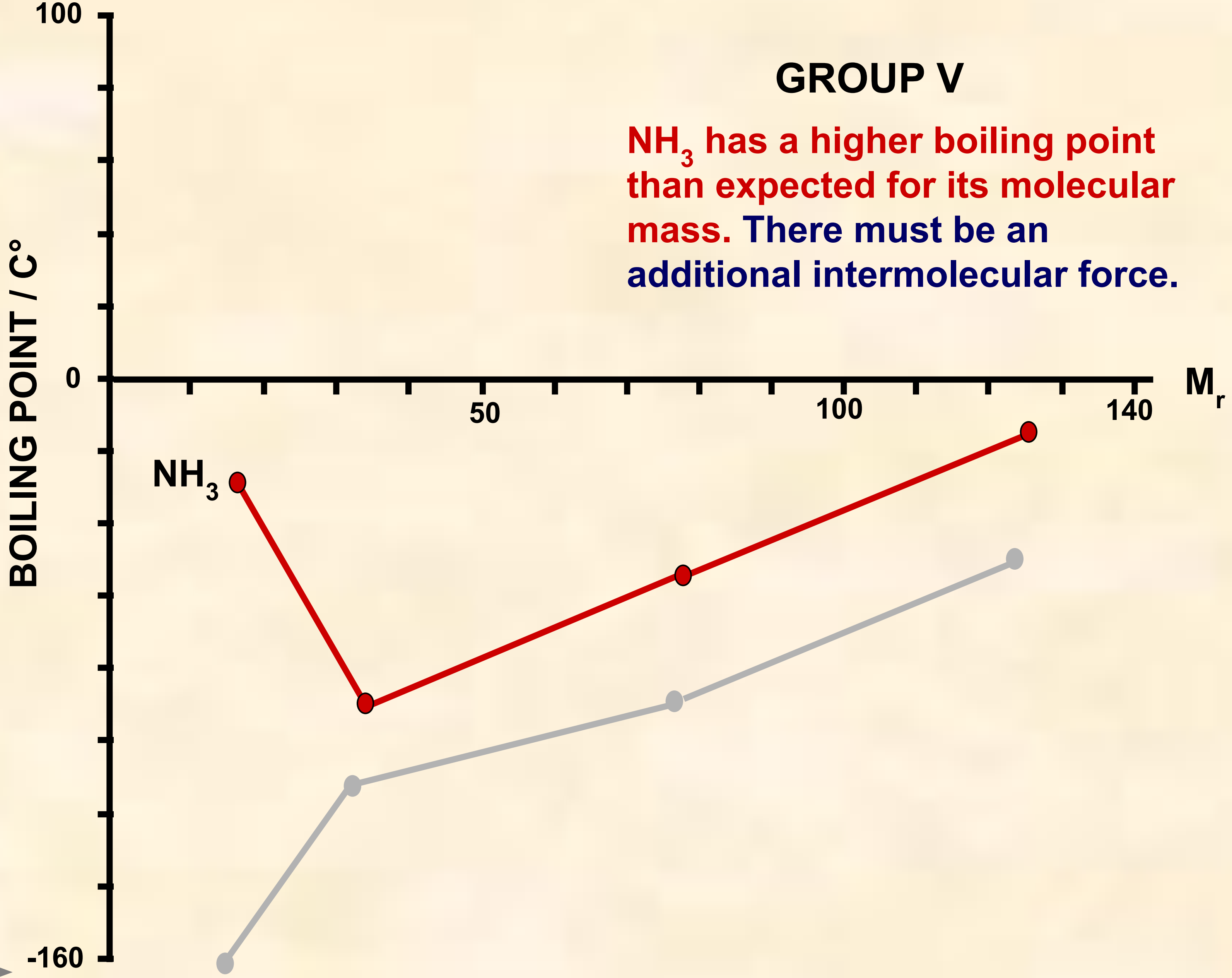


GROUP IV

The boiling points of the hydrides increase with molecular mass. CH_4 has the lowest boiling point as it is the smallest molecule.

Larger molecules have greater intermolecular forces and therefore higher boiling points

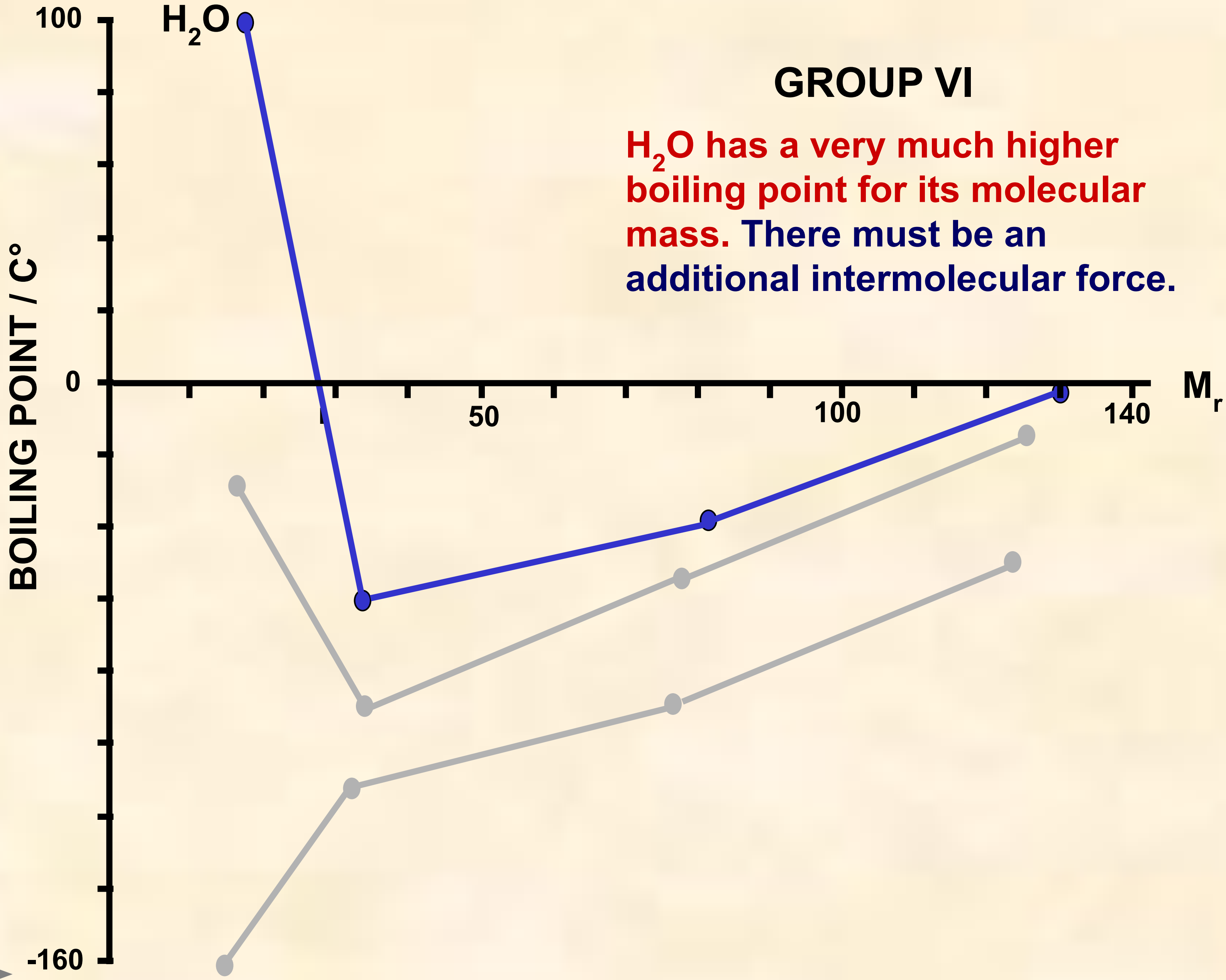
BOILING POINTS OF HYDRIDES



GROUP V

NH_3 has a higher boiling point than expected for its molecular mass. There must be an additional intermolecular force.

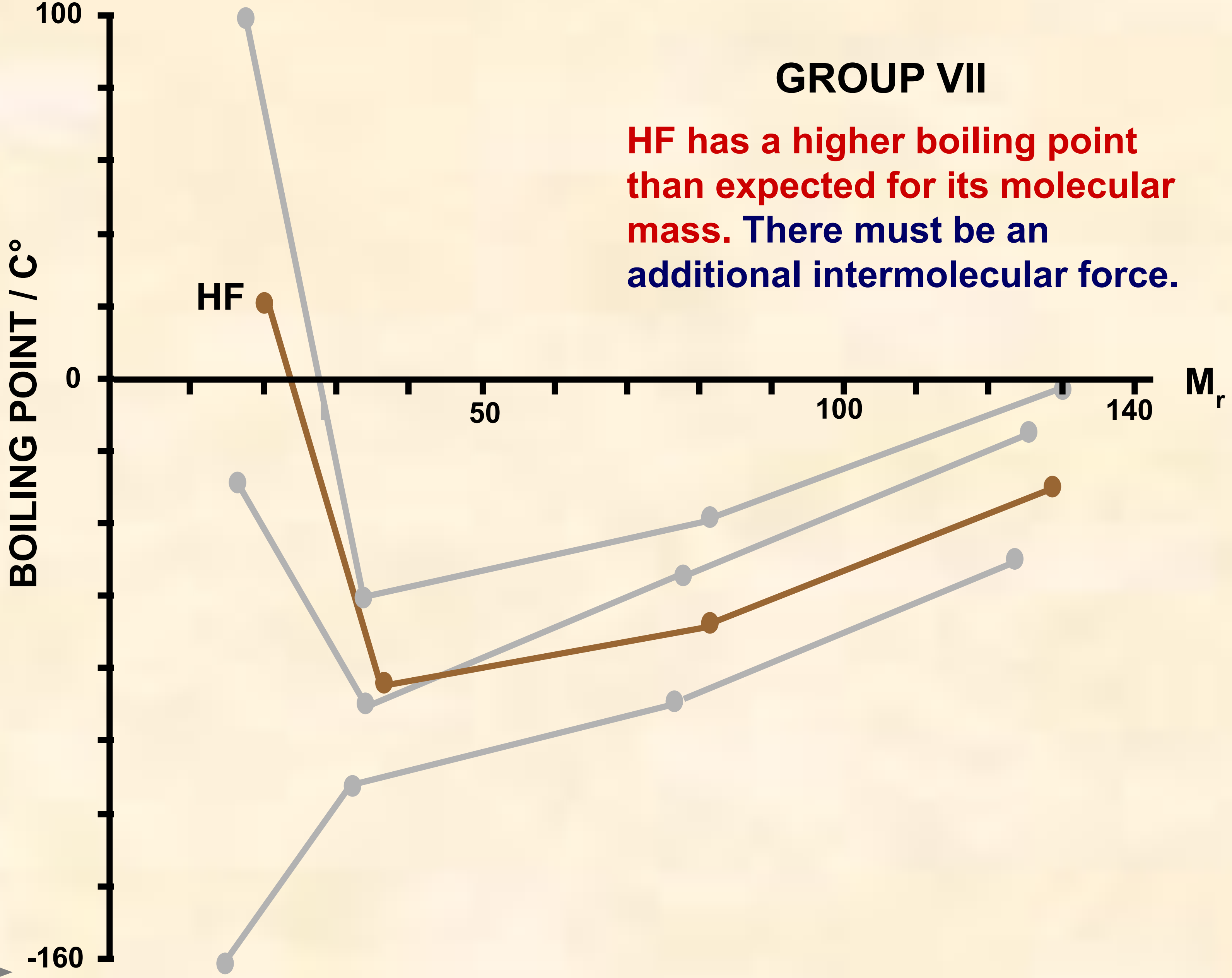
BOILING POINTS OF HYDRIDES



GROUP VI

H_2O has a very much higher boiling point for its molecular mass. There must be an additional intermolecular force.

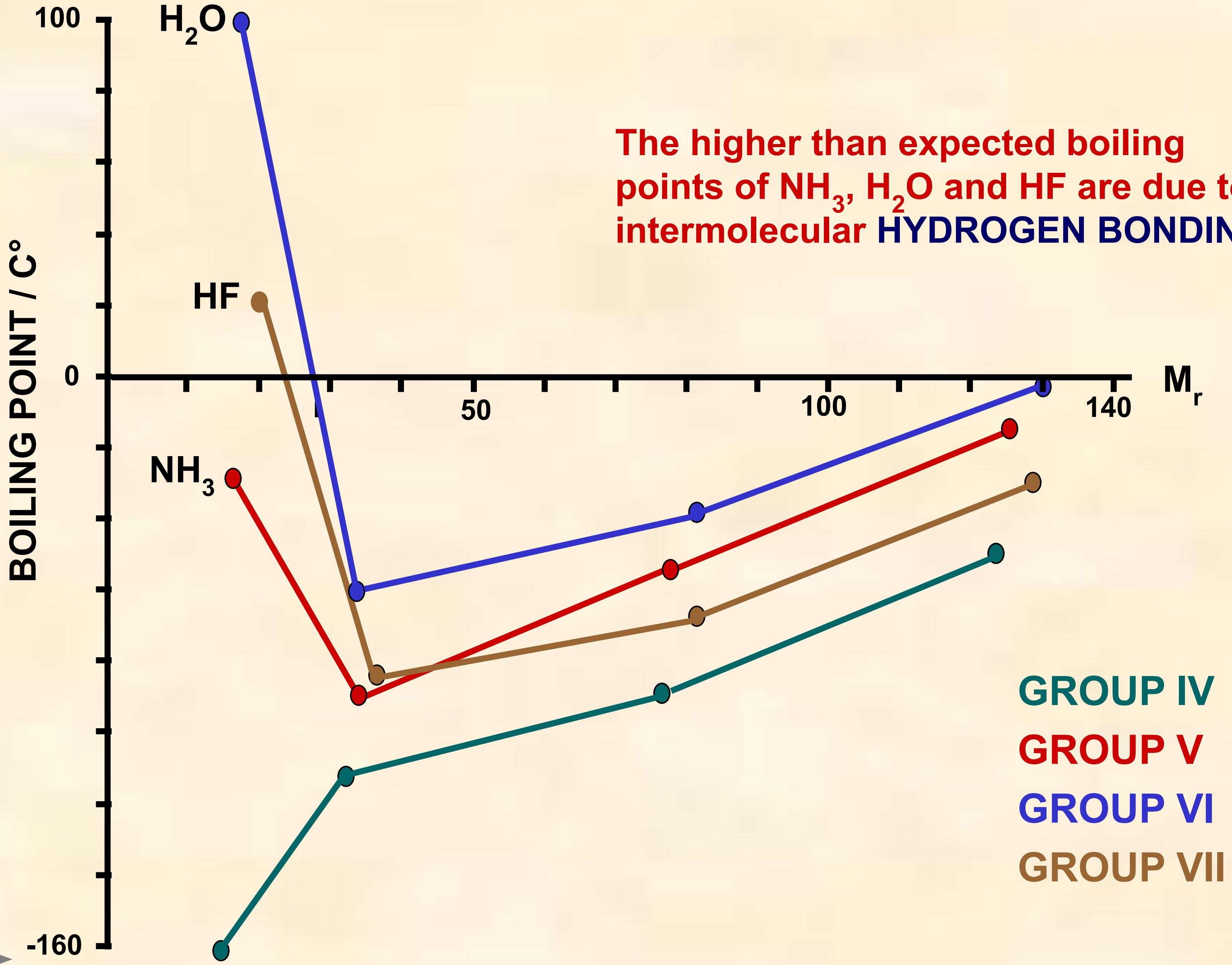
BOILING POINTS OF HYDRIDES



GROUP VII

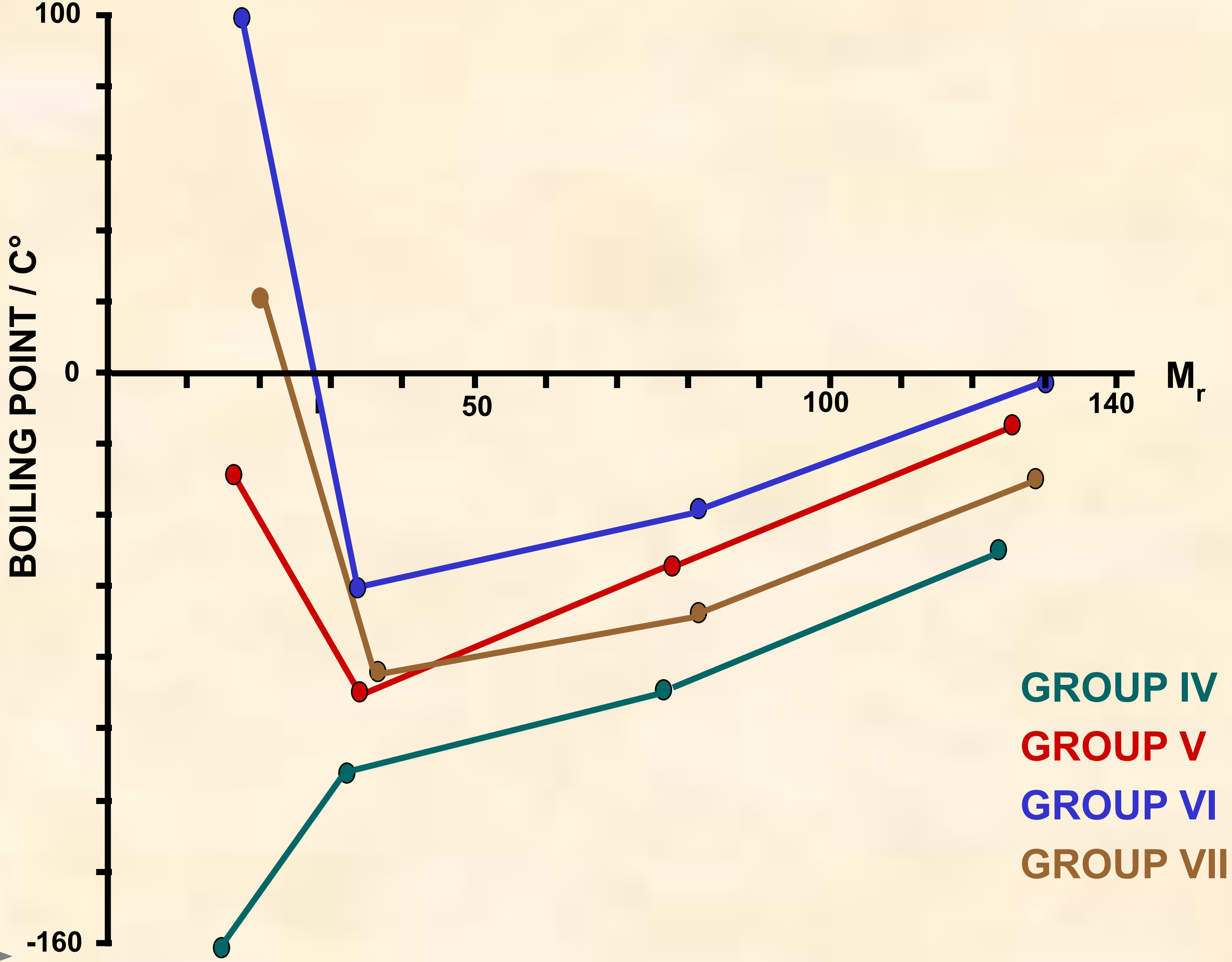
HF has a higher boiling point than expected for its molecular mass. There must be an additional intermolecular force.

BOILING POINTS OF HYDRIDES



The higher than expected boiling points of NH_3, H_2O and HF are due to intermolecular **HYDROGEN BONDING**

BOILING POINTS OF HYDRIDES



HYDROGEN BONDING

- an extension of dipole-dipole interaction
- gives rise to even higher boiling points
- bonds between H and the three most electronegative elements, F, O and N are extremely polar
- because of the small sizes of H, F, N and O the partial charges are concentrated in a small volume thus leading to a high charge density
- makes the intermolecular attractions greater and leads to even higher boiling points

HYDROGEN BONDING - ICE

each water molecule is hydrogen-bonded to 4 others in a tetrahedral formation

ice has a “diamond-like” structure

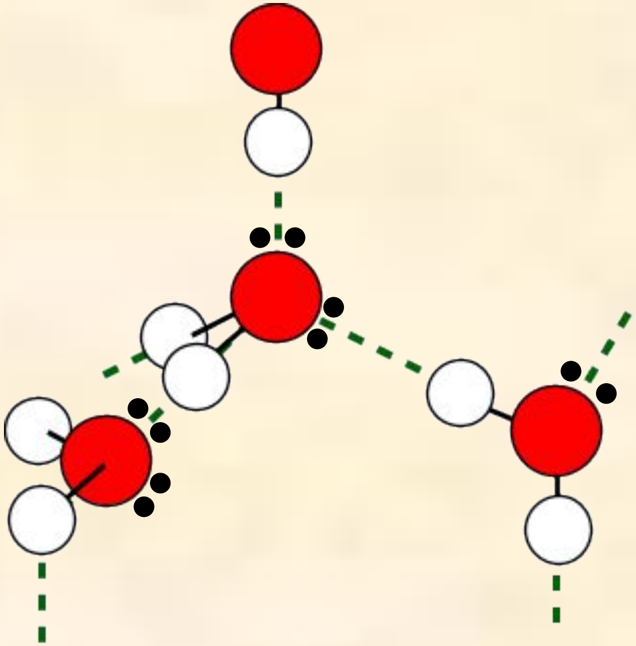
volume is larger than the liquid making it

when ice melts, the structure collapses slightly and the molecules come closer; they then move a little further apart as they get more energy as they warm up

this is why…

a) water has a maximum density at 4°C

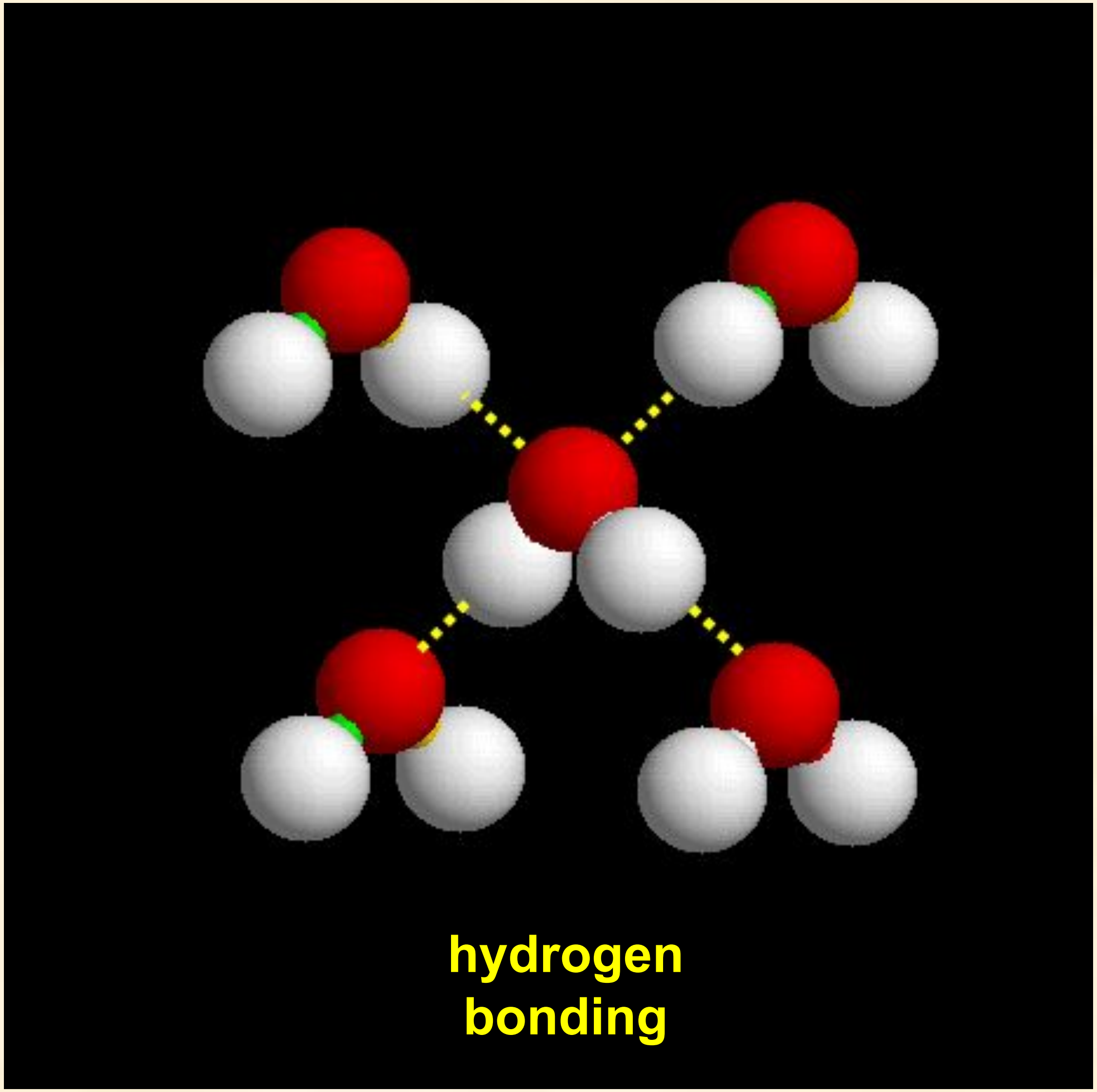
b) ice floats.



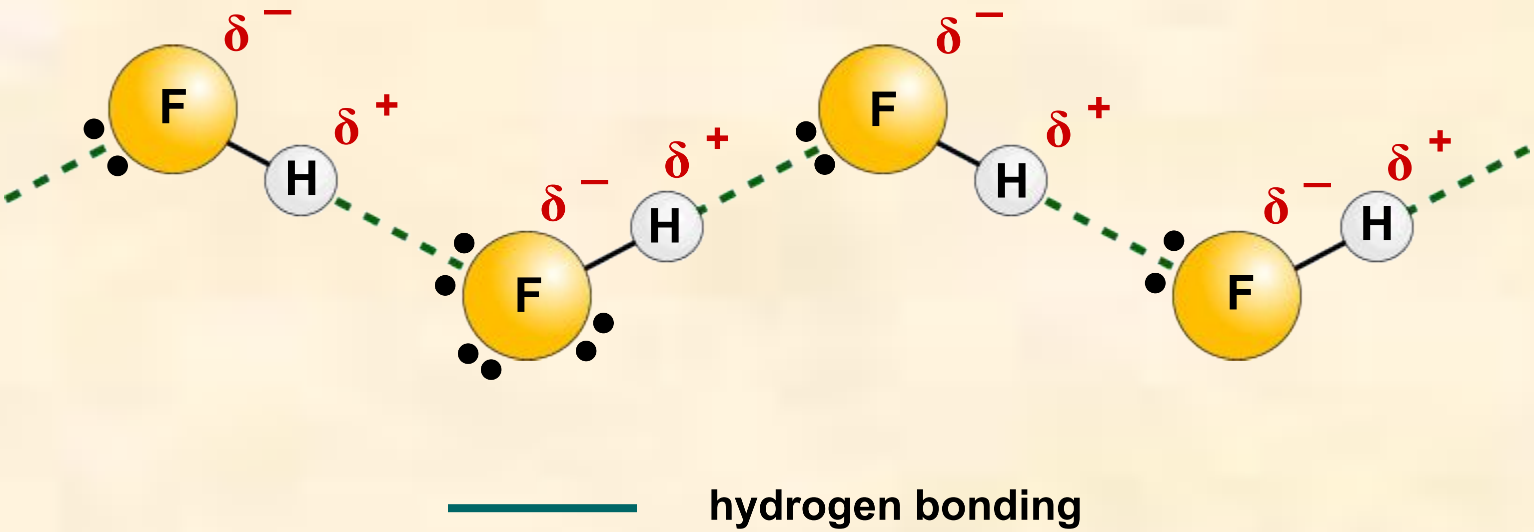
—— hydrogen bonding

●● lone pair

HYDROGEN BONDING - ICE



HYDROGEN BONDING - HF



Hydrogen fluoride has a much higher boiling point than one would expect for a molecule with a relative molecular mass of 20

Fluorine has the highest electronegativity of all and is a small atom so the bonding with hydrogen is extremely polar

DATIVE COVALENT (CO-ORDINATE) BONDING

A dative covalent bond differs from covalent bond only in its formation

Both electrons of the shared pair are provided by one species (donor) and it shares the electrons with the acceptor

Donor species will have lone pairs in their outer shells

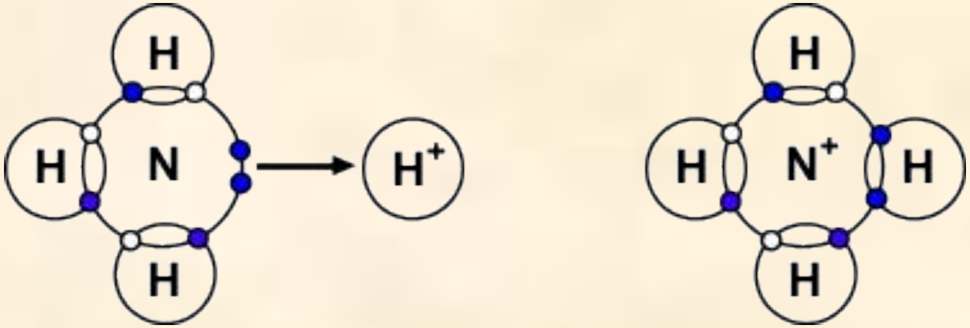
Acceptor species will be short of their “octet” or maximum.

Lewis base	a lone pair donor
Lewis acid	a lone pair acceptor

Ammonium ion, NH_4^+

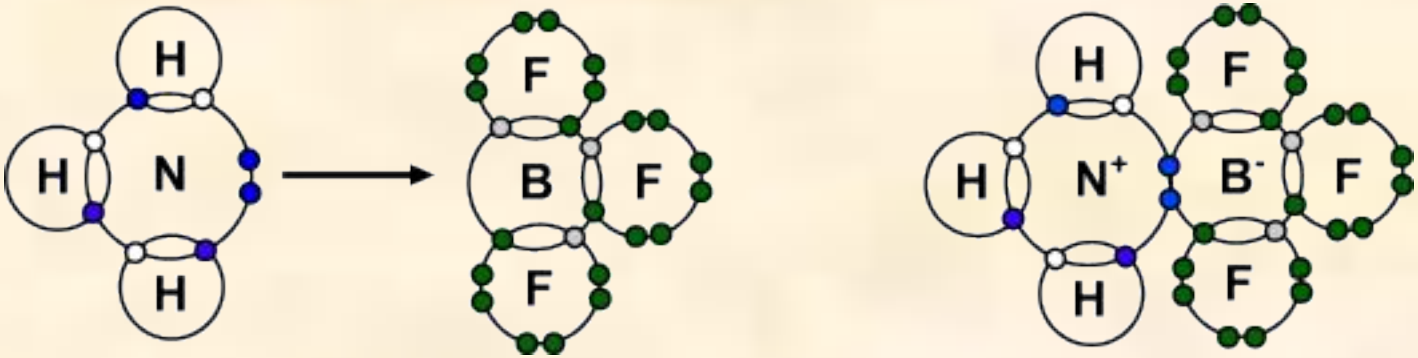
The lone pair on N is used to share with the hydrogen ion which needs two electrons to fill its outer shell.

The N now has a +ive charge as - it is now sharing rather than owning two electrons.



Boron trifluoride-ammonia NH_3BF_3

Boron has an incomplete shell in BF_3 and can accept a share of a pair of electrons donated by ammonia. The B becomes -ive as it is now shares a pair of electrons (i.e. it is up one electron) it didn't have before.



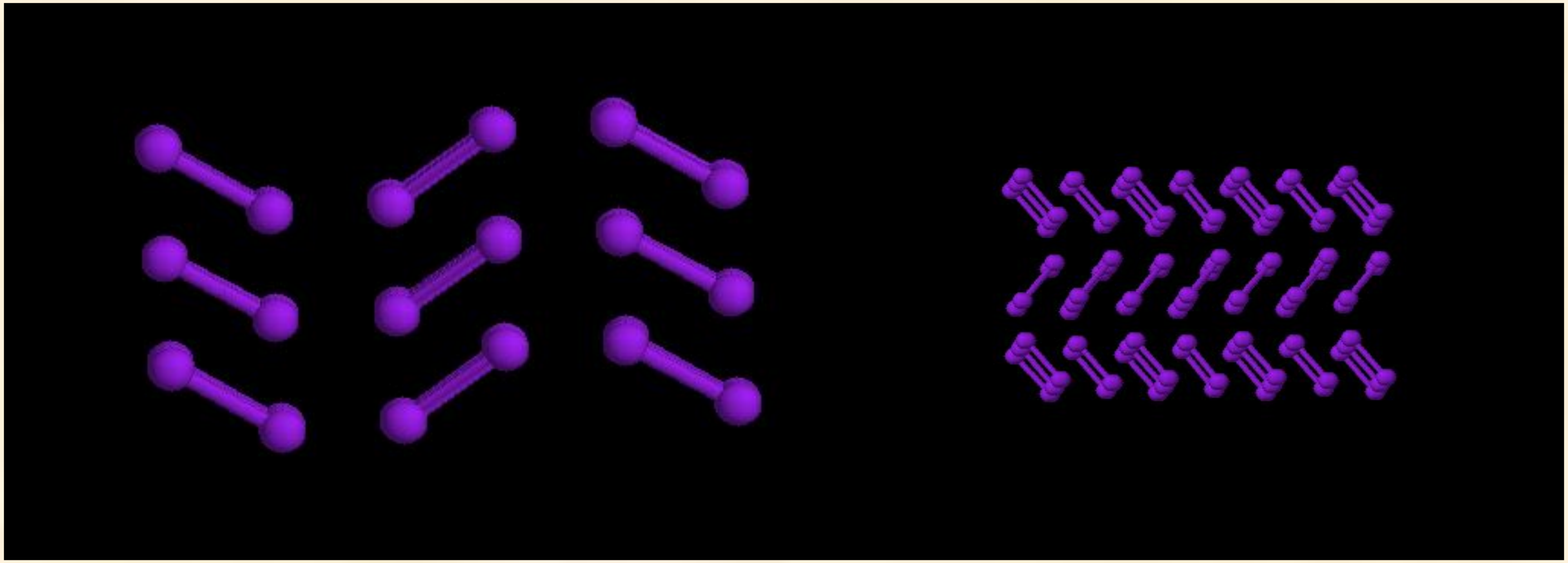
MOLECULAR

SOLIDS

MOLECULAR SOLIDS

IODINE

At room temperature and pressure, iodine is a greyish solid. However it doesn't need to be warmed much in order to produce a purple vapour. This is because iodine is composed of diatomic molecules (I_2) which exist in an ordered molecular crystal in the solid state. Each molecule is independent of the others, only being attracted by van der Waals' forces. Therefore, little energy is required to separate the iodine molecules.



COVALENT NETWORKS

GIANT MOLECULES

MACROMOLECULES

They all mean the same!

GIANT (MACRO) MOLECULES

DIAMOND, GRAPHITE and SILICA

Many atoms joined together in a regular array
by a large number of covalent bonds

GENERAL PROPERTIES

MELTING POINT **Very high**
structure is made up of a large number of covalent bonds,
all of which need to be broken if atoms are to be separated

ELECTRICAL **Don’t conduct electricity - have no mobile ions or electrons**
but... Graphite conducts electricity

STRENGTH **Hard - exists in a rigid tetrahedral structure**
Diamond and silica (SiO_2)... but
Graphite is soft

GIANT (MACRO) MOLECULES

DIAMOND

MELTING POINT VERY HIGH

many covalent bonds must be broken to separate atoms

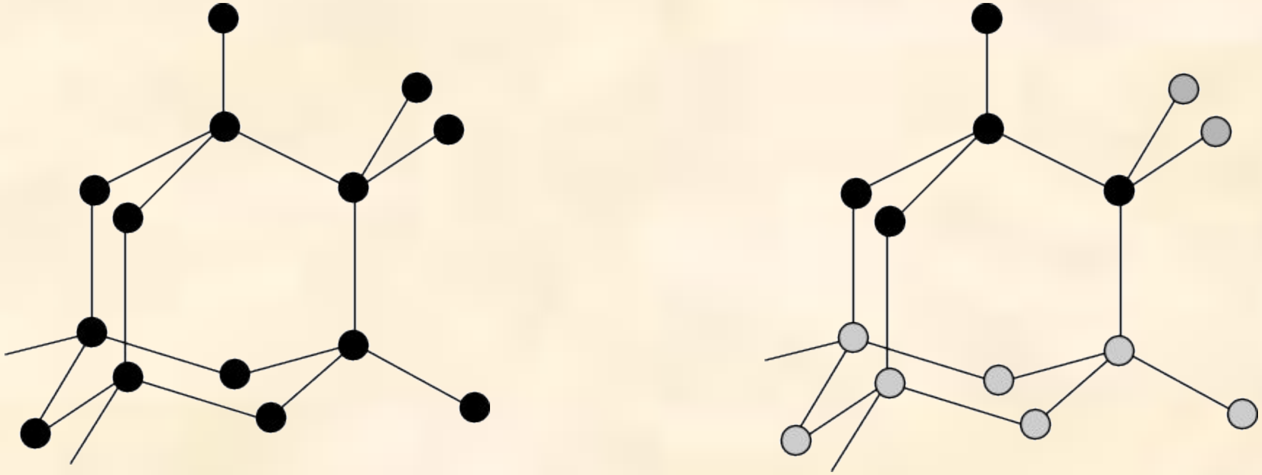
STRENGTH STRONG

each carbon is joined to four others in a rigid structure

Coordination Number = 4

ELECTRICAL NON-CONDUCTOR

No free electrons - all 4 carbon electrons used for bonding



GIANT (MACRO) MOLECULES

GRAPHITE

MELTING POINT VERY HIGH

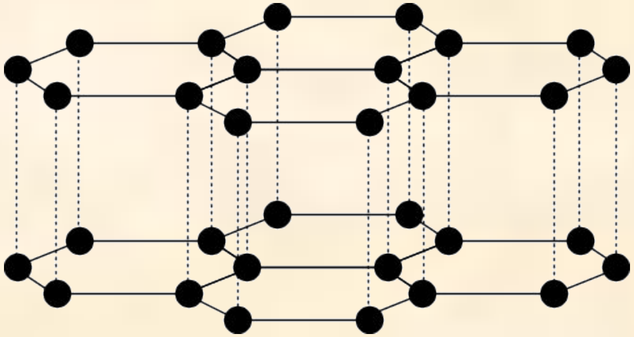
many covalent bonds must be broken to separate atoms

STRENGTH SOFT

each carbon is joined to three others in a layered structure
Coordination Number = 3
layers are held by weak van der Waals' forces
can slide over each other

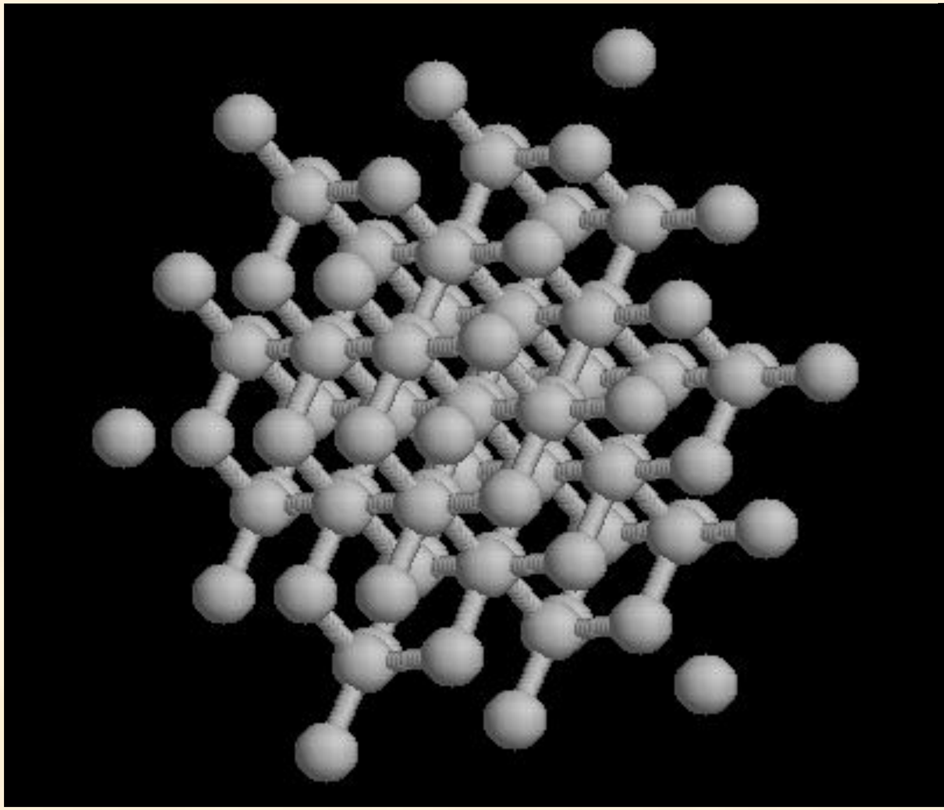
ELECTRICAL CONDUCTOR

Only three carbon electrons are used for bonding which leaves the fourth to move freely along layers

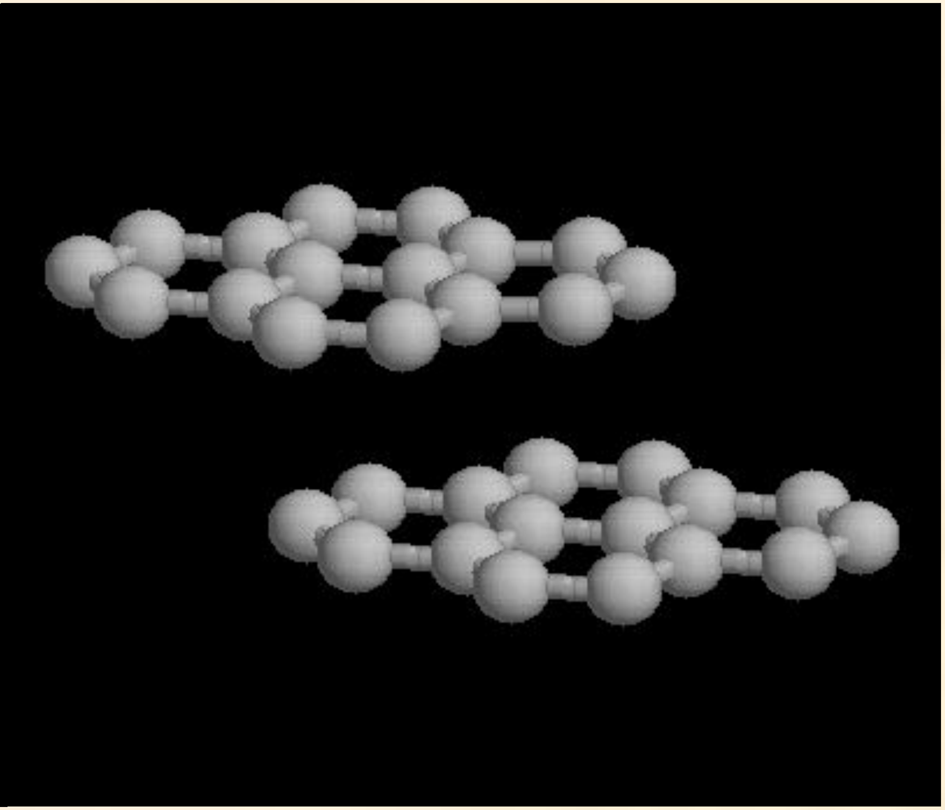


layers can slide over each other
used as a lubricant and in pencils

GIANT (MACRO) MOLECULES



DIAMOND



GRAPHITE

GIANT (MACRO) MOLECULES

SILICA

MELTING POINT VERY HIGH

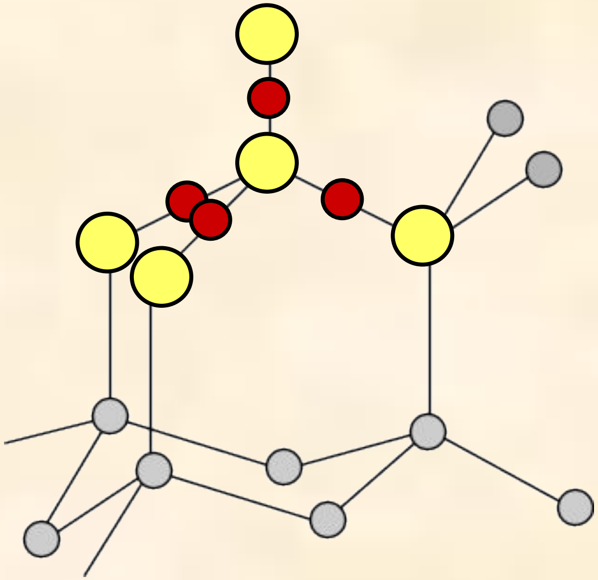
many covalent bonds must be broken to separate atoms

STRENGTH STRONG

each silicon atom is joined to four oxygens - C No. = 4

each oxygen atom are joined to two silicons - C No = 2

ELECTRICAL NON-CONDUCTOR - no mobile electrons



METALLIC
BONDING

METALLIC BONDING

Involves a lattice of positive ions surrounded by delocalised electrons

Metal atoms achieve stability by “off-loading” electrons to attain the electronic structure of the nearest noble gas. These electrons join up to form a mobile cloud which prevents the newly-formed positive ions from flying apart due to repulsion between similar charges.

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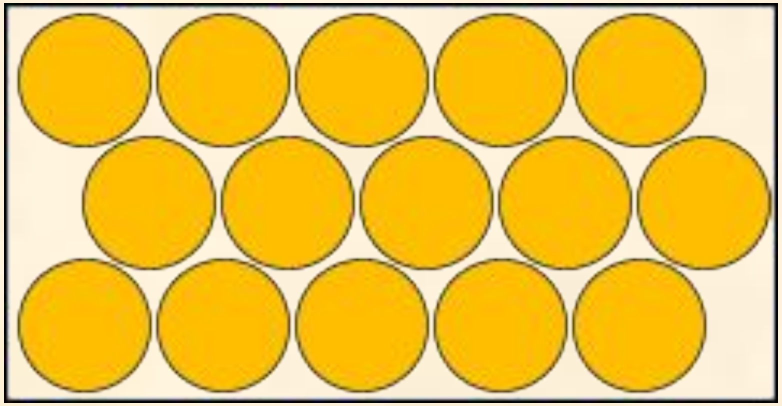
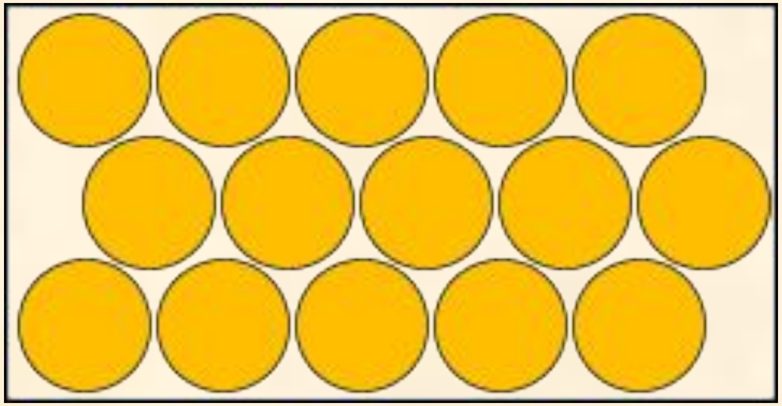


Atoms arrange in regular close packed 3-dimensional crystal lattices.

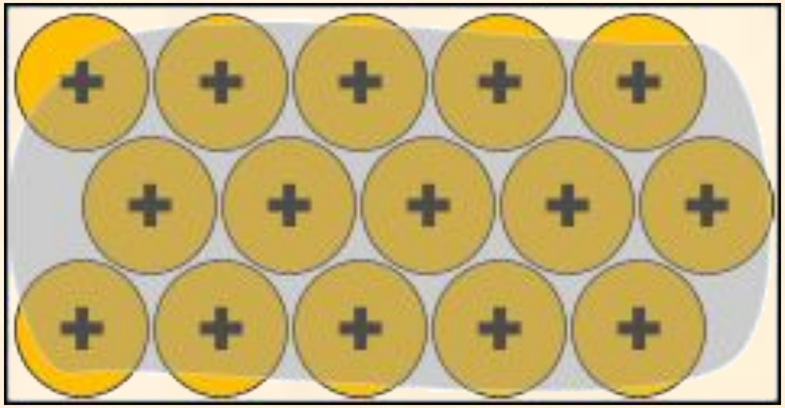
METALLIC BONDING

Involves a lattice of positive ions surrounded by delocalised electrons

Metal atoms achieve stability by "off-loading" electrons to attain the electronic structure of the nearest noble gas. These electrons join up to form a mobile cloud which prevents the newly-formed positive ions from flying apart due to repulsion between similar charges.



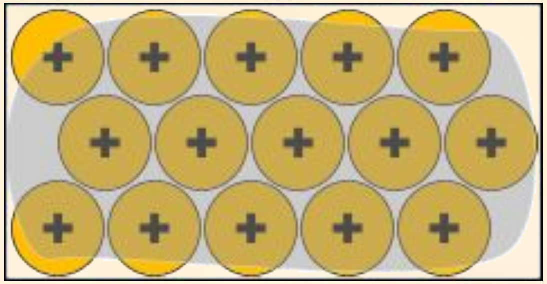
Atoms arrange in regular close packed 3-dimensional crystal lattices.



The outer shell electrons of each atom leave to join a mobile "cloud" or "sea" of electrons which can roam throughout the metal. The electron cloud binds the newly-formed positive ions together.

METALLIC BOND STRENGTH

Depends on the number of outer electrons donated to the cloud and the size of the metal atom/ion.



Na

The strength of the metallic bonding in sodium is relatively weak because each atom donates one electron to the cloud.

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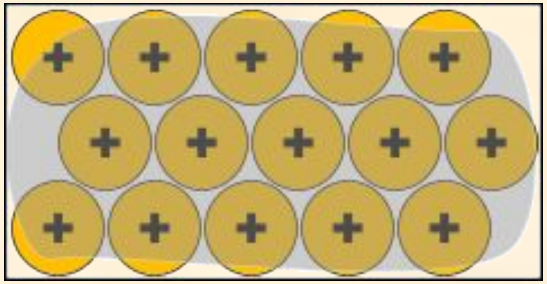

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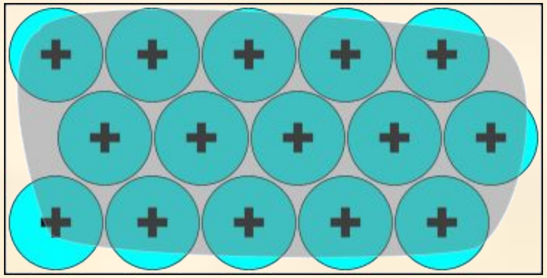

K

The metallic bonding in potassium is weaker than in sodium because the resulting ion is larger and the electron cloud has a bigger volume to cover so is less effective at holding the ions together.

METALLIC BOND STRENGTH

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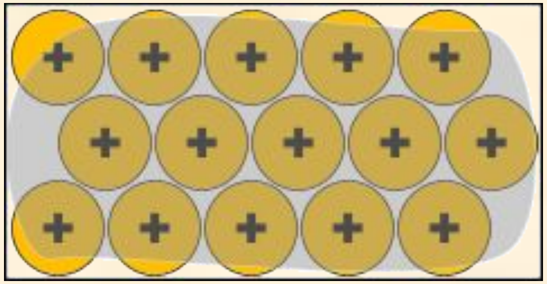

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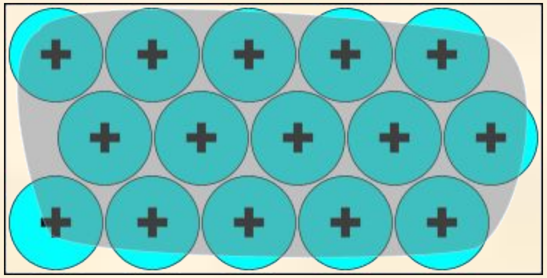

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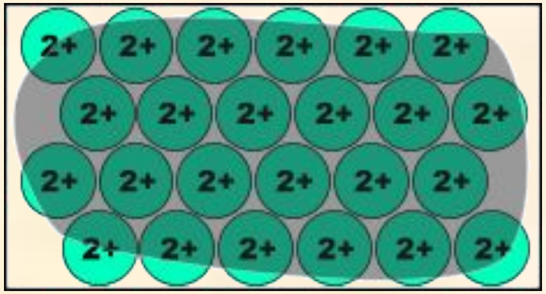


Mg

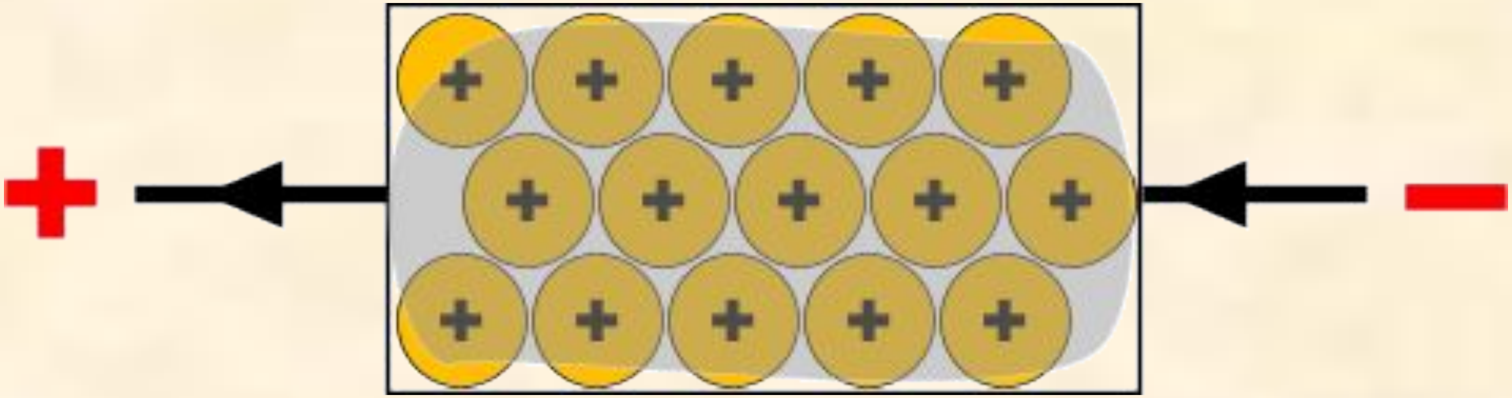
The metallic bonding in magnesium is stronger than in sodium because each atom has donated two electrons to the cloud. The greater the electron density holds the ions together more strongly.

METALLIC PROPERTIES

Metals are excellent conductors of electricity

For a substance to conduct electricity it must have mobile ions or electrons.

Because the ELECTRON CLOUD IS MOBILE, electrons are free to move throughout its structure. Electrons attracted to the positive end are replaced by those entering from the negative end.



MOBILE ELECTRON CLOUD ALLOWS THE CONDUCTION OF ELECTRICITY

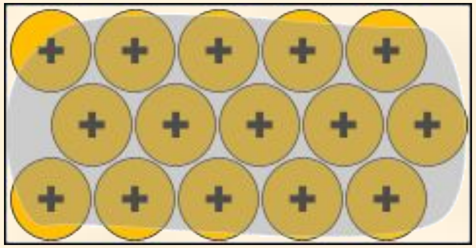
METALLIC PROPERTIES

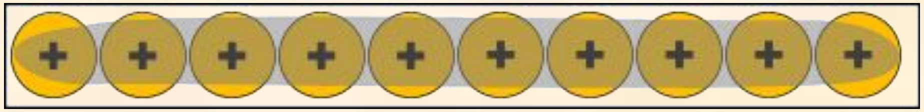
Metals can have their shapes changed relatively easily

MALLEABLE CAN BE HAMMERED INTO SHEETS

DUCTILE CAN BE DRAWN INTO RODS AND WIRES

As the metal is beaten into another shape the delocalised electron cloud continues to bind the "ions" together.





Some metals, such as gold, can be hammered into sheets thin enough to be translucent.

METALLIC PROPERTIES

HIGH MELTING POINTS

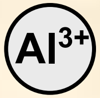
Melting point is a measure of how easy it is to separate individual particles. In metals it is a measure of how strong the electron cloud holds the + ions.

The ease of separation of ions depends on the...

ELECTRON DENSITY OF THE CLOUD

IONIC / ATOMIC SIZE

PERIODS	Na (2,8,1)	<	Mg (2,8,2)	<	Al (2,8,3)
m.pt	98°C		650°C		659°C
b.pt	890°C		1110°C		2470°C



MELTING POINT INCREASES ACROSS THE PERIOD

THE ELECTRON CLOUD DENSITY INCREASES DUE TO THE GREATER NUMBER OF ELECTRONS DONATED PER ATOM. AS A RESULT THE IONS ARE HELD MORE STRONGLY.

METALLIC PROPERTIES

HIGH MELTING POINTS

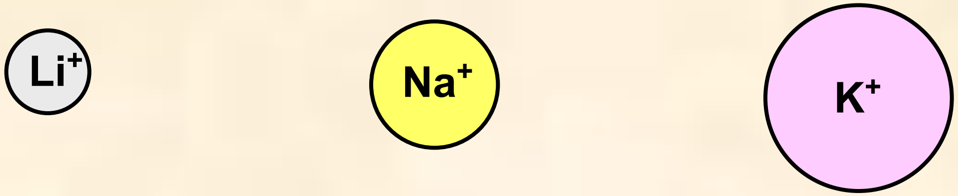
Melting point is a measure of how easy it is to separate individual particles. In metals it is a measure of how strong the electron cloud holds the + ions.

The ease of separation of ions depends on the...

ELECTRON DENSITY OF THE CLOUD

IONIC / ATOMIC SIZE

GROUPS	**Li (2,1)**	**<**	**Na (2,8,1)**	**<**	**K (2,8,8,1)**
m.pt	**181°C**		**98°C**		**63°C**
b.pt	**1313°C**		**890°C**		**774°C**



MELTING POINT INCREASES DOWN A GROUP

IONIC RADIUS INCREASES DOWN THE GROUP. AS THE IONS GET BIGGER THE ELECTRON CLOUD BECOMES LESS EFFECTIVE HOLDING THEM TOGETHER SO THEY ARE EASIER TO SEPARATE.

REVISION CHECK

What should you be able to do?

Recall the different types of physical and chemical bonding

Understand how ionic, covalent, dative covalent and metallic bonding arise

Recall the different forms of covalent structures

Understand how the physical properties depend on structure and bonding

Understand how different types of physical bond have different strengths

Recall and explain the variation in the boiling points of hydrides

Balance ionic equations

Construct diagrams to represent covalent bonding

CAN YOU DO ALL OF THESE?

YES **NO**

You need to go over the relevant topic(s) again

Click on the button to return to the menu

WELL DONE!

Try some past paper questions

AN INTRODUCTION TO
BONDING

THE END