With 2017 NCEA Exam included

2018 Version

Chemistry AS 91164

C2.4 Bonding, Structure and Energy

Achievement Criteria

AS 91164
C2.4

***Bonding and structure* are limited to:**

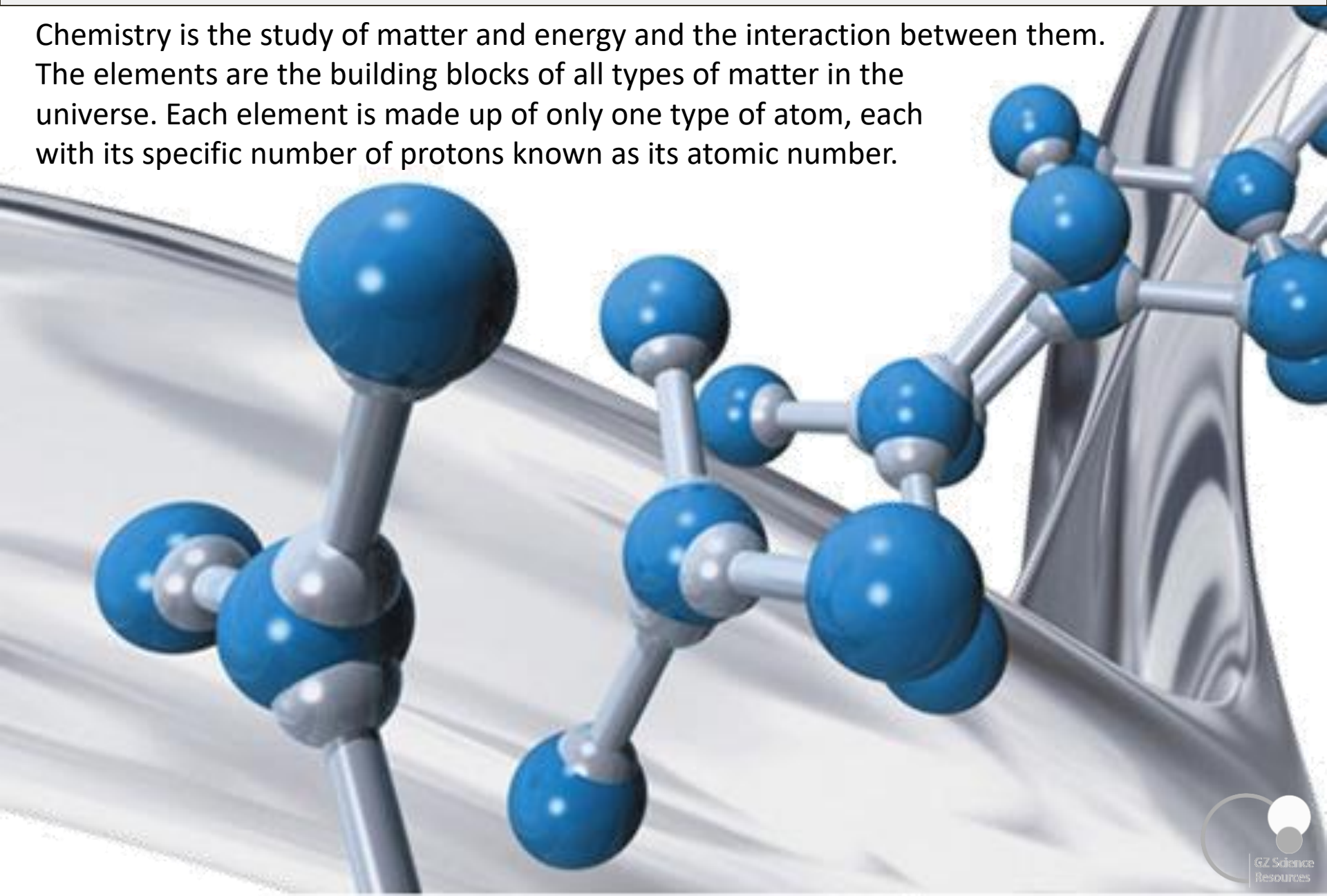
- Lewis structures, shape and polarity of simple molecules.
- Intermolecular forces (the distinction between the different types of intermolecular forces is not required)
- Ionic, covalent and metallic bonding
- Molecular, ionic, metallic and covalent network substances
- Properties are limited to hardness (including malleability and ductility), electrical conductivity, melting and boiling points and solubility.

***Energy changes* are limited to:**

- exothermic and endothermic reactions including energy (enthalpy) changes associated with differing amounts of substances and changes of state and enthalpy changes associated with the making and breaking of chemical bonds
- calculations of energy changes using $\Delta_r H$ and reaction stoichiometry, and bond enthalpy

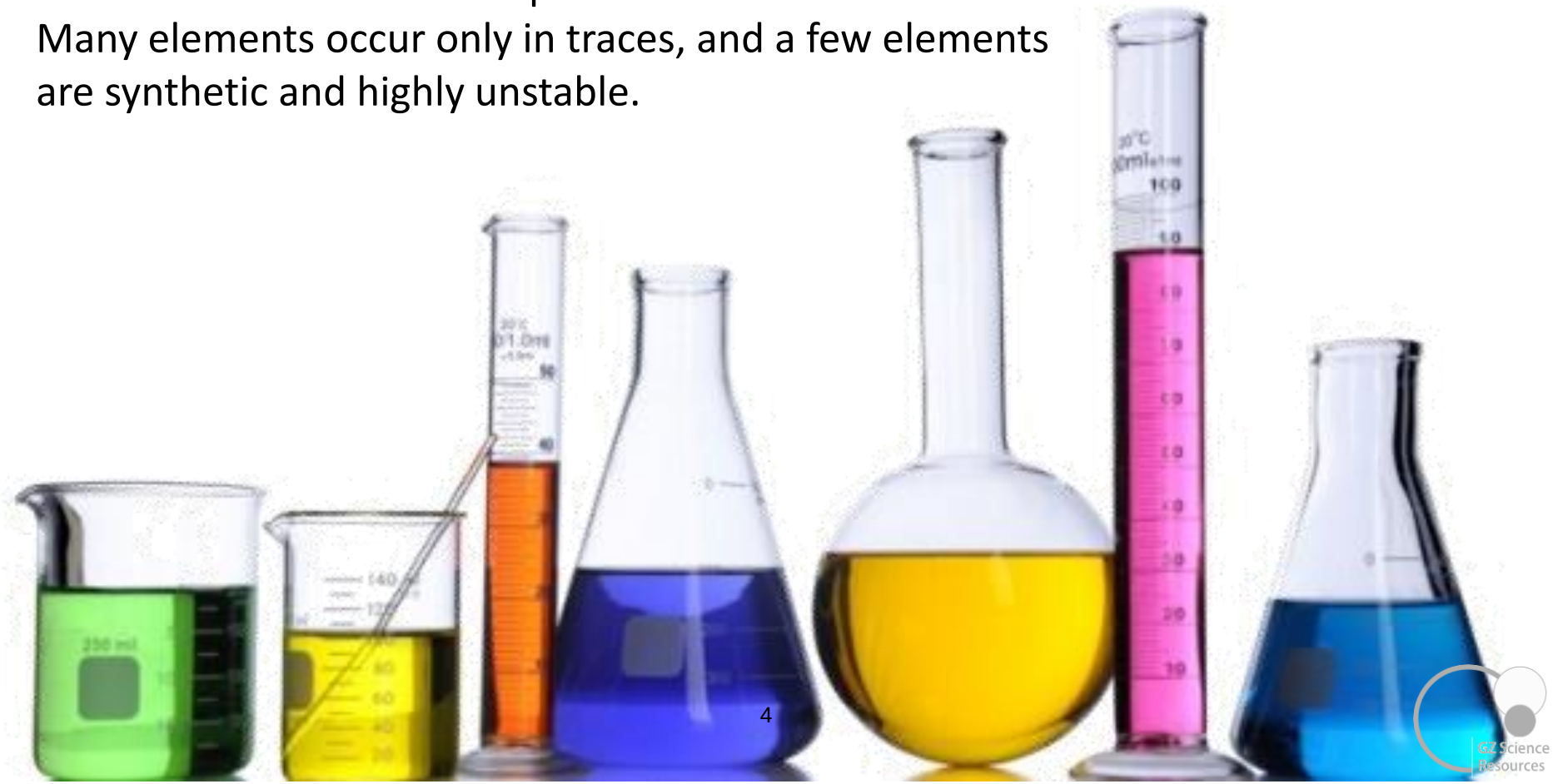
Introduction

Chemistry is the study of matter and energy and the interaction between them. The elements are the building blocks of all types of matter in the universe. Each element is made up of only one type of atom, each with its specific number of protons known as its atomic number.

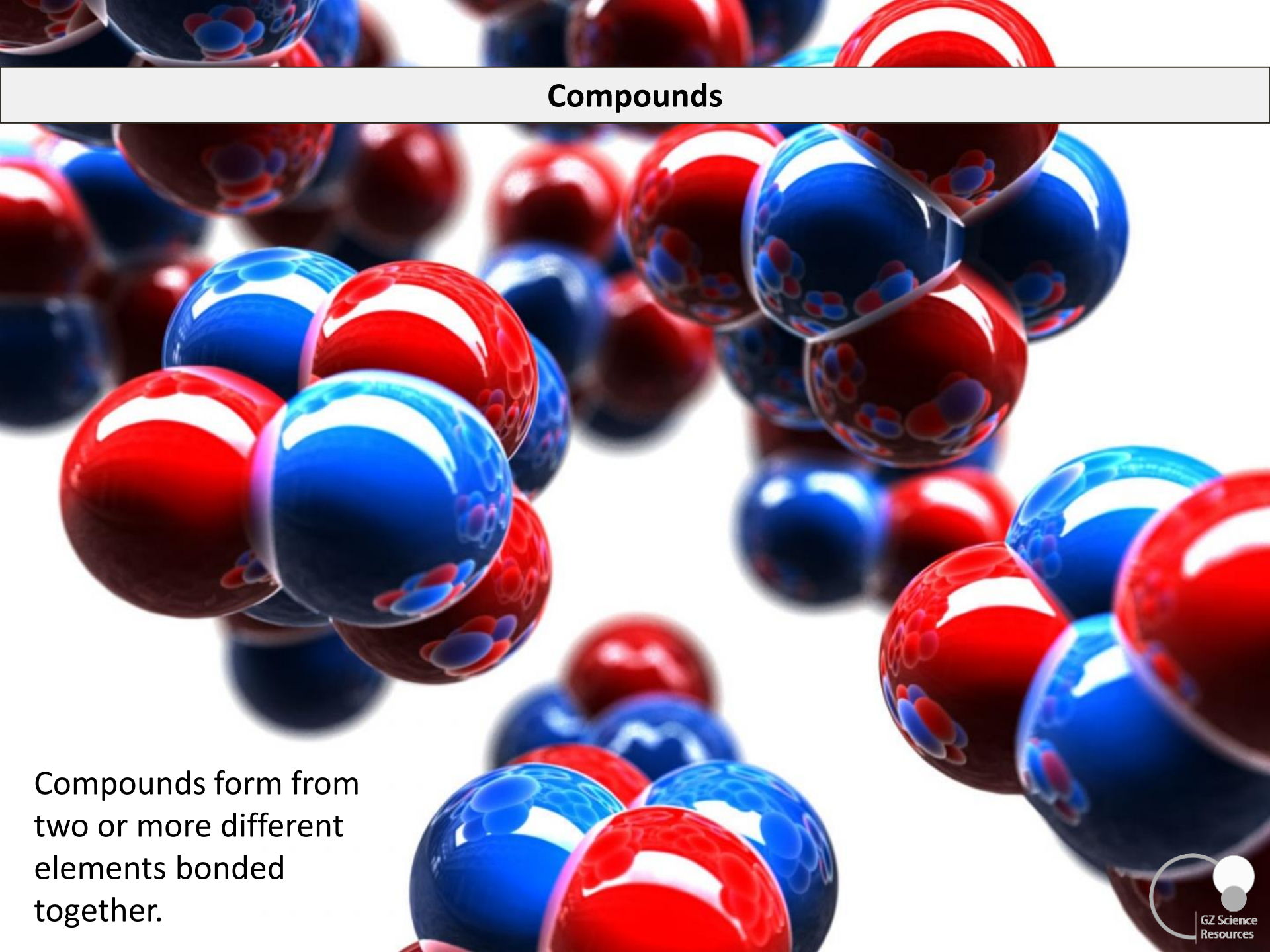


Introduction

A large amount of energy is required to break an atom down into smaller particles. The elements occur in widely varying quantities on earth. The ten most abundant elements make up 98% of the mass of earth. Many elements occur only in traces, and a few elements are synthetic and highly unstable.



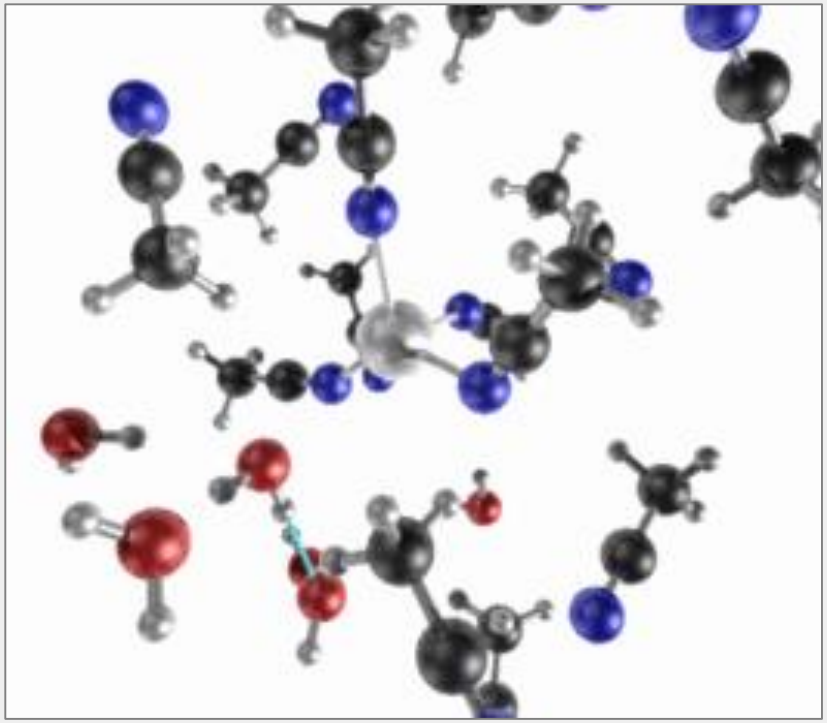
Compounds



Compounds form from two or more different elements bonded together.

Background Knowledge

Compounds



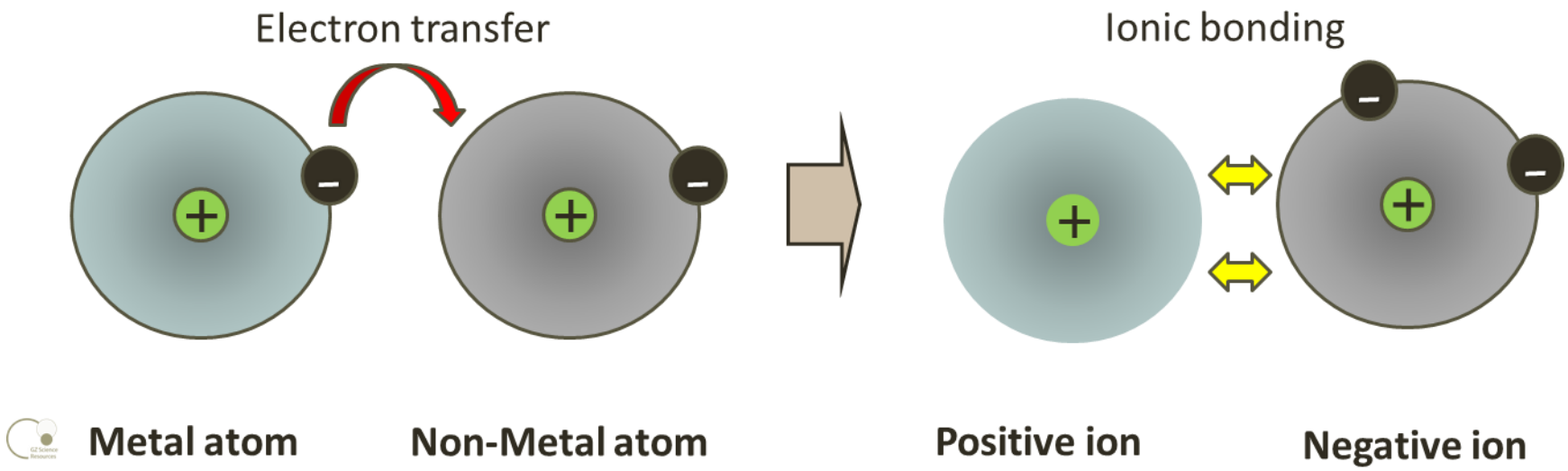
The compounds are often more stable than the elements they originated from and may release this extra energy in the form of heat and/or light when bonding together.

There are two main types of bonding holding atoms together in a compound; Ionic and Covalent.

Ionic Bonding

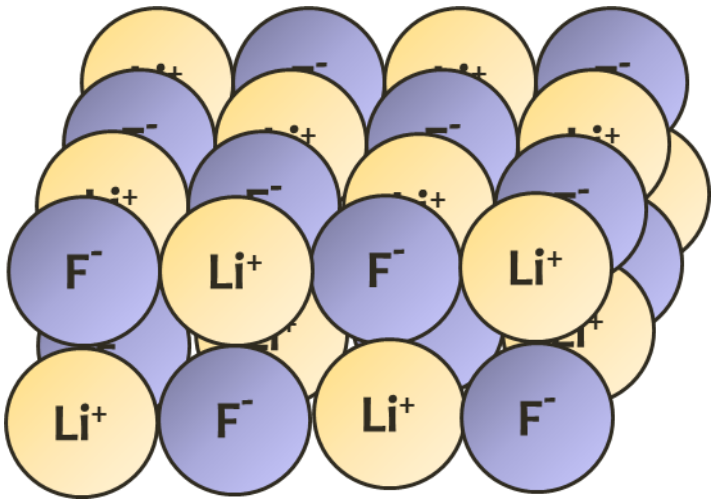
Ionic Bonding is where one atom completely takes valence electrons from another to form ions and the resulting negative and positive ions are held together with electrostatic attraction. This type of bonding occurs when a **metal** and **non-metal** react and there is a **transfer of electrons** to form ions.

The ions then combine in a set ratio to form a neutral compound with negative and positive charges balanced out.



Ionic compounds are the product of chemical reactions between metal and non-metal ions

Some compounds are ionic compounds, since they are made up of cations and anions.



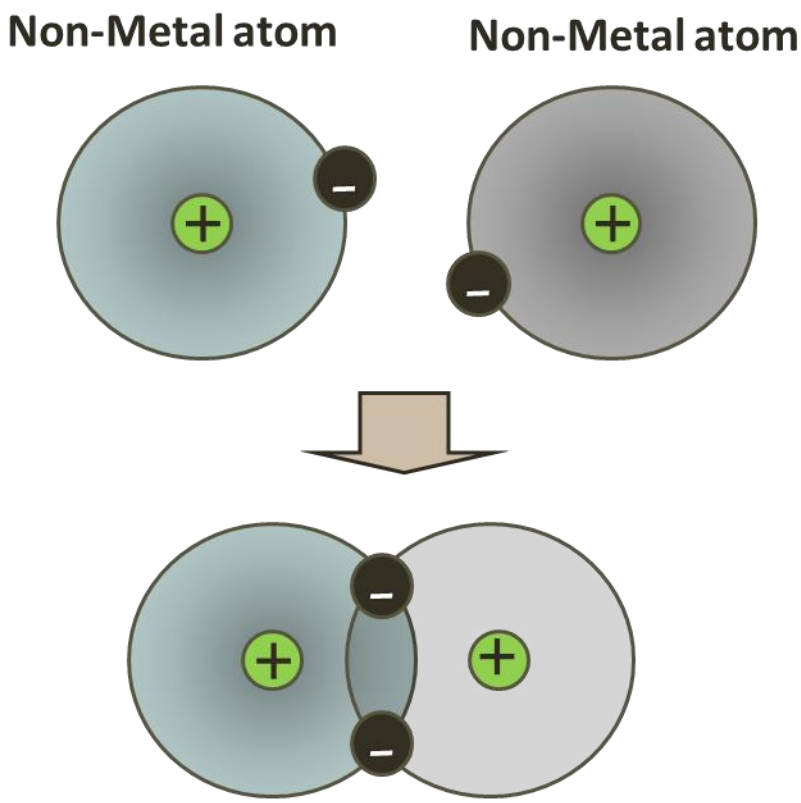
The Anion (F^-) takes an electron from the Cation (Li^+) so their outer energy levels have a stable 8 electrons each.

Anions and Cations have a strong <u>electrostatic attraction</u> for each other so they bond together as a compound.

Compounds are neutral substances. For ionic compounds, the charges of the positive ions are balanced by the charges of the negative ions.

Covalent Bonding

Covalent Bonding occurs when electrons are 'shared' between neighbouring atoms. This often occurs when two or more **non-metals** react. No ions are formed and there is no transfer of electrons. The compound formed is neutral with no charge.



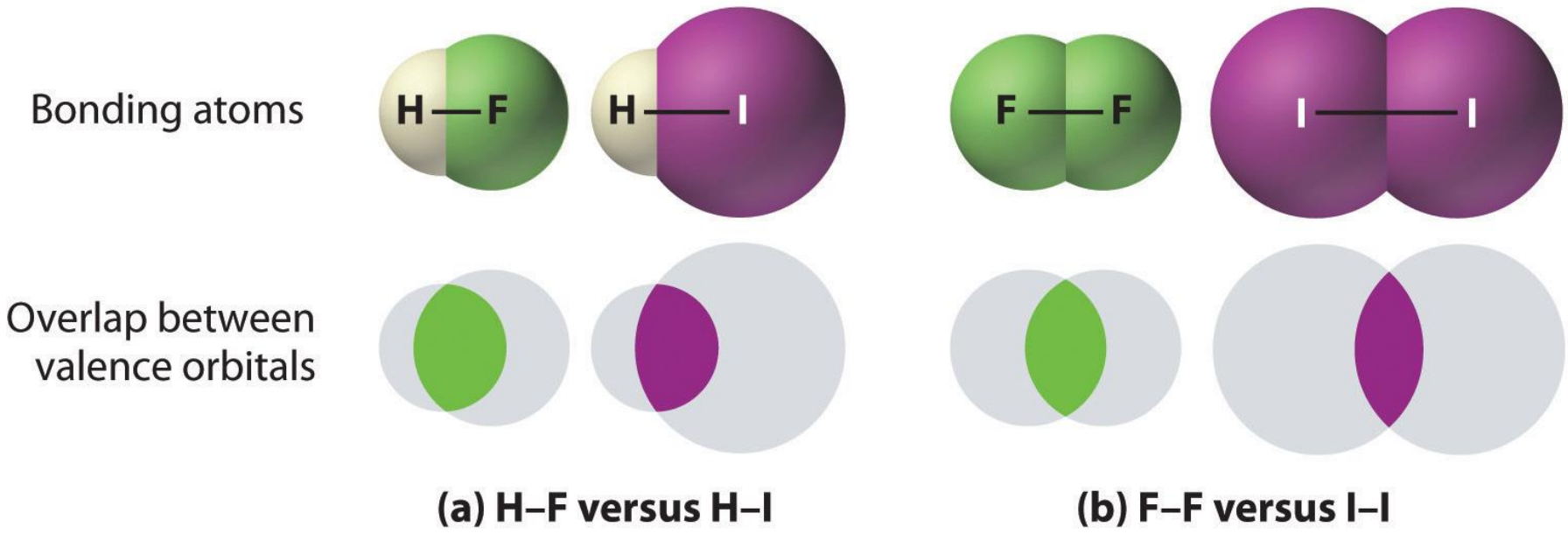
covalent bonding

The valance electrons (electrons in outside energy level) are involved in bonding. These electrons orbit in pairs. The negative charge of the electron pair will attract the positively positive nucleus of other atoms, and this holds the atoms together in a molecule.

Covalent Bonding

extra info

All covalent bonds are strong. That is they require a large amount of energy to 'break' the bond. However, some covalent bonds are stronger than others. The greater the overlap of valence orbitals (the area the valence electrons orbit the nucleus) the stronger the bond.

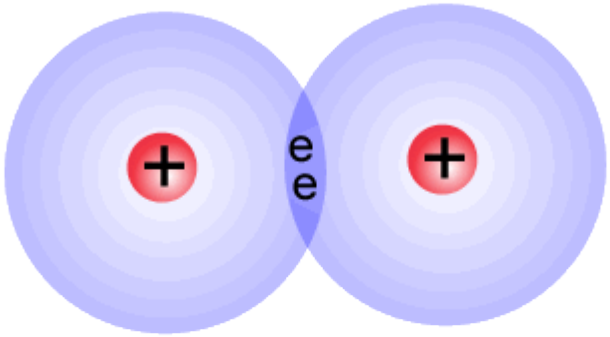


Covalent Bonding

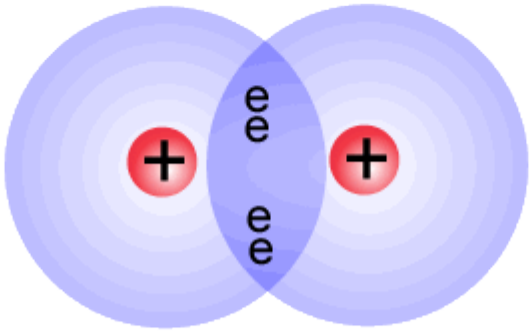
The electron-pair must lie between the nuclei for the attraction to outweigh the repulsion of the two nuclei. This ‘sharing’ of electrons between atoms creates a covalent bond – giving both atoms the stability of a full outer shell. Covalent bonds are normally formed between pairs of non-metallic atoms.

Some covalent bonds involve only one pair of electrons and are known as single bonds. Other covalent bonds can involve two pairs of electrons; double bonds and three pairs of electrons; triple bonds.

Only one pair of electrons holding the nuclei together



Two pair of electrons hold the nuclei tighter and closer



Background Knowledge

Naming compounds

Lavoisier – French Chemist 1789

Lavoisier devised a system of naming compounds based on their chemical composition. If the compound is formed between a Metal **cation** (+ve) and a Non-Metal **anion** (-ve), then the compound name joins the two names together with the metal name first. Names of the ions need to be remembered.

Sodium

+

hydroxide

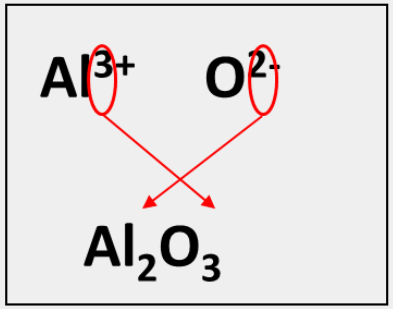
→

Sodium hydroxide

Background Knowledge

Chemical compound formula

1. Write down the ions (with charges) that react to form the compound. Cation comes before Anion.



2. Cross and drop the charge numbers.
3. Place brackets around a compound ion.
4. If the numbers are both the same remove.
5. If any of the numbers are a 1 they are removed
6. Remove any brackets if not followed by a number

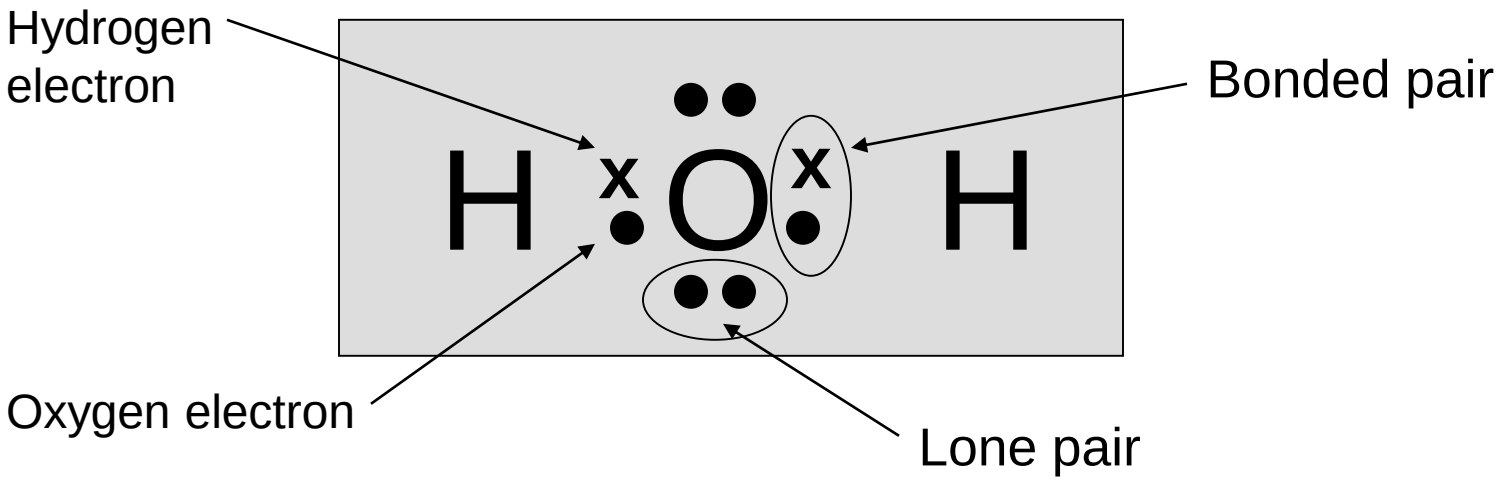
GZ Science Resources

Drawing molecular compounds – Lewis Structures

G Lewis – American Chemist 1916

G Lewis devised a system of drawing covalent molecules showing arrangement of atoms and valence electrons – both those involved in bonding and those that are not (called lone pairs). Electrons in inner shells are not involved in bonding. These diagrams are called **Lewis structures**. (or diagrams) The Lewis structure is drawn so that each atom has eight electrons associated with it (except for hydrogen, which has two). This is the *octet rule*.

Lewis structure of H_2O (water)



Drawing Lewis Structures

1. Calculate valence electrons of all atoms. If the molecule is an ion, then subtract the charge from the total electrons and place the charge outside of square brackets of the Lewis structure at the end. *Example carbon dioxide.*

C = 4

O = 6

O = 6

16

CO_2

2. Write down number of pairs of electrons.

16 / 2 = 8 pairs

3. Place atom with least filled valence shell in the centre with the other atoms arranged around the outside (periphery)

O C O

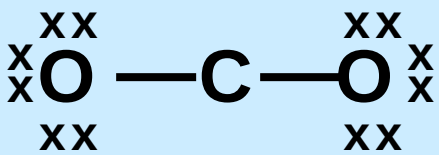
Lewis Structures

4. Bond all atoms together (either ● or **x** or — = one pair of electrons)

O — C — O

8 pairs – 2 pairs = 6 pairs remaining

5. Place remaining e- pairs around the periphery atoms so each has 4 pairs (including bond pair) around it.



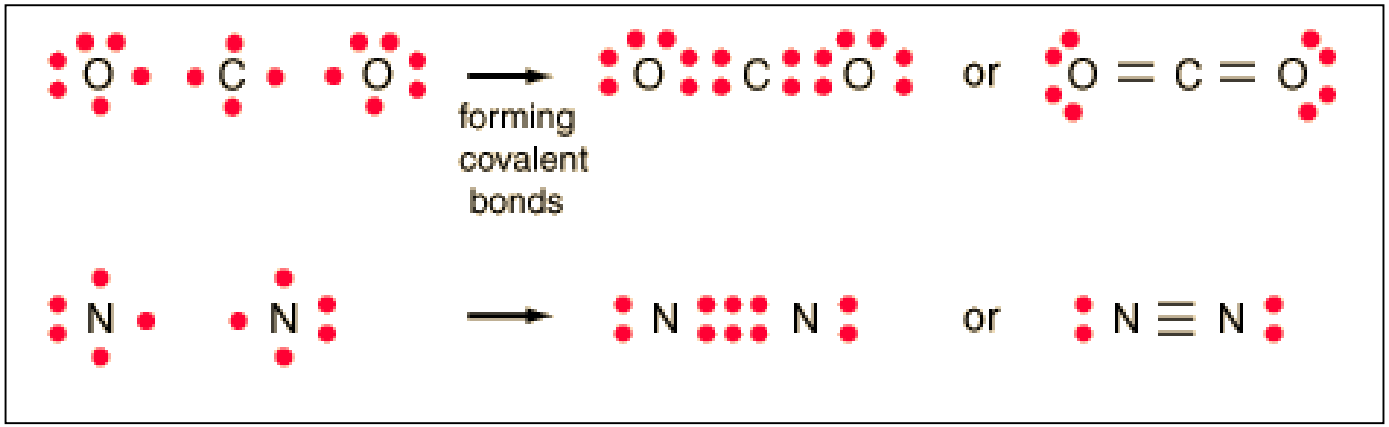
6 pairs – 6 pairs = 0 pairs remaining

6. If there any remaining pairs place them around the outside of the central atom.

7. Rearrange lone pairs (pairs not bonded) into bonded pairs if the central atom does not have 4 pairs around it.

Lewis Structures and Diagrams

- The number of covalent bonds an atom forms is called its *valence*.
- Some atoms have *fixed valence*. E.g.: H = 1, C = 4, F = 1. (most halogens = 1)
- Some atoms have *variable valence*. For example:
  O = 2 (sometimes 3), B, N = 3 (sometimes 4).
- an atom bonded to only one other atom is *peripheral* (monovalent atoms such as H and F are always peripheral).
- an atom bonded to two or more other atoms is *central*.
- Often, the formula is written to indicate connectivity. For example: HCN = H bonded to C, C bonded to N, H and N are not bonded.



Rule of orbitals – exceptions to the rule

If there are extra Lone Pairs of electrons left after all of the periphery atoms are filled in accordance with the *octet rule* then they are placed around the central atom(s) according to the **Rule of Orbitals.**

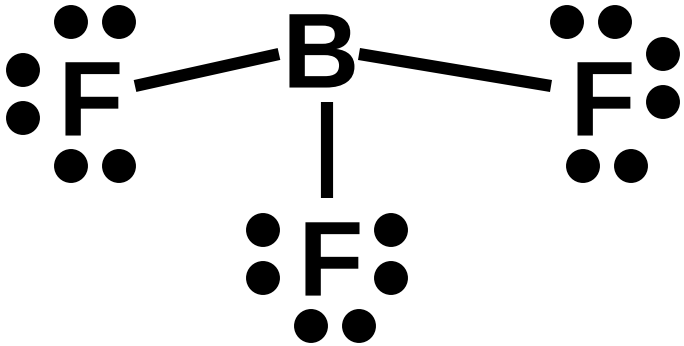
The Rule of Orbitals: the total number of lone pairs and bond pairs (LP+BP) associated with an atom cannot exceed the number of Valence Shell Orbitals (VSO = n^2, where n is the row of the Periodic Table in which that atom resides).

n = 1 (H): <u>maximum</u> VSE pairs (LP+BP) = VSO = 1;

n = 2 (B, C, N, O, F): <u>maximum</u> VSE pairs (LP+BP) = VSO = 4 ("octet rule")

n = 3 ((Al, Si, P, S, Cl): <u>maximum</u> VSE pairs (LP+BP) = VSO = 9 etc.

Boron and Beryllium often are found with only **3** (or 2) lone + bonded pairs of electrons around them

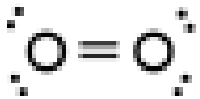
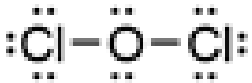
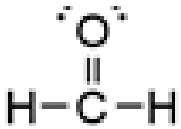


NCEA 2015 Lewis Structures

Achieved Question

Q 1a: Draw the Lewis Structure for each of the following molecules

Draw the Lewis structure (electron dot diagram) for each of the following molecules.

Molecule	O_2	OCl_2	CH_2O
Lewis structure			

Look further through the exam paper for hints on drawing

In the past the Lewis structure is a "stand alone" question – which case there is an Achieved point given for the majority correct.

In past years the following molecules have been used:

HCN CH_2Br_2 AsH_3 CH_4 H_2O N_2 PCl_3 CO_2 H_2S

These have one central atom (except N_2) and 2 or more outside attached atoms

Valence-shell electron-pair repulsion (VSEPR) theory

Sidgwick and Powell – 1940

Sidgewick and Powell devised a theory to predict the shapes molecule formed. It is based on the following ideas:

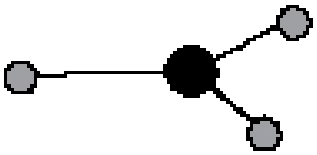
Electron pairs form regions of negative charge

Negative charges repel each other

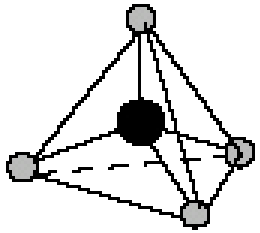
Regions of negative charge will be spaced as far apart as possible around a central atom.

This theory is called **valence-shell electron-pair repulsion (VSEPR) theory**

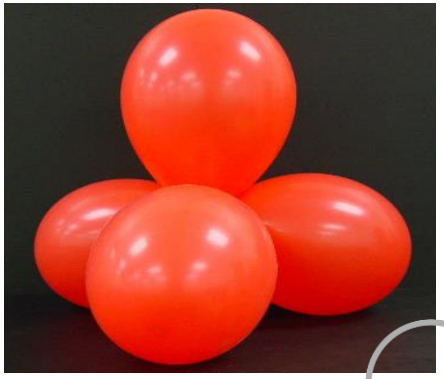
The shapes of molecules are determined by the way the regions of negative charge are arranged around the central atom in the molecule. A region may consist of one lone pair of electrons <u>or</u> one bonded pair <u>or</u> two bonded pairs <u>or</u> three bonded pairs. All of these electron arrangements occupy the same region of space



Three pairs



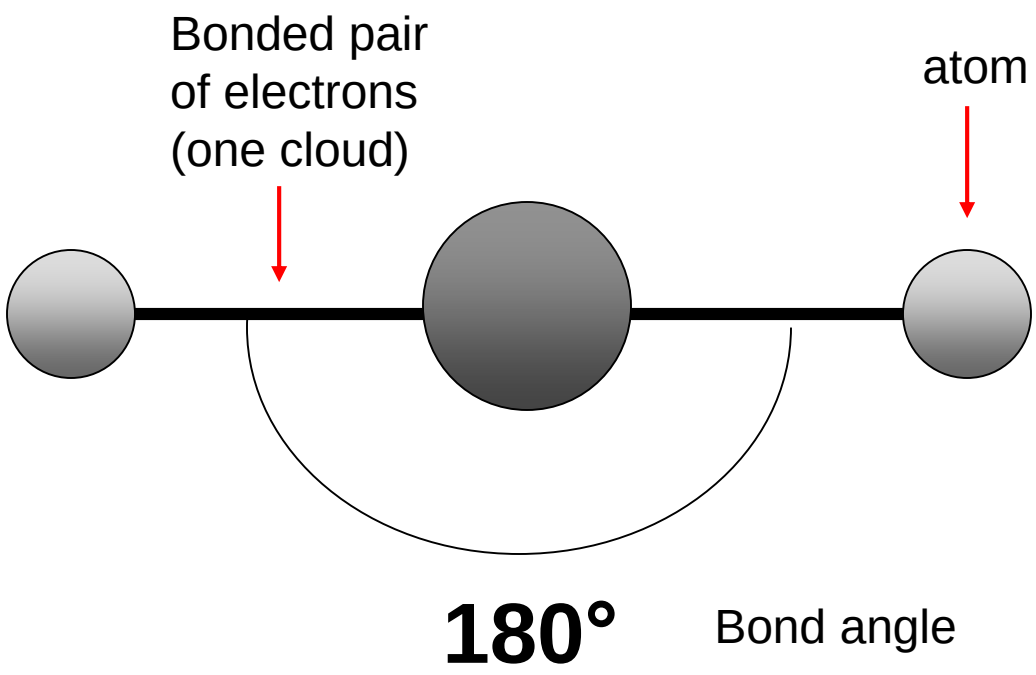
Four pairs



Molecular Shapes – two regions of negative charge

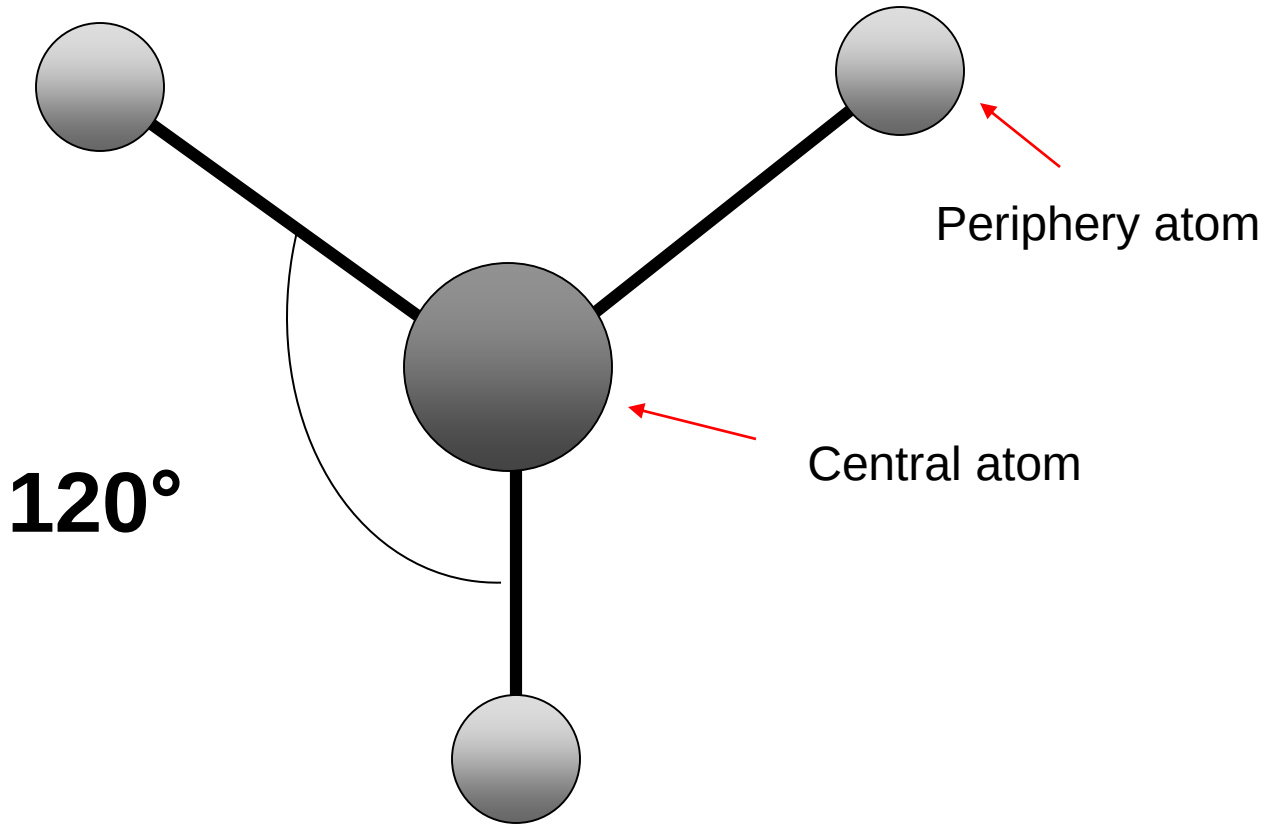
Since regions of electrons are negatively charged, they repel each other as far apart as possible (**VSEPR)**. Two negative regions arrange themselves on opposite sides of the central atom.
The bond angle will be 180°.

The shape name is **linear.**



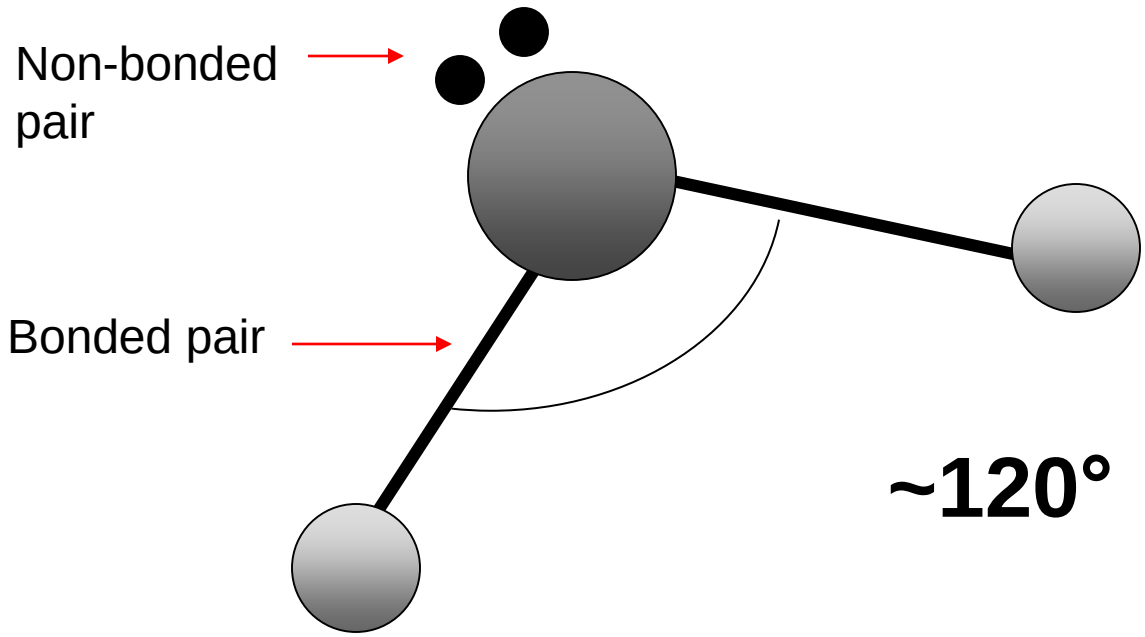
Molecular Shapes – three two regions of negative charge (0 non-bonded pairs)

Three regions of negative charge will cause a bond angle of 120° as they repel each other.
All the atoms still lie on a flat plane (like a sheet of paper).
The shape is **trigonal planar**. (or triangular planar)



Molecular Shapes – two two regions of negative charge (1 non-bonded pair)

The regions of negative charge repel to a trigonal planar shape. The bond angle between the remaining pairs is approximately 120° .
The final shape formed by the atoms is called **bent**.

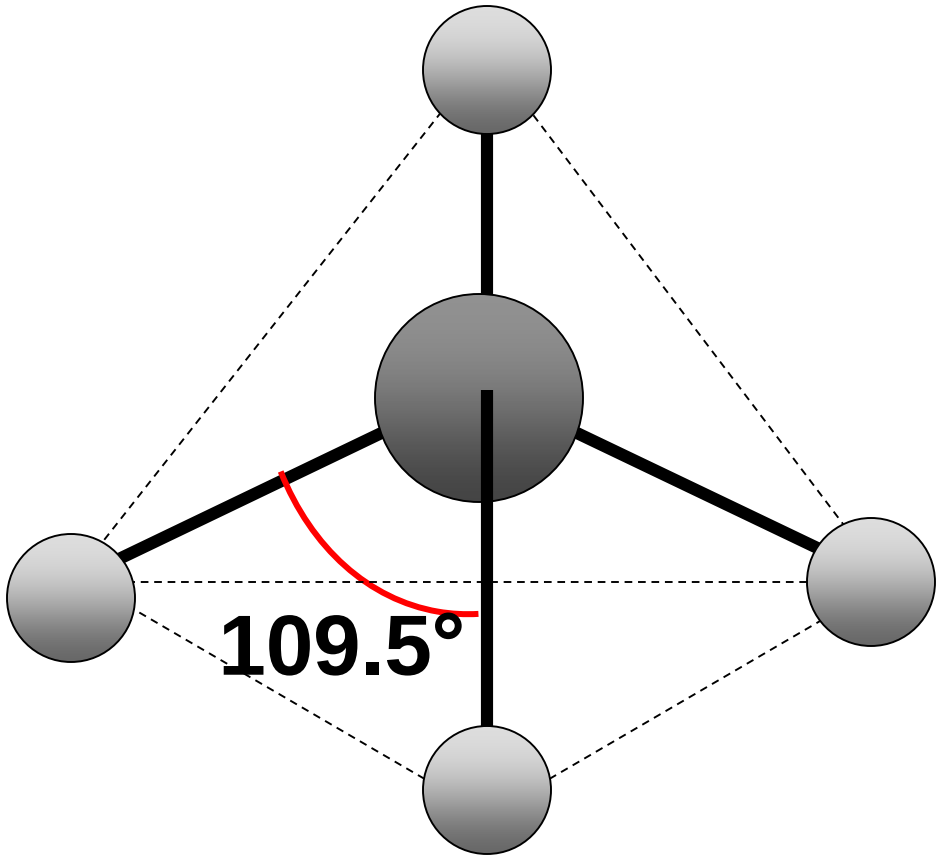


extra info

When one of the regions of electrons is a *non-bonding pair* it will have a slightly greater push to the bonded pairs. This is because the non-bonding pair is only orbiting around one positive nucleus and their negative charge is less 'neutralised' than if they had another nucleus to orbit around.

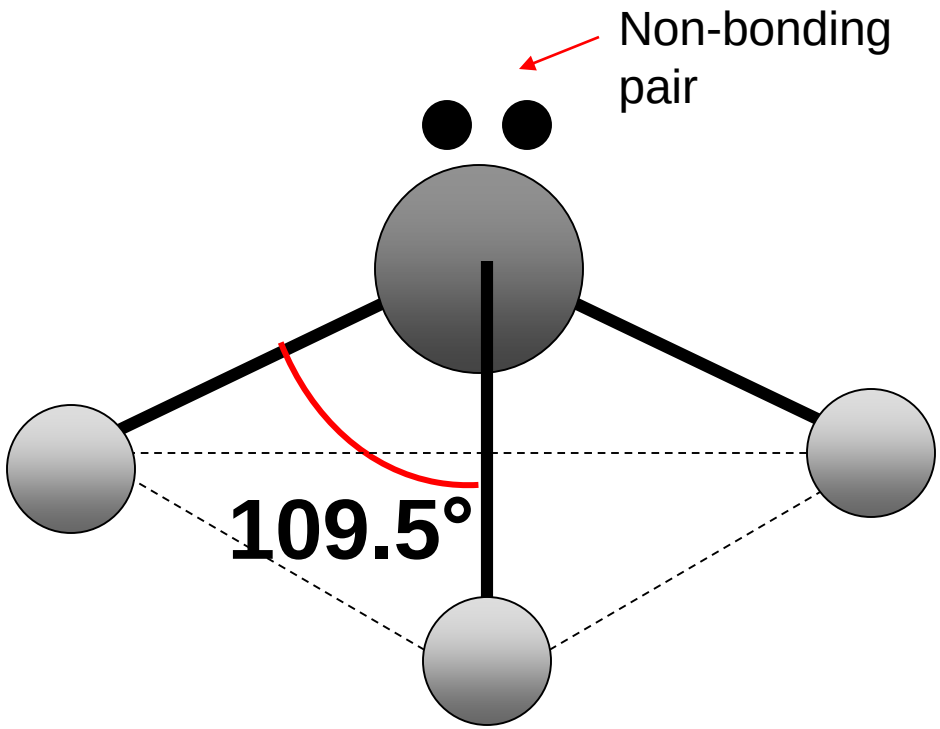
Molecular Shapes – four two regions of negative charge (0 non-bonding pairs)

When four regions of negative charge are around a central atom, they repel each other into a 3-dimensional shape. The bond angle is now 109.5°. This is because it is a 3-dimensional sphere divided into 4 rather than a circle.
This shape is **tetrahedral.**



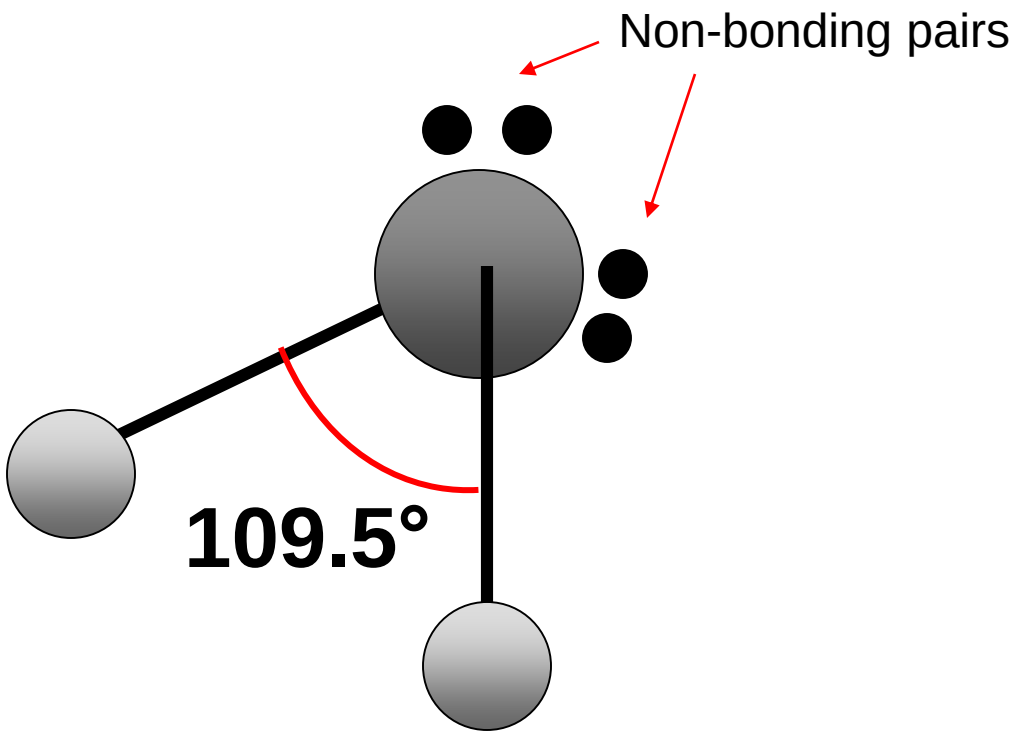
Molecular Shapes – four two regions of negative charge (1 non-bonding pair)

The four regions of negative charge still occupy a 3-dimensional tetrahedral shape.
(The lone pair, however, exerts a stronger repulsion to the remaining bonded pairs).
The bond angle is approximately 109.5°.
The final shape the bonded atoms form is a **trigonal pyramid** (or a triangular pyramid)



Molecular Shapes – four two negative regions of charge (2 non-bonding pairs)

The 4 regions of negative charge repel each other to a (warped) tetrahedral shape. However, the two non-bonding pairs create a much stronger repulsion than one non-bonding pair and the bond angle between the remaining bonded pairs is smaller again at approximately 109.5° (compared to 120 ° of the bent shape with only 3 regions of negative charge). The final shape the bonded atoms form is called **Bent**.



Determining Molecular Shapes

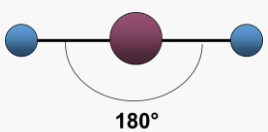
Electron regions of negative charge (bonding or non-bonding pairs)

GZ Science Resources

2 regions

No non-bonding pair

Linear

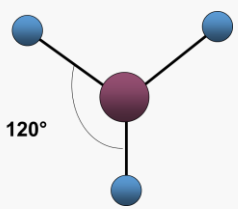


CO_2 — 180°

3 regions

No non-bonding pair

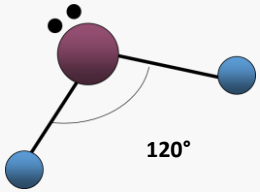
Trigonal Planar



BF_3 — 120°

1 non-bonding pair

Bent

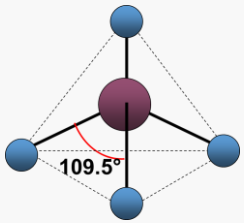


SO_2 — 120°

4 regions

No non-bonding pair

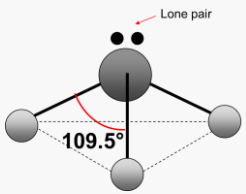
Tetrahedral



CH_4 — 109.5°

1 non-bonding pair

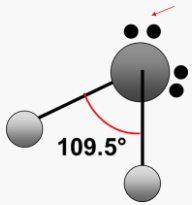
Trigonal Pyramid



NH_3 — 109.5°

2 non-bonding pairs

Bent

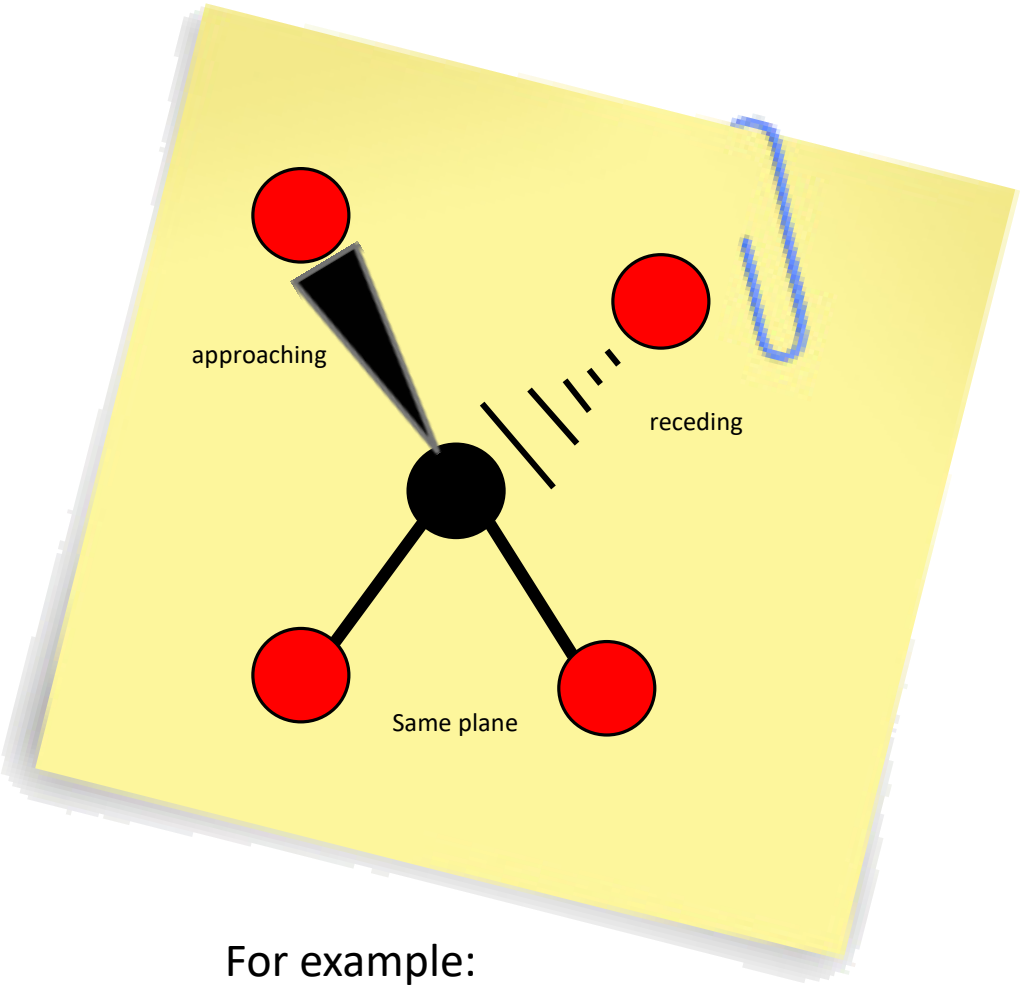


H_2O — 109.5°

Note: make sure the question asks you to draw a shape and **not** a Lewis structure

Drawing Molecular Shapes

1. Atom on **same plane** as central atom – straight solid line
2. Atom **receding** from central atom – lines starting large and getting smaller
3. Atom **approaching** from central atom – solid triangle starting small and getting larger



For example:
Tetrahedral shape

Discussing shapes questions – NCEA example

Explain why the shape of the CO_2 molecule is linear but the shape of H_2O is bent?

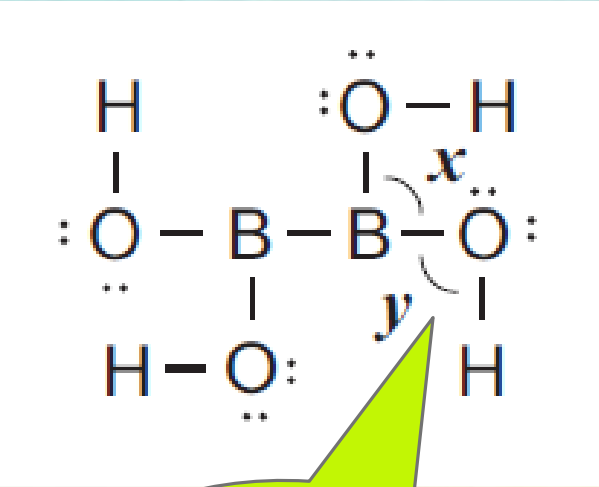
1. The C (central atom) of CO_2 has 2 regions of negative charge around it in the form of double bonds connected to an O atom. (draw Lewis structure)
2. Each of the regions of negative charge repel each other the furthest away from each other in 3 dimensional space into a linear shape.
3. There are no lone pairs so the final CO_2 molecule therefore also forms a *linear shape*

1. The O molecule (central atom) of H_2O has 4 regions of negative charge around it in the form of two single bonds connected to a H atom and two lone pairs. (draw Lewis structure)
2. Each of the regions of negative charge repel each other the furthest away from each other in 3 dimensional space and form a tetrahedral shape.
3. However with only 2 of the regions bonded to atoms the final shape the H_2O molecule forms is a *bent shape*

NCEA 2014 Molecule shapes and bond angle

Achieved Question

Question 1b (i) : The Lewis structure for a molecule containing atoms of boron, oxygen, and hydrogen, is shown below.
The following table describes the **shapes around two of the atoms** in the molecule above.
Complete the table with the approximate bond angles ***x*** and ***y***.



Central atom	Shape formed by bonds around the central atom	Approximate bond angle
B	trigonal planar	$x =$ 120°
O	bent	$y =$ 109.5°

Even though this is an unknown molecule focus on each atom and the number of regions of charge around each

Bond angles

2 regions = 180°

3 regions = 120°

4 regions = 109.5°

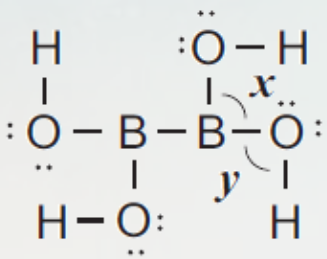
NCEA 2014 Molecule shapes and bond angle

Excellence Question

Question 1b (ii): The bond angles ***x*** and ***y*** in the molecule above are different. Elaborate on why the bond angles are different.

In your answer you should include:

- factors which determine the shape around the:
- **B** atom for bond angle ***x***
- **O** atom for bond angle ***y***
- reference to the arrangement of electrons around the **B** and **O** atoms.



Answer 1b (ii): The B atom has **three regions of electron density** around it. These are all bonding regions. The regions of electron density are arranged to minimise repulsion / are arranged as **far apart as possible** from each other in a **trigonal planar** arrangement. This is why the bond angle is 120°

The O atom has **four regions of electron density** around it. The regions of electron density are arranged to minimise repulsion / are arranged **as far apart** as possible from each other in a **tetrahedral** arrangement / two of these are bonding (and two are non-bonding). This is why the bond angle is 109.5°.

NCEA 2015 Molecule shapes and bond angle

Achieved Question

Question 1b: Carbon atoms can bond with different atoms to form many diff. compounds. The following table shows the Lewis structure for two molecules containing carbon as the central atom, CCl_4 and $COCl_2$. These molecules have different bond angles and shapes.

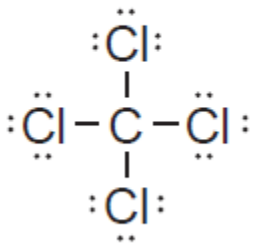
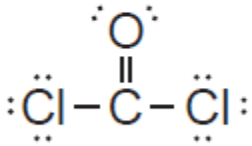
Evaluate the Lewis structure of each molecule to determine why they have different bond angles and shapes.

In your answer you should include:

- The approximate bond angle in each molecule
- The shape of each molecule
- Factors that determine the shape and bond angle for each molecule.

Lewis structures are given in the shape and bond angle questions

Make sure you clearly state the correct shape and bond angle **beside each Lewis structure** - before starting the discussion. Use the formula as the "name" for each molecule – you do not have to know the name unless it is given.

M	CCl_4	$COCl_2$
Lewis structure		

NCEA 2015 Molecule shapes and bond angle

Excellence Question

Answer 1b: In each CCl_4 molecule, there are four negative electron clouds / regions around the central C atom. These repel each other as far away from each other as possible in a tetrahedral (base) arrangement, resulting in a 109.5° bond angle. All of these regions of electrons are bonding, without any non-bonding regions, so the shape of the molecule is tetrahedral.

In each $COCl_2$ molecule, there are three negative electron clouds / regions around the central C atom. These repel as far away from each other as possible in a triangular / trigonal planar (base) shape, resulting in a 120° bond angle. All of these regions of electrons are bonding, without any non-bonding regions, so the shape of the molecule is trigonal planar.

Bond angles

2 regions = 180°

3 regions = 120°

4 regions = 109.5°

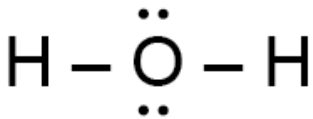
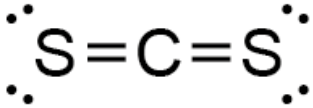
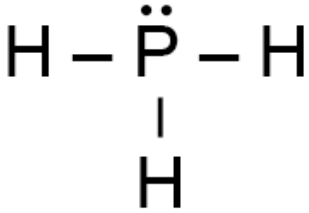
Discuss each molecule separately using the same steps

State number of electron regions around central atom
State base shape they repel to and the angle
State number of bonding/non-bonding regions
State final shape (will be the same as base shape if no lone pairs)

NCEA 2016 Molecule shapes and bond angle (part ONE)

Achieved Question

Question 3a (i) : Draw the Lewis structure (electron dot diagram) for each of the following molecules, and name their shapes.

Molecule	H_2O	CS_2	PH_3
Lewis structure			
Name of shape	bent	linear	Trigonal pyramid
Approximate bond angle around the central atom	109.5°	180°	109.5°

NCEA 2016 Molecule shapes and bond angle (part TWO)

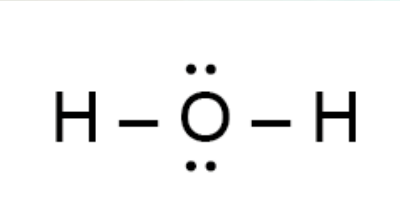
Excellence Question

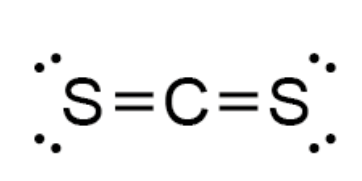
Question 3a (ii): Compare and contrast the shapes and bond angles of H_2O, CS_2 and PH_3.

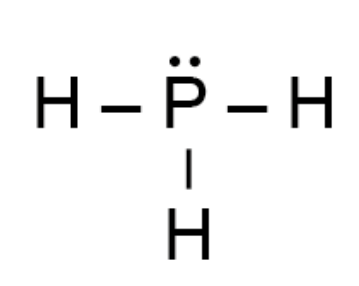
Answer 3a (ii): Bond angle is determined by the number of electron clouds / areas of negative charge around the central atom, which are arranged to minimise repulsion / are arranged as far apart from each other as possible (maximum separation).

Both H_2O and PH_3 have 4 electron clouds / areas of negative charge around the central atom, so the bond angle is that of a tetrahedral arrangement of 109.5°, whereas there are only 2 electron clouds / areas of negative charge around the central atom in CS_2, which means minimum repulsion is at 180°, resulting in CS_2's shape being linear.

The shapes of H_2O and PH_3 differ despite having the same tetrahedral arrangement because water has two non-bonding pairs of electrons around the central atom, while phosphine only has one non-bonding pair. The resulting shapes are bent or v-shaped for H_2O, while PH_3 is trigonal pyramid



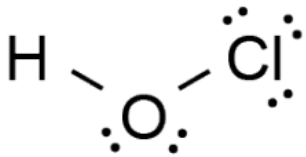
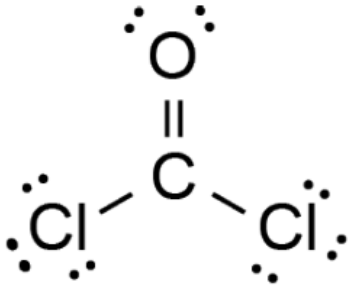
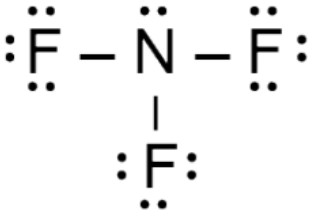




NCEA 2017 Molecule shapes and bond angle (part ONE)

Achieved Question

Question 2a (i) : Draw the Lewis structure (electron dot diagram) for each of the following molecules, and name their shapes.

Molecule	HOCl	$COCl_2$	NF_3
Lewis structure			
Name of shape	bent / v-shaped	trigonal planar	trigonal pyramid
Approximate bond angle around the central atom	109.5°	120°	109.5°

NCEA 2017 Molecule shapes and bond angle (part TWO)

Excellence Question

Question 2a (ii): Justify the shapes and bond angles of HOCl and $COCl_2$.

Answer 2a (ii): Bond angle is determined by the number of electron density regions around the central atom, which are arranged into a position to minimise repulsion / are arranged as far apart from each other as possible (maximum separation).
HOCl has 4 electron density regions / areas of negative charge around the central O atom.
This means the electron density regions around the central atom is arranged with maximum separation in a tetrahedral shape with a bond angle of 109.5°, to minimise (electron-electron) repulsion. Due to the presence of two non-bonding pairs of electrons / regions (or two bonding regions) on the central O atom, HOCl has an actual shape that is bent / v-shaped / angular.
$COCl_2$ has only 3 electron density regions / areas of negative charge around its central C atom so the electron density regions around the central atom is arranged with maximum separation in a trigonal planar shape with a bond angle of 120°, to minimise (electron-electron) repulsion. Since $COCl_2$ has only bonding electron pairs (no non-bonding pairs) on its central atom, the actual shape is trigonal planar (with bond angles of 120°).

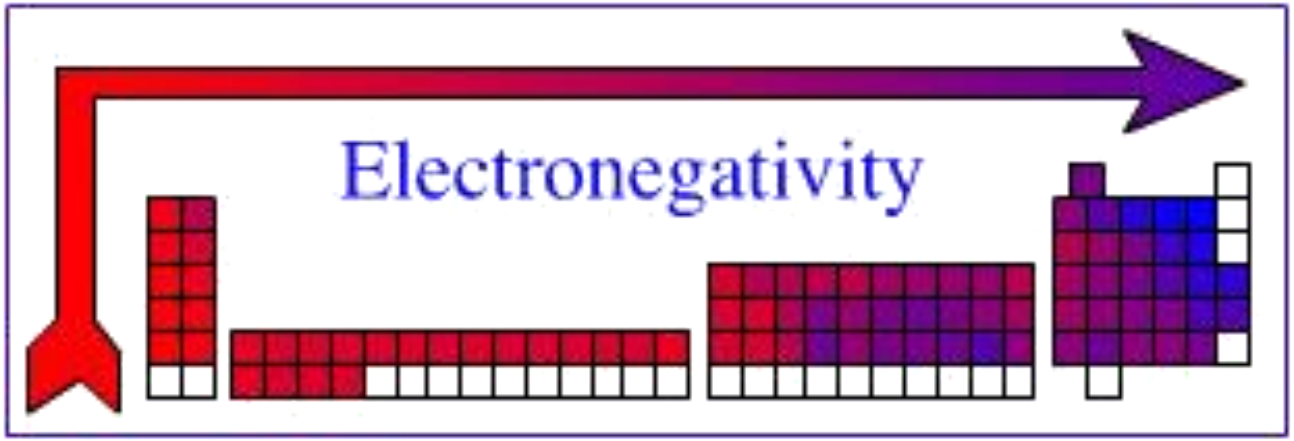
Electronegativity

Electronegativity is the attraction that an atom has towards electrons from another atom. The greater the electronegativity the stronger the pull it has towards other electrons.

Trends in the periodic table

- The larger the nucleus (with the positive protons) the stronger the electronegativity, this means it increases from left to right.

- The further the valence electrons are from the nucleus the less the electronegativity, therefore the electronegativity decreases down a group.



Electronegativity

extra info

We use a Pauling scale to determine electronegativity. The scale starts close to 0 – with minimal electronegativity and goes up to 4 with the highest electronegativity. Most of the Inert gases do not have a value because of there no reactivity with other atoms.

Electronegativity

0.7 → 4

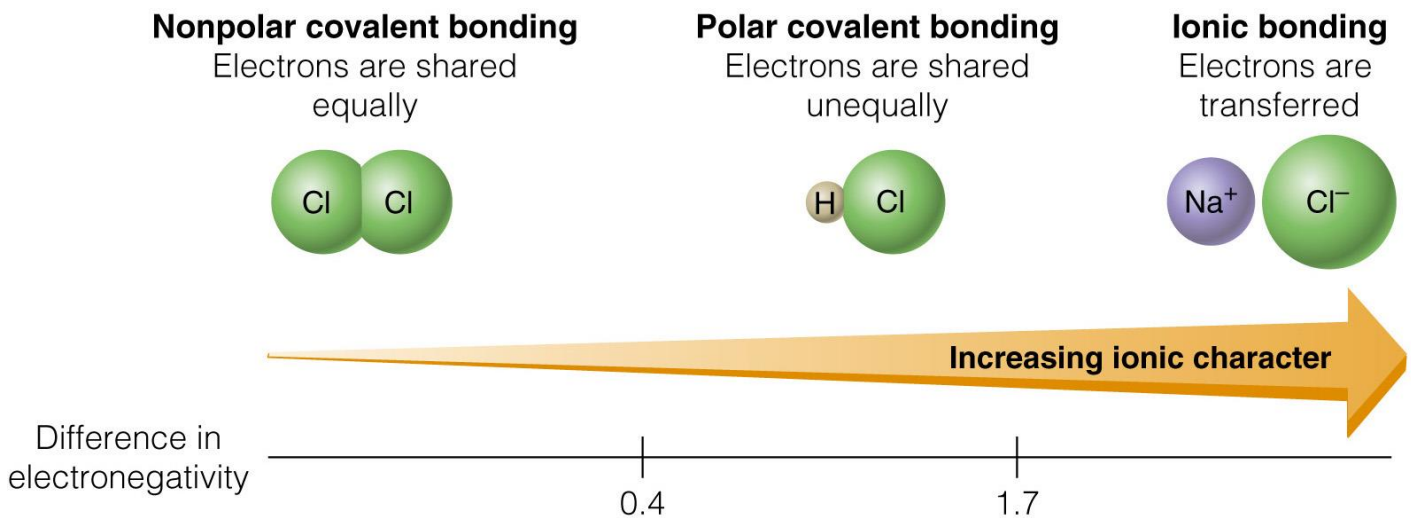
Pauling scale

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H 2.1																	He --
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne --
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar --
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn --
Fr 0.7	Ra 0.9	Ac 1.1	Rf --	Db --	Sg --	Bh --	Hs --	Mt --	Uun --	Uuu --	Uub --	113 --	Uuq --	115 --	116 --	117 --	118 --

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.2	Sm 1.2	Eu 1.1	Gd 1.2	Tb 1.2	Dy 1.2	Ho 1.2	Er 1.2	Tm 1.2	Yb 1.2	Lu 1.3
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.5	Lr --

Ionic – covalent bond continuum due to electronegativity

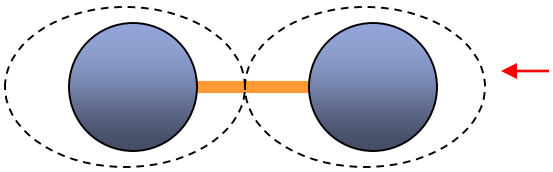
Bond types between atoms can depend on the **electronegativity** of the atoms. Rather than discrete categories, molecules fall along a continuum



If there is little difference in electronegativity between two atoms then they tend to form a covalent bond with no polarity difference. A greater electronegativity difference creates a polar bond with uneven “sharing” of valance electrons.
If the electronegativity is even greater then there will be a complete transfer of electron from one atom (Metal) to another atom (non-metal) and ions will form that are held together with an ionic bond.

Polarity

If two identical atoms are bonded together then they have exactly the same amount of attraction to the shared electrons in the bonded pair. This is because their electronegativity is the same. This becomes a **non-polar molecule** with **non-polar bonds**. Example - Iodine molecule I_2



Path of electrons

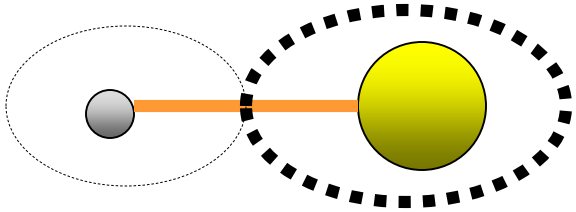
Evenly shared

If two different types of atoms are bonded together then they will exert different levels of attraction for the orbiting electrons. That is because they may have different numbers of electron shells and different numbers of protons in their nucleus. This will cause an electronegativity difference. These bonds become **polar bonds**. Example – hydrochloric acid HCl

Slightly positive

Electrons orbit less

δ+

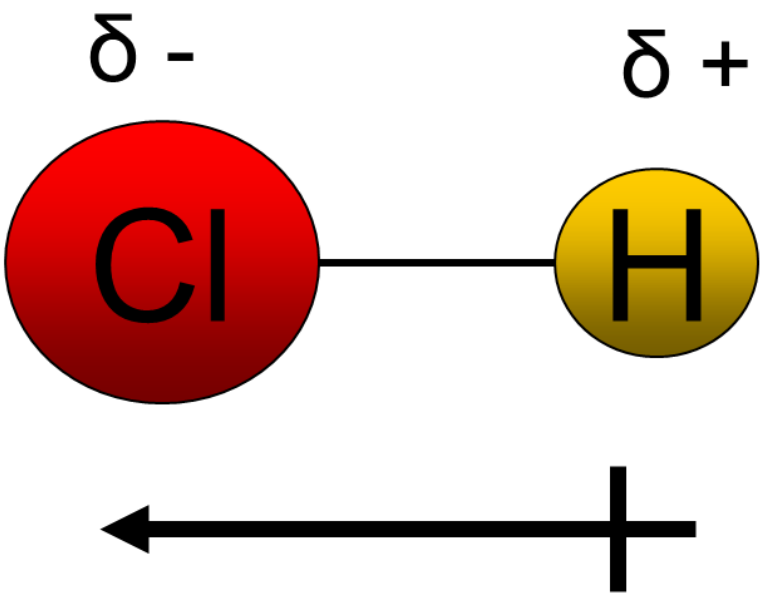


δ-

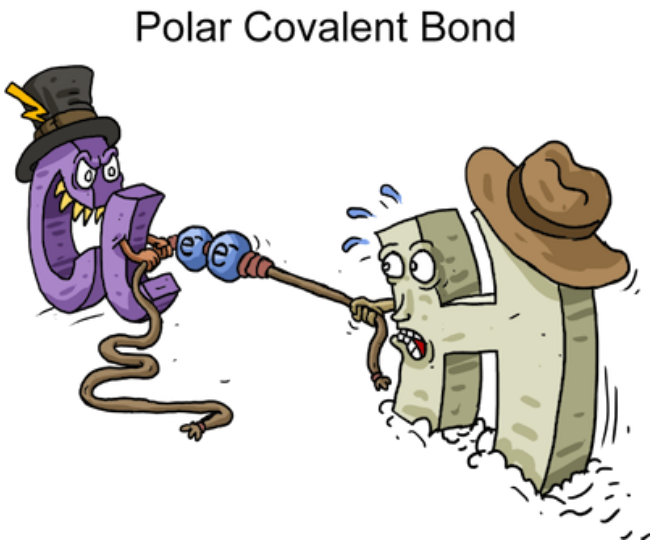
Slightly negative

Electrons orbit more

Polarity

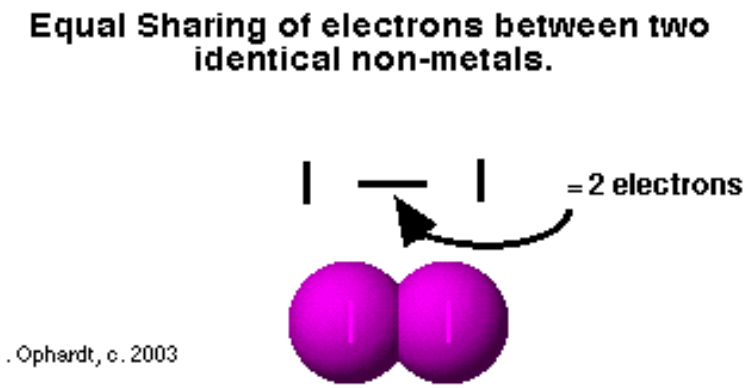


Polarity may also be shown as an arrow, with a cross, +ve, at the tail. The arrow head is the –ve end.

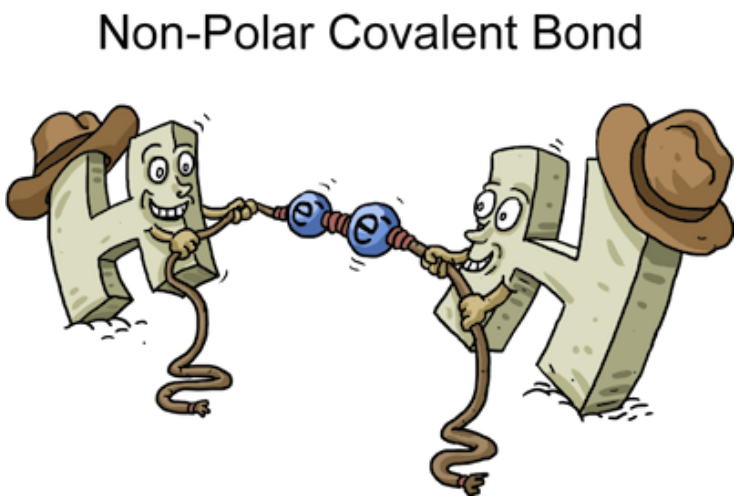


Polarity

If two bonded atoms are the same, the bond is said to be **non-polar**. i.e. I_2. The whole molecule is also non-polar because there is no electronegativity difference and the valence electrons orbit each atom evenly.



If two different atoms are bonded they form a **polar bond** which creates a **dipole**, as there is an electronegativity difference and the valence electrons spend more time around the atom with the higher electronegativity value (that atom becomes slightly negative) The atom that the valence electrons spend less time around becomes slightly positive.



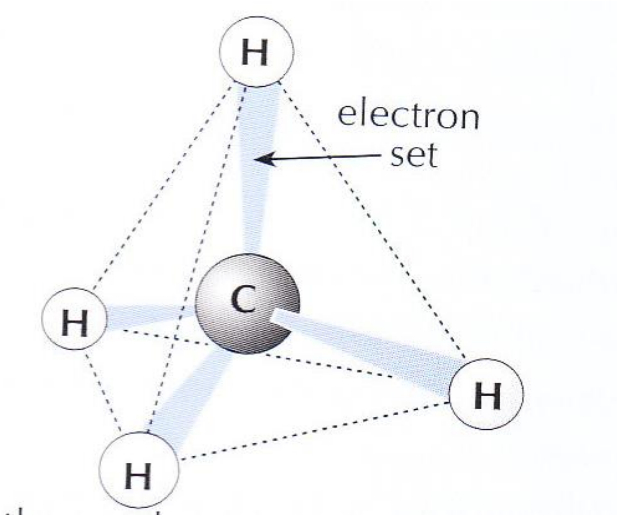
Symmetry and Polarity

The **overall polarity** of a molecule with polar bonds depends upon whether the molecule is symmetrical or not.

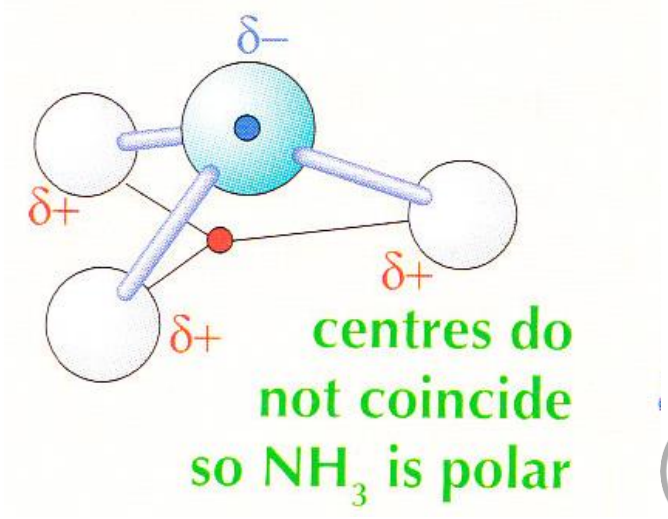
A symmetrical molecule (one where the centres of peripheral atoms coincide) becomes a **non-polar molecule** – as the bond dipoles cancel out.

An unsymmetrical molecule (one where the centre of peripheral atoms do not coincide) is a **polar molecule**, – as the bond dipoles **do not** cancel out.

Non-Polar molecule



Polar molecule



Answering NCEA Polarity Questions

Example: Explain why molecules x (CCl_4) and y (NCl_3) are polar and non-polar?

Polar molecule	Non-polar molecule
1. molecule (NCl_3) is polar (state which one) 2. (NCl_3) contains polar bonds and therefore forms dipoles due to electronegativity difference of N and Cl. 3. over the whole molecule the atoms are not distributed symmetrically in 3 dimensions because its shape is (state which one) and has lone pairs of electrons 4. Dipoles do not cancel each other out and the whole molecule is polar.	1. molecule (CCl_4) is non-polar (state which one) 2. (CCl_4) contains polar bonds and therefore form dipoles due to electronegativity difference of C and Cl. Cl attracts more electrons than C because it has a bigger atomic number than C but with the same number of shells 3. over the whole molecule the atoms are distributed symmetrically in 3 dimensions because its shape is (state which one) 4. Dipoles cancel each other out and the whole molecule is non-polar.

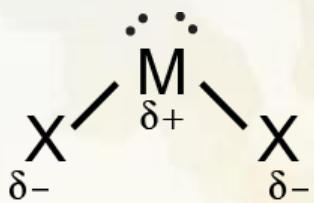
NCEA 2013 Molecule polarity

Excellence Question

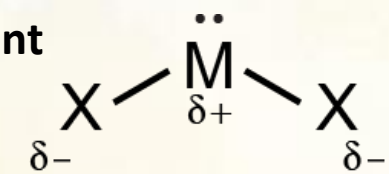
Question 1c (ii): Elements M and X form a compound MX_2. Atoms of element X have a higher electronegativity value than atoms of element M, therefore the M–X bonds are polar. Depending on what elements M and X are, molecules of the compound formed will be **polar** or **non-polar**.
State the most likely shape(s) of the molecule if it is **Polar** and if it is **Non-polar:**
Justify your answer and draw diagrams of the possible molecules with dipoles labelled.

If MX_2 is **polar**, this indicates that the polar M–X bonds are not spread symmetrically around the central M atom. There must be either **three or four regions of negative charge** with only two bonded atoms therefore the shape must be bent.

Polar: bent

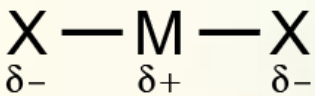


Three regions of negative charge:



Four regions of negative charge:

Non-polar: linear



Two regions of negative charge:

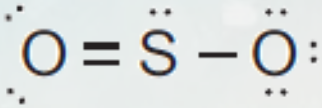
If MX_2 is **non-polar** this means that the polar M-X bonds are spread symmetrically around the central M atom. There must be only **two regions of negative charge** around the M atom, both bonded by X atoms in a linear shape.

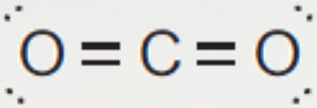
NCEA 2014 Molecule Polarity

Excellence Question

Question 1c: Molecules can be described as being polar or non-polar.
The following diagrams show the Lewis structures for two molecules, **SO_2** and **CO_2**.
Identify the polarity

- Justify your choice





When given a multi- choice question **never** leave it blank

State polarity of molecule first
State polarity of bonds (name atoms)
Link symmetry to dipoles cancelling out (or vice versa)

Answer 1c: SO_2 molecule is polar.
CO_2 molecule is non-polar.
The S–O / S=O bond is polar due to the **difference in electronegativity** between S and O atoms. The bonds are arranged **asymmetrically** in a bent shape around the central S atom; therefore the (bond) **dipoles do not cancel** and the molecule is polar.

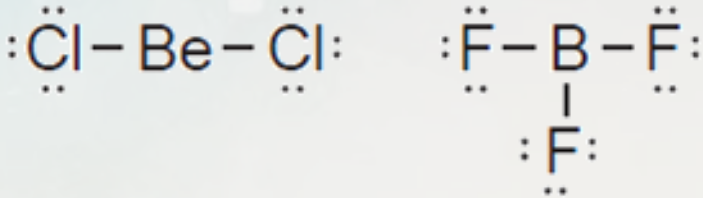
The C=O bond is polar due to the **difference in electronegativity** between C and O atoms. The bonds are arranged **symmetrically** in a linear shape around the central C atom; therefore the (bond) **dipoles cancel** and the molecule is non-polar.

NCEA 2015 Molecule Polarity

Excellence Question

Question 1c: $BeCl_2$ and BF_3 are unusual molecules because there are not enough electrons for the central atoms, Be and N, to have a full valance shell. Their Lewis structures are shown below. Both Molecules have the same polarity.

- Identify the polarity
- Justify your choice



When given a multi- choice question **never** leave it blank

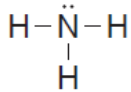
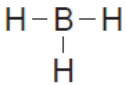
State polarity of molecule first
State polarity of bonds (name atoms)
Link symmetry to dipoles cancelling out (or vice versa)

Answer 1c: Both molecules are non-polar.
The Be-Cl bond is polar because Cl is more electronegative than Be / the atoms have different electronegativities.
Since both the bonds are the same and arranged symmetrically around the central atom, in a linear arrangement, the bond dipoles cancel out, resulting in a non-polar molecule.
The B-F bond is polar because F is more electronegative than B / the atoms have different electronegativities. Since all three bonds are the same and arranged symmetrically around the central atom, in a trigonal planar arrangement, the bond dipoles cancel out, resulting in another non-polar molecule.

NCEA 2016 Molecule Polarity

Excellence Question

Question 3b: The Lewis structures for two molecules are shown.

Molecule	 Ammonia	 Borane
Polarity of molecule	polar	non-polar

Ammonia, NH_3, is polar, and borane, BH_3, is non-polar. Justify this statement.

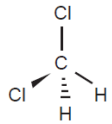
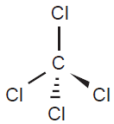
State polarity of molecule first
State polarity of bonds (name atoms)
Link symmetry to dipoles cancelling out (or vice versa)

Answer 3b: Each N-H bond in NH_3 is polar / forms a dipole because the N and H atoms have different **electronegativities.** The shape of the molecule (due to the presence of one non-bonding electron pair) is trigonal pyramidal which is asymmetrical, so the dipoles / bond polarities do not cancel. The **resulting NH_3 molecule is polar.**
Each B-H bond in BH_3 is polar / forms a dipole because the B and H atoms have different **electronegativities.** The shape of the molecule is trigonal planar which is symmetrical, so the dipoles / bond polarities cancel. The **resulting BH_3 molecule is non-polar.**

NCEA 2017 Molecule Polarity

Excellence Question

Question 2b: Three-dimensional diagrams for two molecules are shown below.

Molecule		
Name	Dichloromethane	Tetrachloromethane
Polarity of molecule	polar	Non-polar

(i) In the boxes above, identify the polarity of each molecule, by writing either **polar** or **non-polar.**

(ii) Justify your choices.

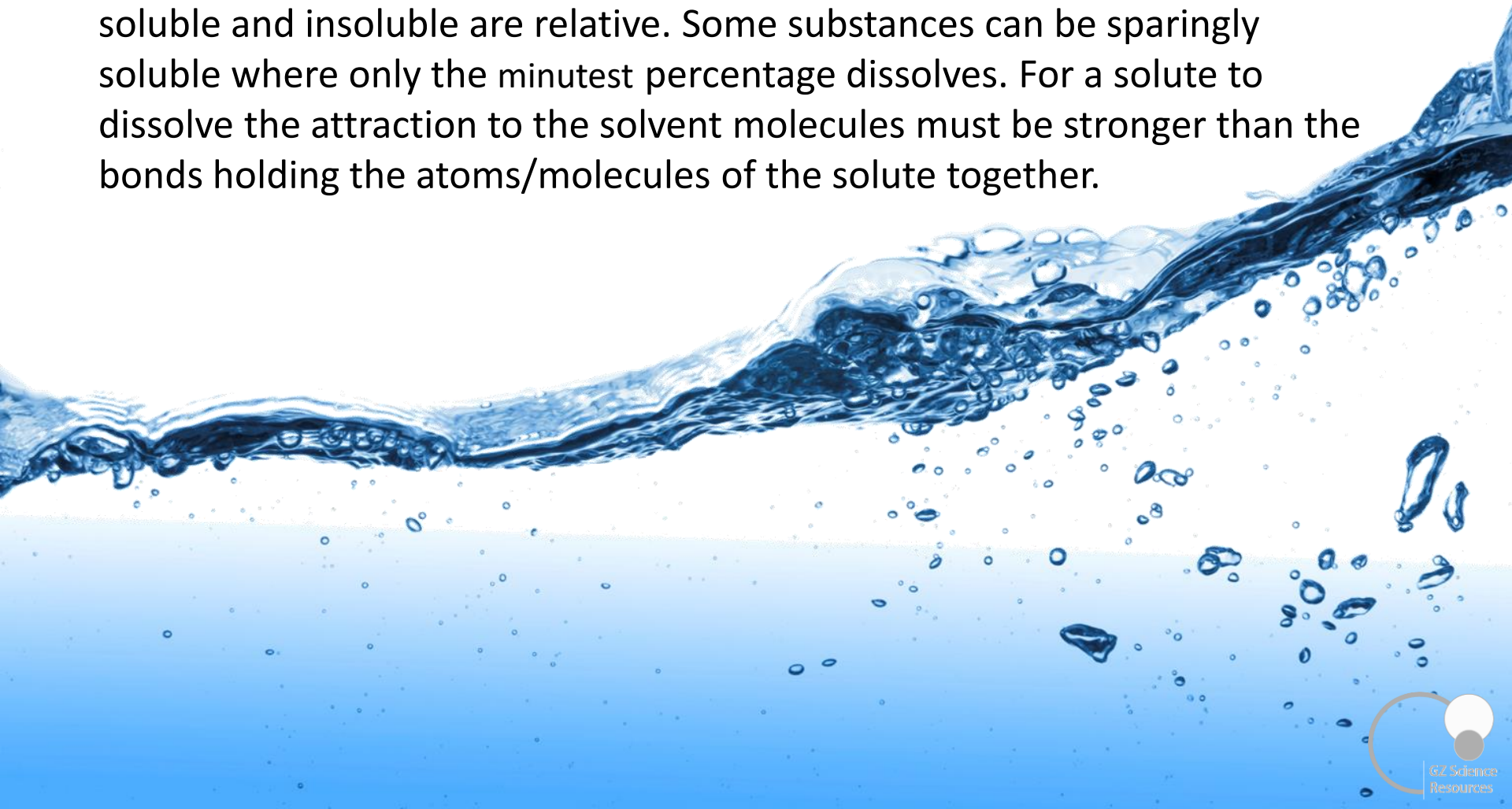
State polarity of molecule first
State polarity of bonds (name atoms)
Link symmetry to dipoles cancelling out (or vice versa)

In CCl_4, the four C–Cl bonds are polar, i.e. have a dipole, due to the difference in electronegativity between C and Cl. These (equally sized) dipoles are arranged in a symmetric **tetrahedral** shape, resulting in the dipoles / bond polarities cancelling each other out, so CCl_4 is non-polar.

In CH_2Cl_2, there are two types of bond, C–H and C–Cl, each polar with dipoles due to the difference in electronegativity between C and H and C and Cl. These dipoles have different polarities / sizes as H and Cl have different electronegativities. (Despite the symmetric tetrahedral arrangement) the different (sized) dipoles / bond polarities do not cancel each other out, so CH_2Cl_2 is polar.

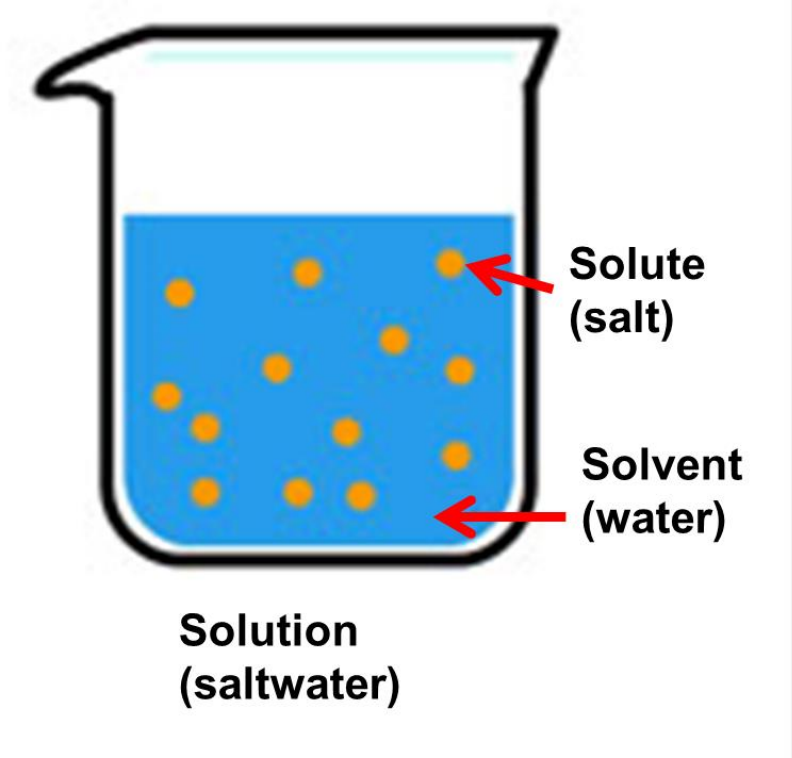
Solubility

The solubility of a substance is the **amount** of that substance that will dissolve in a given amount of solvent. Solubility is a quantitative term. Solubility will vary depending on the solvent and the solute. The terms soluble and insoluble are relative. Some substances can be sparingly soluble where only the minutest percentage dissolves. For a solute to dissolve the attraction to the solvent molecules must be stronger than the bonds holding the atoms/molecules of the solute together.



Background Knowledge

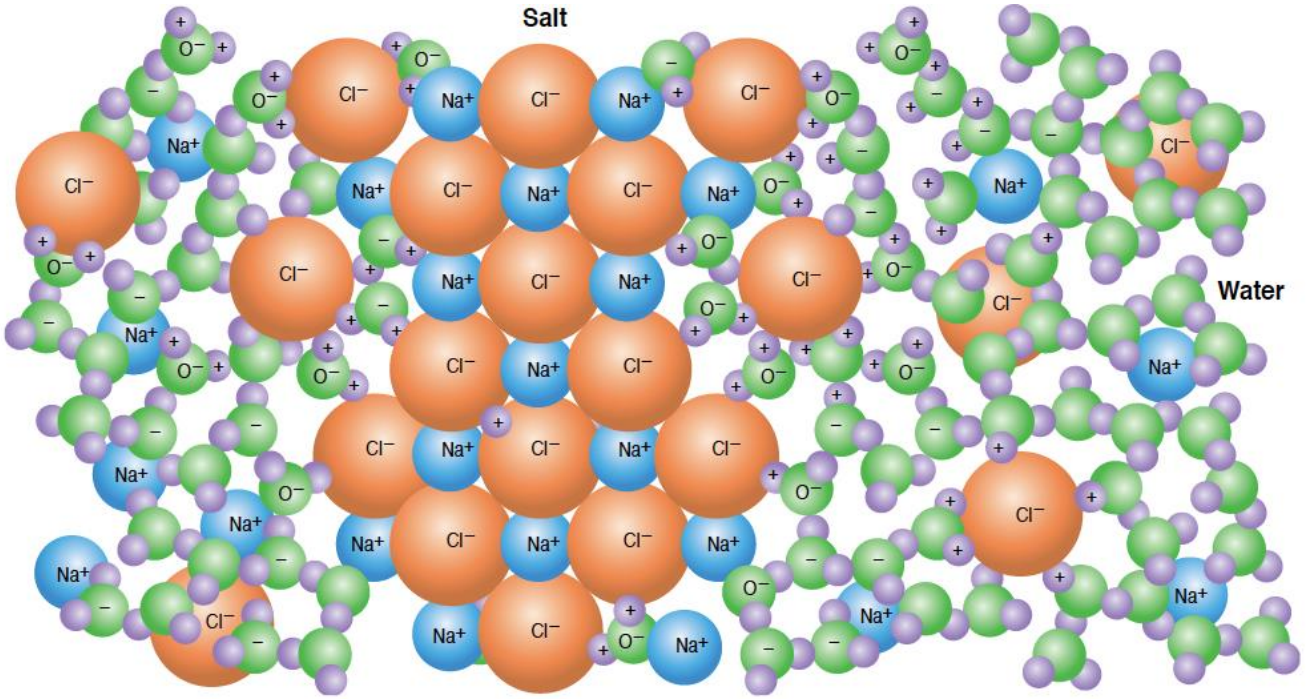
Solutions



A **solution** is made up of a **solvent** and a **solute**. A solvent is a substance such as water that is able to dissolve a solute. The solvent 'pulls apart' the bonds that hold the solute together and the solute particles **diffuse** (spread randomly by hitting into each other) throughout the solvent to create a solution. The solution is a **mixture** with evenly spread solvent and solute particles. These particles can be physically separated by **evaporation**.

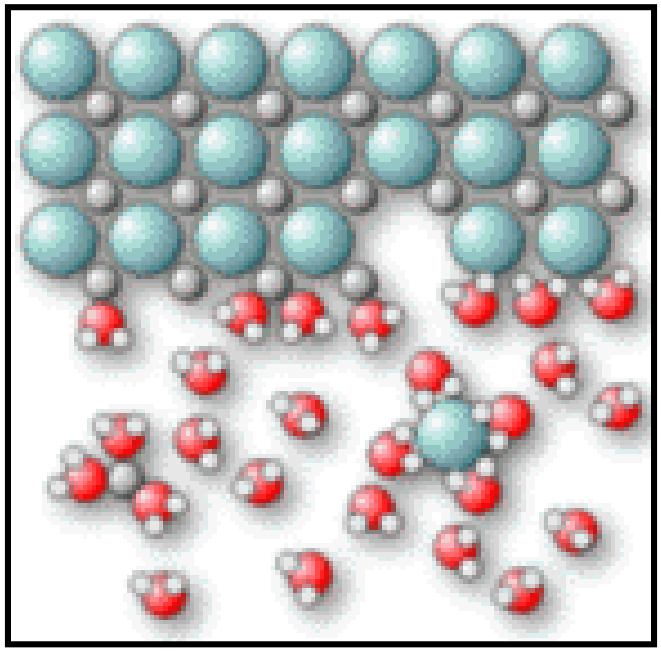
Solutions form when a solute is dissolved in a solvent

When a solid mixes into a liquid and can longer be seen it has **dissolved**. The liquid is called the **solvent** and it pulls apart the bonds between the solid particles, called the **solute**, and they **diffuse**. A **solution** is then created when the solvent particles (often water) are mixed up with the broken apart solute particles.



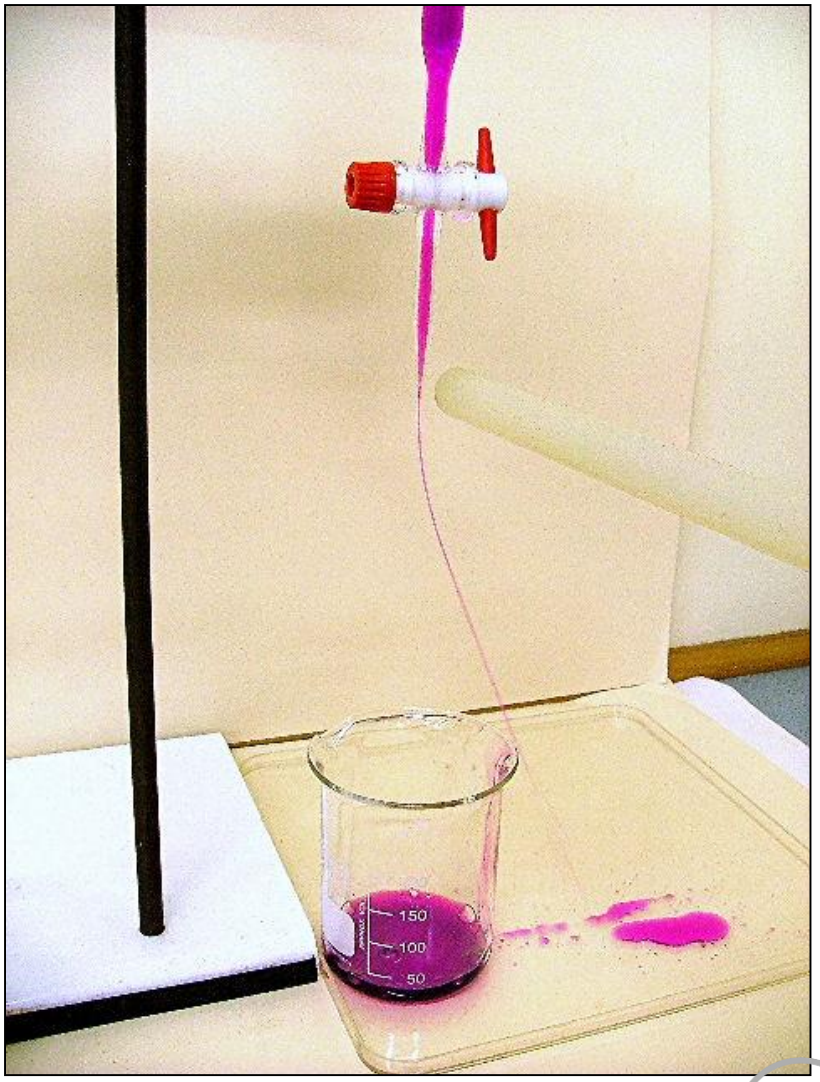
Solubility

For a solute to dissolve, the solvent particles must form bonds with the solute particles that are of similar strength, to the bonds between the solute particles.
Water, being polar attracts ions because they are charged and so dissolves many ionic substances.



Polar Solvents

The water molecule has two polar bonds. Due to the **asymmetry** of the molecule, their dipoles reinforce making the oxygen side of the molecule partially negative (δ-) and the hydrogen side partially positive (δ+). Such molecules are called **'polar'**. Polarity causes a stream of water molecules to attract to a charged plastic pen.



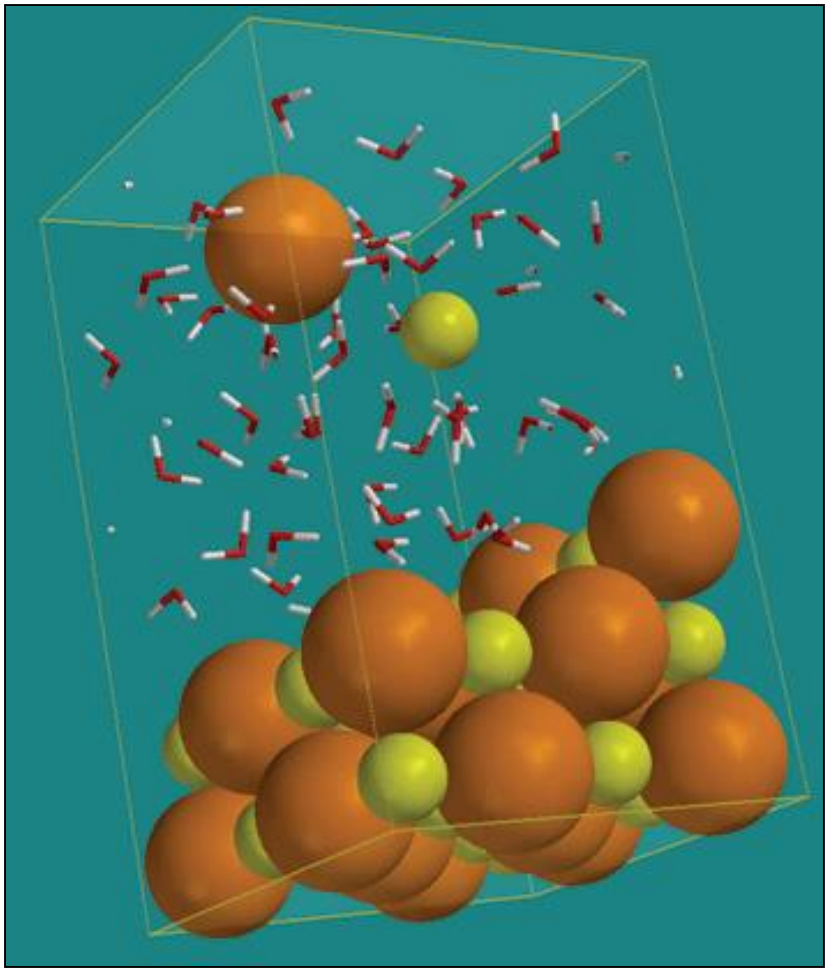
The separation of charge in the molecule (δ+ - δ-) is called a **'dipole'**.

Dissolving and Polarity

Polar substances dissolve polar substances.
e.g. Water, being polar attracts the molecules of other polar substances (e.g. HCl) and will dissolve them.

Polar substances will not dissolve non-polar substances.
e.g. Water, (polar) has a stronger attraction to itself than to non-polar molecules (e.g. cyclohexane) and will not dissolve them.

Non-polar substances dissolve non-polar substances.
e.g. Non-polar solvents (like cyclohexane) attract non-polar solutes (like napthalene) by the same weak molecular forces they attract themselves by and so will dissolve non-polar solutes.



Ionic solid dissolving in water

Common Polar and Non-polar molecular substances

Polar	Non-Polar
water methanol ethanol acetic acid hydrogen chloride	cyclohexane benzene hydrocarbons (e.g. petrol) oxygen hydrogen nitrogen iodine

Sample NCEA Questions:

Potassium chloride will not dissolve in non-polar solvents, but will dissolve in water. Explain by relating the property to the structure and bonding within the solid.

Use the structure and bonding in H_2O and SO_2 to explain why SO_2 is soluble in H_2O.

NCEA 2014 Solubility

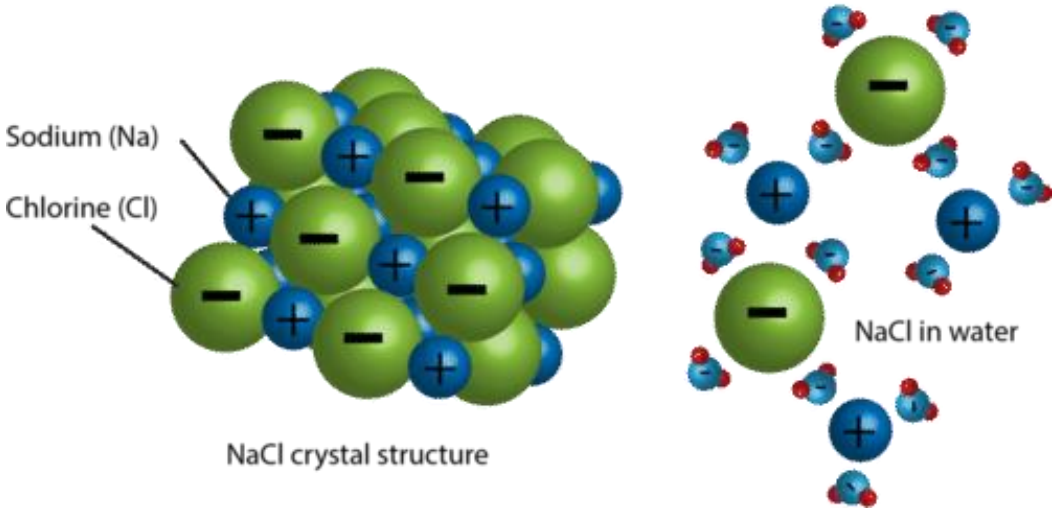
Excellence Question

Question 3b: Use your knowledge of structure and bonding to explain the dissolving process of sodium chloride in water.
Support your answer with an annotated (labelled) diagram.

Only a few ions of each are needed in the diagram. Don't forget to label

Answer 3b : Solubility

When sodium chloride is dissolved in water the attractions between the polar water molecules and between the ions in the salt are replaced by attractions between the water molecules and the ions. The negative charges on the oxygen ends of the water molecules are attracted to the positive Na^+ ions, and the positive hydrogen ends of the water molecules are attracted to the negative Cl^- ions.



NCEA 2017 Solubility

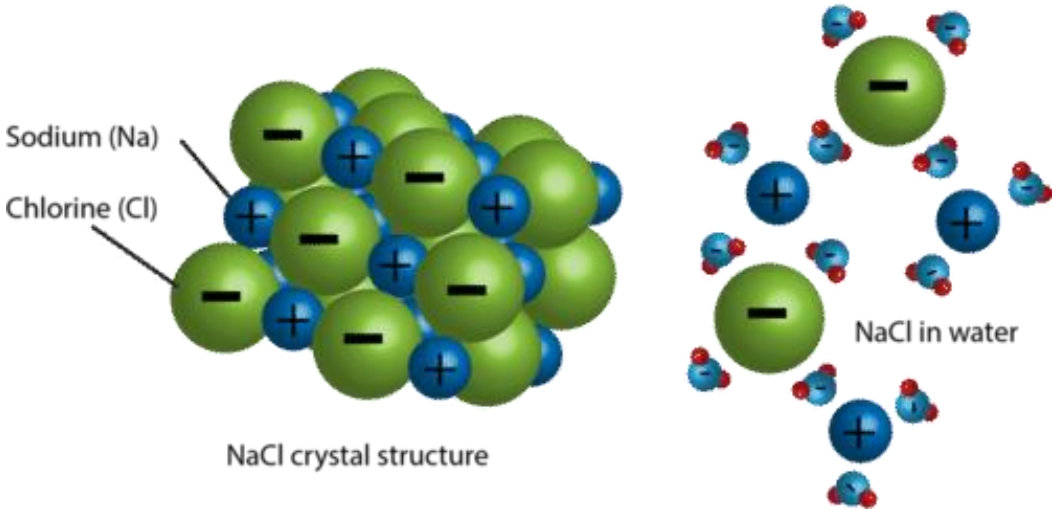
Excellence Question

Question 1b: (iii) Sodium chloride, NaCl, is another compound that is excreted from the body in sweat.
Use your knowledge of structure and bonding to explain the dissolving process of sodium chloride, NaCl, in water.
Support your answer with a labelled diagram.

Only a few ions of each are needed in the diagram. Don't forget to label

Answer 3b : Solubility

Sodium chloride is an ionic substance made up of Na^+ and Cl^- ions arranged in a (3D) lattice and held together by ionic bonds. The δ- O of polar water molecules are attracted to the positive Na^+, while water's δ+ H is attracted to the negative Cl^-, this attraction is sufficiently strong to overcome the attractions between the ions in the salt / crystal /lattice (and between the water molecules in the solvent), dissolving the NaCl.



Groups of substances

Substances are grouped together according to the type of bonds they have between particles, and consequently the structure they form.

This year will cover four groups of substances; Molecular, metallic, ionic and covalent network. The physical properties of these groups will be linked to their structure.

Molecular solids

Non-metals forming molecules

S_2
sulfur

HCl
Hydrogen chloride

I_2
iodine

Ionic solids

Non-metals and metals together forming a ionic compound

KI
Potassium iodide

$CuSO_4$
Copper sulfate

NaCl
Sodium chloride

Metallic solids

Elements that are metals

Fe
iron

Al
aluminium

Cu
copper

Covalent network solids

Carbon and silicon dioxide

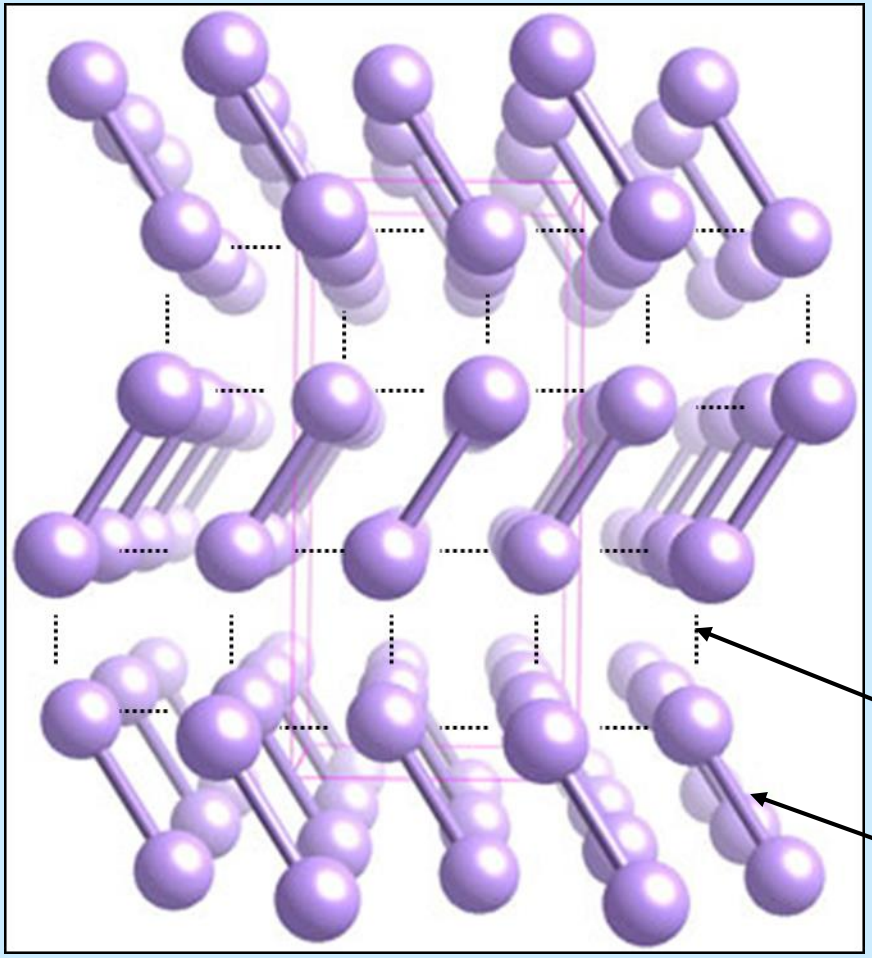
SiO_2
Silicon dioxide

C
diamond

C
graphite

Non-polar Molecular solids

non-metal + non-metal



Molecules are held together by **weak intermolecular bonding** caused by temporary dipoles only. These are induced (created) by electrons randomly spending more time around one nucleus than the other.

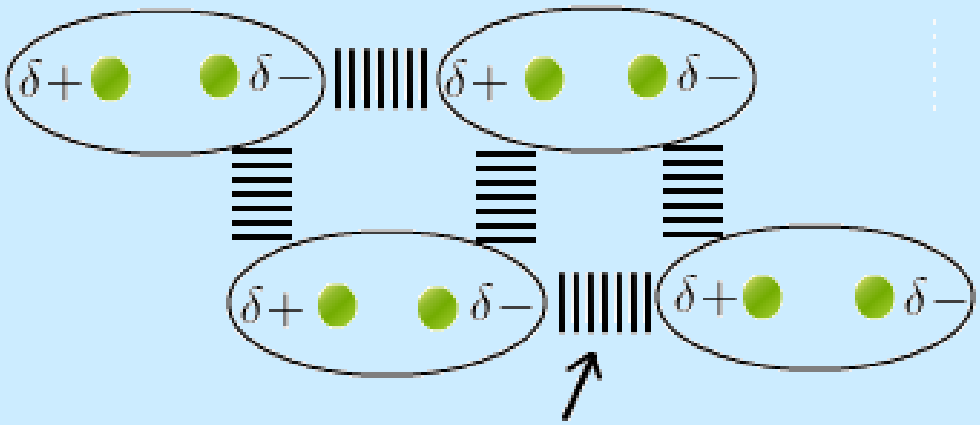
Within the Molecules, the atoms are held together by **strong covalent bonds**.

Weak intermolecular bond

Strong covalent bond

Polar Molecular solids

Polar molecules held together by **weak intermolecular** forces caused by both temporary and permanent dipoles (which tend to be stronger). Permanent dipoles are induced by electrons spending more time around one nucleus in the molecule that has greater electronegativity than the other. The δ –ve end of one molecule is attracted to the δ +ve end of another.



Intermolecular Force between diploles

δ means slightly

Note the distinction:
Intra-molecular Forces: the strong bonding forces within a molecule. i.e. the covalent bonds holding the molecule together.
Inter-molecular Forces: the weak bonding forces between molecules due to the attractions between partial charges. i.e. permanent dipole

Polar Molecular solids - **solubility**

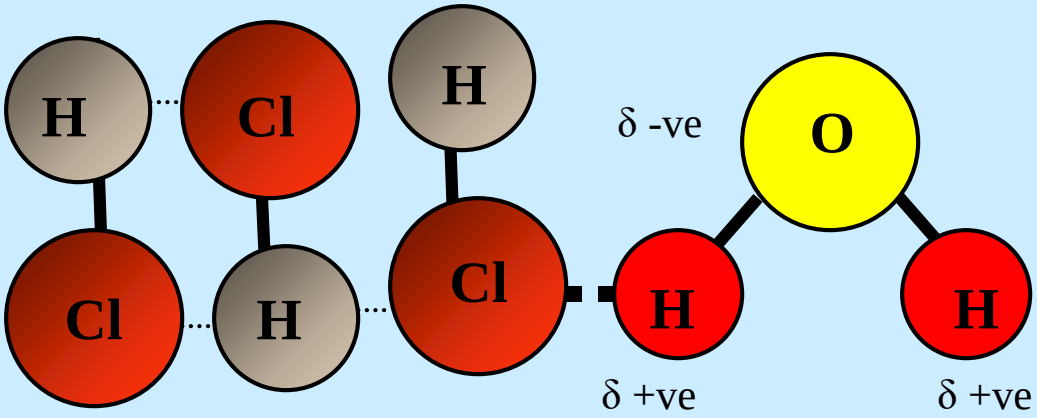
Polar molecular solids are soluble in water

1. Hydrogen chloride (HCl) is a molecular solid

2. Hydrogen chloride is composed of covalently bonded atoms to form molecules

3. These molecules are held together by weak intermolecular forces

4. These molecules are **polar** therefore, the electrostatic attractions of water molecules (which is stronger than the weak intermolecular forces) have sufficient strength to pull the molecules apart hence, hydrogen is soluble



Non-polar Molecular solids - **solubility**

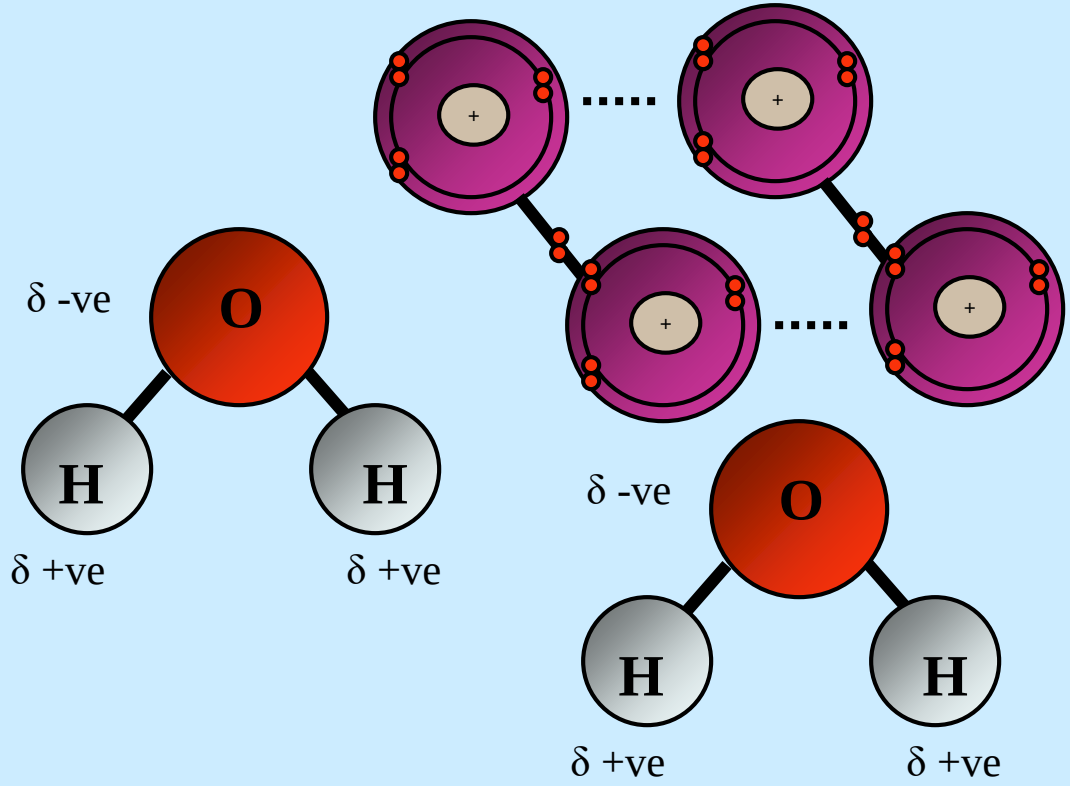
Non-polar molecular solids are insoluble in water

1. Iodine is a molecular solid

2. Iodine is composed of covalently bonded atoms to form molecules

3. These molecules are held together by weak intermolecular forces

4. Iodine is **non-polar** therefore, the electrostatic charges of the water do not have sufficient strength to overcome the weak intermolecular forces holding the molecules together hence, iodine is insoluble



Molecular solids – **Melting point**

molecular solids have a low melting point

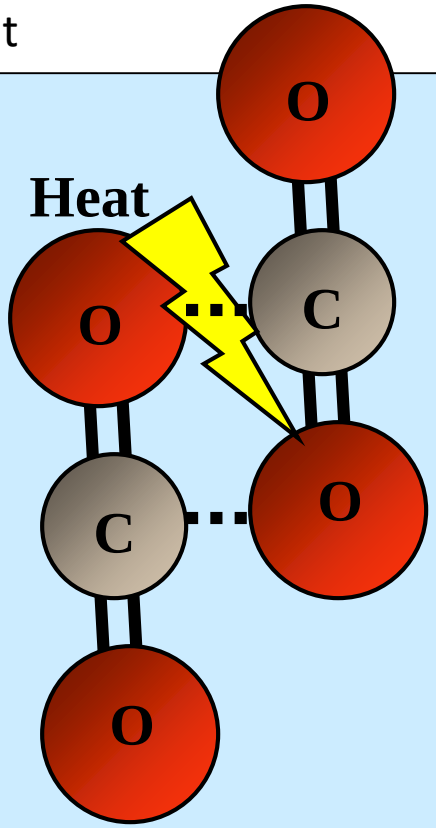
<u>For example:</u>

1. Carbon dioxide is a molecular solid (at low temperatures below -56◦C)

2. Carbon dioxide is composed of covalently bonded atoms to form molecules

3. These molecules are held together by weak intermolecular forces

4. These forces require small amounts of energy to break apart the solid (but not the individual molecules which are held together by strong covalent bonds) therefore, carbon dioxide has low melting point



Many molecular solids are only solid at temperatures well below 0◦C and at room temperature they are gases

Molecular solids - **hardness**

molecular solids tend to be soft

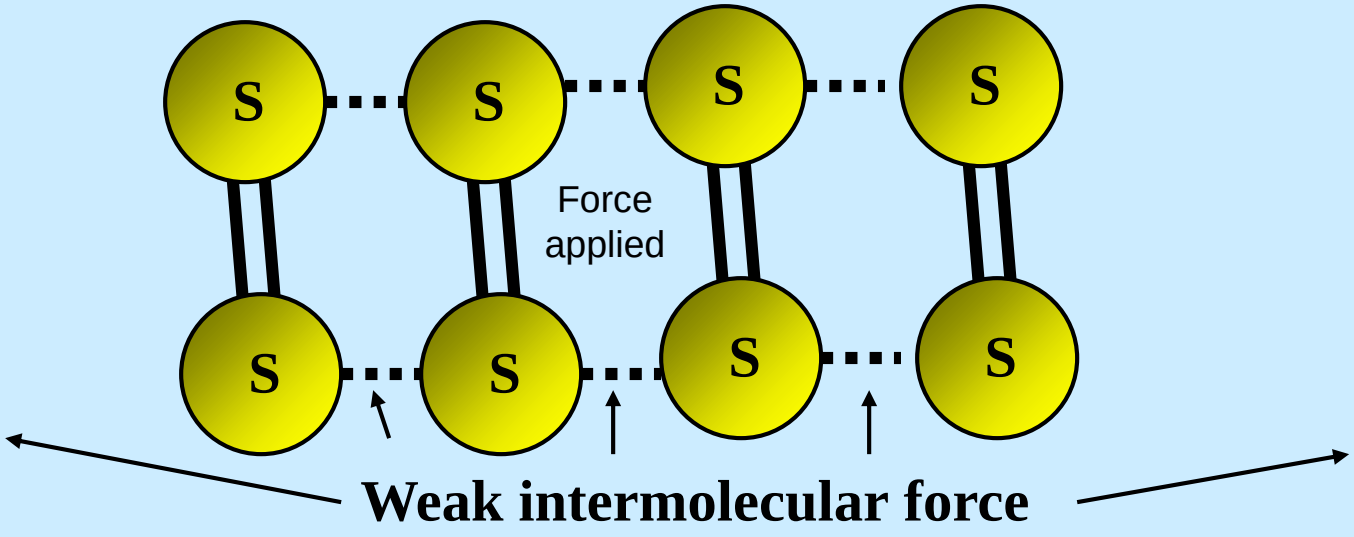
<u>For example:</u>

1. Sulfur is a molecular solid

2. Sulfur is composed of covalently bonded atoms to form molecules

3. These molecules are held together by weak inter molecular forces

4. These forces require small amounts of energy to break apart the solid (but not the individual molecules which are held together by strong covalent bonds) therefore, sulfur is easily broken up



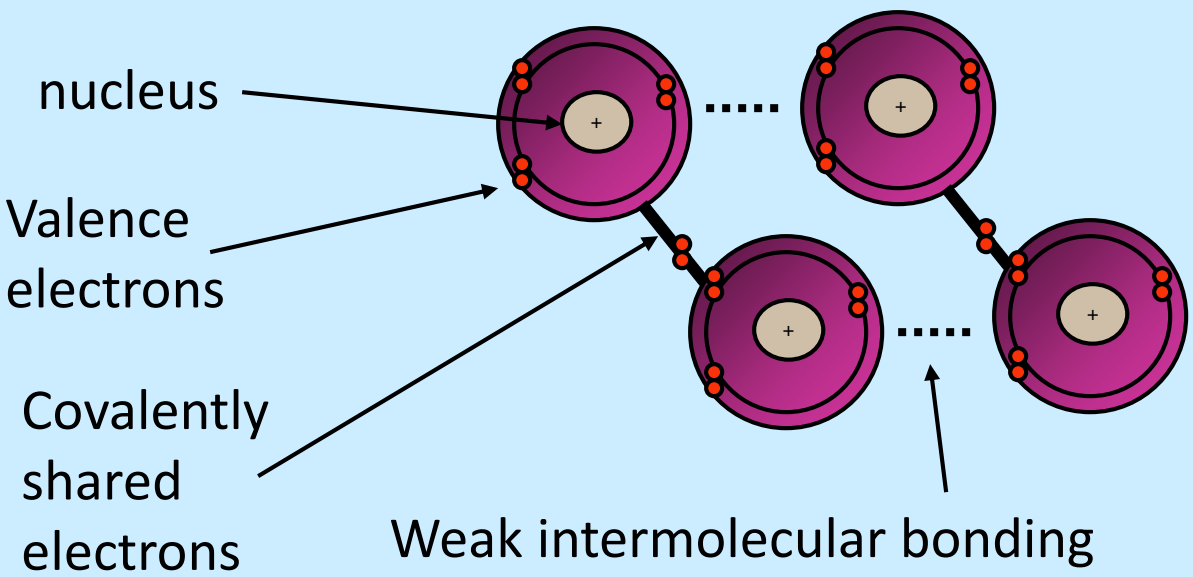
Molecular solids - **Conductivity**

molecular solids do not conduct electricity

In order for a substance to be electrically conductive there must be free moving charged particles

For example:

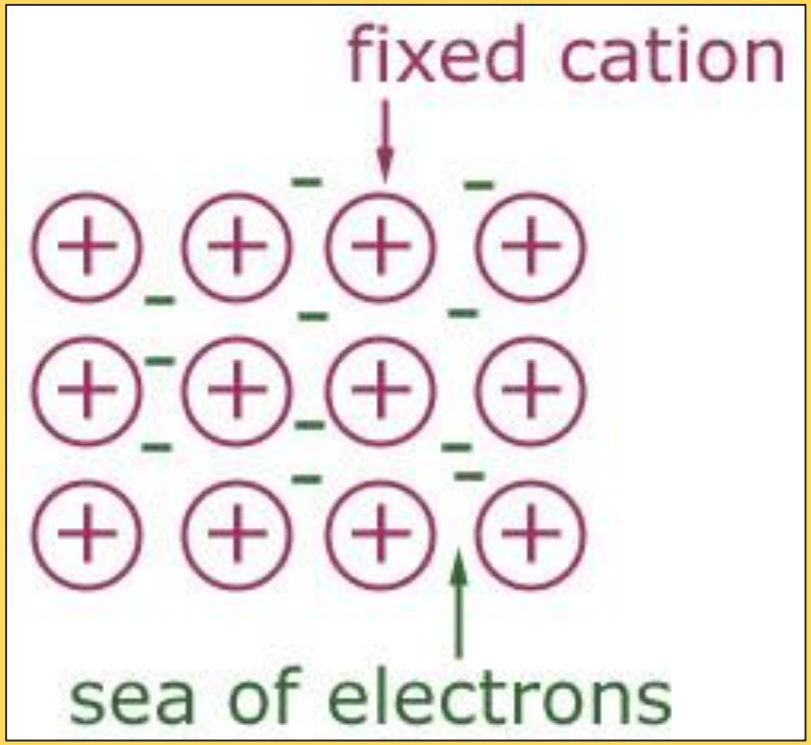
1. Iodine is a molecular solid
2. Iodine is composed of covalently bonded atoms to form molecules; weak intermolecular forces hold these molecules together
3. There are no free moving charges therefore, iodine cannot conduct electricity



Fully occupied valence electrons remain in 'fixed orbit' around nucleus and are not available to carry charge. The molecule is neutral

Metallic Solids - **structure**

Metals atoms are arranged as positive ions held in place in ordered layers by strong attractive **non-directional** bonding, forming a lattice - this gives metals strength.



Metal atoms are held together in a 3–D lattice by metallic bonding in which valence electrons are attracted to the nuclei of neighbouring atoms. The attraction of the metal atoms for the valence electrons is not in any particular direction; therefore metal atoms can move past one another without disrupting the metallic bonding, therefore metal is ductile and malleable.

The atoms are packed tightly together - this makes metals dense

Metallic Solids - **Conductivity**

Metallic solids are conductive

Free moving charged particles are required to carry a charge and for a substance to be electrically conductive

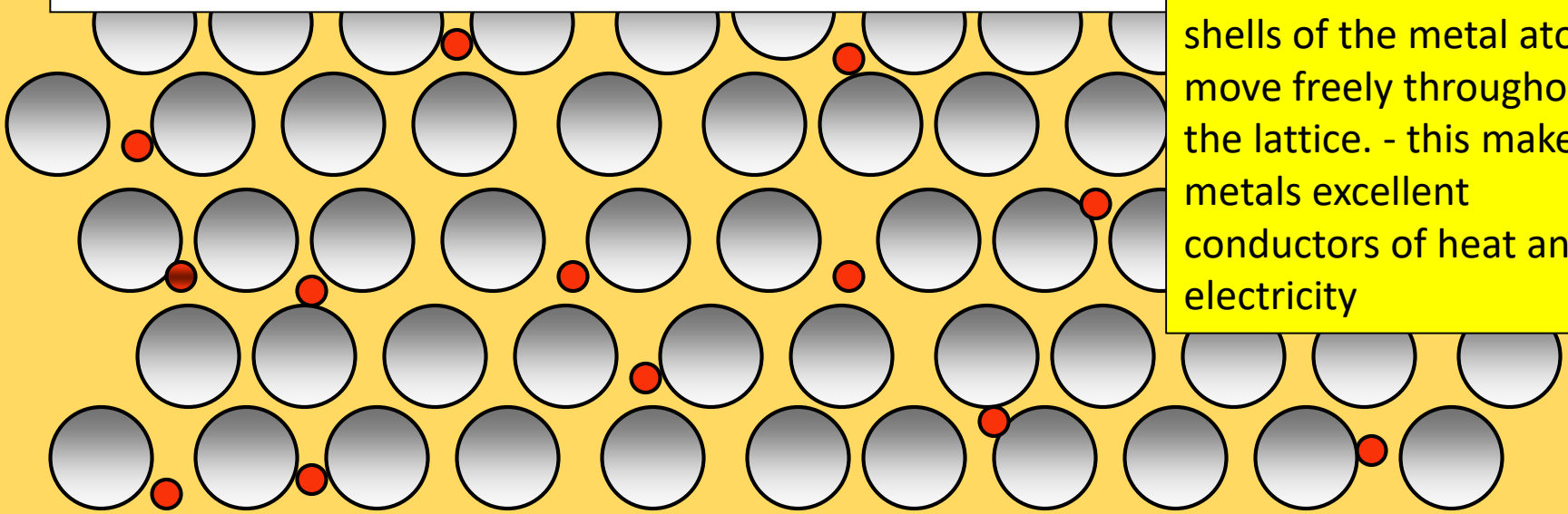
<u>For example:</u>

1. Copper is a metallic solid

2. Copper is arranged as positive ions held in place in ordered layers by strong attractive non-directional forces, in a sea of de-localised electrons

3. Electrons are free moving hence can carry a charge

4. Therefore, copper **can conduct electricity**



Electrons from the outer shells of the metal atoms move freely throughout the lattice. - this makes metals excellent conductors of heat and electricity

Metallic Solids - **Solubility**

Metallic solids are not soluble

<u>For example:</u>

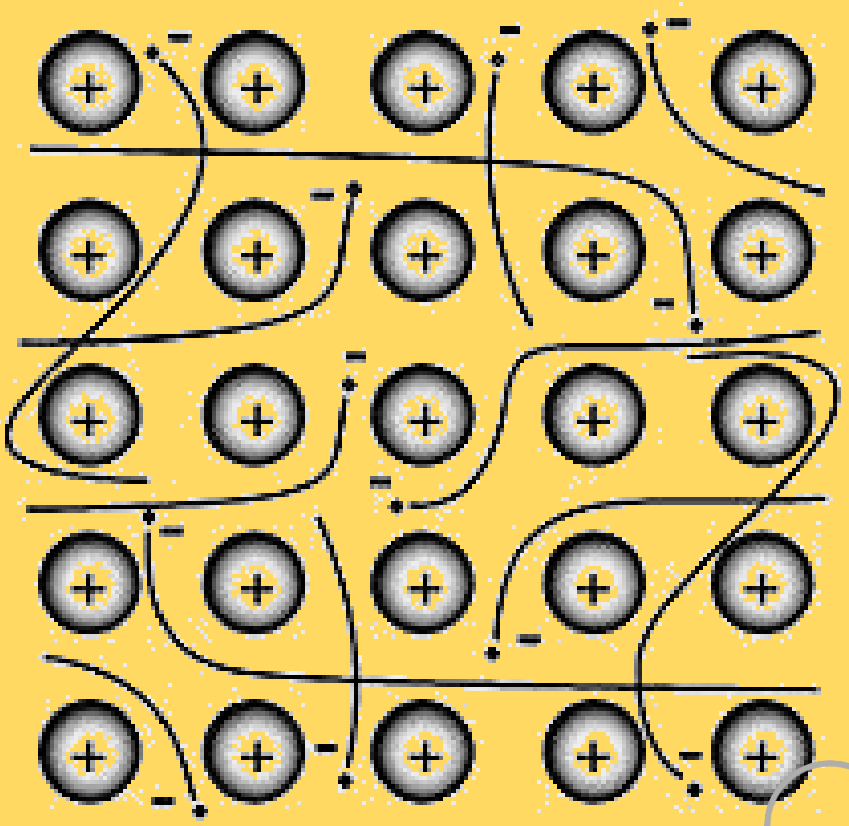
1. Lead is a metallic solid

2. Lead is arranged as positive ions held in place in ordered layers by strong attractive non-directional forces, in a sea of de-localised electrons

3. These forces require a large amount of energy to break therefore the electrostatic attractions of water molecules do not have sufficient strength to pull the atoms apart

4. Therefore, **lead is insoluble**

In order for substance to dissolve in water (a polar liquid) the attraction between the particles in a substance must be less than the attraction towards water molecules



Metallic Solids – **Malleability and ductility**

Metallic solids are Malleable and ductile

Layers of ions can slide over each other without breaking- this makes metals hard and also malleable and ductile

<u>For example:</u>

1. Iron is a metallic solid

2. Iron is arranged as positive ions held in place in ordered layers - **a lattice**, by strong attractive non-directional forces, in a sea of de-localised electrons

3. These forces require large amounts of energy to break apart the solid therefore aluminium is not easily broken up

4. However, when pressure is applied layers can slide over each other, and as the attractive forces are non-directional the metallic particles remain strongly bonded. – This gives the metallic solids the properties of being malleable (moulded into flat sheets) and ductile (drawn out to thin wires)

Metallic Solids – **Melting point**

Metallic solids have a high melting point

<u>For example:</u>

1. Aluminium is a metallic solid

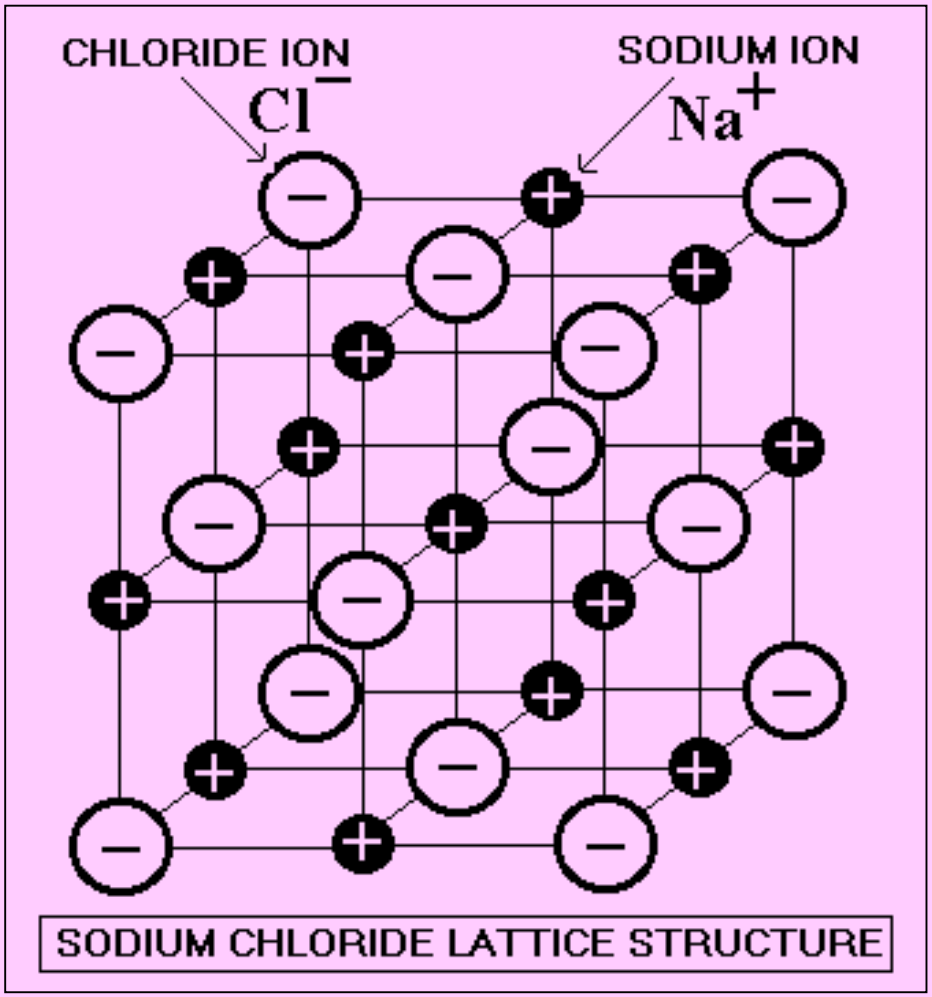
2. Aluminium is arranged as positive ions held in place in ordered layers by strong attractive non-directional forces, in a sea of de-localised electrons

3. These forces require a large amount of energy (high temperature) to break apart the metallic solid therefore the melting point is very high.

The strength of the bonds between particles determines the energy required to break them, and therefore the amount of energy to change a solid into a liquid (the melting point) where the bonds are somewhat broken. Metals in general, have very strong bonds which makes them solid at room temperature (Mercury is the exception)

Three steps to answering structure and physical properties questions.
> The first is state the name of the solid.
> The second is describe the structure of the solid.
> The third is link the structure of the solid to the physical property discussed.

Ionic Solids - **structure**



Metal + Non-Metal

An ionic solid is composed of ions held together by strong directional electrostatic forces (ionic bonding) between +ve (cations) and –ve (anions) ions in a 3-dimensional lattice.

Ionic Solids - **Solubility**

Ionic solids are soluble in solution

In order for substance to dissolve in water (a polar liquid), the attraction between the particles in a substance must be less than the attraction towards water molecules

<u>For example:</u>

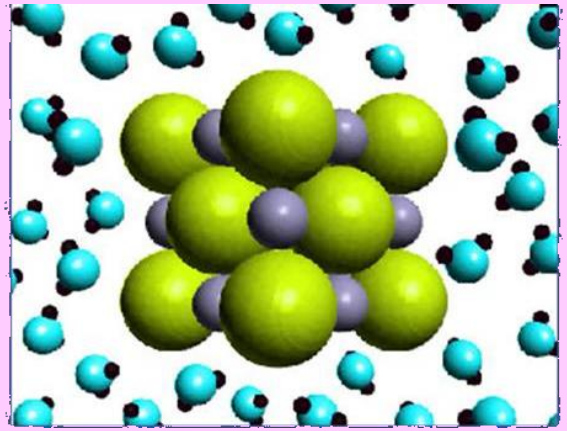
1. Sodium chloride (NaCl) is an ionic solid

2. Sodium chloride is composed of ions held together by strong directional electrostatic attractions between +ve and –ve ions in a lattice

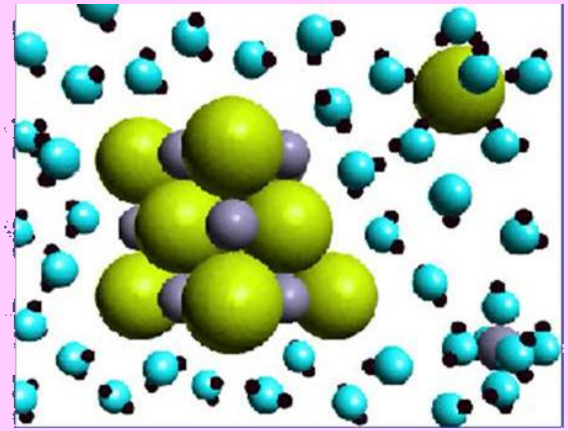
3. The electrostatic attractions of water molecules have sufficient strength to pull the ions apart

4. Therefore, the solid will dissolve and is soluble

NaCl first placed in water



Na^+ and Cl^- ions breaking apart



The positive hydrogen end of water is attracted to the anions and the negative oxygen end of water is attracted to the cations

Ionic Solids - **Conductivity**

Ionic solids are conductive when in solution or molten only

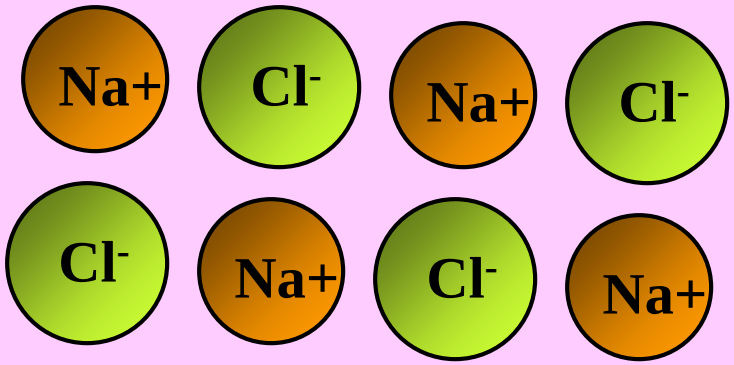
For example:

1. Sodium chloride is an ionic solid

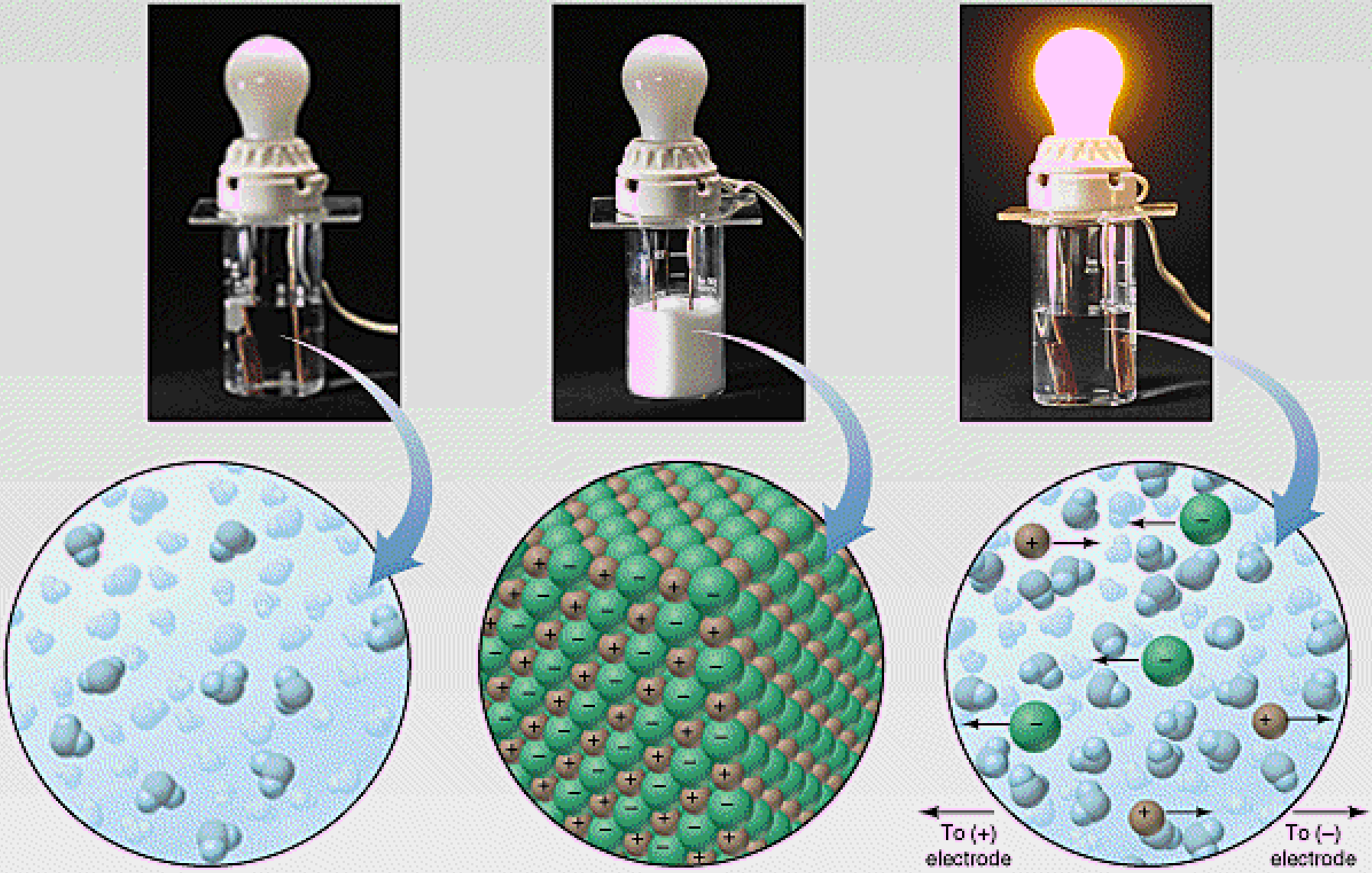
2. Sodium chloride is composed of ions held together by strong directional electrostatic forces between +ve and –ve ions in a 3-d lattice

3. When solid the ions are not free to move therefore it doesn't conduct electricity

4. However, when molten, or dissolved in solution, the bonds are broken and the ions are free to move therefore sodium chloride can conduct electricity



Free moving charged particles are required to carry a charge and for a substance to be electrically conductive



Distilled water does not conduct a current

Positive and negative ions fixed in a solid do not conduct a current

In solution, positive and negative ions move and conduct a current

Ionic Solids – **Hardness and brittleness**

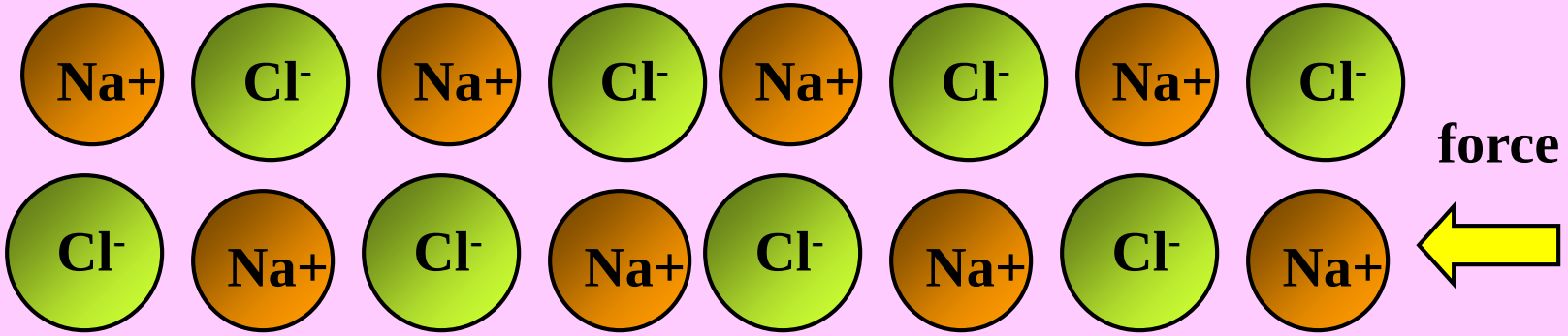
Ionic solids are hard but brittle

For Example:

1. Sodium chloride is an ionic solid

2. Sodium chloride is composed of ions held together by strong directional electrostatic attractions between +ve and –ve ions in a 3-d lattice so requires a lot of energy to break the bonds

3. However, because the ionic bonding is directional, if sideways force is applied and a sheet of the lattice slides then ions of the same charge may come in close contact with each other and repel hence the ionic solid is brittle (and can break into pieces)



Ionic Solids – **Melting Point**

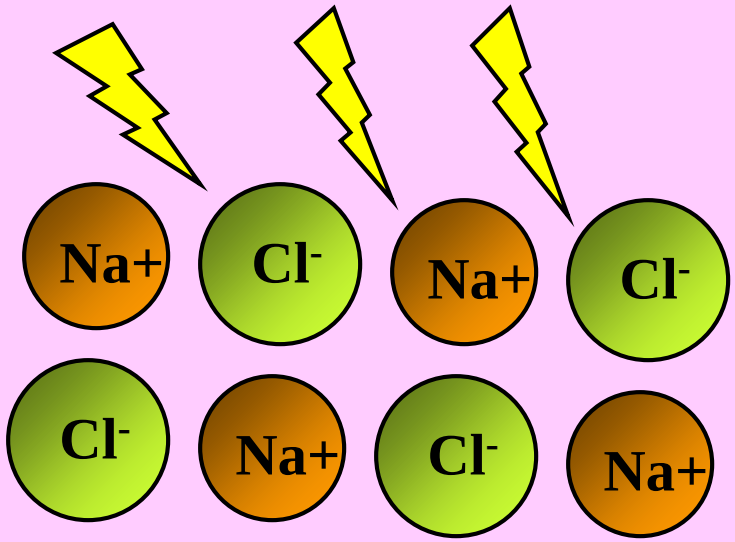
Ionic solids have a high melting point

For example:

1. sodium chloride is an ionic solid

2. Sodium chloride is composed of ions held together by strong directional electrostatic attractions between +ve and –ve ions in a 3-d lattice

3. Because these strong bonds require a large amount of energy to break the ionic solids have a high melting point.



Covalent Network Solids - **structure**

diamond

graphite

Silicon dioxide

All atoms are held together by strong covalent bonds

Diamond is a 3-dimensional covalent network structure where atoms are held together by strong covalent bonds in all planes

Graphite is a covalent network structure that is in 2 dimensional sheets (graphite). Between the layers are free moving electrons from the valance electrons of the carbon atoms.

Silicon dioxide (SiO_2) is a 3-dimensional covalent network structure

Covalent Network (3D) - **Conductivity of diamond**

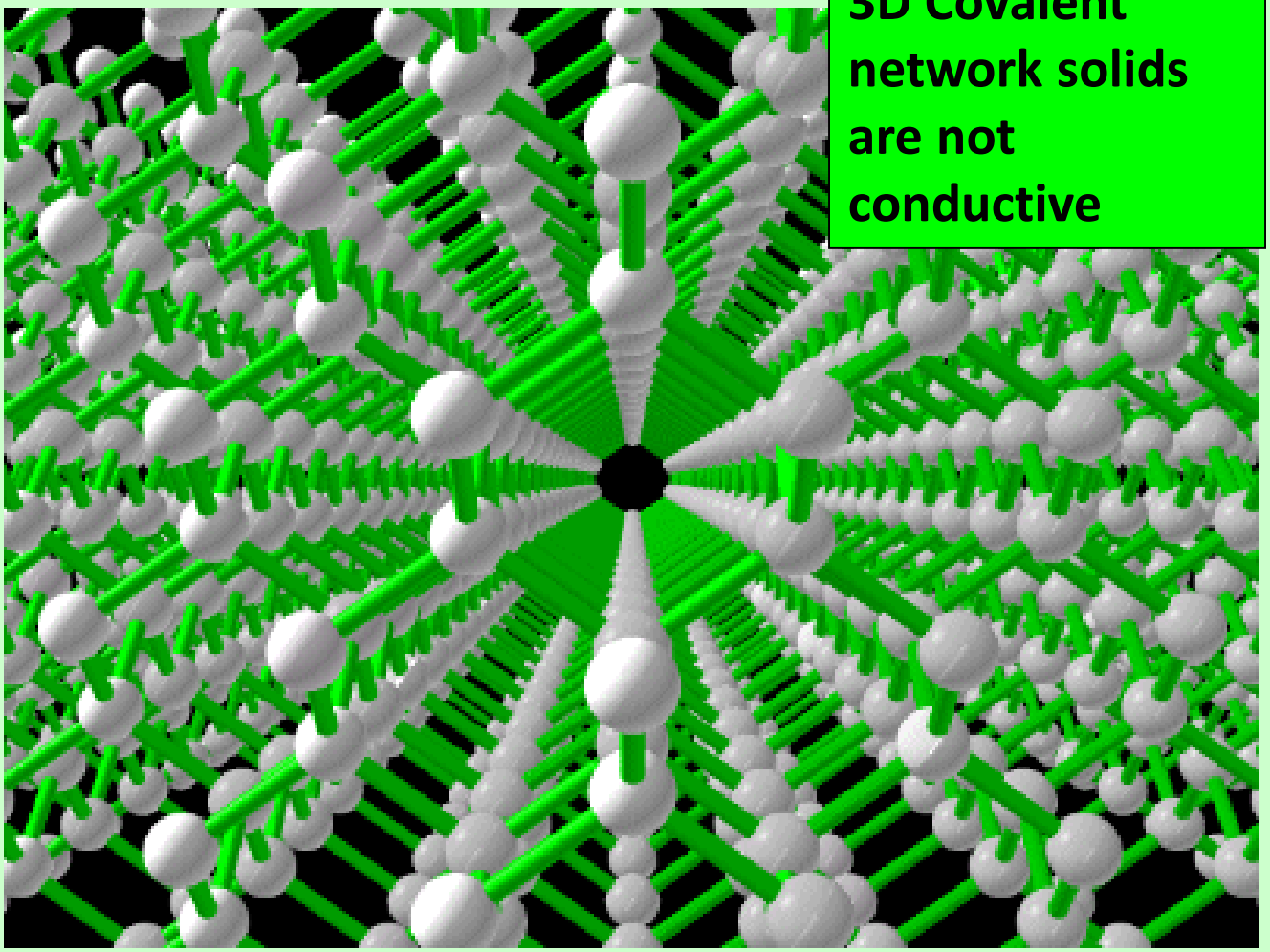
For example:

1. Diamond is a 3-dimensional covalent network structure (diamond)

2. All atoms are held together by strong covalent bonds

3. There is no free moving charged particles

4. Therefore, diamond **cannot conduct electricity**



3D Covalent network solids are not conductive

Covalent Network (2D) - **Conductivity of graphite**

2D Covalent network solids are conductive

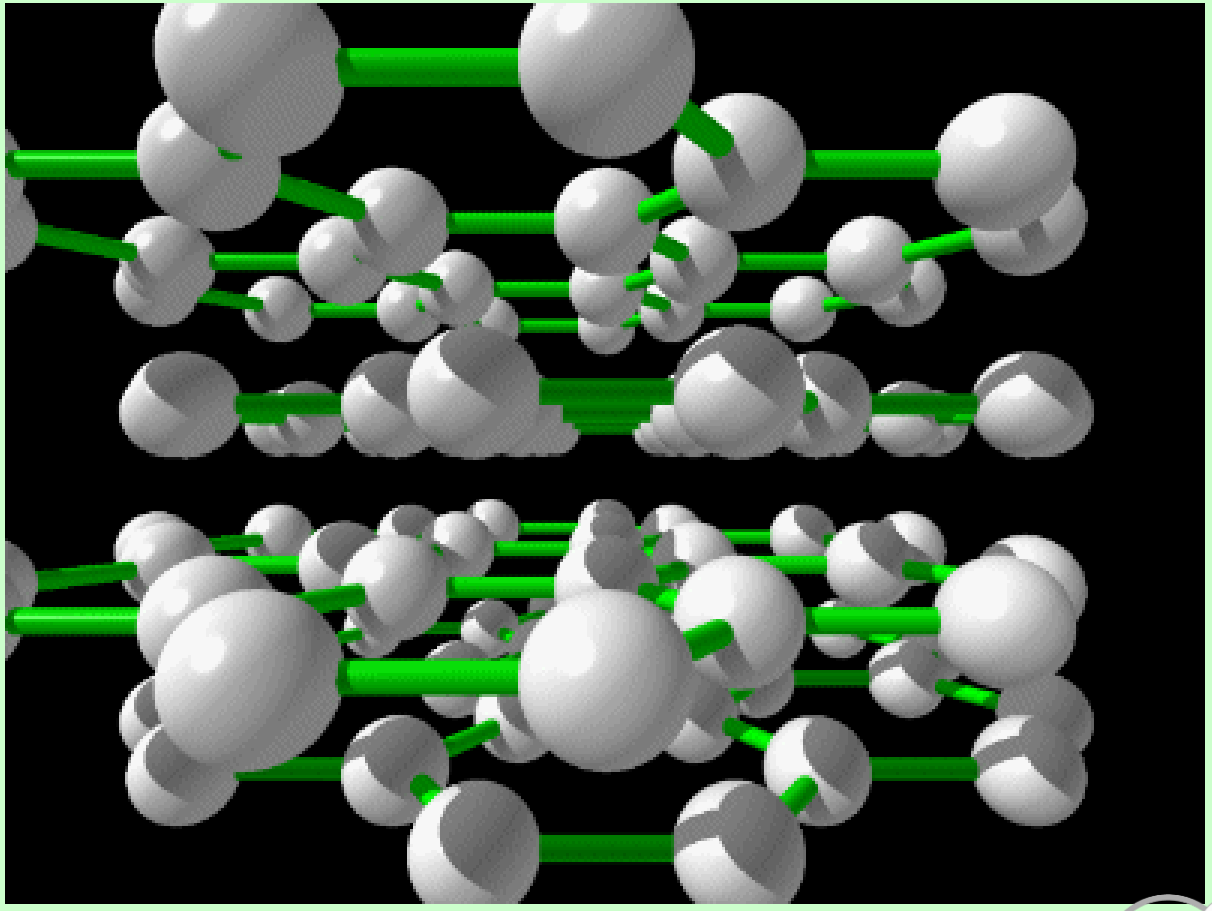
<u>For example:</u>

1. Graphite is a covalent network that is in 2-dimensional sheets

2. Between the layers are free moving electrons from the valance electrons of the carbon atoms.

3. The free moving electrons can carry a current

4. Therefore, graphite **can conduct electricity**



Covalent Network Solids - **Solubility**

Covalent network solids are not soluble

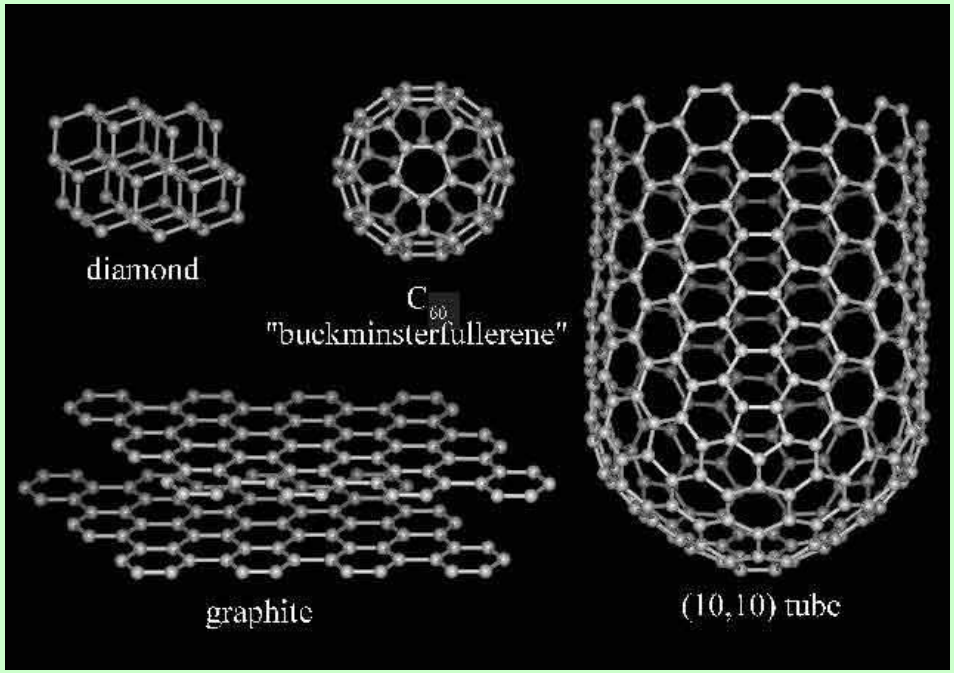
<u>For example:</u>

1. Silicon Dioxide is a 3-dimensional (or 2-dimensional) covalent network structure

2. All atoms are held together by strong covalent bonds

3. These forces require a large amount of energy to break therefore, the electrostatic attractions of water molecules do not have sufficient strength to pull the ions apart

4. Hence, silicon dioxide will not dissolve in water and is insoluble



Covalent Network Solids - **Melting Point**

Covalent network solids have a high melting point

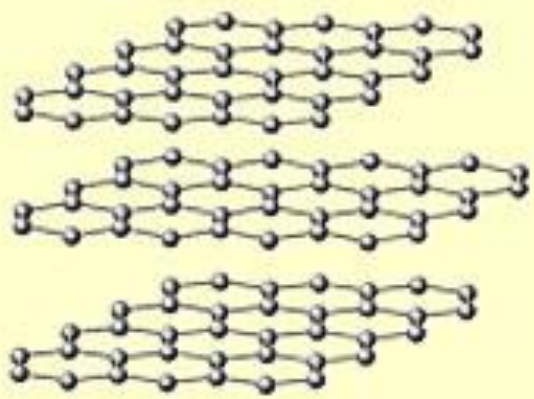
<u>For example:</u>

1 Diamond is a 3-dimensional (or 2-dimensional) covalent network structure

2. All atoms are held together by strong covalent bonds

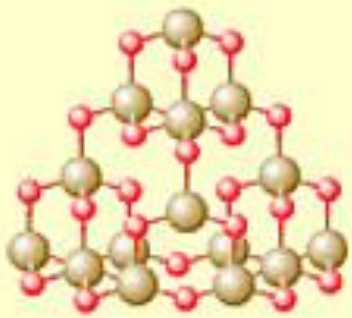
3. These forces require a large amount of energy to break

4. Therefore, diamond has a very high melting point.



graphite

diamond



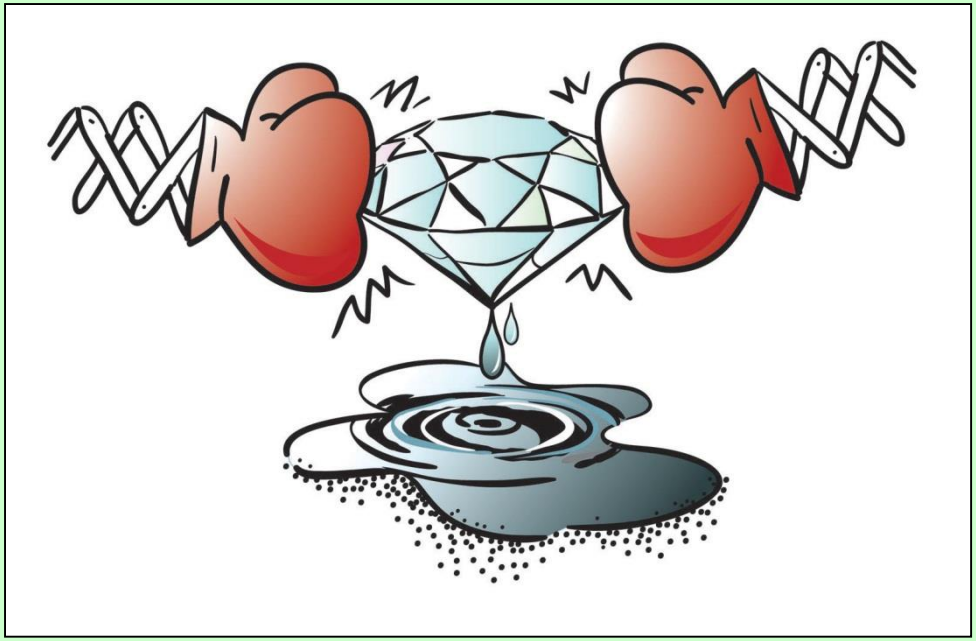
silica

Covalent Network Solids (3-D) - **Hardness**

3-dimensional Covalent network solids are hard

<u>For example:</u>

1. Diamond is a 3-dimensional covalent network structure
2. All atoms are held together by strong covalent bonds
3. These forces require a large amount of energy to break
4. Therefore diamond is very hard.



Covalent Network Solids (2-D) - **Hardness**

2-dimensional Covalent network solids are soft

<u>For example:</u>

1. Graphite is a 2-dimensional covalent network structure

2. Atoms are held together by strong covalent bonds in 2-dimensional layers

3. However, the attractive forces holding the layers together are very weak and are broken easily, so the layers easily slide over one another, but the attraction is not strong enough to hold the layers together

4. Therefore, graphite is considered soft.

Solids Summary

Name of solid substance	Type of particle in solid	Attractive force broken when solid melts	Attractive force between particle – weak or strong (hardness)	Relative melting point	solubility	Electrical conductivity	Malleable
Molecular	molecules	Weak inter molecular	weak	low	Yes if polar No if non-polar	no	no
Metallic	atoms	Metallic bonding	strong	high	no	yes	yes
Ionic	ions	Electrostatic Ionic bonding	strong	high	Yes	Only if molten or in solution	No - brittle
Covalent Network 3-D	atoms	Covalent bonding	strong	high	no	no	no
Covalent Network 2-D	atoms	Covalent bonding	Strong (but weak between layers)	high	no	yes	no

Property	Type of Solid			
	Molecular	Covalent network	Ionic	Metallic
Solubility in Water				
Electrical Conductivity				
Melting Point				
Hardness				

NCEA 2013 Solids

Merit Question

Question 2a: Complete the table below by stating the type of substance, the type of particle, and the bonding (attractive forces) between the particles for each of the substances.

Substance (for example)	Type of substance	Type of particle	Attractive forces between particles
$C_{(s)}$ Graphite	Covalent network	Atom	Covalent (and weak intermolecular forces)
$Cl_{2\ (s)}$ chlorine	Molecular	Molecules	Weak intermolecular forces
$CuCl_{2(s)}$ copper chloride	Ionic	Ion	Ionic bonds / electrostatic attraction
$Cu_{(s)}$ copper	Metal	Atom / cations and electrons	Metallic bonds / electrostatic attraction

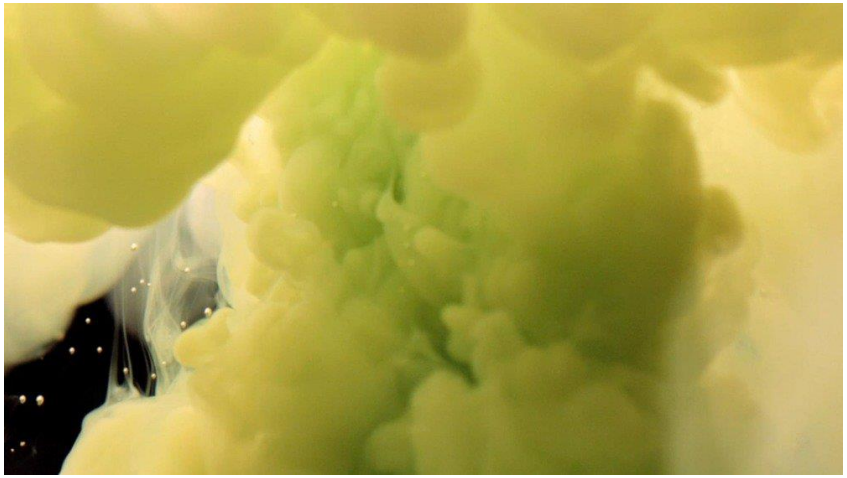
This chart needs to be learnt

There will not necessarily be one example for each group but information from this chart MUST be used in following questions about solids

NCEA 2013 Solids

Merit Question

Question 2b: Explain why chlorine is a gas at room temperature, but copper chloride is a solid at room temperature.
In your answer, you should refer to the particles and the forces between the particles in **both** substances.



Chlorine is a molecular substance composed of chlorine molecules held together by weak intermolecular forces. The weak intermolecular forces do not require much heat energy to break, so the boiling point is low (lower than room temperature); therefore chlorine is a gas at room temperature.

Copper chloride is an ionic substance. It is composed of a lattice of positive copper ions and negative chloride ions held together by electrostatic attraction between these positive and negative ions. These are strong forces, therefore they require considerable energy to disrupt them and melt the copper chloride; hence copper chloride is a solid at room temperature.

NCEA 2013 Solids - (PART ONE)

Excellence Question

Question 2b (ii) : Using your knowledge of structure and bonding, explain why, although both graphite and copper are good conductors of electricity, copper is suitable for electrical wires, but graphite is not.

For a substance to conduct electricity, it must have charged particles which are free to move.

Graphite is a covalent network solid composed of layers of C atoms covalently bonded to three other C atoms. The remaining valence electron is delocalised (ie free to move) between layers; therefore these delocalised electrons are **able to conduct** electricity.

Copper is a metallic substance composed of copper atoms packed together. Valence electrons are loosely held and are attracted to the nuclei of the neighbouring Cu atoms; ie the bonding is non-directional. These delocalised valence electrons are **able to conduct** an electrical current.

NCEA 2013 Solids - (PART ONE)

Excellence Question

Question 2b (ii) : Using your knowledge of structure and bonding, explain why, although both graphite and copper are good conductors of electricity, copper is suitable for electrical wires, but graphite is not.

For a substance to be made into wires, it needs to be stretched or drawn out without breaking.

In **graphite**, the attractive forces holding the layers together are very weak and are broken easily, so the layers easily slide over one another, but the attraction is not strong enough to hold the layers together and allow it to be drawn into wires or although the layers can slide due to weak forces, if graphite was to be made into a wire the very strong covalent bonds within the layers would have to be broken.

Copper metal is malleable and can easily be drawn into wires since, as it is stretched out, the non-directional metallic bonding holds the layers together, allowing it to be stretched without breaking.

NCEA 2014 Solids

Achieved Question

Question 2a: Complete the table below by stating the type of substance, the type of particle, and the type of bonding (attractive forces) between the particles for each of the two substances. Mg (magnesium) and I_2 (iodine)

This chart needs to be learnt

Answer 2a:

	Type of substance	Type of particle	Attractive forces between particles
Mg	Metallic	Atoms / cations and electrons	Metallic bonds / electrostatic attraction between positive ion (cation) and electron
I_2	Molecular	Molecules	Intermolecular forces

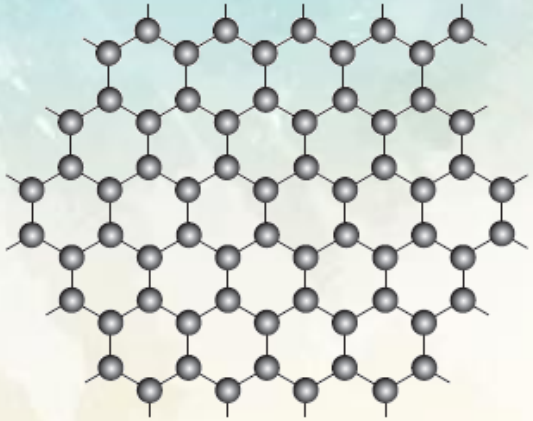
There will not necessarily be one example for each group but information from this chart MUST be used in following questions about solids

Identify Metal first then Ionic (metal + non-metal), next Covalent network (C or SiO_2) and then molecular (non-metal + non-metal)

NCEA 2014 Solids

Excellence Question

Question 2b: Graphene is a new 2-dimensional material made of carbon atoms. Graphene can be described as a 'one-atom-thick' layer of graphite. A diagram of graphene and two of its properties is shown below. Use your knowledge of structure and bonding to explain the two properties of graphene given above.



Properties of graphene:

Melting point: very high

Electrical conductivity: excellent

Answer 2b: Graphene has strong covalent bonds. Because the covalent bonds are strong / there are a large number of covalent bonds, it requires a lot of energy to break these bonds, and therefore the melting point is high.

Each carbon atom is bonded to only three other carbon atoms. Therefore each carbon atom has free / delocalised /valence electron(s), to conduct electricity.

.

NCEA 2014 Solids - (Part ONE)

Excellence Question

Question 2c: Solid Mg and I_2 were tested for three physical properties. The table below shows the results of the tests. Use your knowledge of structure and bonding to explain the results of the tests.

Refer back to particle chart

	Physical property		
Substance tested	Ductile	Soluble in cyclohexane (non-polar solvent)	Conducts electricity
Mg	yes	no	yes
I_2	no	yes	no

Answer 2c: Magnesium atoms are held together in a 3–D lattice by metallic bonding in which valence electrons are attracted to the nuclei of neighbouring atoms.
Iodine molecules are held together by weak intermolecular forces.

REMEMBER: (Properties are limited to hardness (including malleability and ductility), electrical conductivity, melting and boiling points and solubility)

NCEA 2014 Solids - (Part TWO)

Excellence Question

Question 2c: Solid Mg and I_2 were tested for three physical properties. The table below shows the results of the tests. Use your knowledge of structure and bonding to explain the results of the tests.

Refer back to particle chart

	Physical property		
Substance tested	Ductile	Soluble in cyclohexane (non-polar solvent)	Conducts electricity
Mg	yes	no	yes
I_2	no	yes	no

Answer 2c: Ductility

The attraction of the Mg atoms for the valence electrons is not in any particular direction **(non-directional)** ; therefore Mg atoms can move past one another without disrupting the metallic bonding, therefore Mg is **ductile**.

The attractions between iodine molecules are **directional**. If pressure is applied the repulsion between like-charged ions will break the solid (brittle) , therefore I_2 is **not ductile**.

NCEA 2014 Solids - (Part THREE)

Excellence Question

	Physical property		
Substance tested	Ductile	Soluble in cyclohexane (non-polar solvent)	Conducts electricity
Mg	yes	no	yes
I_2	no	yes	no

Refer back to particle chart

Answer 2c: Dissolving in cyclohexane

Magnesium **does not dissolve in cyclohexane** because cyclohexane molecules are not attracted to the magnesium atoms in the metallic lattice.

Iodine **is soluble**, as iodine is a non-polar molecule. The iodine molecules and cyclohexane molecules form weak intermolecular attractions.

Electrical conductivity

Valence electrons of Mg atoms are free to move throughout the structure. This means that magnesium can conduct electricity. There are free moving charge particles.

Iodine does not conduct electricity as it does not contain delocalised electrons (free moving charged particles). Molecules are neutral compounds.

NCEA 2015 Solids

Merit Question

Question 3a: Complete the table below by stating the type of solid, the type of particle, and the attractive forces between the particles in each solid.

This chart needs to be learnt

Answer 3a:

Substance	Type of Substance	Type of particle	Attractive forces between particles
Cu(*s*)	metal / metallic	Atom (or cation and delocalised electrons)	metallic bond
PCl_3(*s*)	molecular	molecule	intermolecular (forces)
SiO_2(*s*)	covalent network	atom	covalent bond
KCl(*s*)	ionic	ion	ionic bond

There will not necessarily be one example for each group but information from this chart MUST be used in following questions about solids

Identify Metal first then Ionic (metal + non-metal), next Covalent network (C or SiO_2) and then molecular (non-metal + non-metal)

NCEA 2015 Solids

Excellence Question

Question 3b: Phosphorus trichloride, PCl_3, is a liquid at room temperature, and does not conduct electricity.
Explain these two observations in terms of the particles, structure, and bonding of PCl_3.

Refer back to particle chart

Answer 3b: Phosphorus trichloride, PCl_3, is a **molecular solid**, made up of non-metal phosphorus and chlorine atoms **covalently** bonded together. The molecules are held together by **weak intermolecular forces**. Since these forces are weak, not much energy is required to overcome them, resulting in low melting / boiling points. (In the case of PCl_3, its melting point is lower than, and its boiling point is higher than room temperature, so it is liquid.)

PCl_3 does not contain free moving ions nor any delocalised / free moving valence electrons, meaning PCl_3 **does not contain any charged particles**. Since free moving charged particles are required to carry electrical current, PCl_3 is unable to conduct electricity.

REMEMBER: (Properties are limited to hardness (including malleability and ductility), electrical conductivity, melting and boiling points and solubility)

NCEA 2015 Solids - (Part ONE)

Excellence Question

Question 3c: Consider each of the solids copper, Cu, silicon dioxide, SiO_2, and potassium chloride, KCl.
Complete the table below by identifying which of these solids have the listed physical properties:

Refer back to particle chart

Physical properties	Solid
The solid is insoluble in water and is malleable.	**METAL - copper**
The solid is soluble in water and is not malleable.	**IONIC – potassium chloride**
The solid is insoluble in water and is not malleable.	**COVALENT NETWORK – silicon dioxide**

REMEMBER: Draw up a quick chart for each physical property and each solid

NCEA 2015 Solids - (Part TWO)

Excellence Question

Answer 3c: Justify TWO of your choices in terms of the particles, structure, and bonding of these solids. You may use diagrams in your justification.

Refer back to particle chart

State the properties and solid type

Cu is **insoluble in water and malleable**.

1. Copper is a metal made up of an array of atoms (or ions) held together by **non-directional forces** between the positive nuclei of the atoms and the delocalised / free moving valence electrons.
2. There is no attraction between the copper atoms and the (polar) water molecules, therefore Cu is **insoluble in water**.
3. Since the attractive forces are non-directional, when pressure is applied, the Cu atoms can move past each other to change shape without the bonds breaking, so Cu **is malleable**. (Note – labelled diagrams can provide replacement evidence).

REMEMBER: discuss solid structure first then each of the properties. You need both properties for full marks.

NCEA 2015 Solids - (Part TWO)

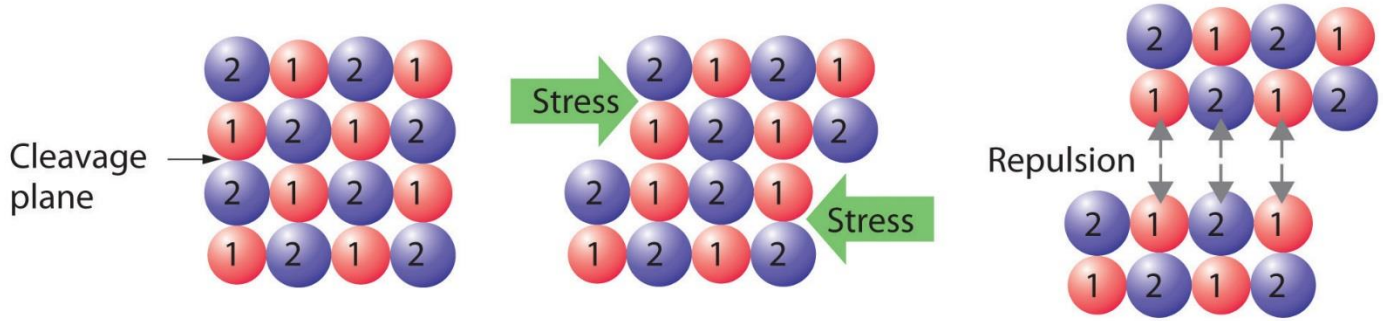
Excellence Question

Question 3c: Justify TWO of your choices in terms of the particles, structure, and bonding of these solids. You may use diagrams in your justification.

Refer back to particle chart

KCl is **soluble in water and not malleable**.

1. KCl is made up of positive K^+ ions, and negative Cl^- ions, ionically bonded in a 3D lattice.
2. When added to water, polar water molecules form electrostatic attractions with the K^+ and Cl^- ions. The partial negative charge, δ^-, on oxygen atoms in water are attracted to the K^+ ions and the partial positive, δ^+, charges on the H's in water are attracted to the Cl^- ions, causing KCl to **dissolve in water**.
3. KCl is **not malleable** because if pressure is applied to an ionic lattice, it forces ions with the same charge next to each other; they repel each other and break the structure.



NCEA 2015 Solids - (Part TWO)

Excellence Question

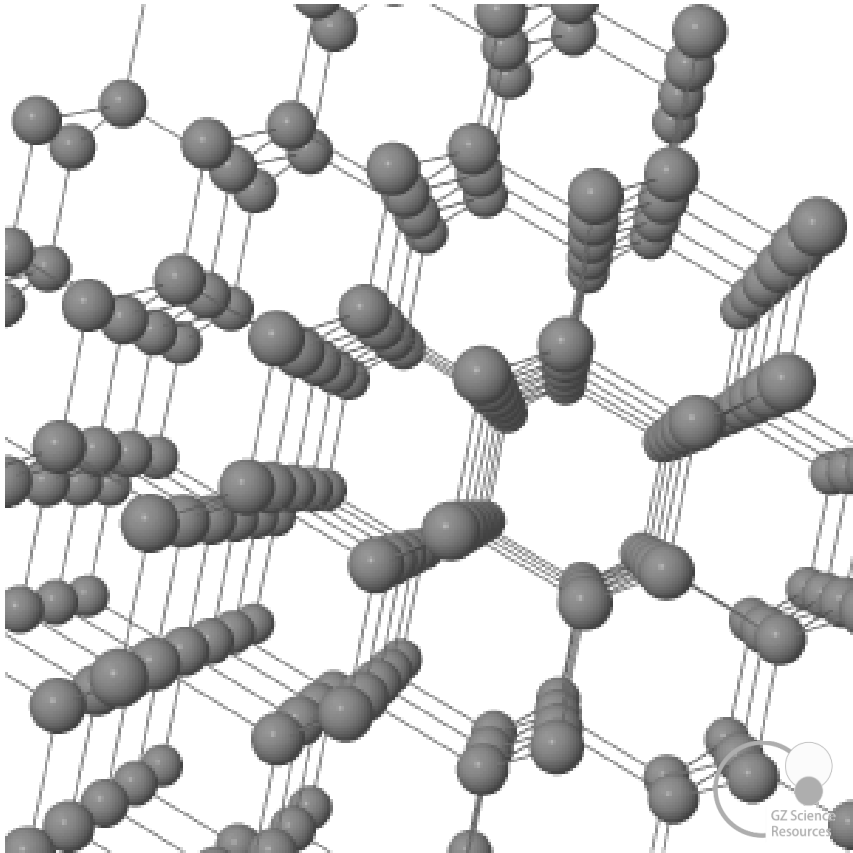
Question 3c: Justify TWO of your choices in terms of the particles, structure, and bonding of these solids. You may use diagrams in your justification.

Refer back to particle chart

SiO_2 is insoluble in water and not malleable.

1. SiO_2 is a **covalent network** made up of atoms covalently bonded together in a 3D lattice structure.
2. (Covalent bonds are strong), Polar water molecules are not strong / insufficiently attracted to the Si and O atoms, therefore **SiO_2 is insoluble** in water.
3. SiO_2 is not malleable because if pressure is applied, the directional / strong covalent bonds have to be broken before the atoms can move.

(Note - labelled diagrams can provide replacement evidence).



NCEA 2016 Solids

Achieved Question

Question 2a: Complete the table below by stating the type of substance, the type of particle, and the attractive forces between the particles in the solid for each substance.

This chart needs to be memorised

Answer 2a:

Substance	Type of substance	Type of particle	Attractive forces between particles
$ZnCl_2(s)$ (zinc chloride)	ionic	ions	ionic
C(s) (graphite)	covalent network	atoms	covalent
$CO_2(s)$ (carbon dioxide / dry ice)	molecular	molecules	intermolecular

There will not necessarily be one example for each group but information from this chart MUST be used in following questions about solids

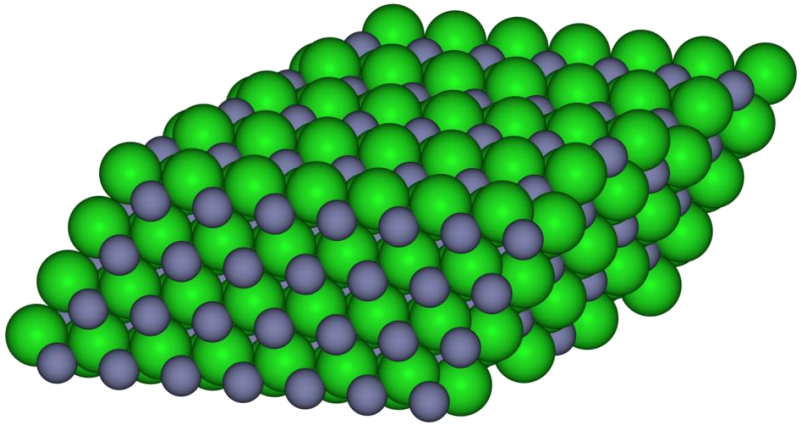
Identify Metal first then Ionic (metal + non-metal), next Covalent network (C or SiO_2) and then molecular (non-metal + non-metal)

NCEA 2016 Solids - (PART ONE)

Excellence Question

Question 2b : Carbon (graphite) conducts electricity when it is solid, whereas zinc chloride, ZnCl2, will not conduct electricity when solid, but will conduct when molten. Justify this statement in terms of the particles, structure, and bonding for both substances.

For a substance to conduct electricity, it must have charged particles which are free to move.



Graphite is a covalent network solid composed of layers of C atoms covalently bonded to three other C atoms. The remaining valence electron is delocalised (ie free to move) between layers; therefore these delocalised electrons are **able to conduct** electricity.

$ZnCl_2$ is an ionic compound that cannot conduct electricity when solid because the ions (charged particles) are fixed in place in a 3D lattice structure and unable to move. When molten, the ionic bonds between the ions break, so the ions are free to move in the molten liquid. With charged particles / ions free to move, $ZnCl_2$ can then conduct electricity.

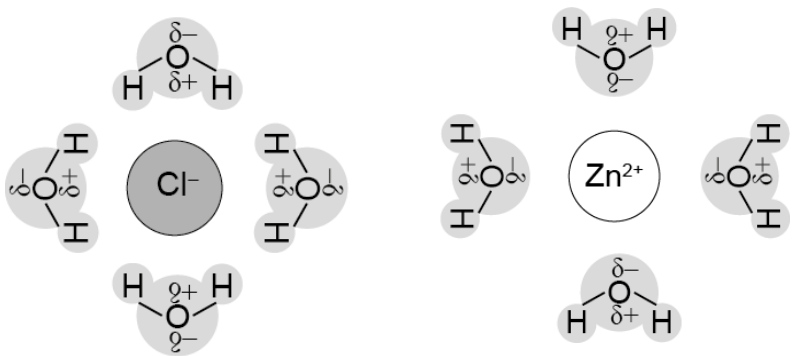
NCEA 2016 Solids - (PART TWO)

Excellence Question

Question 2c : Solid zinc chloride, $ZnCl_{2(s)}$, is soluble in water. Dry ice, $CO_{2(s)}$, is not readily soluble in water. Justify these statements in terms of the particles, structure, and bonding of these substances.

For a substance to be soluble Polar water molecules need to be insufficiently attracted to the particles of the substance placed in

Polar water molecules attract the ions in **zinc chloride's** 3-D lattice strongly enough to separate and dissolve them. The negative charges on the oxygen ends of the water molecules are attracted to the positive Zn^{2+} ions, and the positive hydrogen ends of the water molecules are attracted to the negative Cl^- ions, forming hydrated ions that can spread out through the solution. Therefore it is soluble.



The polar water molecules are unable to interact with the non-polar **carbon dioxide** molecules strongly enough to break the intermolecular forces between the carbon dioxide molecules. Therefore it is not soluble.

NCEA 2017 Solids - (PART ONE)

Merit Question

Question 3a : Complete the table below by stating the type of solid, the type of particle, and the type of bonding (attractive forces) between the particles in each solid.

Solid	Type of solid	Type of particle	Attractive forces between particles
$Al(s)$ (Aluminium)	*metal / metallic*	*atoms (or cations and delocalised valence electrons)*	*metallic (bonds)*
$MgCl_2(s)$ (Magnesium chloride)	*ionic compound*	*ions*	*ionic (bonds)*
$S_8(s)$ (Sulfur)	*molecular*	*molecules*	*intermolecular (bonds)*

NCEA 2017 Solids - (PART TWO)

Excellence Question

Question 3b: Circle the substance which has the lowest melting point.

$Al_{(s)}$ $MgCl_{2(s)}$ $S_{8(s)}$

Justify your choice, referring to the attractive forces between the particles of ALL three substances.

Sulfur has the lowest melting point.
Sulfur is a molecular substance with weak intermolecular forces between the molecules.
These forces do not require much energy to overcome, so they will break at lower temperatures, giving sulfur a lower melting point.
Al is a metal with strong metallic bonds. These attractions require a lot of energy to overcome, so the melting point is higher than sulfur's melting point.
$MgCl_2$ is an ionic compound with strong ionic bonds between the cations and anions.
These bonds also require a lot of energy to overcome, so the melting point is also higher than sulfur's melting point.

NCEA 2017 Solids - (PART THREE)

Excellence Question

Question 3c: Circle the substance which is malleable.

$Al_{(s)}$ $MgCl_{2(s)}$ $S_{8(s)}$

Justify your choice by referring to the structure and bonding of your chosen substance. You may include a diagram or diagrams in your answer.

Aluminium is malleable.
Aluminium is a metal made up of atoms / cations in a sea of electrons which are held together by non-directional metallic bonds in a (3D) lattice. The metallic bonds are non-directional as the (bonding) electrons are delocalised across the lattice / shared by many atoms. When a force (or pressure) is applied the atoms / layers can move without breaking / disrupting these non-directional bonds thus the structure can change shape without breaking the lattice.

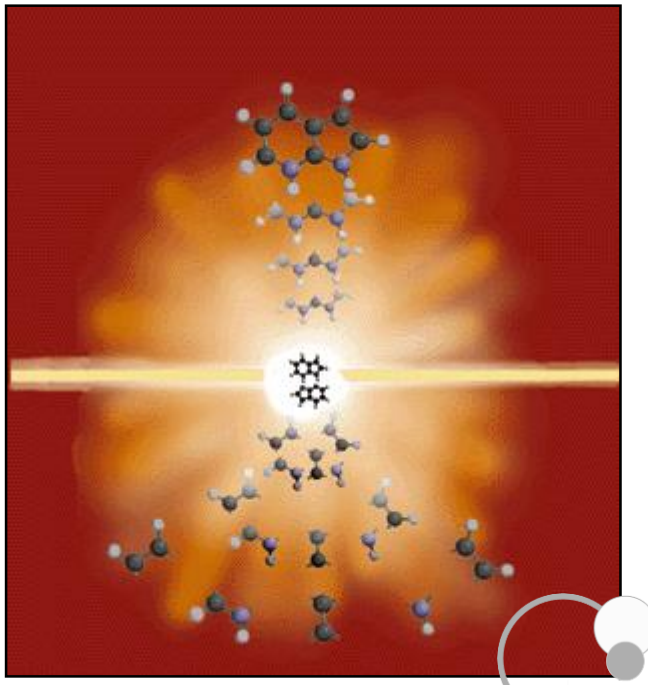
Enthalpy and Enthalpy Change ΔH

Enthalpy (or Heat Content) is the energy in a substance due to kinetic energy of particles and potential energy in chemical bonds

Enthalpy change ΔH is the difference in enthalpy of products H_P and reactants H_R

$$\Delta H = H_P - H_R$$

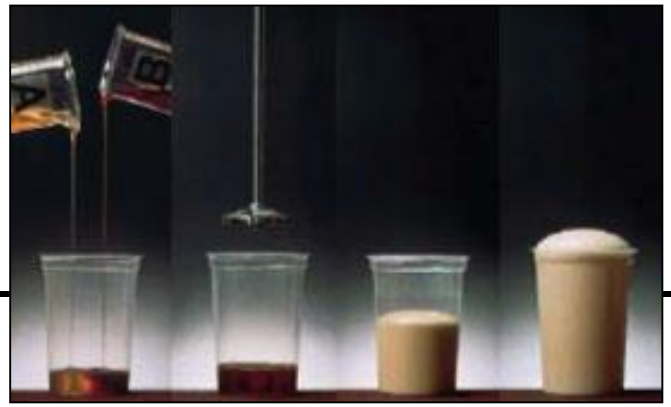
The unit for Enthalpy is kilojoules (kJ)



Enthalpy Change

H_P (products) and H_R (reactants) cannot be measured directly.

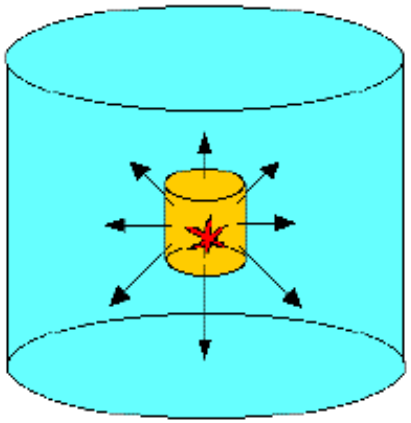
We can measure Enthalpy change (ΔH) by measuring energy;

Released to surroundings (Exothermic Reactions)	Absorbed from surroundings (Endothermic Reactions)
	

Exothermic Reactions

These are reactions where **heat energy** is **released** into the surroundings.

Surroundings gain heat energy.
(increase in temperature)



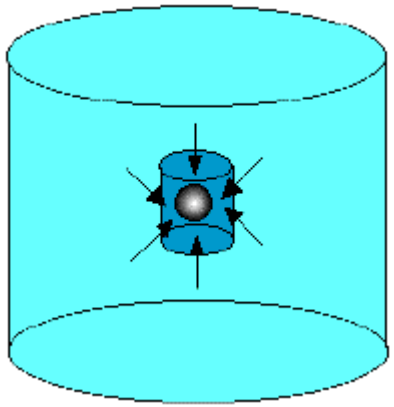
Products will have less energy than reactants.

ΔH is NEGATIVE (-)

Endothermic Reactions

These are reactions where **heat energy** is **absorbed** from the surroundings.

Surroundings lose heat energy.
(Decrease in temperature)



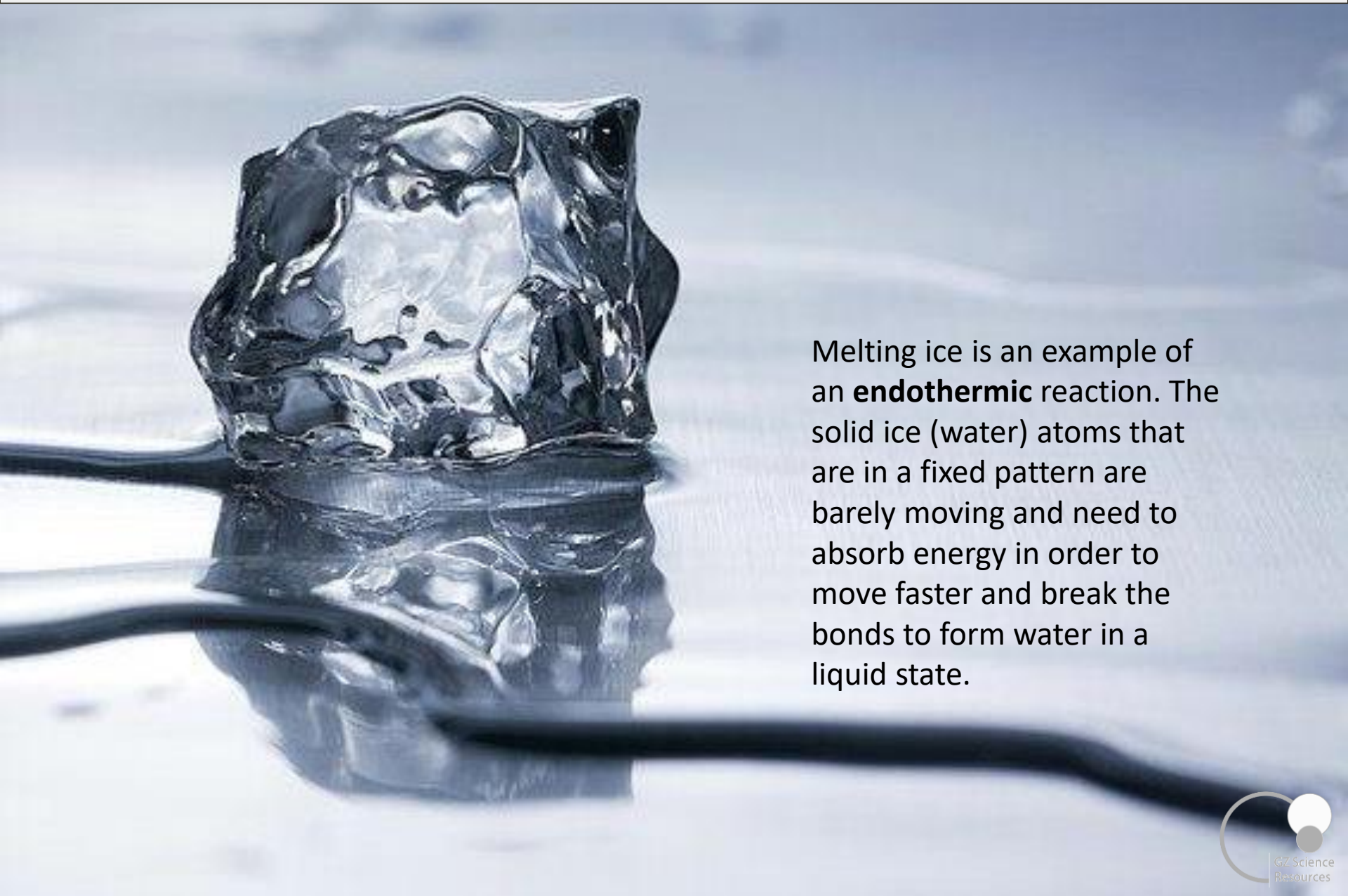
Products will have more energy than reactants.

ΔH is POSITIVE (+)

Exothermic reactions

Any combustion reaction is **exothermic**. The bonds holding the atoms of fuel molecules together (usually consisting of carbon and hydrogen atoms) release a lot of energy in the form of light and heat when they are broken. The total energy holding the bonds together in the products are less than the total energy in the reactions and the difference is released.

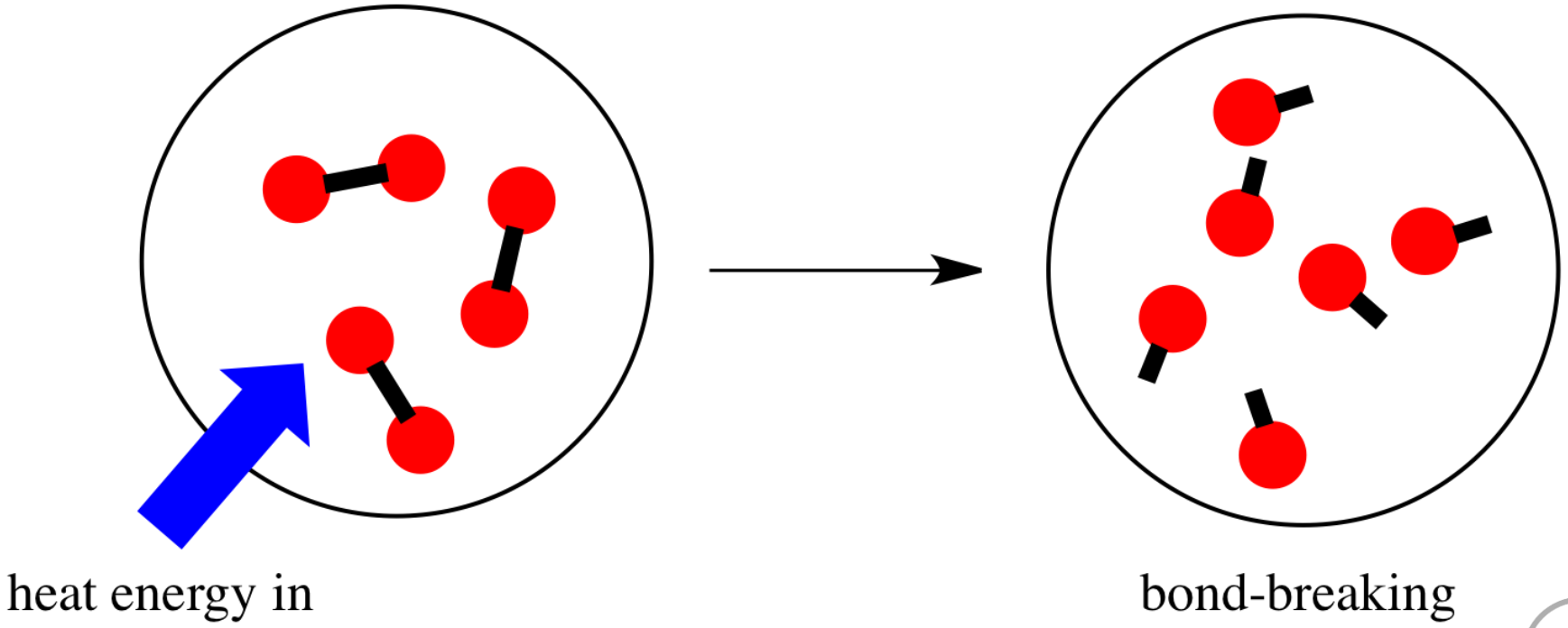
Endothermic reactions



Melting ice is an example of an **endothermic** reaction. The solid ice (water) atoms that are in a fixed pattern are barely moving and need to absorb energy in order to move faster and break the bonds to form water in a liquid state.

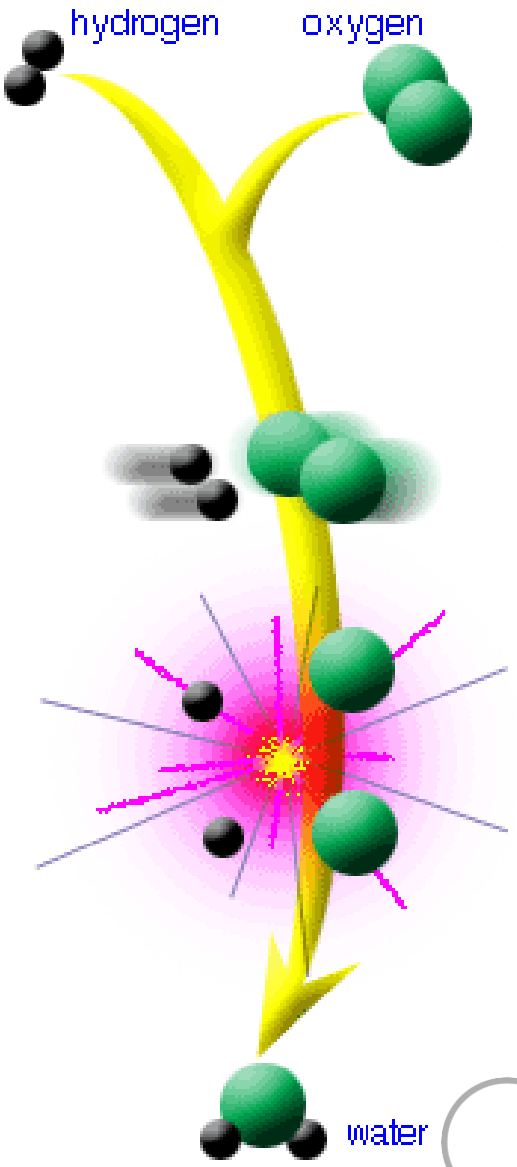
Breaking Bonds - endothermic

Bonds holding atoms and molecules together require the input of energy in order to break them apart therefore breaking of bonds is an endothermic reaction. The input of energy (usually light or heat energy) cause the atoms and molecules to move faster and 'pull away' from each other. Each type of bond has its own specific amount of energy, called bond enthalpy measured in kJ, required to break its bond.



Forming Bonds - exothermic

Bonds forming between atoms and molecules release energy therefore bond forming is an exothermic reaction. Bonds are formed to form a stable molecule.

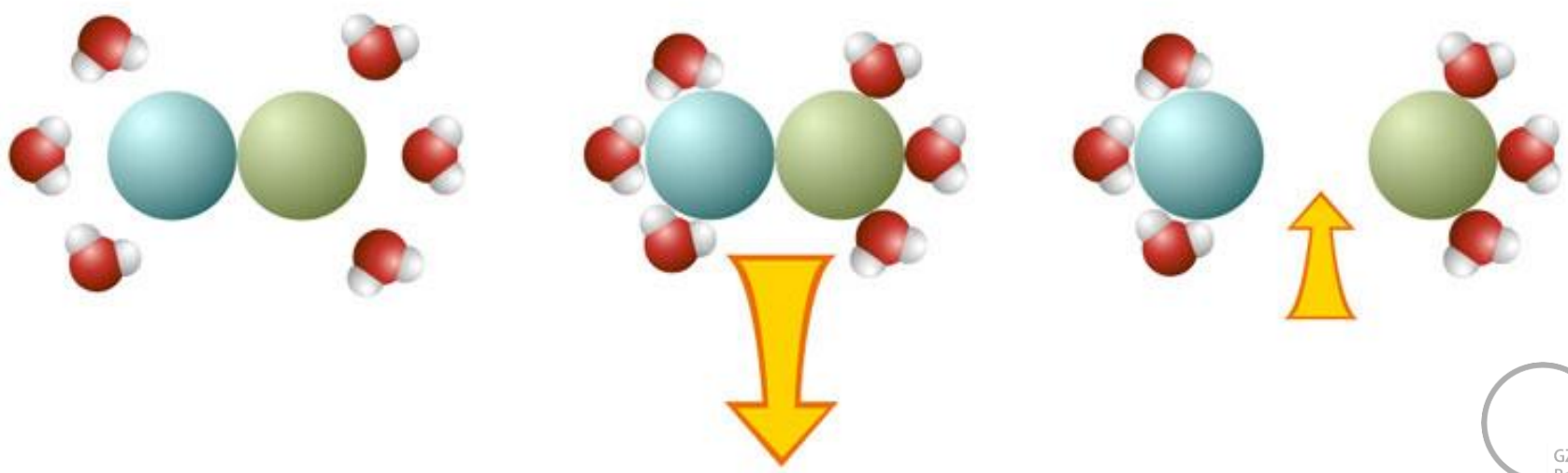


If more energy is required to break the bonds of the reactants than released when the bonds of the products are form then the overall reaction is **endothermic**.
If less energy is required to break the bonds than is released when the bonds of the products are formed then the overall reaction is **exothermic**.

Enthalpy in Dissolving

If more energy is released when water bonds to the solute than it takes to separate the solute, the dissolving is **exothermic** and the temperature increases. An example is adding a strong acid (such as sulfuric acid) or base (such as sodium hydroxide)

However, for some substances dissolving the reaction is **endothermic**, for example Potassium chloride in water. More energy is needed to break the bonds between this ionic salt than released when new bonds are formed with the ions and water molecules.



NCEA 2013 Endothermic and Exothermic

Merit Question

Question 3a: Dissolving ammonium nitrate in a beaker containing water can be represented by the following equation:

$NH_4NO_{3(s)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$ $\Delta_r H° = 25.1$ kJ mol^{-1}

Give the term below that best describes this process and give the description that best describes what you would observe happening to the beaker during this process.

Answer 3a: Endothermic - Gets colder

The process is endothermic since the enthalpy change (ΔrH°) is positive, which indicates that energy is absorbed by the system as the ammonium nitrate dissolves. Since heat energy is absorbed by the system from the surroundings (water & beaker), the water or beaker will get cooler as they lose heat energy.

Heat absorbed is always endothermic

Question 3b: Glucose is an important source of energy in our diet. The equation below shows the combustion of glucose to form carbon dioxide and water.

$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(\ell)}$ $\Delta_r H° = -2\,820$ kJ mol^{-1}

Give the term below that best describes this process and give a reason

Answer 3b: Exothermic

The reaction is exothermic because the enthalpy change (ΔrH°) is negative; indicating that heat energy is produced during the reaction.

NCEA 2014 Endothermic and Exothermic

Achieved Question

Question 3a (i): When solid sodium hydroxide is added to water, the temperature increases.

- Identify the term that best describes this reaction
- Give a reason for your choice

Heat released is always exothermic

Answer 3a (i) : Exothermic, as the temperature increases, which shows energy is being released.

Question 3a(ii): The freezing of water to form ice can be represented by the following equation.

$H_2O_{(l)} \rightarrow H_2O_{(s)}$

- Identify the term that best describes this reaction
- Give a reason for your choice

Answer 3a(ii): Exothermic, weak intermolecular attractions form between the water molecules, this releases energy.

NCEA 2015 Endothermic and Exothermic

Achieved Question

Question 2a: Hand warmers contain a supersaturated solution of sodium ethanoate which, when activated, crystallises and releases heat.

- Identify the term that best describes this reaction
- Give a reason for your choice

Heat released is always exothermic

Answer 2a: Exothermic because the temperature of the solution increases / heat is released / particles slow down / bonds are formed

Question 2b(i): Glucose is made in plants during photosynthesis when carbon dioxide gas, $CO_{2(g)}$, and water, $H_2O_{(l)}$, react to produce glucose, $C_6H_{12}O_{6(aq)}$, and oxygen gas, $O_{2(g)}$. The photosynthesis reaction can be represented by the following equation:

$6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}$ $\Delta_r H° = +2803$ kJ mol^{-1}

Answer 2b(i): Endothermic because the $\Delta_r H^o$ value is positive / it uses the sun's energy

Always attempt these questions as there is only 2 possible answers

NCEA 2016 Endothermic and Exothermic

Achieved Question

Question 1a Instant cold packs are useful for treating sports injuries on the field. They contain salts such as ammonium nitrate, NH_4NO_3. When the packs are activated, the salt dissolves in water, causing the temperature to decrease.

- Identify the term that best describes this reaction
- Give a reason for your choice

Heat absorbed is always endothermic

Answer 1a : Endothermic

The temperature decreased OR heat / energy has been absorbed.

Question 1b: The equation for hydrating anhydrous copper sulfate is as follows:

$CuSO_{4(s)} + 5H_2O_{(l)} \rightarrow CuSO_4.5H_2O_{(s)}$ $\Delta_r H^{\circ} = -78.2$ kJ mol^{-1}

- Identify the term that best describes this reaction
- Give a reason for your choice

Answer 1b: Exothermic.

The enthalpy of the reaction is negative / energy has been released.

NCEA 2016 Endothermic and Exothermic

Excellence Question

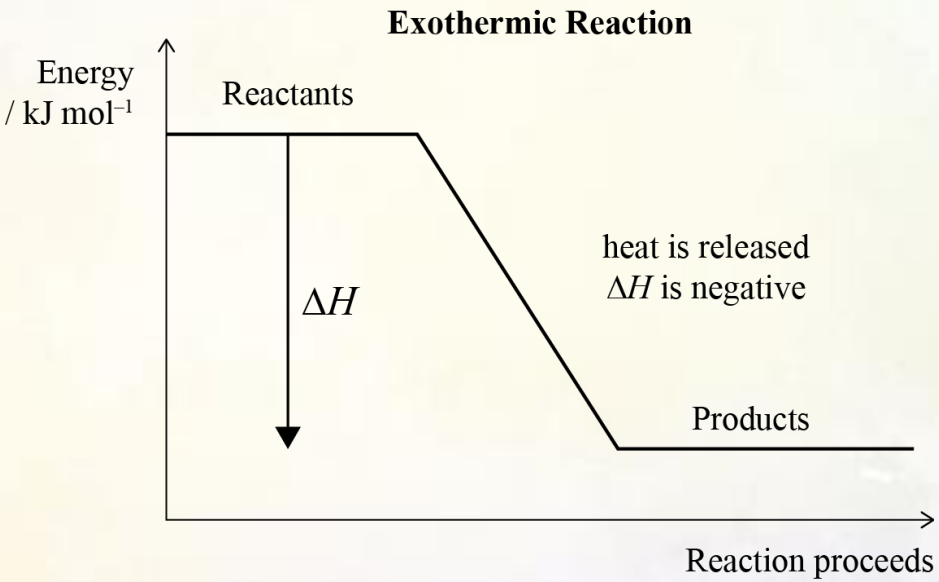
Question 1c (i): Pentane, C_5H_{12}, is a liquid at room temperature. It evaporates at 36.1°C in an endothermic process.
(i) Explain why the evaporation of pentane is an endothermic process.

Solid →liquid →gas is always endothermic

Answer 1c(i) : Energy is required to change pentane from a liquid to a gas. The energy / heat is used to break weak intermolecular forces / bonds / attraction between pentane molecules.

Question 1c(ii) : Draw, including labels, the energy diagram for the combustion of pentane, $C_5H_{12(l)}$.
Pentane combustion: $C_5H_{12(l)} + 8O_{2(g)} \rightarrow 5CO_{2(g)} + 6H_2O_{(l)}$ $\Delta_r H^{\circ} = -3509$ kJ mol^{-1}
Include in your diagram the reactants, products, and change in enthalpy.

Answer 1c(ii)



NCEA 2017 Endothermic and Exothermic

Merit Question

Question 1a: When solid calcium chloride, $CaCl_{2(s)}$, reacts with water, the temperature increases.
Which term that best describes this reaction.

Temperature increase is always exothermic

Exothermic
The temperature increased / energy or heat has been released into the surroundings /
energy is lost from the substance ($CaCl_2$)

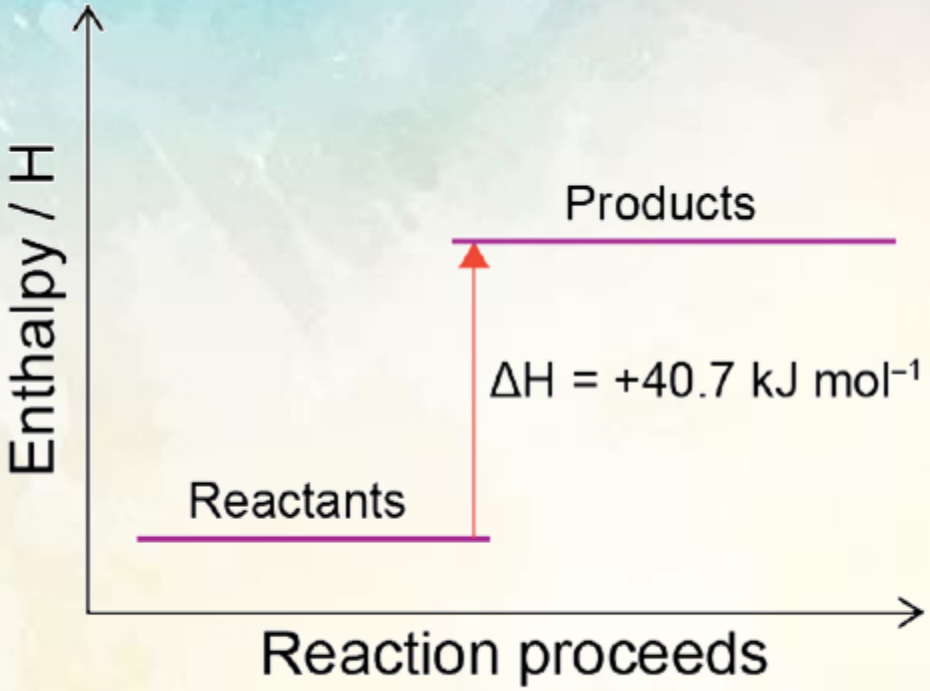
Question 1b (i): When a person sweats, water is lost from the body by evaporation. This is an endothermic process. This evaporation speeds up when a person exercises.
(i) Explain why the evaporation of water in sweat from the body is endothermic, and why exercise increases this evaporation.

The water in sweat is changing state from liquid to gas. It needs to absorb energy to break the forces / bonds between liquid water molecules. It absorbs this from the heat of the body. The temperature of the body increases when exercising, so more water can be evaporated.

NCEA 2017 Endothermic and Exothermic

Excellence Question with b(i)

Question 1b (ii): Draw a labelled enthalpy diagram for the evaporation of water, $H_2O_{(\ell)}$.
$H_2O_{(\ell)} \rightarrow H_2O_{(g)}$ $\Delta_r H° = 40.7$ kJ mol^{-1}



Can show activation energy but not required.

Enthalpy Diagrams

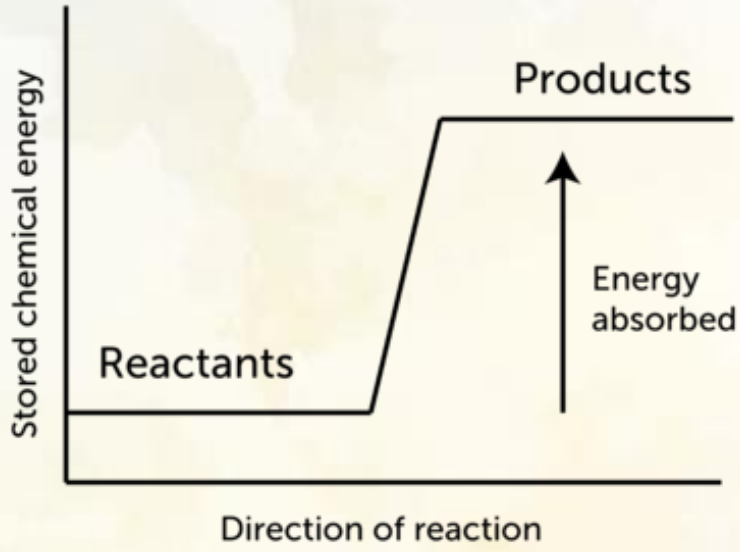
Enthalpy Diagrams can be used to show the relative amounts of enthalpy of the reactants and products in a reaction as well as the direction and relative size of enthalpy change. In endothermic reactions, the enthalpy change will be positive and in exothermic reactions, the enthalpy change will be negative.

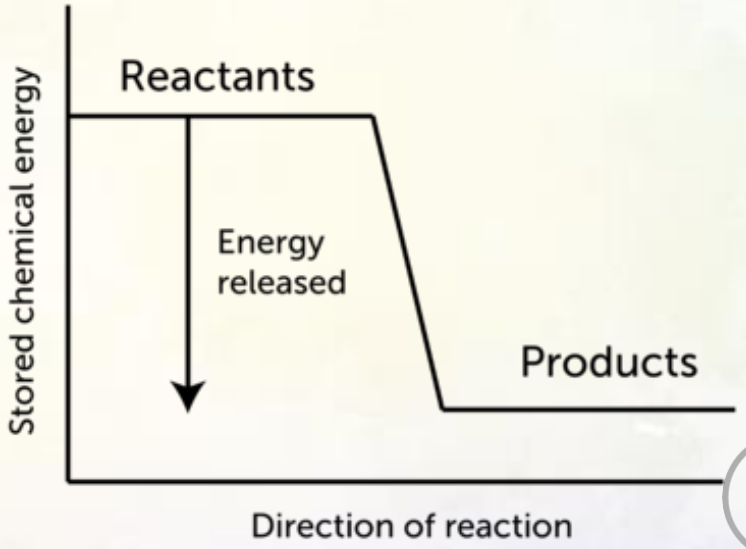
<u>Endothermic Reactions</u>

e.g. Reacting methane with steam at high pressure and temp. Energy is absorbed

<u>Exothermic Reactions</u>

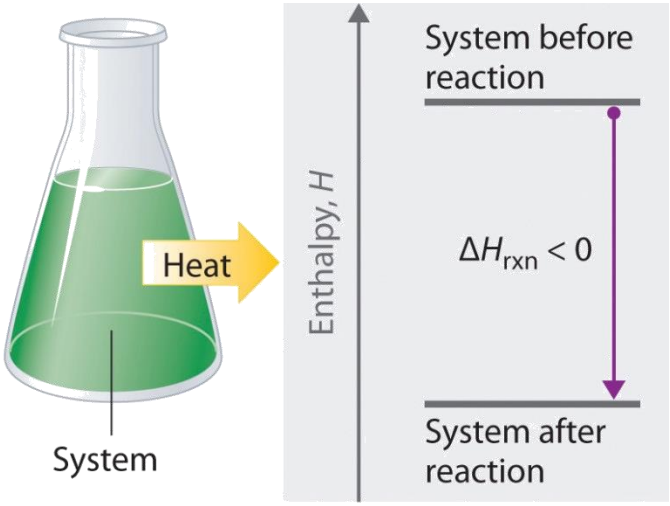
e.g. Burning of methane in air. Energy is released





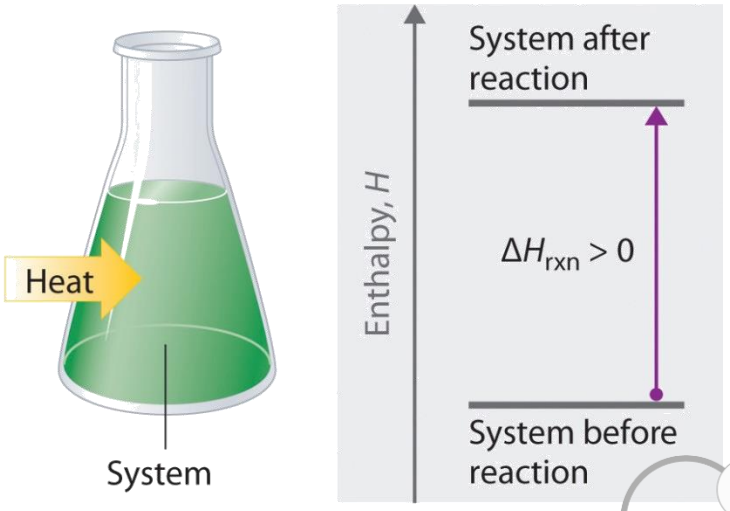
Enthalpy Change

An exothermic reaction will release energy and the products will be at a lower enthalpy level than the reactants.
The reaction system will feel will feel hot to the touch as the energy is released as heat energy.



(a) Exothermic reaction

An endothermic reaction will absorb energy and the products will be at a higher enthalpy than the reactants.
The reaction system will feel cool to the touch as heat energy is taken from the surroundings, including your skin, and used to break bonds in the molecules.

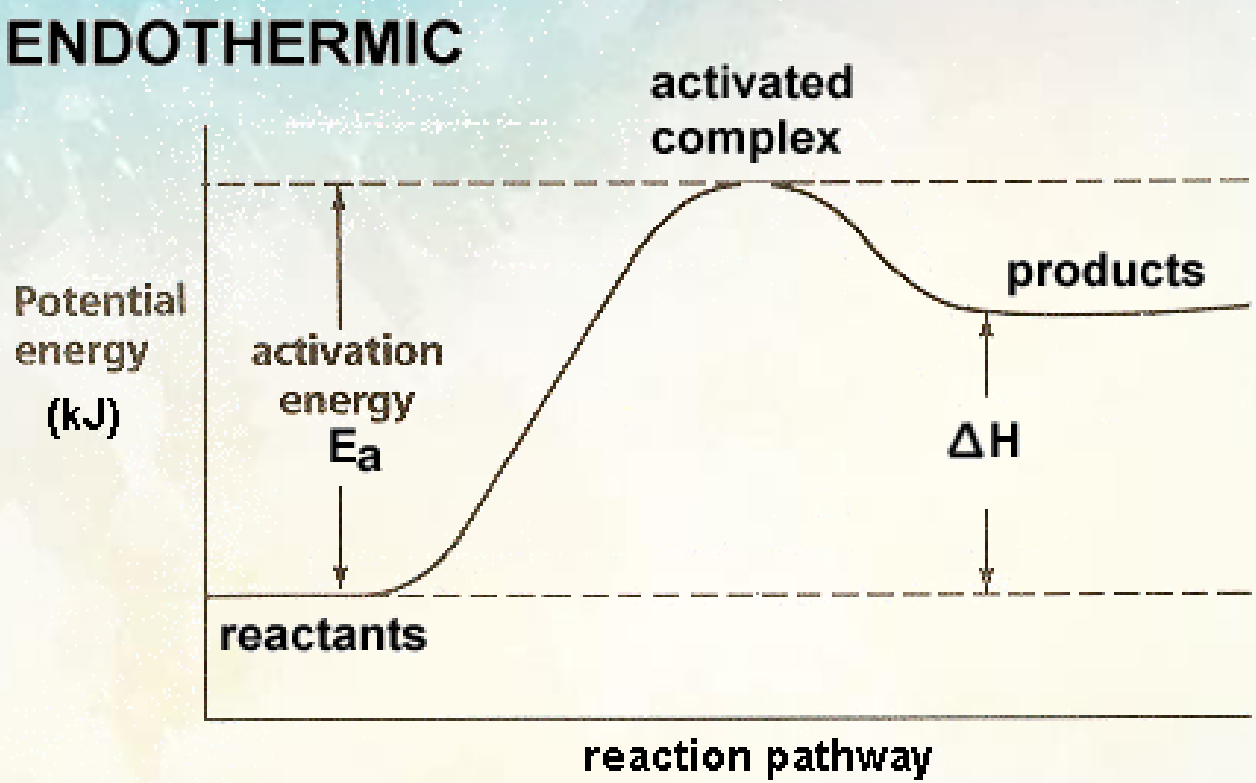


(b) Endothermic reaction

$\Delta_r H$ is the enthalpy of a reaction and is measured in $kJmol^{-1}$

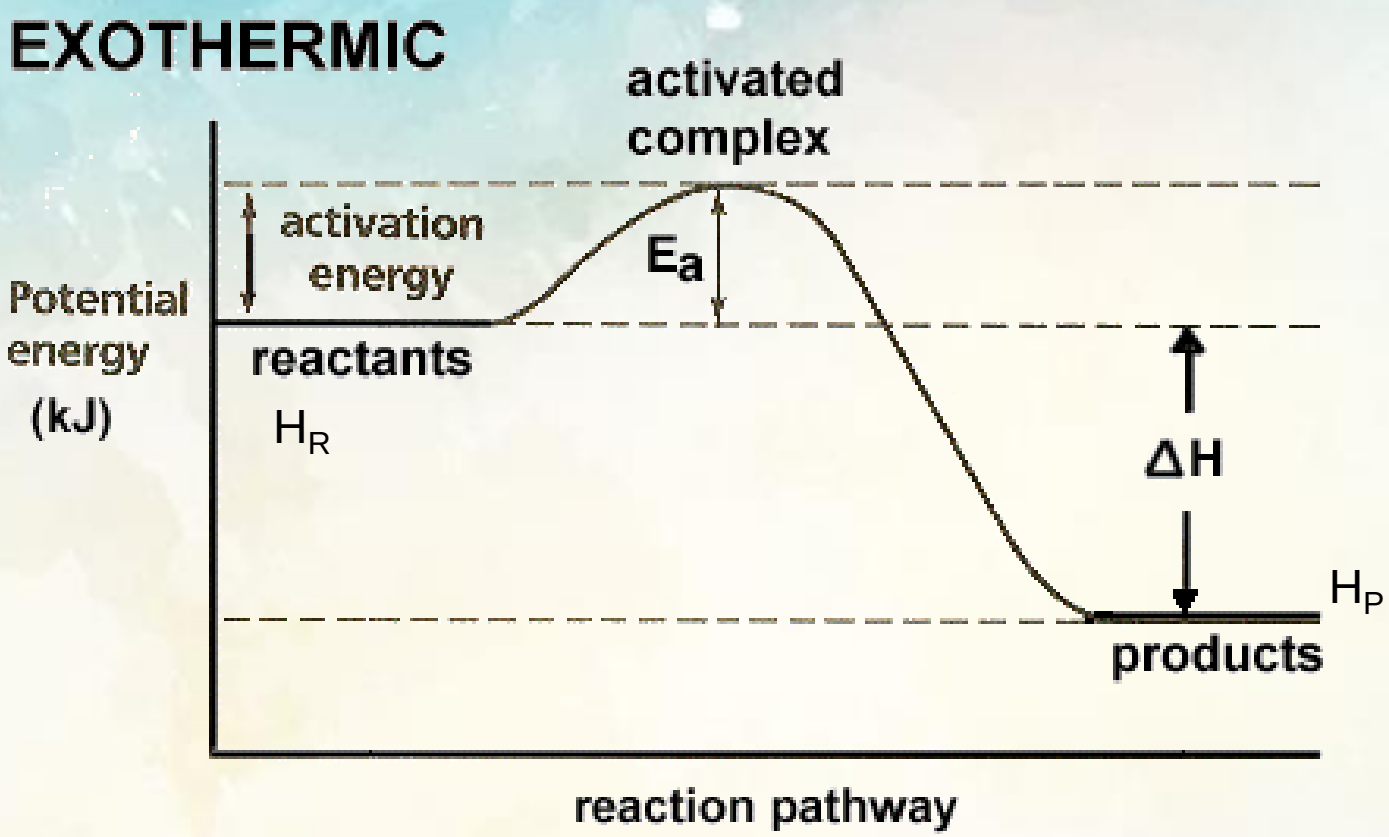
Energy Diagrams

Endothermic Reaction e.g. Reacting methane with steam at high pressure and temp. Energy is absorbed



Energy Diagrams

Exothermic Reaction e.g. concentrated Hydrochloric acid reacting with zinc metal



Standard conditions

Measurements depend on conditions
When measuring an enthalpy change you will get different values under different conditions. For example, the enthalpy change of a particular reaction will be different at different temperatures, different pressures or different concentrations of reactants. The values for enthalpy are given for standard conditions, indicated by the superscript θ

<u>Standard conditions include:</u>
Temperature of 25°C
Atmospheric pressure conditions of 1ATM
Concentration of 1mol per Litre

Thermochemical Calculations and stoichiometry

You can perform stoichiometry calculations using energy changes from thermochemical equations. Using a balanced chemical equation to calculate amounts of reactants and products is called *stoichiometry*. Energy released or absorbed can be calculated per amount of substance.

Exothermic

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

$$\Delta H = -888 kJmol^{-1}$$

This thermochemical equation reads; 888kJ of heat is released when 1 mole of CH_4 reacts with 2 moles of O_2 to produce 1 mole of CO_2 and 2 moles of H_2O

Endothermic

$$CH_4 + H_2O \longrightarrow CO + 3H_2$$

$$\Delta H = 206 kJmol^{-1}$$

This thermochemical reaction reads; 206kJ of heat is absorbed when 1 mole of CH_4 reacts with 1 mole of H_2O to produce 1 mole of CO and 3 moles of H_2.

Use thermochemical equations to find ΔH, n and m. $n = m/M$

n = moles (6.02×10^{23} particles) m = mass (grams) M = Molar Mass ($gmol^{-1}$)

The **mole**, abbreviated **mol**, is an SI unit which measures the number of particles in a specific substance. **One mole** is equal to $6.02214179 \times 10^{23}$ atoms

Thermochemical Equation Example

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ $\Delta_rH = -888 kJmol^{-1}$

Use the equation above to find heat released if 2.5 moles of CH_4 burns.

1 mole of CH_4 releases 888kJ
2.5 moles CH_4 releases **x** kJ
x = 2.5 x 888 = **2220kJ**

An equation is a mole ratio – the number in front of each substance tells you how many moles of that there is to any other substance.
For example there is 1 mole of CH_4 to every 2 moles of O_2
The enthalpy of the equation shows you the amount of energy per unit of substance.
888 = $1CH_4$ 888 = $2O_2$ (444 = $1O_2$)

Thermochemical Equation Example 2

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ $\Delta_rH = -888 kJmol^{-1}$

Calculate the **amount** (in moles) of **H_2O** produced when the reaction above releases **10,000kJ**.

Amount of mols in equation

2 moles H_2O when 888kJ released

x moles H_2O when 10000kJ released

$$x = \frac{2 \times 10000}{888} = 22.5 \text{ moles}$$

Amount energy per unit of substance

Total energy released

An **alternative method** is to find out how much energy is released per mole first

2 moles H_2O = 888kJ

Therefore 1 mole H_2O = 444kJ

10,000kJ/444kJ = 22.5

So 22.5 moles of water are produced at 444kJ to reach 10,000kJ

(22.5 x 444 = 10,000)

Thermochemical Equation Example 3

$CH_4 + H_2O \longrightarrow CO + 3H_2$ $\Delta_rH = 206 kJmol^{-1}$

$M(C) = 12gmol^{-1}$ $M(O) = 16gmol^{-1}$

Calculate the energy required to produce 1kg of CO gas from the reaction above

Step one

moles of CO produced

$M = 1000g$ $M(CO) = 28gmol^{-1}$

$n = m/M$

$n = 1000/28 = 35.7$ moles

1kg = 1000g. Must be converted to grams first

If Molar mass is not given then use the periodic table

The units are kJ **not** $kJmol^{-1}$ as it is total amount not amount per mole.

Step two

1 mole CO produced requires 206kJ (as per the equation above)

35.7 mols CO produced so…..

enthalpy = 35.7 x 206 = **7354kJ**

NCEA 2013 Thermochemical calculations

Excellence Question

Question 3b(ii) : Females who are moderately active need 9 800 kJ of energy per day. Calculate the number of moles of glucose that would provide this daily energy requirement.

$C_6H_{12}O_{6(s)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(\ell)}$ $\Delta_r H° = -2\,820$ kJ mol^{-1}

Answer 3b(ii):

9800 kJ / 2820 kJ mol–1 = 3.48 mol

Question 3c(ii) : The equation below shows the combustion of butane.

$C_4H_{10(g)} + 13/2\ O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$ $M(C_4H_{10}) = 58.1$ g mol^{-1}.

When 100 g of butane undergoes combustion, 4 960 kJ of energy is released.

Calculate the enthalpy change when 1 mole of butane undergoes combustion.

Answer 3c(ii):

$n(C_4H_{10})$ = 100 g / 58.1 g mol^{-1}

= 1.7212 mol

–4960 kJ / 1.7212 mol = –2882 kJ mol^{-1}

Calculating the amount of energy per mole

An equation and n=m/M are required for this type of thermochemical calculation

NCEA 2013 Thermochemical calculations

Excellence Question

Question 3d: The iron oxides Fe_3O_4 and Fe_2O_3 react with aluminium as shown below.

$3Fe_3O_{4(s)} + 8Al_{(s)} \rightarrow 4Al_2O_{3(s)} + 9Fe_{(s)}$ $\Delta_r H° = -3\ 348$ kJ mol^{-1}

$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(s)}$ $\Delta_r H° = -851$ kJ mol^{-1}

Justify which iron oxide, Fe_3O_4 or Fe_2O_3, will produce more heat energy when 2.00 kg of iron is formed during the reaction with aluminium.

Your answer should include calculations of the heat energy produced for the given mass of iron formed.

M(Fe) = 55.9 g mol^{-1}.

n(Fe) = 2000 g / 55.9 g mol–1 = 35.78 mol

Fe_3O_4:

3348 kJ / 9 = 372 kJ mol–1

372 kJ mol–1 × 35.78 mol = 13 310.16 kJ

= (–)1.33 ×10^4 kJ

Fe_2O_3:

851 kJ / 2 = 425.5 kJ mol–1

425.5 kJ mol–1 × 35.78 mol = 15 224.4 kJ

= (–)1.52 × 10^4 kJ

Complete each calculation one after the other

Don't forget to make a comparison statement

Therefore Fe_2O_3 produces more heat energy when 2 kg iron is formed.

NCEA 2014 Thermochemical calculations

Excellence Question

Question 3c: Methanol and ethanol can both be used as fuels. Their combustion reactions can be represented by the following equations:

Methanol combustion: $2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$ $\Delta_r H^o = -1450$ kJ mol^{-1}

Ethanol combustion: $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$ $\Delta_r H^o = -1370$ kJ mol^{-1}

Justify which fuel, methanol or ethanol, will produce more heat energy when 345 g of each fuel is combusted in excess oxygen.

$M(CH_3OH) = 32.0$ g mol^{-1}

$M(C_2H_5OH) = 46.0$ g mol^{-1}

Answer 3c :

$n(CH_3OH) = m / M = 345 / 32 = 10.78$ mol

$n(C_2H_5OH) = m / M = 345 / 46 = 7.50$ mol

Complete each calculation one after the other

2 mol CH_3OH release 1 450 kJ of energy

1 mol CH_3OH releases 725 kJ of energy

10.78 mol CH_3OH releases 725 kJ x 10.78mol

= 7 816 kJ of energy

Don't forget to make a comparison statement

1 mol C_2H_5OH releases 1 370 kJ of energy

7.5 mol C_2H_5OH releases 1 370 kJ x 7.50mol = 10 275 kJ of energy

Therefore C_2H_5OH releases more energy when 345 g of fuel are combusted.

NCEA 2015 Thermochemical calculations

Excellence Question

Question 2b(ii) : Calculate how much energy is absorbed or released in the photosynthesis reaction if 19.8 g of carbon dioxide gas, $CO_{2(g)}$, reacts completely with excess water, $H_2O_{(l)}$, to form glucose, $C_6H_{12}O_{6(aq)}$, and oxygen gas, $O_{2(g)}$.
Show your working and include appropriate units in your answer.
$M(CO_2) = 44.0$ g mol^{-1}
$\mathbf{6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}\ \Delta_r H^\circ = +2803\ kJ\ mol^{-1}}$

An equation and n=m/M are required for this type of thermochemical calculation

Answer 2b(ii):

$n(CO_2) = m/M \quad = 19.8/44.0 \quad = 0.450$mol

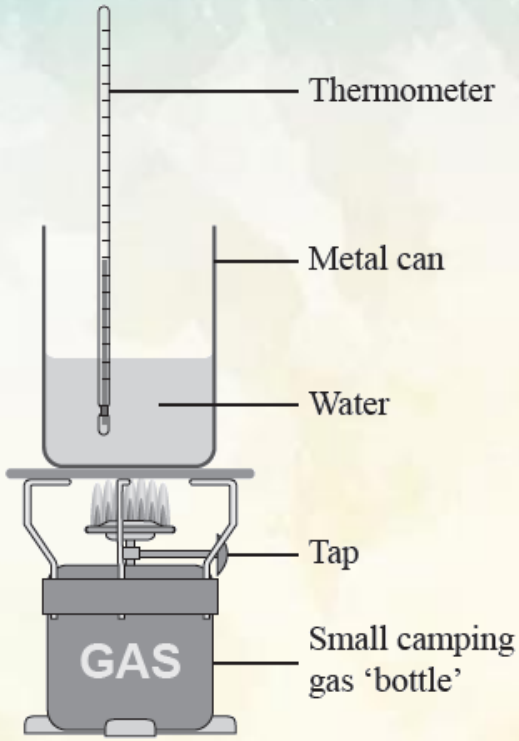
Convert mass into mols

Since 6 moles of CO_2 reacting requires 2803 kJ of energy
then 1 mole of CO_2 reacting requires 2803/6 = 467.2 kJ of energy
and 0.450 moles of CO_2 requires 467.2 x 0.450 = 210 kJ of energy absorbed.

Units are in kJ and the + sign indicates energy is absorbed

Thermochemical Experimental data

Enthalpy change ($\Delta_r H$) for a reaction, can be collected from a thermochemical investigation. Values required are masses of water in which the reaction takes place, temperature change (in °C) and the specific heat capacity value for water. (These calculations will be done in Level 3)



Often questions will ask why a particular investigation set up does not provide the same thermochemical data as the accepted enthalpy change.

Reasons can include errors such as:

1. Some energy is used to heat the metal can and the air surrounding the experiment / the experiment was not conducted in a closed system
2. Incomplete combustion of butane.
3. Some butane may have escaped before being ignited.
4. The butane in the gas canister was impure.
5. Some water evaporated
6. Some energy was converted to light and sound
7. Not carried out under standard conditions

Therefore, not all of the energy released by the combustion of butane was transferred to heating the water

NCEA 2015 Thermochemical calculations - (Part ONE)

Excellence Question

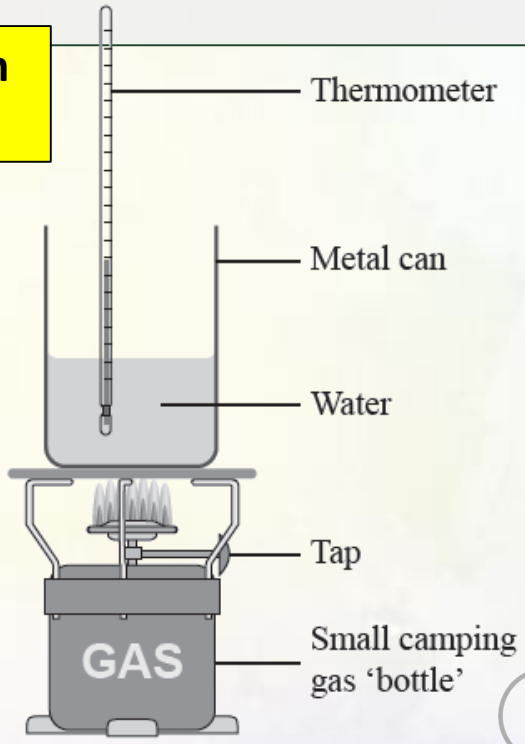
Question 2c: A small camp stove containing butane gas, $C_4H_{10(g)}$, is used to heat some water, as shown in the diagram below. A student measures the temperature change in the water and calculates that when 3.65 g of butane is combusted, 106 kJ of heat is released.

The reaction for the combustion of butane is shown in the equation below.

$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(l)}$

(i) Calculate the enthalpy change ($\Delta r\ H$) for this reaction, based on the above measurements. $M(C_4H_{10}) = 58.0$ g mol^{-1}

Values are all given to 3 sgf



Answer 2c:

$n(C_4H_{10}) = m/M \quad = 3.65/58.0 \quad = 0.0629$mol

If 0.0629 moles of C_4H_{10} releases 106 kJ of energy

Then 1 mole of C_4H_{10} releases 106/0.0629 = 1685 kJ of energy

And 2 moles of C_4H_{10} releases 1685 x 2 = 3370 kJ of energy (3368) ($\Delta_r H = -3370$ kJ mol^{-1})

Must show sign

NCEA 2015 Thermochemical calculations - (part TWO)

Excellence Question

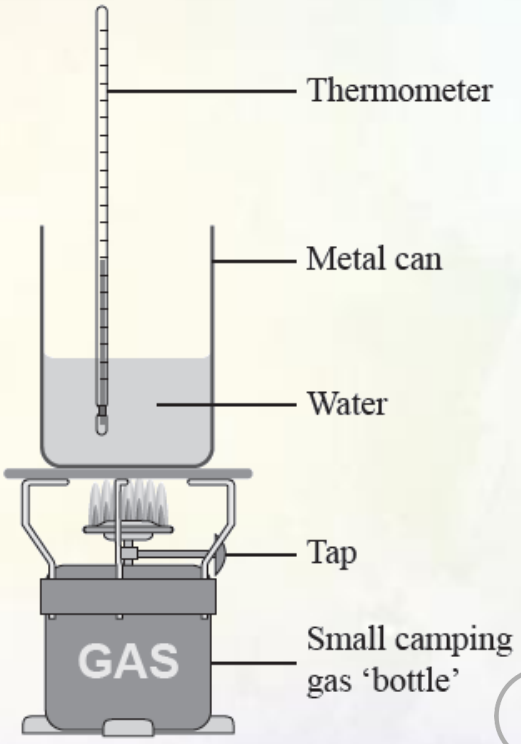
Question 2c: (ii) The accepted enthalpy change for the combustion reaction of butane gas, $C_4H_{10(g)}$, is $\Delta r\ H = -5754$ kJ mol^{-1}.
Explain why the result you calculated in part (c)(i) is different to the accepted value.
In your answer, you should include at least TWO reasons.

TWO errors required for Excellence

Answer 2c: The results from this experiment are less than the accepted results, due to errors in the experimental design.
The errors could include:

1. Some energy is used to heat the metal can and the air surrounding the experiment / the experiment was not conducted in a closed system
2. Incomplete combustion of butane.
3. Some butane may have escaped before being ignited.
4. The butane in the gas canister was impure.
5. Some water evaporated
6. Some energy was converted to light and sound
7. Not carried out under standard conditions

Therefore, not all of the energy released by the combustion of butane was transferred to heating the water.



NCEA 2015 Thermochemical calculations - (part THREE)

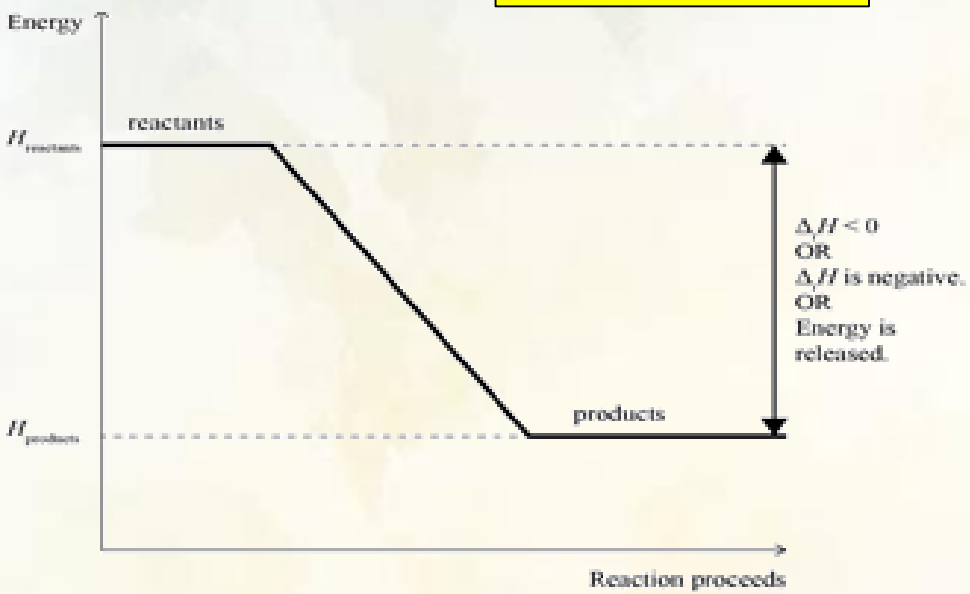
Excellence Question

Question 2c (iii): Complete, including labels, the energy diagram for the combustion of butane gas showing reactants, products, and the change in enthalpy.

Question 2c (iv): Butane gas is a useful fuel because when it undergoes combustion, energy is released.
Explain why energy is released in this reaction, in terms of making and breaking bonds.

Both required to be correct for Excellence

Answer 2c (iii):



Answer 2c (iv): When butane undergoes combustion, heat is released, so it is an exothermic reaction.

Bond-making is an exothermic process / releases energy and bond-breaking is endothermic / requires energy. For the overall reaction in the combustion of butane to release energy, more energy is given out as bonds are made (when the products, CO_2 and H_2O are formed) than the energy being used to break the bonds (in the reactants, C_4H_{10} and O_2).

NCEA 2016 Thermochemical calculations

Excellence Question

Question 1c(iii): Hexane, C_6H_{14}, like pentane, will combust (burn) in sufficient oxygen to produce carbon dioxide gas and water. Pentane combustion: $\Delta_r H^{\circ} = -3509$ kJ mol^{-1}
Hexane combustion: $2C_6H_{14(l)} + 19O_{2(g)} \rightarrow 12CO_{2(g)} + 14H_2O_{(l)}$ $\Delta_r H^{\circ} = -8316$ kJ mol^{-1}
Justify which alkane – pentane or hexane – will produce more heat energy when 125 g of each fuel is combusted in sufficient oxygen.
$M(C_5H_{12}) = 72.0$ g mol^{-1} $M(C_6H_{14}) = 86.0$ g mol^{-1}

An equation and n=m/M are required for this type of thermochemical calculation

Answer 1c(iii):

n (pentane) = 125 g / 72.0 g mol^{-1} = 1.74 mol
n (hexane) = 125 g / 86.0 g mol^{-1} = 1.45 mol
If 1 mole of pentane releases 3509 kJ energy, then 1.74 mol of pentane:
1.74 × 3509 = 6106 kJ energy released.

Convert mass into mols

If 2 moles of hexane release 8316 kJ energy, then 1 mole of hexane releases 4158 kJ energy.
So 1.45 mol of hexane 1.45 × 4158 = 6029 kJ energy releases.

Make sure both fuels are compared for Excellence

So pentane releases more energy (77.0 kJ) than hexane, per 125 g of fuel.

NCEA 2017 Thermochemical calculations

Excellence Question

Question 1c: Thermite reactions occur when a metal oxide reacts with a metal powder. The equations for two thermite reactions are given below:

Reaction 1: $Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow 2Fe_{(s)} + Al_2O_{3(s)}$ $\Delta_rH° = -852$ kJ mol^{-1}

Reaction 2: $3CuO_{(s)} + 2Al_{(s)} \rightarrow 3Cu_{(s)} + Al_2O_{3(s)}$) $\Delta_rH° = -1520$ kJ mol^{-1}

Use calculations to determine which metal oxide, iron(III) oxide, $Fe_2O_{3(s)}$, or copper(II) oxide, $CuO_{(s)}$, will produce more heat energy when 50.0 g of each metal oxide is reacted with aluminium powder, $Al_{(s)}$.

$M(Fe_2O_3) = 160$ g mol^{-1} $M(CuO) = 79.6$ g mol^{-1}

An equation and n=m/M are required for this type of thermochemical calculation

$n(Fe_2O_3) = \frac{50.0 \text{ g}}{160 \text{ g mol}^{-1}} = 0.313$mol

$n(CuO) = \frac{50.0 \text{ g}}{79.6 \text{ g mol}^{-1}} = 0.628$ mol

Convert mass into mols

Reaction 1: If 1 mole of Fe_2O_3 releases 852 kJ energy

0.313 mol × 852 kJ mol^{-1} = 266 kJ energy released

Reaction 2: If 3 mole of CuO releases 1520 kJ energy

Then 1 mole of CuO releases 507 kJ energy

0.628 mol × 507 kJ mol–1 = 318 kJ energy released

So 50.0 g CuO releases more energy than 50.0 g Fe_2O_3

OR

CuO releases more energy (52 kJ) than Fe_2O_3

(**Reaction 2** releases more energy.)

Make sure both substances are compared for Excellence

Bond Enthalpy

Bond enthalpy (also known as bond energy) is defined as the amount of energy required to break one mole of the stated bond.

The high values for bond enthalpy explains why some substances are very resistant to chemical attack and form very stable molecules

In a polyatomic (more than one atom) molecule, the bond strength between a given pair of atoms can vary slightly from one compound to another. The value given for bond enthalpy is the average of all these variations.

A multiple bond (double/triple) is always stronger than a single bond because more electrons bind the multiple bonded atoms together.
The table below shows some common average bond enthalpies.

Bond enthalpy /kJ mol^{-1}		Bond enthalpy /kJ mol^{-1}		C≡C thalpy /kJ mol^{-1}	
H - H	436	C - H	412	C = C	612
H - O	463	C - Cl	338	C o C	837
H - N	388	C - F	484	C = O	743
H - Cl	431	C - O	360	O = O	496
H - F	565	C - C	348	N ≡ N	944
F - F	158	O – O	146		
Cl - Cl	242				

Bond Enthalpy calculations

Bonds Broken – Endothermic **Bonds formed – Exothermic**

$\Delta_r H°$= ∑ (energy of bonds broken) - ∑(energy of bonds formed)

Note: Bond energies calculated for gases. Convert using $\Delta_{vap}H°$ or $\Delta_{sub}H°$ if in solid or liquid state.

$$CO_{(g)} + H_2O_{(g)} \rightarrow H_{2(g)} + CO_{2(g)} \qquad \Delta_r H° = -1.0\ Kjmol^{-1}$$

Bonds Broken		Bonds formed	
C≡O	995kJ	C=O x 2	2(743)kJ
H-O x 2	2(463)kJ	H-H	436kJ
	1921kJ		1922kJ

$\Delta_r H°$= 1921kJmol^{-1} – 1922kJmol^{-1}

$\Delta_r H°$= -1.0 kJmol^{-1}

The equation can also be arranged to calculate unknown bond energy

Bond Enthalpy

Bonds Broken – Endothermic **Bonds formed – Exothermic**

$\Delta_r H° = \sum$ (energy of bonds broken) - $\sum$(energy of bonds formed)

Reactants: Draw lewis diagrams to calculate the number and type of bond

Multiply the bond energy given by the number of bonds

Products: Draw lewis diagrams to calculate the number and type of bond

$CO_{(g)} + H_2O_{(g)} \rightarrow H_{2(g)} + CO_{2(g)}$ $\Delta_r H° = -1.0\ Kjmol^{-1}$

Bonds Broken		Bonds formed	
C≡O	995kJ	C=O x 2	2(743)kJ
H-O x 2	2(463)kJ	H-H	436kJ
	1921kJ		1922kJ

Total the bond energy for reactant molecules

Total the bond energy for product molecules

$\Delta_r H° = 1921 kJmol^{-1} - 1922 kJmol^{-1}$

$\Delta_r H° = -1.0\ kJmol^{-1}$

The equation can also be arranged to calculate unknown bond energy

bonds broken (reactants) minus bonds formed (product) = total enthalpy

Using Bond Enthalpy to calculate $\Delta_r H°$

Bond enthalpy is the change in enthalpy when the covalent bond, **in a gaseous molecule,** is broken. It is always a positive value because bond breaking always requires an input of energy.
Making bonds releases energy so generally speaking the more bonds a substance can form the more stable it will be.

The strength of a covalent bond depends on the **electrostatic attraction** between the positive nuclei and the shared electron pair. The larger the atomic radius of an atom (which increases down a group) the further the shared electron pair from the positive nucleus – which creates decreasing electrostatic attraction. Therefore the weaker the covalent bond and the lower the value of bond enthalpy**.**

The stronger a covalent bond, the higher the value of the bond enthalpy. The units are kJ mol^{-1}

Draw Lewis structures if not given

NCEA 2013 Bond Enthalpy

Excellence Question

Question 2c: Chlorine reacts with methane to form chloromethane and hydrogen chloride, as shown in the equation below.

$CH_{4(g)} + Cl_{2(g)} \rightarrow CH_3Cl_{(g)} + HCl_{(g)}$

Use the following bond enthalpies to calculate Δr *H*° for this reaction.

$\Delta_r H^o = \Sigma$Bond energies(bonds broken) – ΣBond energies(bonds formed)

Bond energies always given – use units on chart to remind you

Bonds broken	
C-H x 4	1656
Cl–Cl	242
	1898

Bonds formed	
C–Cl	324
C–H	3 × 1242
H-Cl	431
	1997

= 1898 – 1997 = –99 kJ mol^{-1}

Alternative calculation that includes the breaking and reforming of only the bonds involved in the reaction

(656 – 755 = 99kJ mol^{-1})

Bond	Bond enthalpy /kJ mol^{-1}
H–Cl	431
C–H	414
C–Cl	324
Cl–Cl	242

NCEA 2014 Bond Enthalpy

Excellence Question

Question 1d: Hydrogen gas, $H_{2(g)}$, reacts with oxygen gas, $O_{2(g)}$, as shown by the following equation

$$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(g)} \quad \Delta_r H^o = \mathbf{-242\ kJ\ mol^{-1}}$$

Given the average bond enthalpies in the table below, calculate the average bond enthalpy of the **O – H** bond in H_2O.

$\Delta_r H^o = \Sigma$Bond energies(bonds broken) – ΣBond energies(bonds formed)

Bond energies always given – use units on chart to remind you

Bond	Average bond enthalpy / kJ mol^{-1}
H–H	436
O=O	498

Bonds broken
H–H = 436
½ × O=O = ½ × 498
Total = 685 kJ mol^{-1}

Bonds formed
2 × O–H
Total = ?

–242 kJ mol^{-1} = 685kJ - ?
? = 685 – (– 242)
= 927 kJ mol^{-1}

Swap total and ?

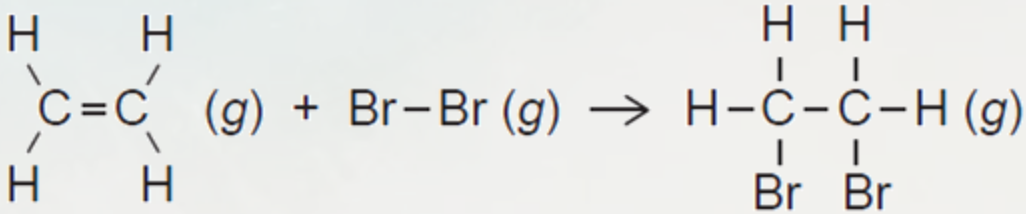
2 × O–H = 927 kJ mol^{-1}
but ONE **O–H** = 464 (463.5) kJ mol^{-1}

NCEA 2015 Bond Enthalpy

Excellence Question

Question 1d: Ethene gas, $C_2H_{4\,(g)}$, reacts with bromine gas, $Br_{2(g)}$, as shown in the equation below.
Calculate the enthalpy change, $\Delta_r H°$, for the reaction between ethane and bromine gases, given the average bond enthalpies in the table below. Show your working and include appropriate units in your answers.

Draw Lewis structures if not given



$\Delta_r H^o$ = ΣBond energies(bonds broken) – ΣBond energies(bonds formed)

Bond energies always given – use units on chart to remind you

Bond	Average bond enthalpy/kJ mol^{-1}
Br–Br	193
C–C	346
C=C	614
C–Br	285
C–H	414

Bonds broken		Bonds formed	
C=C	614	C–C	346
Br–Br	193	C–Br	2 × 285
	807		916

= 807 – 916 = –109 kJ mol^{-1}

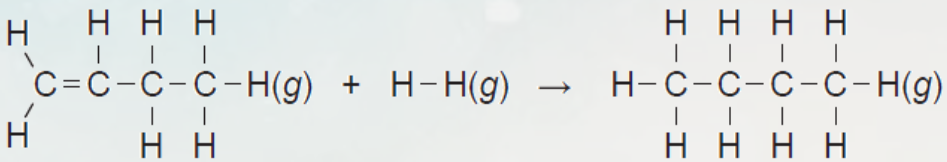
Alternative calculation that includes the breaking and reforming of four C-H bonds
(or 2463 – 2572)

Draw Lewis structures if not given

NCEA 2016 Bond Enthalpy

Excellence Question

Question 3c: Calculate the enthalpy change, $\Delta_r H°$, for the reaction of but-1-ene gas, $C_4H_{8(g)}$, with hydrogen gas, $H_{2(g)}$, to form butane gas, $C_4H_{10(g)}$.
Use the average bond enthalpies given in the table below.



Bond energies always given – use units on chart to remind you

Bond	Average bond enthalpy / kJ mol^{-1}
C=C	614
C–C	346
C–H	414
H–H	436

Bond breaking

C=C	614
C–C × 2	692
C–H × 8	3312
H–H	436
	5054 kJ mol^{-1}

Bond making

C–C × 3	1038
C–H × 10	4140
	5178 kJ mol^{-1}

$\Delta_r H°$ = Bond breaking – bond making

$\Delta_r H°$ = 5054 kJ mol^{-1} – 5178 kJ mol^{-1}

$\Delta_r H°$ = –124 kJ mol^{-1}

Alternative calculation that includes the breaking and reforming of only the bonds involved in the reaction

$\Delta_r H°$ = 1050 – 1174 = –124 kJ mol^{-1}

NCEA 2017 Bond Enthalpy

Excellence Question

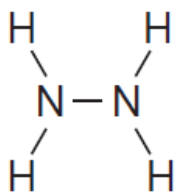
Question 2c: Hydrazine, N_2H_4, is used as rocket fuel.

Use calculations to determine which of **Reaction 1** or **Reaction 2** releases more energy.

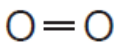
Reaction 1: $N_2H_{4(g)} + O_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$

Reaction 2: $N_2H_{4(g)} + 2F_{2(g)} \rightarrow N_{2(g)} + 4HF_{(g)}$

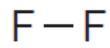
The structure of each chemical species is shown in the box below. Show your working and include appropriate units in your answer.



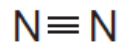
hydrazine



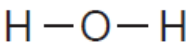
oxygen



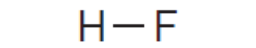
fluorine



nitrogen



water



hydrogen fluoride

Use the average bond enthalpies given in the table below.

Bond	Average Bond enthalpy /kJ mol^{-1}	Bond	Average Bond enthalpy /kJ mol^{-1}
H–H	436	N–N	158
H–F	567	F–F	159
N–H	391	O=O	498
O–H	463	N≡N	945

NCEA 2017 Bond Enthalpy

Excellence Question

Question 2c: Hydrazine, N_2H_4, is used as rocket fuel.
Use calculations to determine which of **Reaction 1** or **Reaction 2** releases more energy.

Reaction 1
Hydrazine and oxygen

Bond breaking		Bond making	
N–N	158	N≡N	945
N–H × 4	1564	O–H × 4	1852
O=O	498		2797
	2220		

Bond breaking – bond making
2220 – 2797 = –577 kJ mol^{-1}

Reaction 2
Hydrazine and Fluorine

N–N	158	N≡N	945
N–H × 4	1564	H–F × 4	2268
F–F × 2	318		3213
	2040		

Bond breaking – bond making
2040 – 3213 = –1173 kJ mol^{-1} (or -1170 kJ mol_{-1})
Reaction 2 releases more energy than **Reaction 1** (by 596 kJ mol_{-1}).