

With 2019 NCEA  
Exam included



Part One

# NCEA Chemistry 3.4

## Particles, Substances and Thermochemistry AS 91390



A red seal with a scalloped edge, containing the text "AS 91390" and "C3.4" in white.

AS 91390  
C3.4

This achievement standard involves describing properties of atoms, molecules, and ions, and thermochemical principles.

*Properties of particles* will be limited to:

- ☐ electron configuration of atoms and ions of the first 36 elements (using *s, p, d* notation)
- ☐ periodic trends in atomic radius, ionisation energy, and electronegativity, and comparison of atomic and ionic radii
- ☐ Lewis structures and shapes (up to six electron pairs about the central atom for molecules and polyatomic ions, including those with multiple bonds), polarity of molecules
- ☐ attractive forces between atoms, ions, and molecules. These will include ionic bonds, covalent bonds, and intermolecular attractions due to temporary dipoles and permanent dipoles (including hydrogen bonding).

*Properties of substances* will be limited to:

- ☐ melting point, boiling point and solubility
- ☐ enthalpy and entropy changes associated with spontaneity in chemical reactions (entropy calculations are not required)

*Thermochemical principles* will include

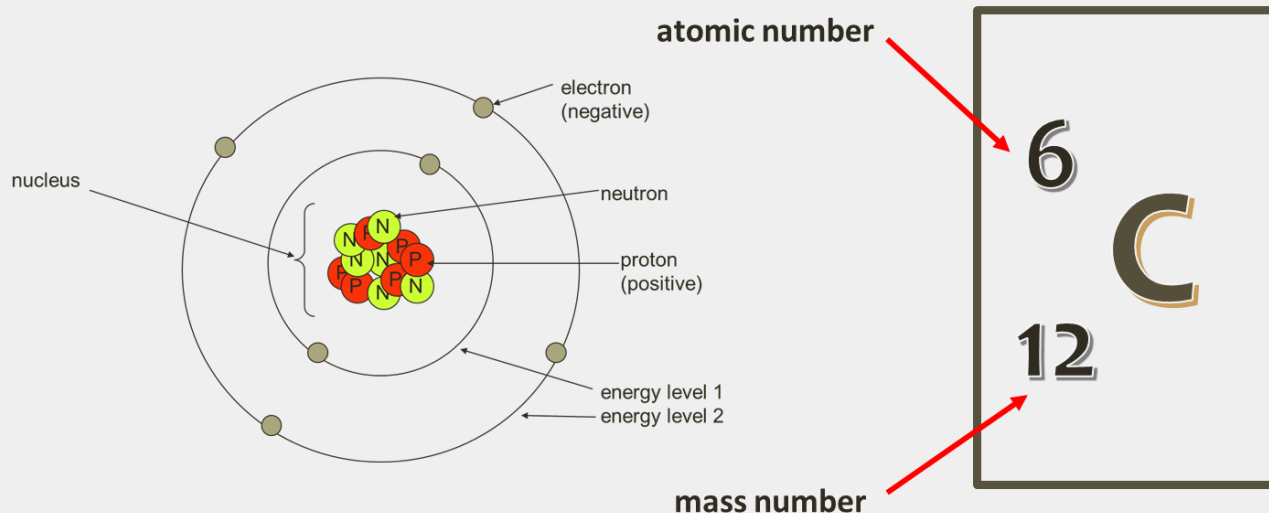
- ☐ specific heat capacity
- ☐  $\Delta_c H^\circ$ ,  $\Delta_f H^\circ$ ,  $\Delta_r H^\circ$ ,  $\Delta_{vap} H^\circ$ ,  $\Delta_{sub} H^\circ$ , and  $\Delta_{fus} H^\circ$
- ☐ Hess's Law including application of  $\Delta_r H = \sum \Delta_f H(\text{products}) - \sum \Delta_f H(\text{reactants})$  and related calculations.

## Atomic and Mass number

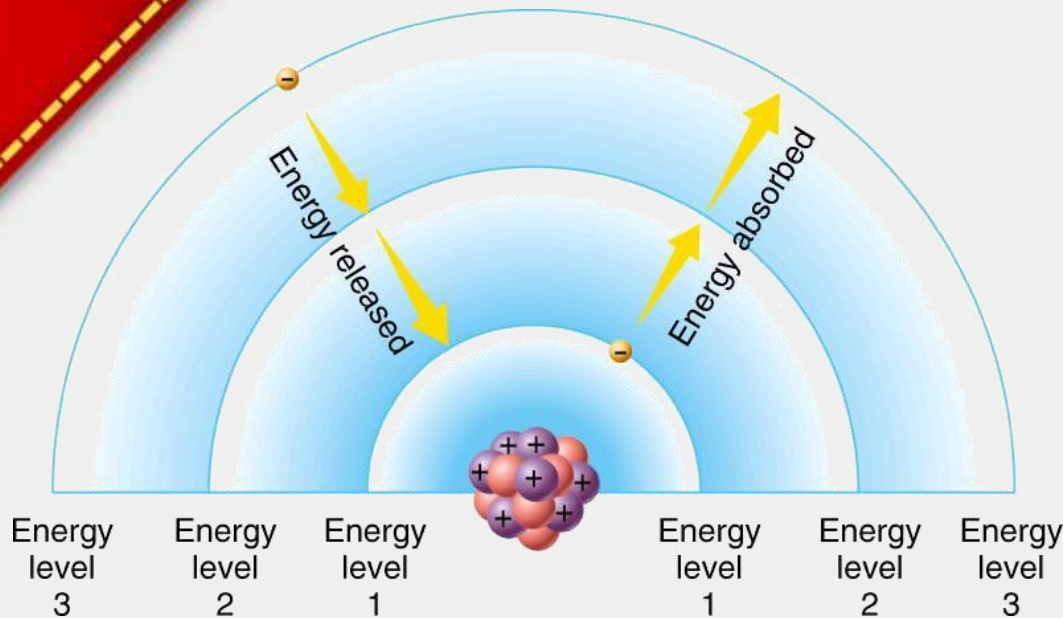
The **atomic number** is unique for each element, representing the number of protons in an atom's nucleus.

In a uncharged atom, this also equals the number of electrons. The periodic table is arranged in order of an element's atomic number.

The **mass number** represents the average number of protons and neutrons in an atom's nucleus, of all possible isotopes with each particle equating to 1 mass unit (1amu)



## Electrons orbit a nucleus of an atom at fixed energy levels.



An electron can be located in a fixed energy level around a nucleus of an atom.  
A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons because a ball can only rest on each step, not between steps.

An electron can move up an energy level if energy is absorbed, such as light or heat. When the electron moves back down to its lowest possible level the energy is released again, often in the form of light.



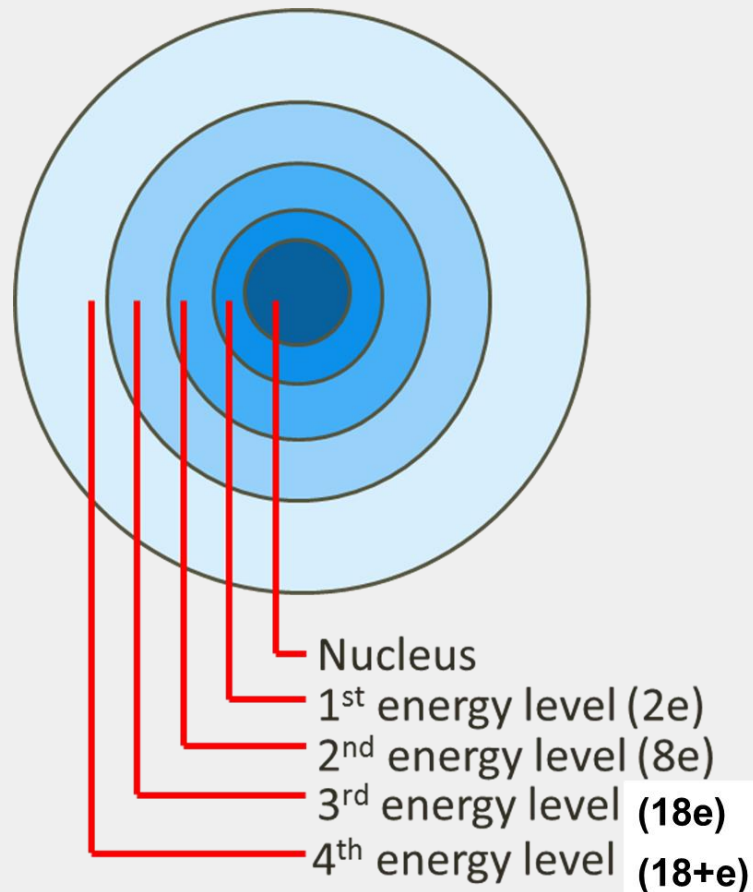
# Background Knowledge



The electrons in an atom are arranged in a series of energy levels.

Electrons move or 'orbit' around the nucleus in **energy levels** or shells. The energy levels further away from the nucleus are able to fit more electrons.

The first energy level is filled first, followed by the second and so on until all the electrons (the same number of protons in an atom) have been used.

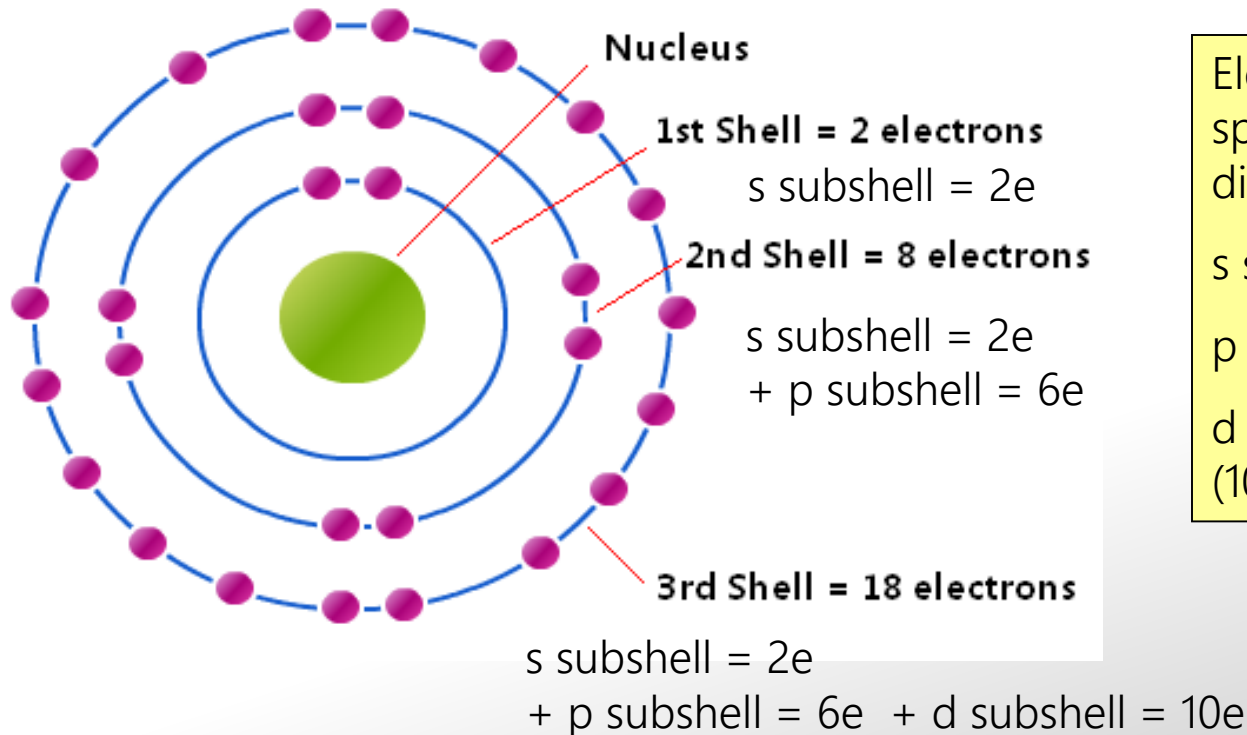


You need to draw the SPD configurations of the first 36 elements

# Electron orbitals

The major energy levels (shell) split into **subshells**. Each sub shell level made up of one or more **atomic orbitals**.

An **Orbital** is an area of space with high probability of finding a particular electron pair. Electrons orbit in pairs spinning opposite directions.



Electrons orbit in pairs  
spinning opposite  
directions

s subshell = 1 orbital (2e)

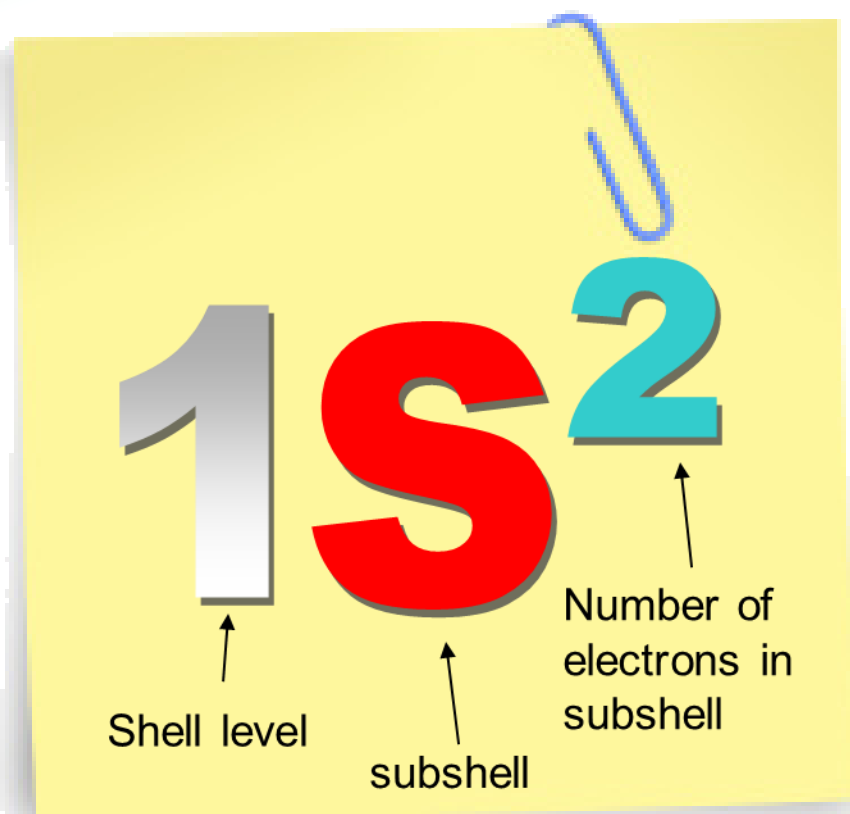
p subshell = 3 orbitals (6e)

d subshell = 5 orbitals  
(10e)

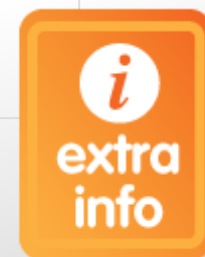
# Electron Configuration notation

Even though the subshells are filled according to energy level, the notation is often written out in numerical order.

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^2, 4p^6, \dots$

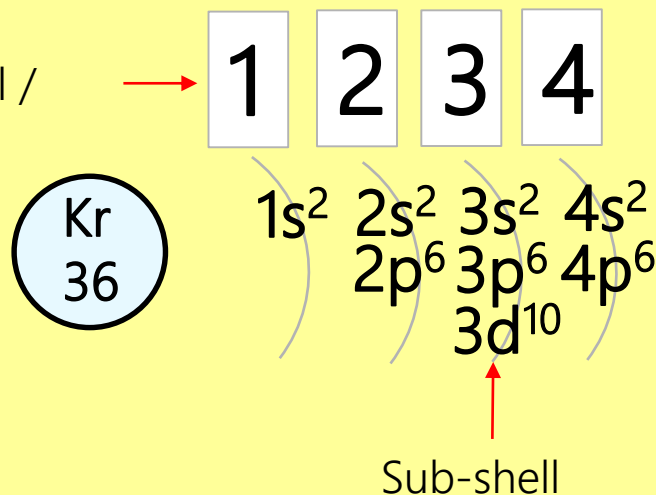


Atoms and ions with the same electron configuration are said to be **isoelectronic**.



# Energy levels and subshells

Major energy level / shell



The sub shells are numbered according to the major shell/energy level they are in.

Major Energy Level	No. of orbitals of a given type			Total electrons
<b>n</b>	<b>s</b>	<b>p</b>	<b>d</b>	
1	1			2
2	1	3		8
3	1	3	5	18
4	1	3	5	18+

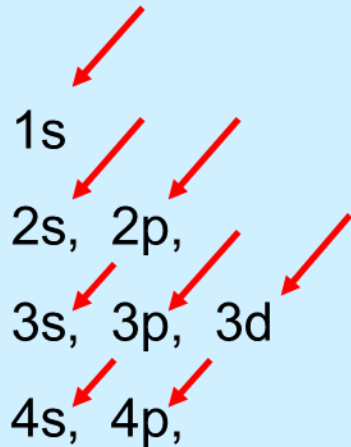


# Order of fill

The arrangement of the electrons around the nucleus is known as its configuration.

The **Aufbau principle** states that electrons always fill the shells with the lowest energy level first.

## Order of fill



1. All of the orbitals hold up to 2 electrons which make up a pair (s1,p3,d5).
2. The electron pairs spin in opposite directions indicated with  $\uparrow\downarrow$  arrows
3. Electrons fill up orbitals with lower energy levels first  
 $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p$
4. The most stable (lower energy) arrangement is for the sublevels to be filled with a single electron first (with the same direction spin) before pairing electrons. (**Hund's rule**) Place electrons singly first showing upwards arrow.



# Condensed Electron Configuration

As it is the valence shell that is of the most interest, the inner shells up to that point may be shown by the symbol for the previous inert gas element (neon)

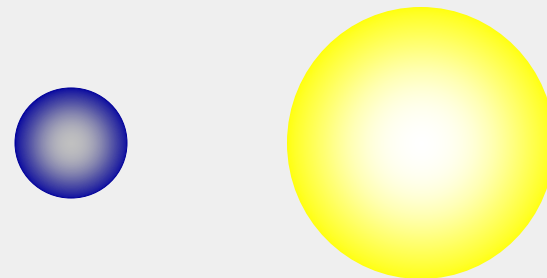
e.g. Cl [Ne]  $3s^2 3p^5$  rather than  $1s^2 2s^2 2p^6 3s^2 3p^5$

Element	Electron Configuration
Hydrogen	$1s^1$
Helium	$1s^2$
Lithium	$1s^2 2s^1$
Beryllium	$1s^2 2s^2$
Boron	$1s^2 2s^2 2p^1$
Carbon	$1s^2 2s^2 2p^2$
Nitrogen	$1s^2 2s^2 2p^3$
Oxygen	$1s^2 2s^2 2p^4$
Fluorine	$1s^2 2s^2 2p^5$
Neon	$1s^2 2s^2 2p^6$
Sodium	$1s^2 2s^2 2p^6 3s^1$
Magnesium	$1s^2 2s^2 2p^6 3s^2$
Aluminium	$1s^2 2s^2 2p^6 3s^2 3p^1$
Silicon	$1s^2 2s^2 2p^6 3s^2 3p^2$
Phosphorus	$1s^2 2s^2 2p^6 3s^2 3p^3$
Sulfur	$1s^2 2s^2 2p^6 3s^2 3p^4$
Chlorine	$1s^2 2s^2 2p^6 3s^2 3p^5$
Argon	$1s^2 2s^2 2p^6 3s^2 3p^6$

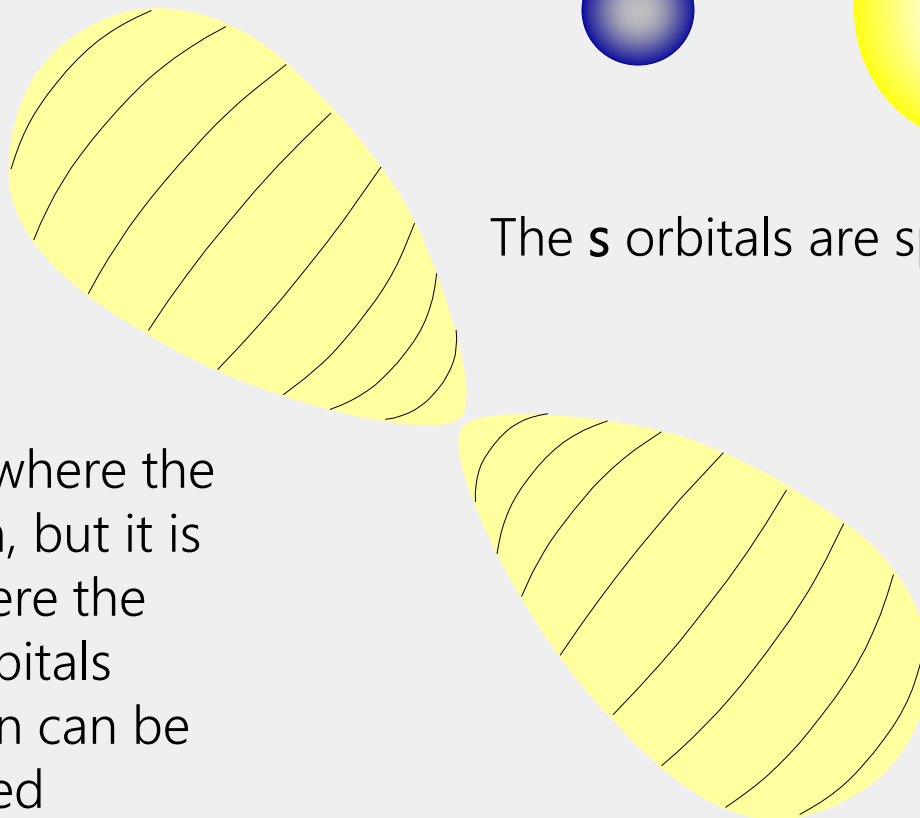
# Background Knowledge



## Electron orbital shapes



The **s** orbitals are spheres



Research has determined where the orbitals are inside an atom, but it is never known precisely where the electrons are inside the orbitals. The area where an electron can be found, the orbital, is defined mathematically, but we can see it as a specific shape in 3-dimensional space...

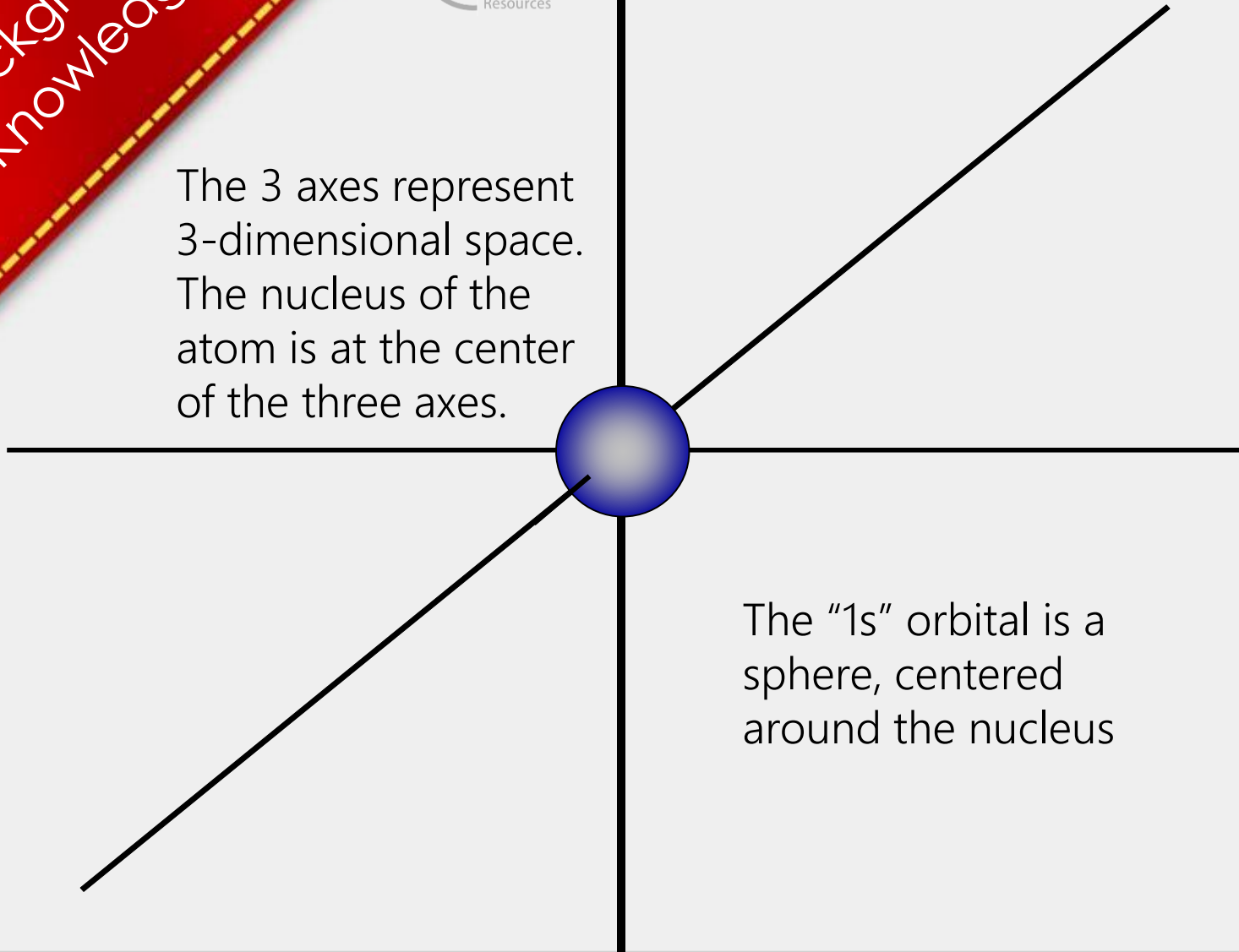
The **p** orbitals are tear shaped

# Background Knowledge



## Electron orbital shapes

The 3 axes represent 3-dimensional space. The nucleus of the atom is at the center of the three axes.

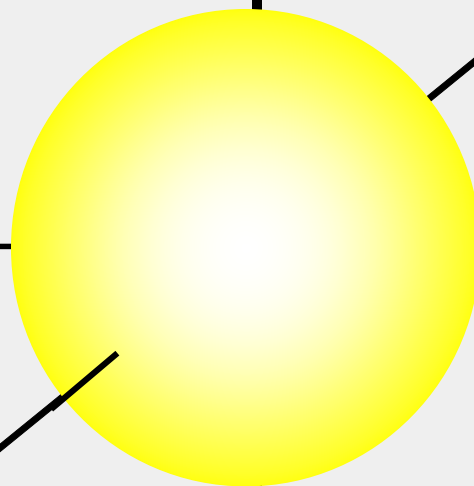


The "1s" orbital is a sphere, centered around the nucleus



## Electron orbitals shapes

The 2s orbital is also  
a sphere.

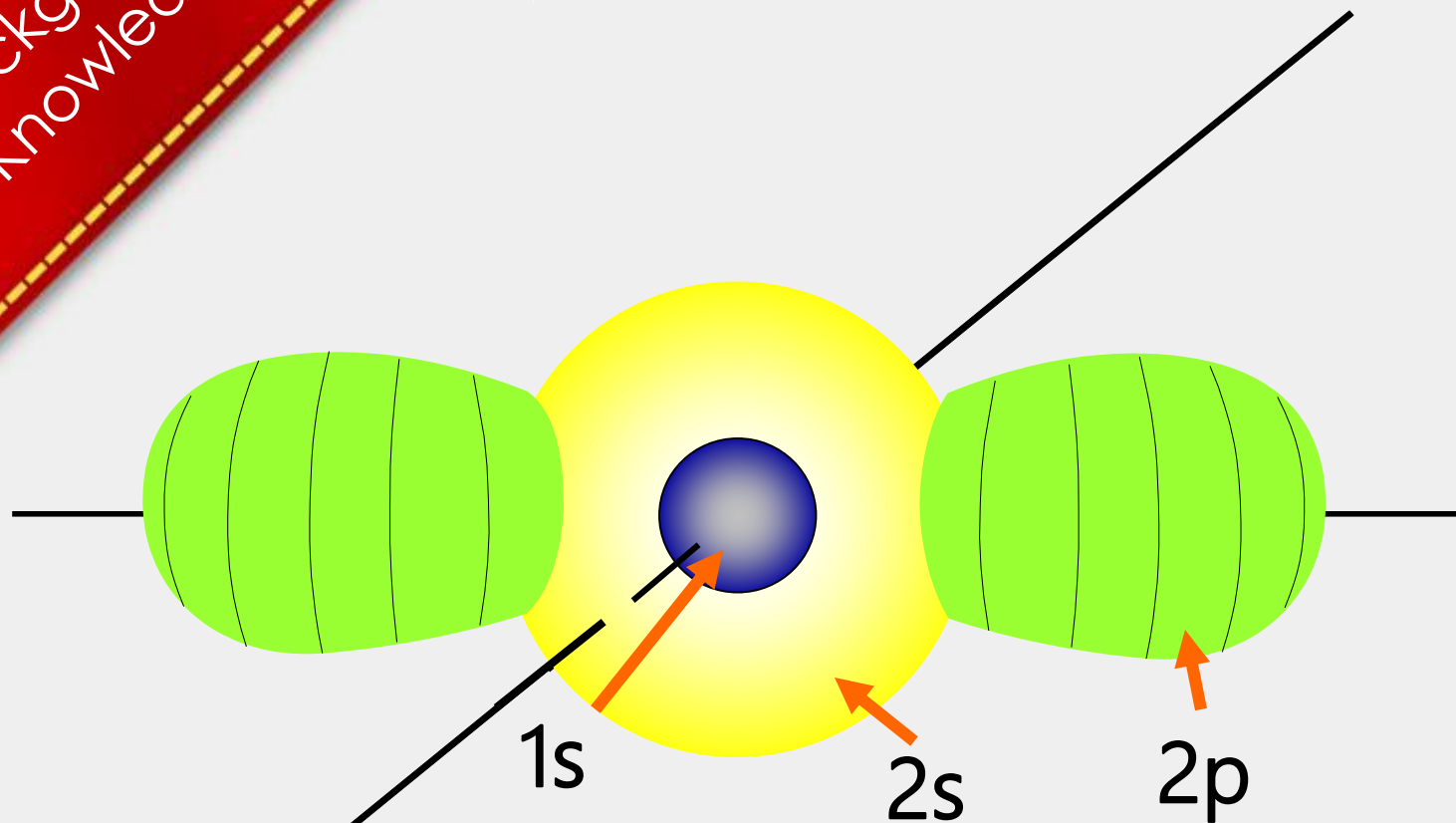


The 2s electrons have a  
higher energy than the 1s  
electrons. Therefore, the 2s  
electrons are generally more  
distant from the nucleus,  
making the 2s orbital larger  
than the 1s orbital.





## Electron orbital shapes



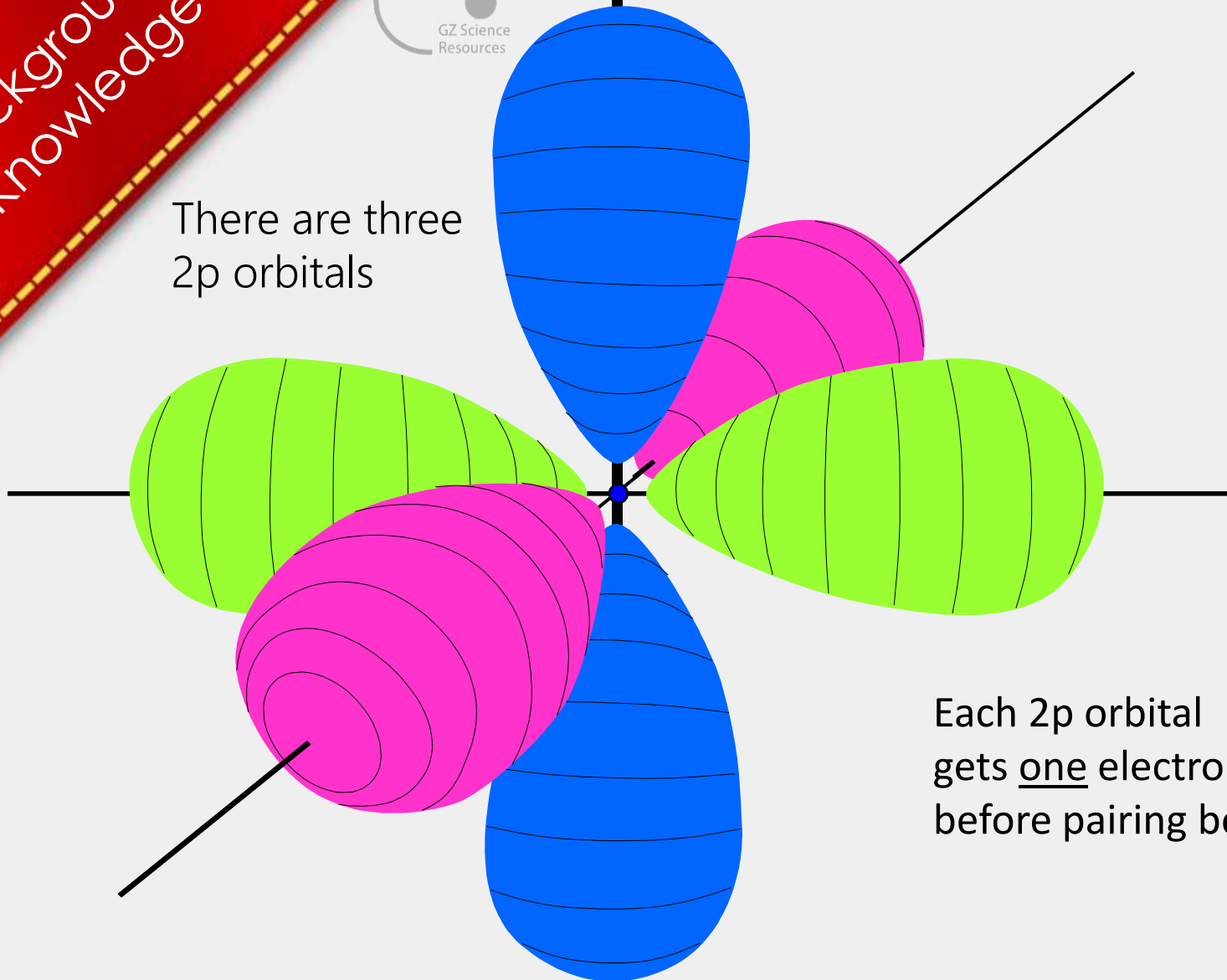
Once the **1s** orbital is filled, the **2s** orbital begins to fill. Once the **2s** orbital is filled, the **2p** orbitals begin to fill

# Background Knowledge

GZ Science Resources

## Electron orbital shapes

There are three 2p orbitals

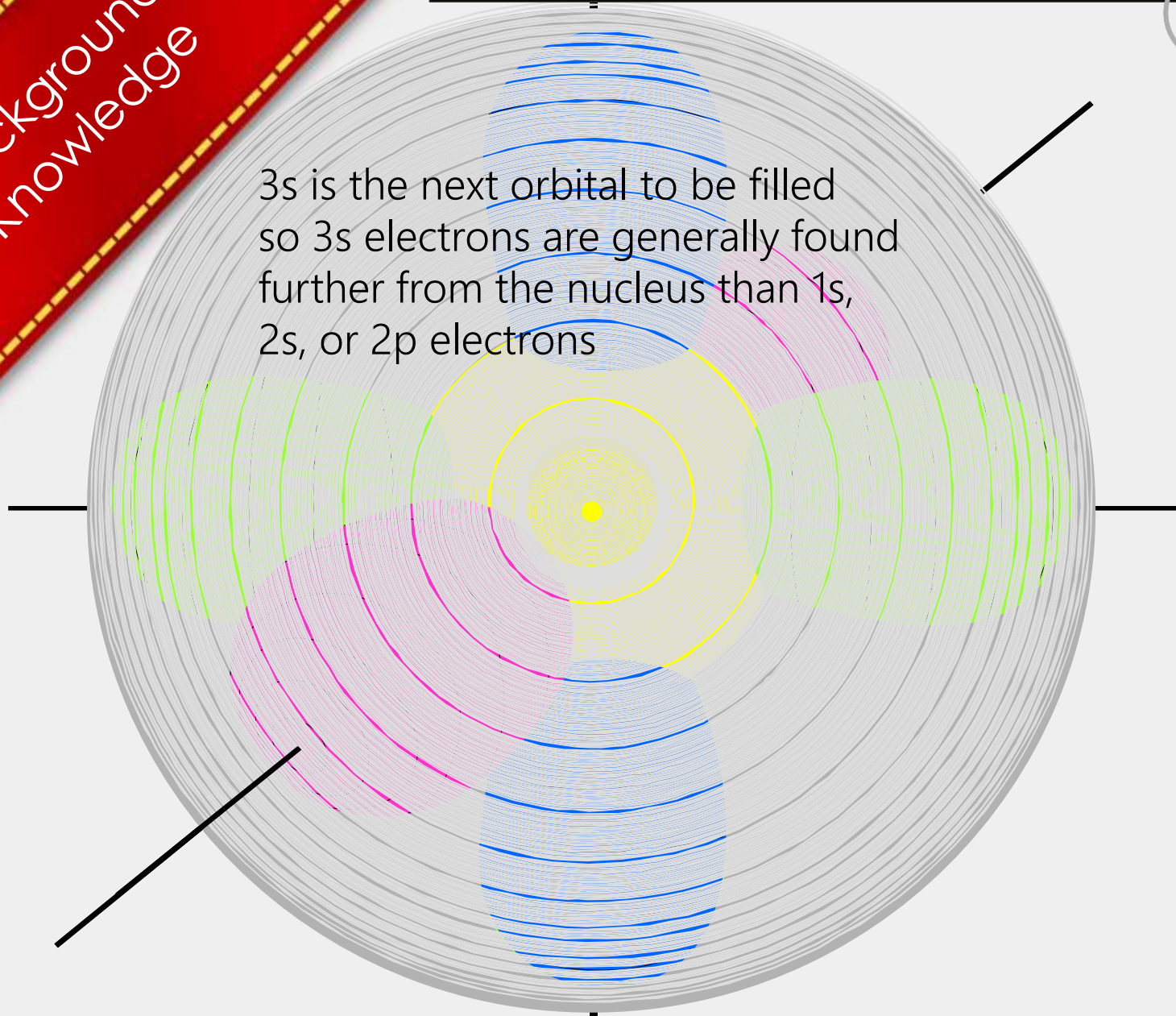


Each 2p orbital gets one electron before pairing begins

# Electron orbital shapes

Background  
Knowledge

3s is the next orbital to be filled  
so 3s electrons are generally found  
further from the nucleus than 1s,  
2s, or 2p electrons



## NCEA 2013 s,p,d configurations

Achieved  
Question

Q 1a: Complete the following table

Symbol	Electron configuration
Se	Se: $[\text{Ar}]3d^{10}4s^24p^4$ or $4s^23d^{10}4p^4$
V	V: $[\text{Ar}]3d^34s^2$ or $4s^23d^3$
$\text{V}^{3+}$	$\text{V}^{3+}$ : $[\text{Ar}]3d^2$ where $[\text{Ar}]$ : $1s^22s^22p^63s^23p^6$

Shorthand is acceptable  
spd are drawn in order not fill

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al   Cl    $\text{Cu}^{2+}$    Sc   Zn   K   Cr    $\text{Cr}^{3+}$    As   Se   V    $\text{V}^{3+}$   
 $\text{Ca}^{2+}$     $\text{Mn}^{2+}$

# NCEA 2014 s,p,d configurations

Merit  
Question

Q 1a: Complete the following table

Symbol	Electron configuration
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ [Ar]
Cr	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ [Ar] $3d^5 4s^1$
As	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$ [Ar] $3d^{10} 4s^2 4p^3$

Shorthand is acceptable  
spd are drawn in order not fill

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al Cl  $\text{Cu}^{2+}$  Sc Zn K Cr  $\text{Cr}^{3+}$  As Se V  $\text{V}^{3+}$   
 $\text{Ca}^{2+}$   $\text{Mn}^{2+}$



# NCEA 2015 s,p,d configurations

Achieved  
Question

Q 1a: Complete the following table

Symbol	Electron configuration		
Al	Al	$= [\text{Ne}] 3s^2 3p^1$	where $[\text{Ne}] = 1s^2 2s^2 2p^6$
$\text{Cu}^{2+}$	$\text{Cu}^{2+}$	$= [\text{Ar}] 3d^9$	
Sc	Sc	$= [\text{Ar}] 3d^1 4s^2$	where $[\text{Ar}] = 1s^2 2s^2 2p^6 3s^2 3p^6$

Shorthand is acceptable  
spd are drawn in order not fill

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al   Cl    $\text{Cu}^{2+}$    Sc   Zn   K   Cr    $\text{Cr}^{3+}$    As   Se   V    $\text{V}^{3+}$   
 $\text{Ca}^{2+}$     $\text{Mn}^{2+}$

## NCEA 2016 s,p,d configurations

Achieved  
Question

Q 1a: Complete the following table

Symbol	Electron configuration
Cl	Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$
Zn	Zn: $[\text{Ar}] 3d^{10} 4s^2$
$\text{Cr}^{3+}$	$\text{Cr}^{3+}$ : $[\text{Ar}] 3d^3$

Shorthand is acceptable  
spd are drawn in order not fill

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al   Cl    $\text{Cu}^{2+}$    Sc   Zn   K   Cr    $\text{Cr}^{3+}$    As   Se   V    $\text{V}^{3+}$   
 $\text{Ca}^{2+}$     $\text{Mn}^{2+}$

# NCEA 2017 s,p,d configurations

Achieved  
Question

Q 1a: Complete the following table

Symbol of particle	Electron configuration (use <i>s</i> , <i>p</i> , <i>d</i> notation)	Charge	Atomic number
Cl	Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$	0	17
$Ca^{2+}$	$1s^2 2s^2 2p^6 3s^2 3p^6$	+2	20
$Mn^{2+}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$	+2	25

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al   Cl    $Cu^{2+}$    Sc   Zn   K   Cr    $Cr^{3+}$    As   Se   V    $V^{3+}$   
 $Ca^{2+}$     $Mn^{2+}$

## NCEA 2018 s,p,d configurations

Achieved  
Question

Q 1a: Complete the following table

Symbol	Electron configuration (use <i>s</i> , <i>p</i> , <i>d</i> notation)
V	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$
$\text{Cu}^+$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$
$\text{Br}^-$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

In the past 5 years the spd configuration is a "stand alone" question – which case there is an Achieved point given for the majority correct. (or sometimes Merit if all correct)

In past years the following elements or ions have been used:

Al   Cl    $\text{Cu}^{2+}$    Sc   Zn   K   Cr    $\text{Cr}^{3+}$    As   Se   V    $\text{V}^{3+}$   
 $\text{Ca}^{2+}$     $\text{Mn}^{2+}$

## NCEA 2019 s,p,d configurations

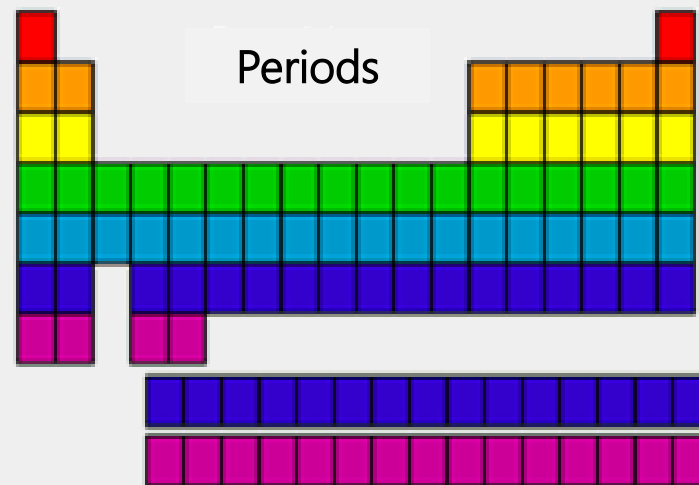
Achieved  
Question

Question 1a: Complete the following table.

Symbol	Electron configuration (use <i>s</i> , <i>p</i> , <i>d</i> notation)
Cr	Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ or [Ar] $4s^1 3d^5$
Fe <sup>3+</sup>	Fe <sup>3+</sup> : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ or [Ar] $3d^5$
Ge	Ge: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$ or [Ar] $3d^{10} 4s^2 4p^2$



The rows (across) of a periodic table are called periods.



Elements in the same **period** all have the same number of shells of electrons in their atoms

# Periodic Table

1

Atomic Number

1

Name

1

Symbol

1.0

Atomic mass

gas

liquid

solid

synthetic

Metals

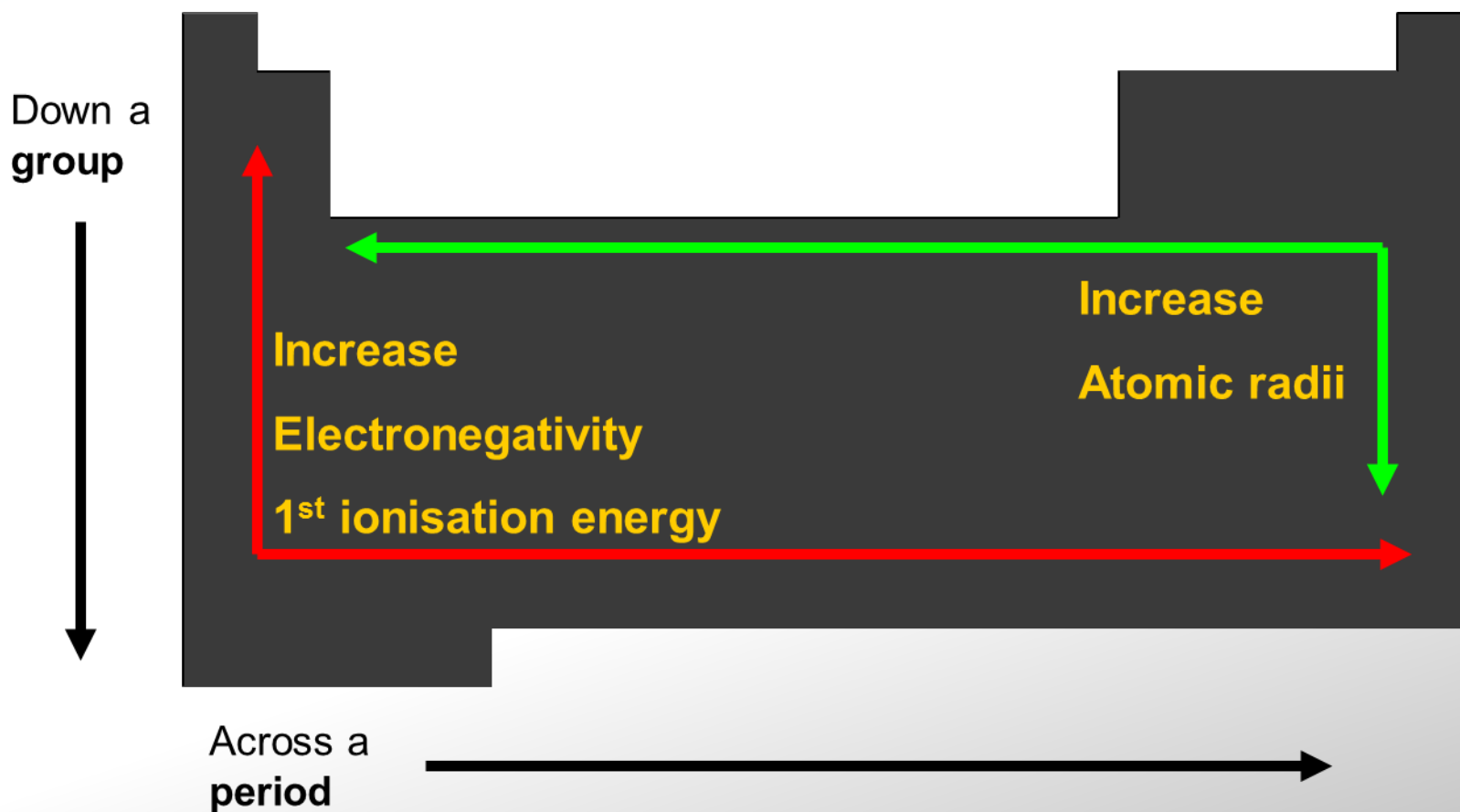
Semi-Metals

Non-Metals

1	2											13	14	15	16	17	18
1 H Hydrogen 1.0																	2 He Helium 4.0
2 Li Lithium 6.9	4 Be Beryllium 9.0											5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.2
11 Na Sodium 23.0	12 Mg Magnesium 24.3											13 Al Aluminum 27.0	14 Si Silicon 28.1	15 P Phosphorus 31.0	16 S Sulfur 32.0	17 Cl Chlorine 35.5	18 Ar Argon 40.0
Group																	
3	4	5	6	7	8	9	10	11	12								
19 K Potassium 39.1	20 Ca Calcium 40.1	21 Sc Scandium 45.0	22 Ti Titanium 47.9	23 V Vanadium 50.9	24 Cr Chromium 52.0	25 Mn Manganese 54.9	26 Fe Iron 55.9	27 Co Cobalt 58.9	28 Ni Nickel 58.7	29 Cu Copper 63.6	30 Zn Zinc 65.4	31 Ga Gallium 69.7	32 Ge Germanium 72.6	33 As Arsenic 74.9	34 Se Selenium 78.9	35 Br Bromine 79.9	36 Kr Krypton 83.8
37 Rb Rubidium 85.5	38 Sr Strontium 87.6	39 Y Yttrium 88.9	40 Zr Zirconium 91.2	41 Nb Niobium 92.9	42 Mo Molybdenum 95.9	43 Tc Technetium 98	44 Ru Ruthenium 101	45 Rh Rhodium 103	46 Pd Palladium 106	47 Ag Silver 108	48 Cd Cadmium 112	49 In Indium 115	50 Sn Tin 119	51 Sb Antimony 122	52 Te Tellurium 128	53 I Iodine 127	54 Xe Xenon 131
55 Cs Caesium 133	56 Ba Barium 137		72 Hf Hafnium 179	73 Ta Tantalum 178	74 W Tungsten 184	75 Re Rhenium 186	76 Os Osmium 190	77 Ir Iridium 192	78 Pt Platinum 195	79 Au Gold 197	80 Hg Mercury 201	81 Tl Thallium 204	82 Pb Lead 207	83 Bi Bismuth 209	84 Po Polonium 210	85 At Astatine 210	86 Rn Radon 222
87 Fr Francium 223	88 Ra Radium 226		104 Rf Rutherfordium 261	105 Db Dubnium 262	106 Sg Seaborgium 263	107 Bh Bohrium 262	108 Hs Hassium 265	109 Mt Meitnerium 266	110 Ds Darmstadtium 266	111 Rg Roentgenium 280	112 Cn Copernicium 285	113 Nh Nihonium 286	114 Fl Flerovium 289	115 Mc Moscovium 289	116 Lv Livermorium 293	117 Ts Tennesse 294	118 Og Oganesson 294
Alkali Metals		Alkaline Earth	Transition Metals									Basic Metals			Halogens		Inert Gases
Lanthanides		57 La Lanthanum 139	58 Ce Cerium 140	59 Pr Praseodymium 141	60 Nd Neodymium 144	61 Pm Promethium 147	62 Sm Samarium 150	63 Eu Europium 152	64 Gd Gadolinium 157	65 Tb Terbium 159	66 Dy Dysprosium 163	67 Ho Holmium 165	68 Er Erbium 167	69 Tm Thulium 169	70 Yb Ytterbium 173	71 Lu Lutetium 175	
Actinides		89 Ac Actinium 227	90 Th Thorium 232	91 Pa Protactinium 231	92 U Uranium 238	93 Np Neptunium 237	94 Pu Plutonium 239	95 Am Americium 241	96 Cm Curium 247	97 Bk Berkelium 249	98 Cf Californium 251	99 Es Einsteinium 254	100 Fm Fermium 257	101 Md Mendelevium 258	102 No Nobelium 256	103 Lr Lawrencium 262	

## Periodic trends

Periodic trends can be seen in atoms and ions as you move across a period and down a group. The trends investigated will include Atomic and ionic radii (distance from the centre of the nucleus to the valence energy level), electronegativity and the 1<sup>st</sup> ionization energy



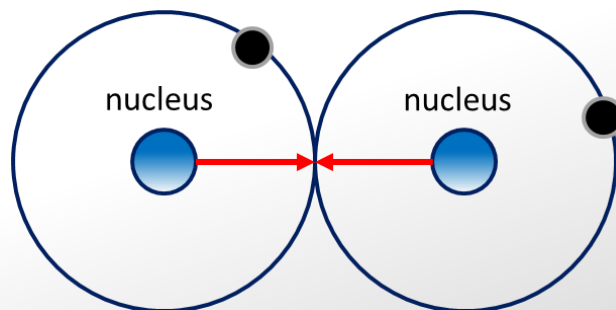
# Atomic Radii

Atomic Radii is affected by two main factors :

(1) **nuclear charge (number of protons)** : The stronger the 'pull' the protons have to the electrons with electrostatic attraction, then the smaller the size of the atom radii

(2) **number of energy levels**: The greater the number of energy levels, the larger the atomic radii. The internal energy levels "shield" and reduce electrostatic attraction of the valence electrons to the protons. Addition of another electron does not result in a fractional decrease in the electrostatic attraction to any **given** electron, but it does increase the electron-electron repulsion, so an **overall decrease in Nett attractive force**. [The most significant factor is nuclear charge – atoms are neutral and as the number of protons increases, the number of electrons increases at the same rate. The increased numbers of electrons in the valence shell, also have an increased number of equivalent protons, and they will be pulled in tighter and therefore show a smaller atomic radii.]

A 3<sup>rd</sup> factor that is also important: The electrons repel each other (**electron-electron repulsion**) and spread out to maximum distance in each energy level.



Atomic radius is half the distance between the nuclei of two bonded atoms

## Atomic Radii Trends

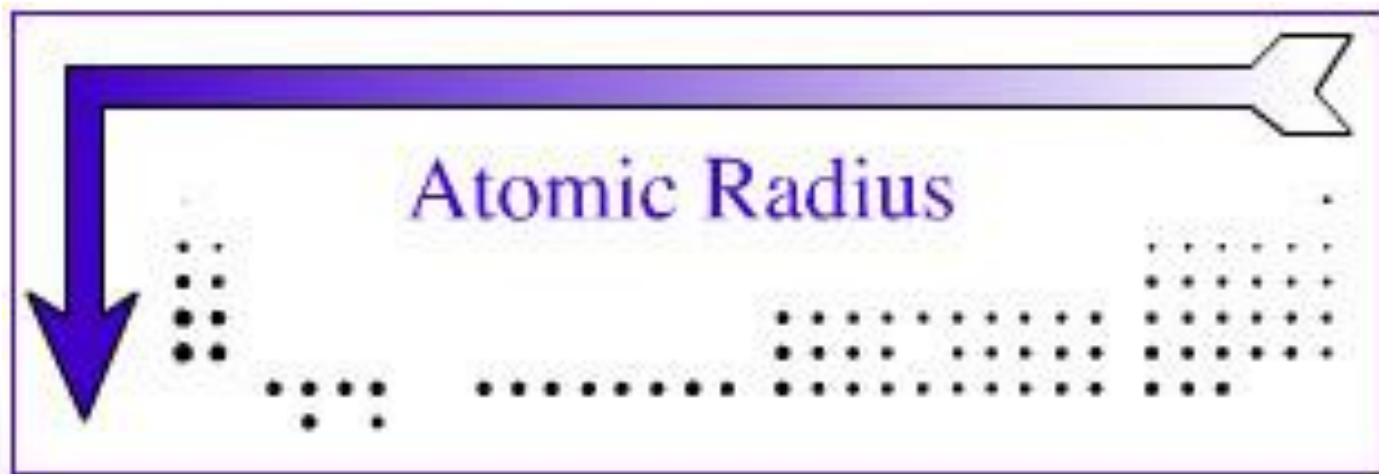
Addition of another electron does increase the electron-electron repulsion, but this is overcome with increased nuclear charge which effectively increases the **net attractive force**



Across the period the Atomic radii **decreases**

e.g. Li → Ne in **period 2**

As the nuclear charge increases across the period, so does the electrostatic attraction and so outer electrons are pulled closer to nucleus. The electron repulsions are balanced by the nuclear charge attractions, and as the nuclear charge gets larger, so the electrons get closer together. The net attractive electrostatic attraction is increased as the nuclear charge increases.





# Atomic Radii trends

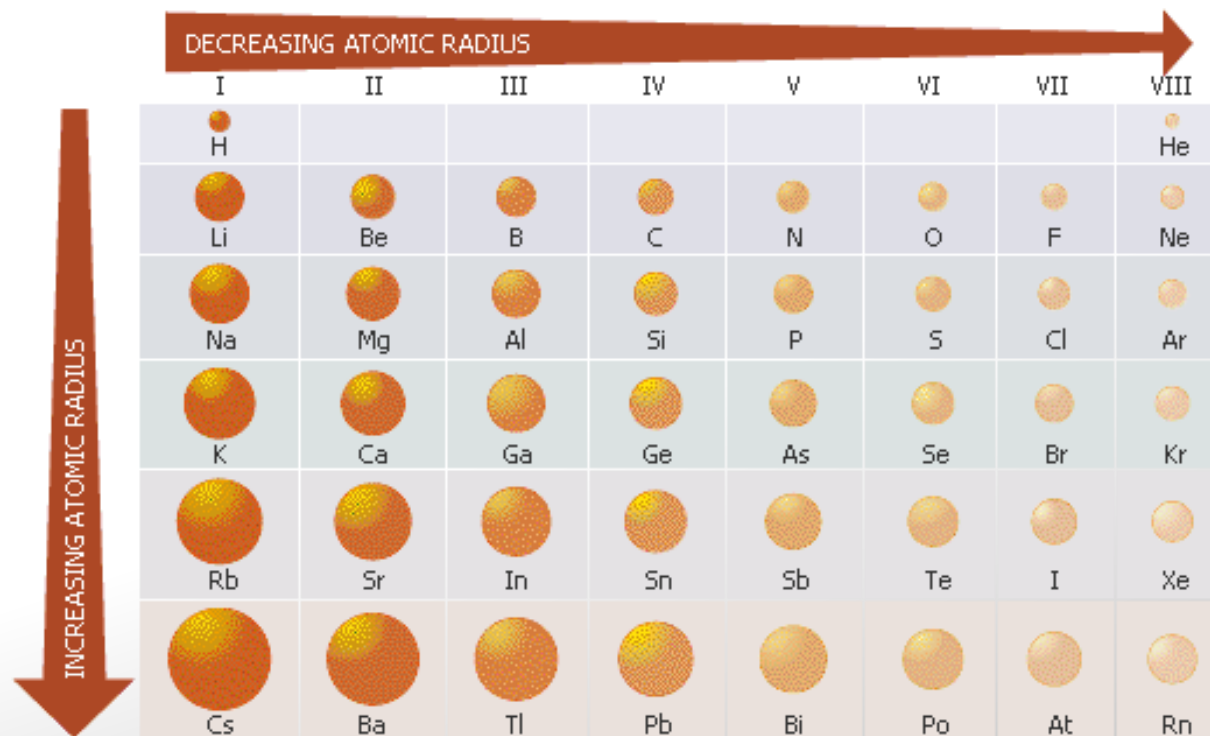
Down the group the Atomic radii **increases**

Addition of another electron does not result in a fractional decrease in the electrostatic attraction "shielding" to any given electron, but it does increase the electron-electron repulsion which effectively reduces the **nett attractive force**

**i**  
**extra**  
**info**

e.g. Li → Fr in **group one**

Electrons are being added to successive shells and both charge on nucleus and electron repulsion increase in step to "cancel each other out". However, successive shells are further from the nucleus therefore, there is a subtle increase in atomic radii and an overall decrease in Net electrostatic attraction.



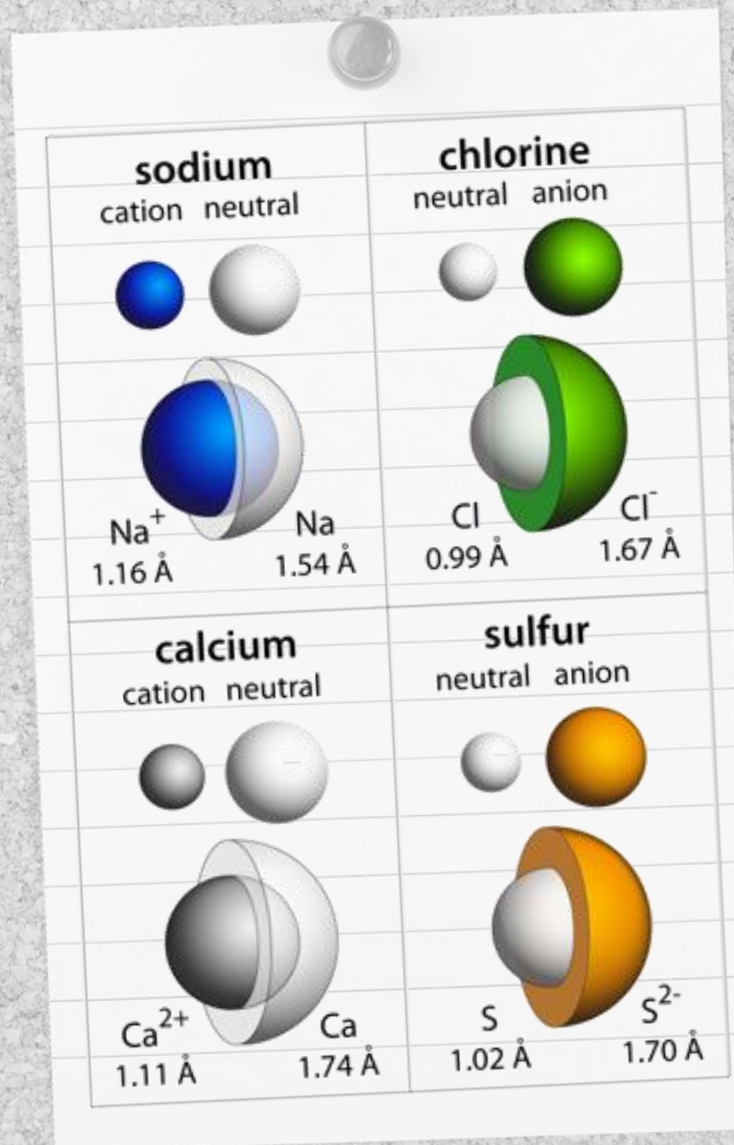
# Ionic Radii

**Cations (metal ions) are smaller in radii than their atoms**

The outside energy level of electrons are removed but the nuclear charge (number of protons) remains the same creating **smaller radii than the atom**

**Anions (non-metal ions) are larger in radii than their atoms**

Extra electrons are added to the outside valance shell that have to be accommodated for and there is no change to the nuclear charge. Electron-electron repulsion spreads the electrons out further creating **larger radii than the atom**

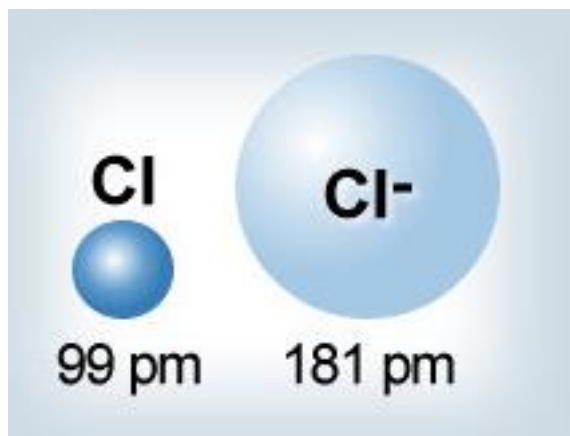
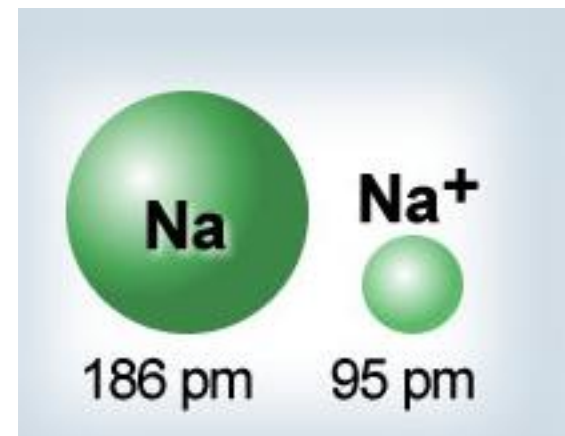


There is less of a trend seen across a periodic table of ions

## Ionic Radii trends

### Cations i.e. Na to $\text{Na}^+$

The inter-electron repulsion experienced by the electron cloud of the cation is **less** than the neutral atom (less electrons than protons), and since both species have the same amount of nuclear charge, the net electrostatic attractive force on the electron cloud in the cation is greater than the neutral atom resulting in a smaller cation size.



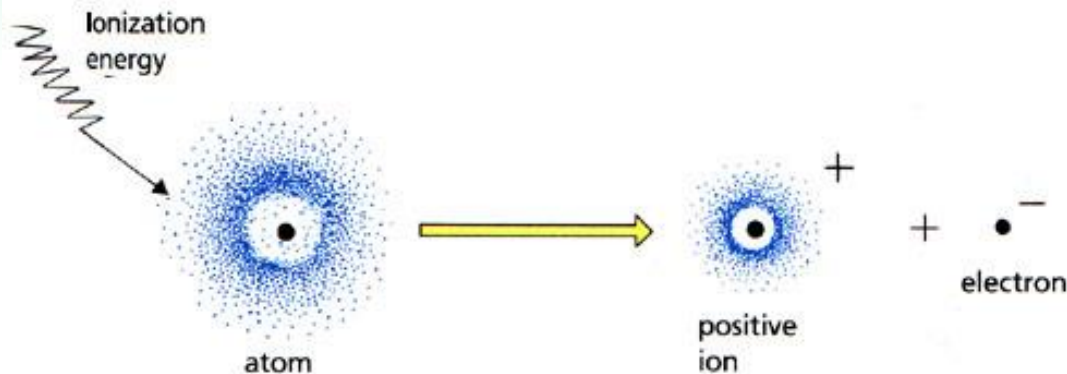
### Anions i.e. Cl to $\text{Cl}^-$

The inter-electron repulsion experienced by the electron cloud of the anion is **greater** than the neutral atom (more electrons than protons), and since both species have the same amount of nuclear charge, the net electrostatic attractive force on the electron cloud in the anion is less than the neutral atom resulting in a larger anion size.



# 1<sup>st</sup> ionisation energy

The first ionisation energy (1<sup>st</sup> I.E.) is the energy required to remove one mole of electrons from the outside valance shell of 1 mole of atoms in a gaseous state.



If the ionisation energy is high, that means it takes a lot of energy to remove the outermost electron. If the ionisation energy is low, that means it takes only a small amount of energy to remove the outermost electron.

Ionisation energy is affected by two factors:

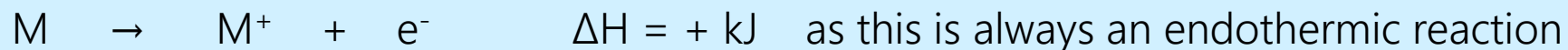
- (1) **Nuclear charge:** As NC increases, there is a stronger pull to the electrons by electrostatic attraction → increased 1<sup>st</sup> I.E. with increased NC
- (2) **Number of energy levels:** Electrons in a lower energy level are much closer to the nucleus and thus have much stronger net electrostatic attraction to it. Electrons in a lower energy level shell have electron repulsion but are closer together. Electrons in higher energy level shells are experience less net electrostatic attraction to the nucleus, as they are further away → decreased 1<sup>st</sup> I.E with increased number energy levels

# 1<sup>st</sup> ionisation energy

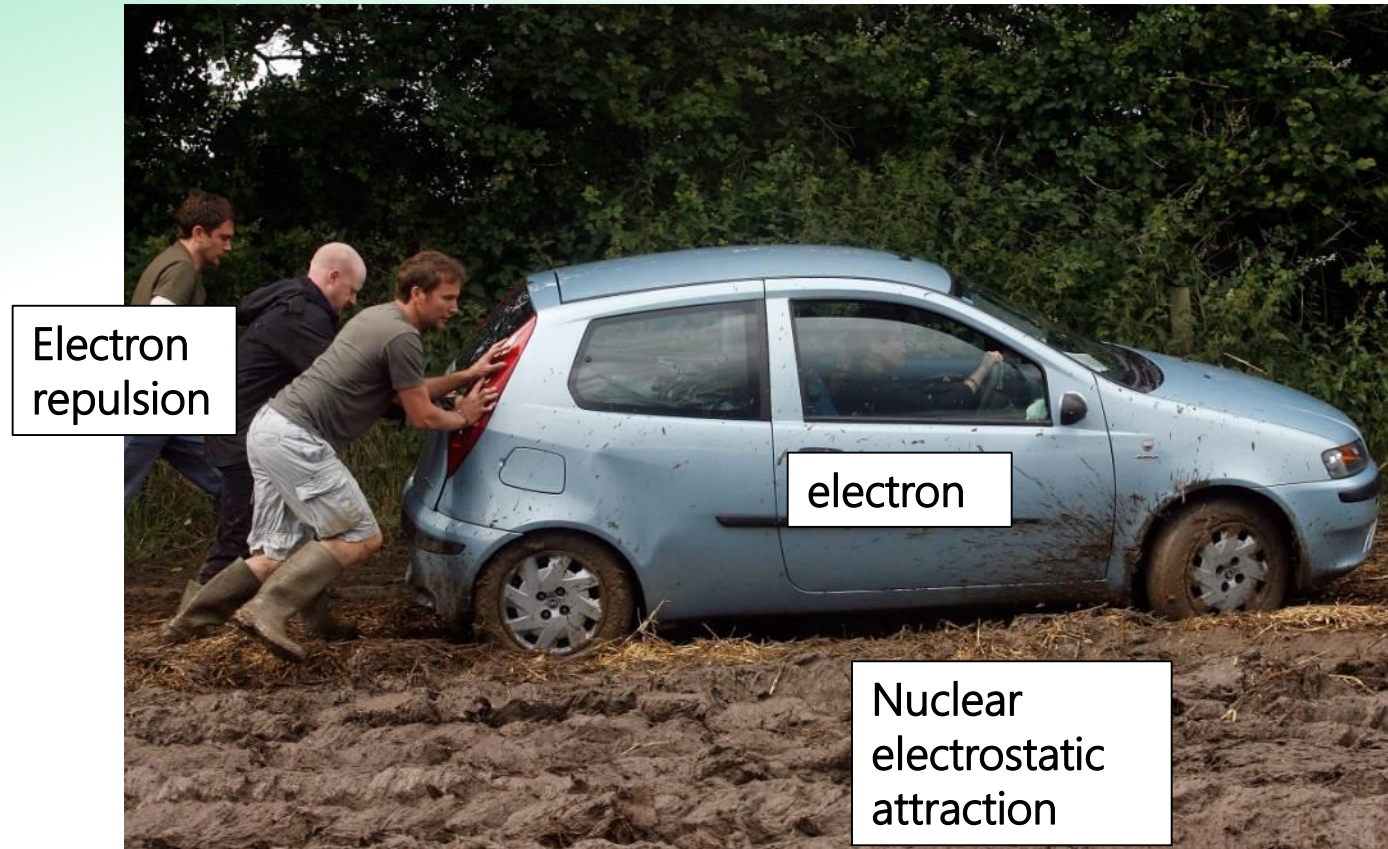
1																	18							
1	<b>H</b> 1312.0																	<b>He</b> 2372.3						
2	<b>Li</b> 520.2	<b>Be</b> 899.5																	<b>B</b> 800.6	<b>C</b> 1086.5	<b>N</b> 1402.3	<b>O</b> 1313.9	<b>F</b> 1681.0	<b>Ne</b> 2080.7
3	<b>Na</b> 495.8	<b>Mg</b> 737.7																	<b>Al</b> 577.5	<b>Si</b> 786.5	<b>P</b> 1011.8	<b>S</b> 999.6	<b>Cl</b> 1251.2	<b>Ar</b> 1520.6
4	<b>K</b> 418.8	<b>Ca</b> 589.8	<b>Sc</b> 633.1	<b>Ti</b> 658.8	<b>V</b> 650.9	<b>Cr</b> 652.9	<b>Mn</b> 717.3	<b>Fe</b> 762.5	<b>Co</b> 760.4	<b>Ni</b> 737.1	<b>Cu</b> 745.5	<b>Zn</b> 906.4	<b>Ga</b> 578.8	<b>Ge</b> 762.2	<b>As</b> 944.5	<b>Se</b> 941.0	<b>Br</b> 1139.9	<b>Kr</b> 1350.8						
5	<b>Rb</b> 403.0	<b>Sr</b> 549.5	<b>Y</b> 599.9	<b>Zr</b> 640.1	<b>Nb</b> 652.1	<b>Mo</b> 684.3	<b>Tc</b> 702	<b>Ru</b> 710.2	<b>Rh</b> 719.7	<b>Pd</b> 804.4	<b>Ag</b> 731.0	<b>Cd</b> 867.8	<b>In</b> 558.3	<b>Sn</b> 708.6	<b>Sb</b> 830.6	<b>Te</b> 869.3	<b>I</b> 1008.4	<b>Xe</b> 1170.3						
6	<b>Cs</b> 375.7	<b>Ba</b> 502.9	<b>La</b> 538.1	<b>Hf</b> 658.5	<b>Ta</b> 728.4	<b>W</b> 758.8	<b>Re</b> 755.8	<b>Os</b> 814.2	<b>Ir</b> 865.2	<b>Pt</b> 864.4	<b>Au</b> 890.1	<b>Hg</b> 1007.1	<b>Tl</b> 589.4	<b>Pb</b> 715.6	<b>Bi</b> 703.0	<b>Po</b> 812.1	<b>At</b>	<b>Rn</b> 1037.1						
7	<b>Fr</b> 393.0	<b>Ra</b> 509.3	<b>Ac</b> 498.8	<b>Rf</b> 580	<b>Db</b>	<b>Sg</b>	<b>Bh</b>	<b>Hs</b>	<b>Mt</b>	<b>Ds</b>	<b>Rg</b>	<b>Uub</b>	<b>Uut</b>	<b>Uuq</b>	<b>Uup</b>									
Lanthanides				<b>Ce</b> 534.4	<b>Pr</b> 528.1	<b>Nd</b> 533.1	<b>Pm</b> 538.6	<b>Sm</b> 544.5	<b>Eu</b> 547.1	<b>Gd</b> 593.4	<b>Tb</b> 565.8	<b>Dy</b> 573.0	<b>Ho</b> 581.0	<b>Er</b> 589.3	<b>Tm</b> 596.7	<b>Yb</b> 603.4	<b>Lu</b> 523.5							
Actinides				<b>Th</b> 608.5	<b>Pa</b> 568	<b>U</b> 597.6	<b>Np</b> 604.5	<b>Pu</b> 581.4	<b>Am</b> 576.4	<b>Cm</b> 578.1	<b>Bk</b> 598.0	<b>Cf</b> 606.1	<b>Es</b> 619	<b>Fm</b> 627	<b>Md</b> 635	<b>No</b> 642	<b>Lr</b> 472.8							

In general Atoms that lose electrons easily have low ionisation energy and atoms that gain electrons have high ionisation energy

An Ionisation equation can be written as: (M representing an atom)



# 1<sup>st</sup> ionisation energy Analogy



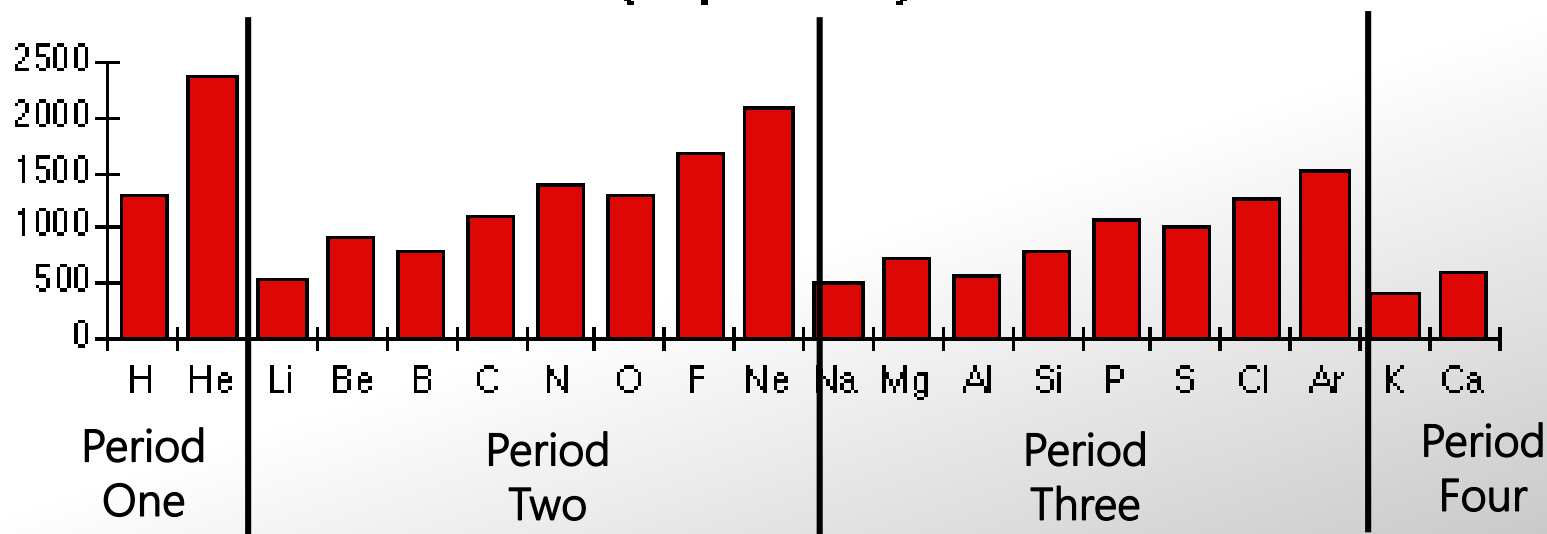
In order to remove an electron from an atom you need to overcome the nuclear attraction of its protons holding it around the nucleus. This can be shown by the mud that the car is stuck in. The more mud (nuclear attraction) the more energy to remove the car (electron). However, other electrons in the atom are repelling the electron to be removed – so the more people pushing the car (electron repulsion) the easier it is to extract the car (electron)

## 1<sup>st</sup> ionisation energy trends

Across a period the 1<sup>st</sup> ionisation energy **increases**

As the **nuclear charge increases**, the attraction between the nucleus and the electrons increases and it requires more energy to remove an electron from the outermost energy level and that means there is a higher ionisation energy. As you go across the periodic table, nuclear charge is the most important consideration. Therefore, going across the periodic table, there should be an increase in ionisation energy because of the increasing nuclear charge.

**First ionisation energies from hydrogen to calcium  
(kJ per mole)**





## 1<sup>st</sup> ionisation energy trends

All bonding is electrostatic and is characterised by **Coulombs Law** ( $F = kQq/r^2$ ) which is an inverse square law so distance from nucleus really does matter



Down a group the 1<sup>st</sup> ionisation energy **decreases**

Going down the table, the effect of increased nuclear charge is weighed against the effect of **increased electron repulsion**, and the number of energy levels becomes the predominant factor. With more energy levels, the outermost electrons (the valence electrons) are **further from the nucleus** and are not so strongly attracted to the nucleus, and therefore there is a reduction in net electrostatic attraction. Thus, the ionisation energy of the elements decreases as you go down the periodic table because it is easier to remove the electrons. The more stable elements have higher ionisation energies.



The more stable elements have higher ionisation energies.

## Sample questions – Ionisation energy (Part A)

For each of the following sets of atoms, decide which has the highest and lowest ionisation energies and why.

### S, Si, Mg (period 3)

All are in the same period and use the same number of energy levels. Mg has the lowest I.E. because it has the lowest effective nuclear charge. S has the highest I.E. because it has the highest effective nuclear charge.

### Mg, Ca, Ba (group 2)

All are in the same group and have the same effective nuclear charge. Mg has the highest I.E. because it uses the smallest number of energy levels. Ba has the lowest I.E. because it uses the largest number of energy levels.

### F, Cl, Br (group 17)

All are in the same group and have the same effective nuclear charge. F has the highest I.E. because it uses the smallest number of energy levels. Br has the lowest I.E. because it uses the largest number of energy levels.

## Sample questions – Ionisation energy (Part B)

For each of the following sets of atoms, decide which has the highest and lowest ionisation energies and why.

**Where trends conflict e.g. IE down a group, the increased radius factor will outweigh the increased nuclear charge because it is a squared function.**

**Ne, Cu, Ba**

All are in different groups and periods, so both factors must be considered. Fortunately both factors reinforce one another. Ba has the lowest I.E. because it has the lowest effective nuclear charge and uses the highest number of energy levels. Ne has the highest I.E. because it has the highest effective nuclear charge and uses the lowest number of energy levels.

**N, P, Si**

Si has the lowest I.E. because it has the lowest effective nuclear charge and is tied (with P) for using the most energy levels. N has the highest I.E. because it uses the fewest energy levels and is tied (with P) for having the highest effective nuclear charge.

# Electronegativity

Electronegativity is the **tendency of an atom to attract bonding electrons from another atom**. Higher electronegativity values mean a higher tendency to attract electrons. Atoms with high E.N. are strong oxidants (gain electrons).

Electronegativity is affected by two factors:

**Nuclear charge:** As an atoms nuclear charge increases, there is a stronger pull on electrons of another atom by electrostatic attraction.

**Number of energy levels:** the more energy levels an atom has the lower the net electrostatic attraction and the radii of the atom is larger. Because this then creates a bigger distance between 'neighbouring' atoms, Electrons from other atoms experience less electrostatic attraction to the nucleus of another atom. Therefore an atom in the same group has less electronegativity than an atom above it with less energy levels( Even though it has more nuclear charge)

## Metals

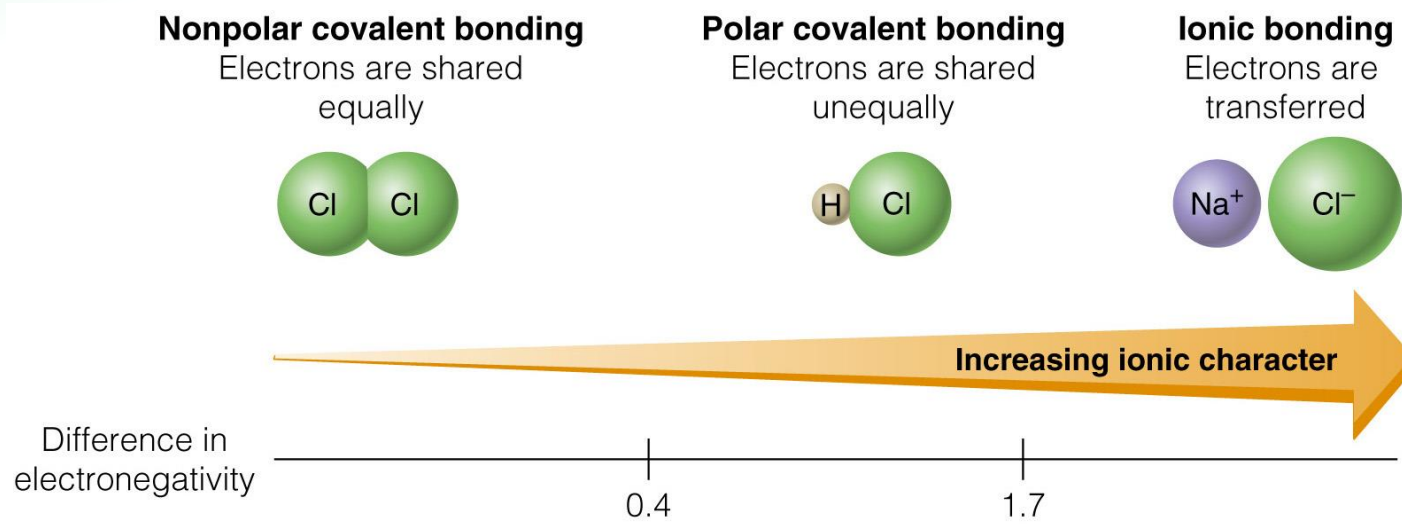
- ☐ tend to have lower E.N.
- ☐ Lose electrons to form cations
- ☐ Usually have fewer electrons in valance shell than  $\frac{1}{2}$  full

## Non-Metals

- ☐ tend to have higher E.N.
- ☐ gain electrons to form anions
- ☐ Usually have more electrons in valance shell than  $\frac{1}{2}$  full

# 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

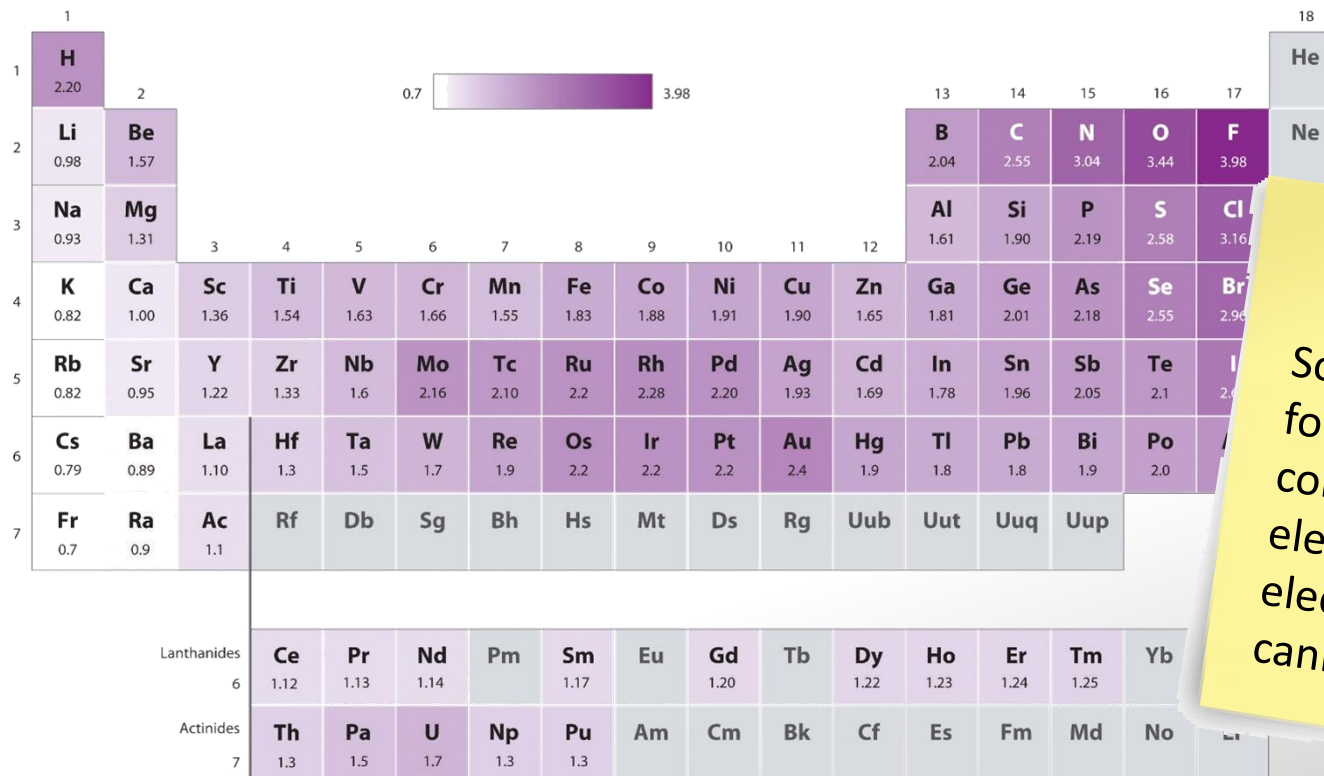


If there is little difference in electronegativity between two atoms then they tend to form a covalent bond with no polarity difference. A greater electronegativity difference creates a polar bond with uneven “sharing” of valance electrons. If the electronegativity is even greater then there will be 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.

# Electronegativity trends

Across a Period the electronegativity **increases**

e.g.  $\text{Li} \rightarrow \text{Ne}$ : The atoms have increased “pulling power” as the nuclear charge is increasing. Electrons are held tighter to the nucleus and there is a greater net electrostatic attraction. This allows another atom to be closer and it has a stronger attraction to electrons from that atom, so electronegativity increases.



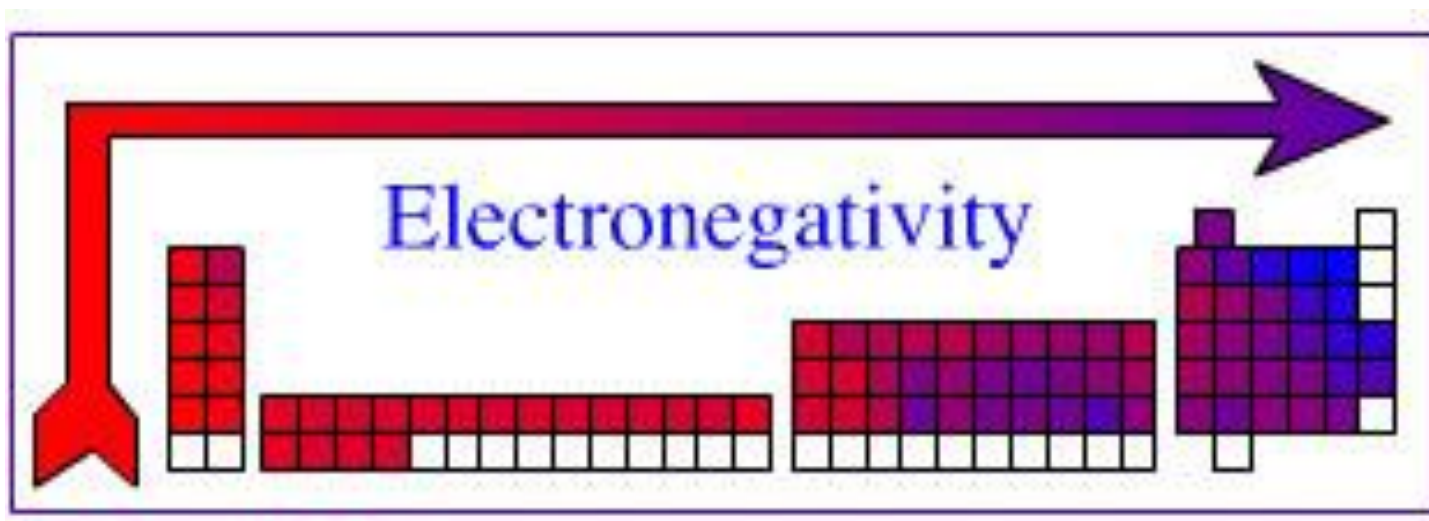
Some elements do not form stable covalent compounds with other elements; so their electronegativities cannot be measured

## Electronegativity trends

Down a group the electronegativity **decreases**

e.g.  $\text{Li} \rightarrow \text{Fr}$

Both nuclear attraction and electron repulsion increase in step, but with an overall decrease in Net electrostatic attraction. However, as successive shells increase atomic radii, then the electrostatic attraction of the nucleus to other atoms' electrons decreases, so atoms have less electronegativity as you move down a group.





## Periodic trends Summary

	Electronegativity	1 <sup>st</sup> ionisation energy	Atomic radii
Across a Period	The greater the nuclear charge, the easier it is to obtain more electrons from other atoms > more p+ to pull with <b>INCREASES</b>	As the nuclear charge is larger it requires more energy to remove an electron as they are held tighter (and closer) to the nucleus <b>INCREASES</b>	Across a period the energy level numbers stay the same but nuclear charge increases pulling more at the valence electrons <b>DECREASES</b>
Down a group	The larger the number of energy levels the less net electrostatic attraction, the further the neighbouring nucleus so the less ability an atoms has to remove the valance electrons of another atom <b>DECREASES</b>	As the energy level numbers increase the easier it is to remove electrons as the valence electrons are further from the 'pull' of the protons and less net electrostatic attraction (and greater radius) <b>DECREASES</b>	Down a group the energy levels become further away from the nucleus therefore the valence electrons repel more, less effected by protons, taking up more space and increasing the size <b>INCREASES</b>

## NCEA 2013 Periodic trends - (PART ONE)

Excellence  
Question

**Question 1b:** Discuss the data for each of the following pairs of particles.

Atom	Electronegativity
O	3.44
Se	2.55

$\text{Cl}^-$  has an extra electron in its outermost/same energy level. This causes increased repulsion between electrons in the valence shell, so the electrons move further apart. This makes  $\text{Cl}^-$  bigger than Cl. Both Cl and  $\text{Cl}^-$  have the same number of protons/attractive force of the nucleus remains the same.

Se has more shells/electrons in energy levels further from the nucleus than O, with increased shielding from inner shells. This means there is a weaker electrostatic attraction between the nucleus and the bonded electrons, so Se has a lower electronegativity than O.

Atom or ion	Radius/pm
Cl	99
$\text{Cl}^-$	181

## NCEA 2013 Periodic trends - (PART TWO)

Excellence  
Question

**Question 1b:** Discuss the data for each of the following pairs of particles.

Atom	First ionisation energy / $\text{kJ mol}^{-1}$
Li	526
Cl	1 257

Cl has more protons than Li. Therefore there is a greater attraction between the nucleus and outer electrons/electrons held more tightly so it is harder to remove an electron from Cl than Li.

Even though the valence electrons of Cl are in the 3rd energy level/has an extra energy level the extra shielding is not as significant as the effect of the increased nuclear charge, so Cl has a higher first ionisation energy than Li.

## NCEA 2014 Periodic trends

**Question 1c:** The following table shows the electron configurations of four atoms, He, B, N, and Ne. Arrange these atoms in order of increasing first ionisation energy by writing the symbol of the appropriate atom in the boxes below.

Atom	He	B	N	Ne
Electron configuration	$1s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^6$

Achieved  
Question



B



N



Ne



He

lowest  
ionisation energy

highest  
ionisation energy

## NCEA 2015 Periodic Trends

Excellence  
Question

**Question 1b:** Define the terms electronegativity and first ionisation energy.

**Answer 1b:** Electronegativity is the ability of an atom in a compound to **attract electrons to itself**.

First ionisation energy is the minimum energy required to **remove one mole of electrons from one mole of gaseous atoms**.

It is well worth learning the definitions of the periodic trend well as these two correct together were worth an Excellence grade

## NCEA 2014 Periodic Trends

Excellence  
Question

**Question 1b:** Explain the difference between the radii of the K atom and the  $K^+$  ion.

**Answer 1b:** The  $K^+$  ion has a smaller radius than the K atom, as the ion has lost an electron from the valence/outer energy level, and therefore has fewer shells. This results in greater attraction between the nucleus and the valence electrons, as the outer electrons are now closer to the nucleus. There is less repulsion between the remaining electrons. Both species have the same number of protons / amount of nuclear charge.

**Question 1c:** The following table shows the first ionisation energy values for elements in the third period of the periodic table. Justify the periodic trend of first ionisation energies shown by the data in the table above, and relate this to the expected trend in atomic radii across the third period.

Element	First ionisation energy / $\text{kJ mol}^{-1}$
Na	502
Al	584
Si	793
Ar	1 527

Both periodic trends are influenced by nuclear charge and the number of shells / distance, the ionisation energy increases while the atomic radii decrease.

The larger the ionisation energy the more strongly the valence electrons are held. Thus atomic radii across Period 3 decrease

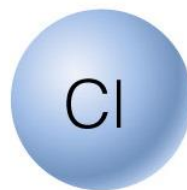
**Answer 1c:** First ionisation energy increases from 502 in Na to 1527 in Ar. There is an increase in the number of protons and thus the nuclear charge / attractive force of the nucleus. As the electrons are added to the same energy level, there is no increase in repulsion between energy levels. The nuclei with a greater number of protons have a stronger electrostatic attraction for the valence electrons in the third shell, thus the first ionisation energy increases across a period.

## NCEA 2016 Periodic trends

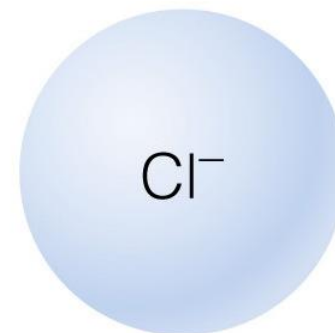
Merit  
Question

**Question 1b (i)** : Explain why the radius of the Cl atom and the radius of the Cl<sup>-</sup> ion are different.

	Radius (pm)
<b>Cl atom</b>	99
<b>Cl<sup>-</sup> ion</b>	181



Chlorine atom  
17 protons  
17 electrons  
**99 pm radius**



Chloride ion  
17 protons  
18 electrons  
**181 pm radius**

**Answer 1b (i)** : The Cl atom gains one electron to complete its valence shell to form the Cl<sup>-</sup> ion, the nuclear charge remains the same. The increased inter-electron repulsion in the outer energy level causes the valence electrons to move further from the nucleus, so the Cl<sup>-</sup> ion is larger than the Cl atom.



## NCEA 2016 Periodic trends

Excellence  
Question

**Question 1b (ii) :** Explain the factors influencing the trends in electronegativity and first ionisation energy down a group of the periodic table.

In your answer you should:

- ☐ define both electronegativity and first ionisation energy
- ☐ explain the trend in both electronegativity and first ionisation energy down a group
- ☐ compare the trend in electronegativity and first ionisation energy down a group.

**Electronegativity** decreases down a group. Electronegativity is a measure of how strongly an atom attracts bonding electrons. Although the nucleus will become increasingly positive down a group (number of protons increases), the atomic radius increases down a group as more energy levels are added and shielding / repulsion from inner shells increases. Therefore, the bonding electrons in the valence shell will be further from the positive nucleus, resulting in a weaker electrostatic attraction between the nucleus and the bonding electrons.

**First ionisation energy** is a measure of how easily the first mole of electrons is removed from one mole of gaseous atoms. It becomes easier to remove an electron down a group / first IE decreases down a group as the valence electrons are further from nucleus with greater repulsion / shielding from inner shells, so there is less electrostatic attraction between protons in the nucleus and valence electron to be removed. **For both EN and first IE, the attraction between the positive nucleus and bonding / valence electrons in the outer shell is decreasing down a group, so both EN and first IE decrease down a group.**

## NCEA 2017 Periodic Trends

MERIT  
Question

**Question 1b(i) :** Define the terms electronegativity and first ionisation energy.

**Answer 1b:** Electronegativity is the ability of an atom in a compound to **attract a bonding pair of electrons to itself**.

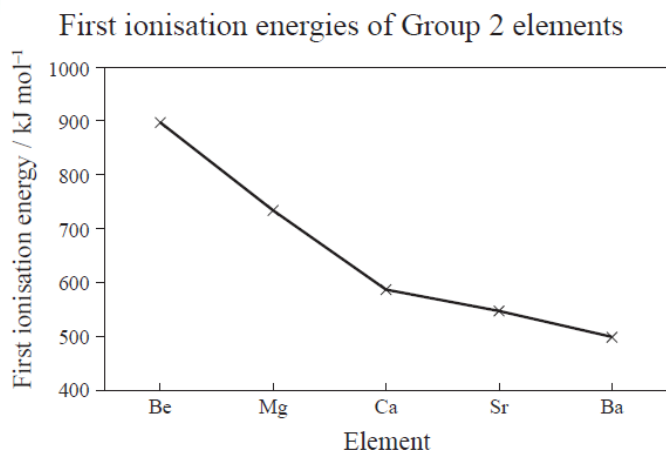
It is well worth learning the definitions of the periodic trend well as these two correct together were worth an Excellence grade

**(ii)** Explain why the electronegativity of chlorine is greater than that of phosphorus

EXCELLENCE  
Question

**Answer 1b (ii):** Electronegativity increases as you go across a period. Both Cl and P are row 3 elements and have valence electrons in their 3rd shell. The electrons are in the same shell so experience the same shielding effect. Chlorine has more protons in its nucleus than phosphorus so its nuclear charge is greater. This means that chlorine will have more attraction for the bonding pair of electrons so its electronegativity is greater.

**Question 1c:** The following graph shows the first ionisation energies of the Group 2 elements from Be to Ba.



(i) Write an equation to show the first ionisation energy for the element calcium.



(ii) Explain the trend shown of first ionisation energies of the Group 2 elements.

The first ionisation energy is the energy required to remove one mole of the most loosely held electrons from one mole of gaseous atoms.

The trend is that the ionisation energy decreases going down the group two elements.

Although the nuclear charge increases due to more protons in the atoms going down a group, it is offset by the increasing distance of the outer electrons from the nucleus as the atomic radius increases due to more energy levels being added.

The full inner energy levels shield the outer electrons from the protons in the nucleus so the electrostatic attraction is less. Additional energy levels result in greater shielding / repulsion between energy levels. The further the outer electron is from the nucleus, the less energy needed to remove it.

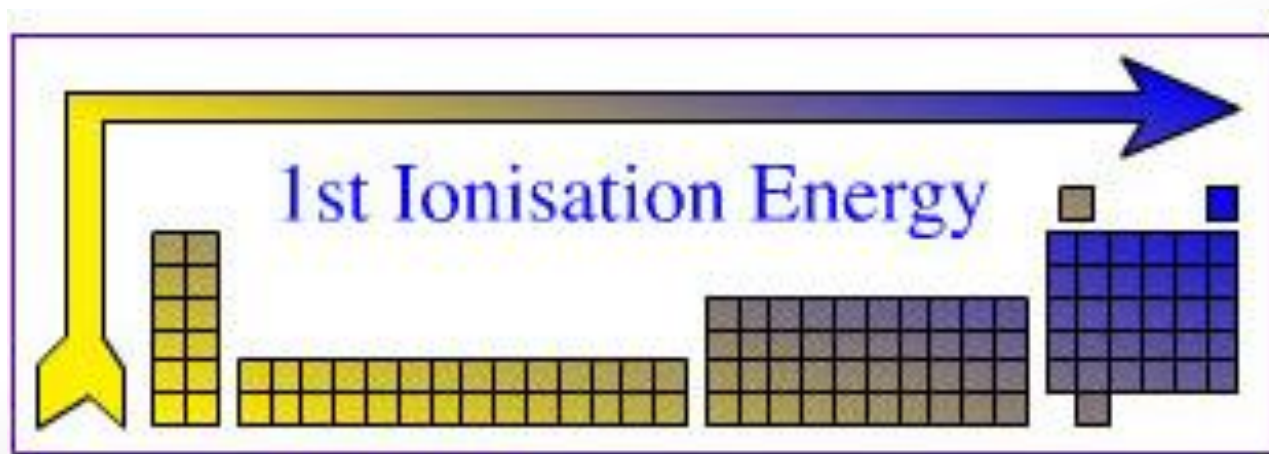
## NCEA 2018 Periodic Trends (part ONE)

**Q 1b:** Explain the factors influencing the trends in first ionisation energy and atomic radius across the second period of the periodic table.

In your answer, you should:

- describe the trends in both first ionisation energy
- explain the factors influencing the trends in first ionisation energy

The first ionisation energy increases across the second period. There is an increase in the number of protons therefore the nuclear charge / attractive force of the nucleus increases. As the electrons are added to the same energy level across the second period, the electrostatic attraction for the valence electrons increases. This means more energy is required to remove an electron from the valence shell.



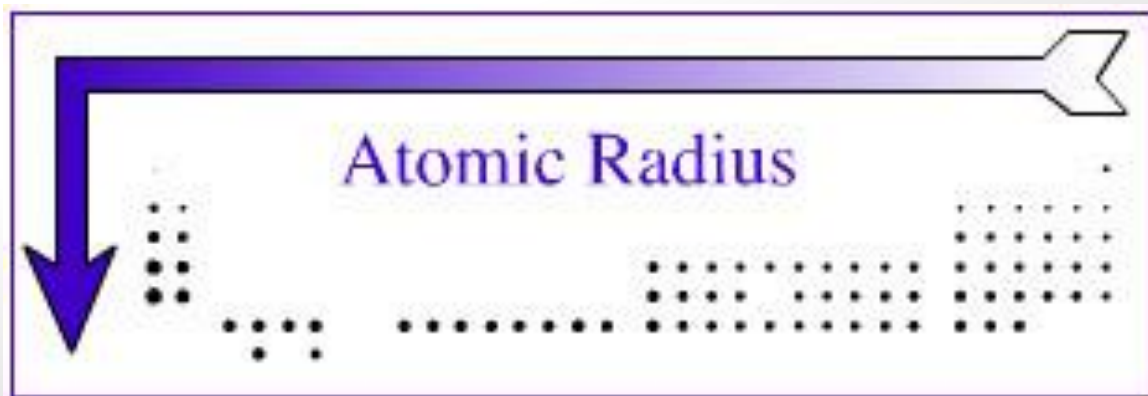
## NCEA 2018 Periodic Trends (part TWO)

**Q 1b:** Explain the factors influencing the trends in first ionisation energy and atomic radius across the second period of the periodic table. In your answer, you should:

- describe the trends in atomic radius across the second period
- explain the factors influencing the trends in atomic radius across the second period
- relate the trend in first ionisation energy to the trend in atomic radius.

The atomic radius decreases across the second period. There is an increase in the number of protons therefore the nuclear charge / attractive force of the nucleus increases. Electrons are added to the same energy level as well. This causes the electrostatic attraction between the positive nucleus and the valence electrons to increase across the period pulling the valence electrons closer to the nucleus, so the atomic radius decreases.

As the ionisation energy increases, the atomic radius decreases, this is due to the same factor of increased nuclear charge due to more protons in the nucleus going across the period whilst electrons are adding to the same energy level. This decreased radius means more energy is required to remove the valence electron due to stronger attractive forces.





## NCEA 2019 Periodic Trends

Merit  
Question

**Question 1c:** (i) Explain why the radii of the S atom and the  $S^{2-}$  ion are different.

Sulfur and the sulfur ion have the same number of proton/nuclear charge but when the sulfur atom gains two electrons to form the sulfur ion, there is increased electron-electron repulsion in the valence energy level. As a result, the electrons move further apart, and therefore the  $S^{2-}$  ion has a larger radius than the S atom.



**Question 1c:** (ii) Justify the difference in electronegativities for oxygen, sodium, and sulfur.

Element	Electronegativity
Oxygen, O	3.44
Sodium, Na	0.93
Sulfur, S	2.58

Electronegativity increases across a period, i.e. from Na to S. Both Na and S have the same number of energy levels and therefore the same shielding/electron-electron repulsion from inner levels. S has more protons/greater nuclear charge and therefore a greater attraction for valence/bonding electrons therefore greater electronegativity than Na.

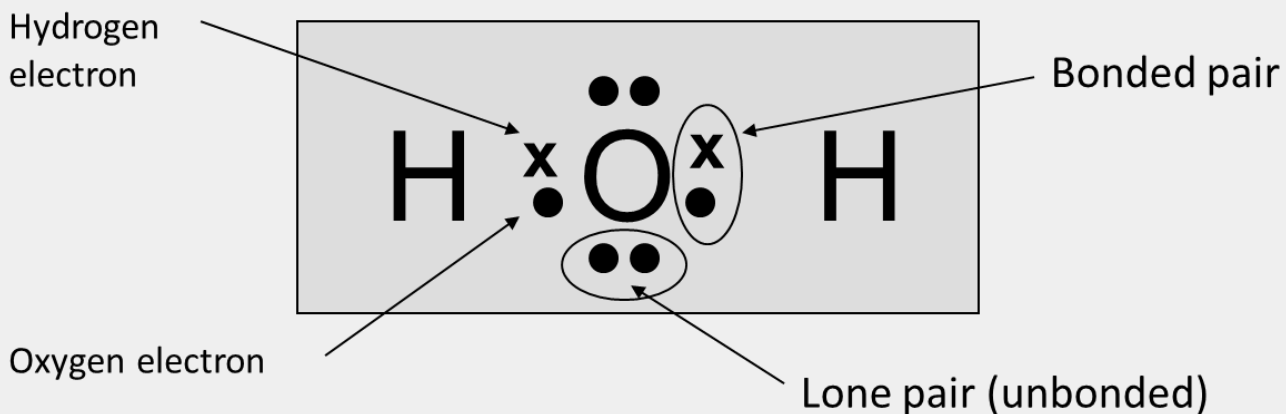
Electronegativity decreases down a group. Sulfur has one more energy level and therefore increased shielding/electron-electron repulsion. Even though S has greater nuclear charge/more protons than O, because the valence electrons are further from the nucleus, electronegativity is lower.



## Drawing Lewis Diagrams

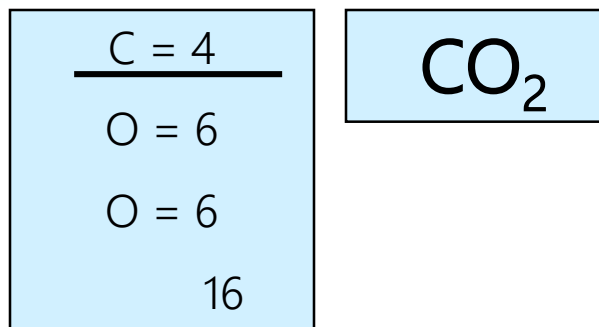
G Lewis devised a system of drawing covalent molecules showing arrangement of atoms and valence electrons – both those involved in bonding and those that are not (called lone pairs). Electrons in inner shells are not involved in bonding. These diagrams are called **Lewis structures (or diagrams)**. The Lewis diagram 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<sub>2</sub>O (water)



## Lewis Diagrams – Steps to drawing

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 diagram at the end. *Example carbon dioxide.*



2. Write down number of pairs of electrons.

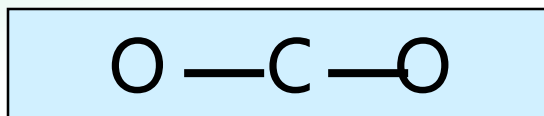
$$16 / 2 = 8 \text{ pairs}$$

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



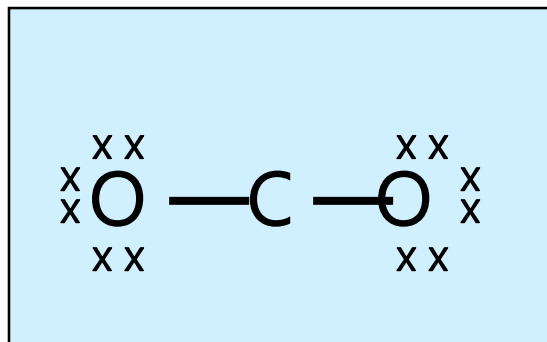
## Lewis Diagrams - steps to drawing

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



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.



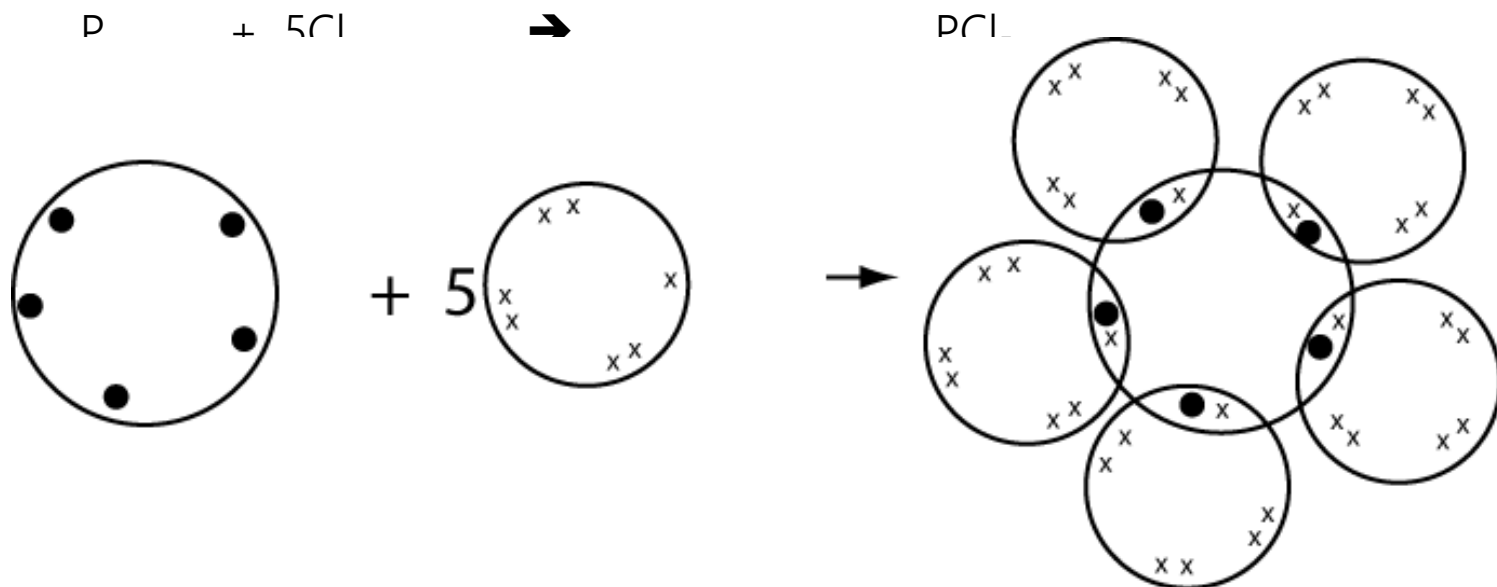
6 pairs – 6 pairs = 0  
pairs remaining

6. If there are any remaining pairs place them around the outside of the central atom.
7. Rearrange non-bonded pairs into bonded pairs if the central atom does not have 4 pairs around it.

## Expanded Octet

Third row elements (e.g., Al, Si, P, S, Cl) often have more than four valence shell orbitals filled with non-bonded Pairs and/or Bond Pairs; this is called "**expanded Octet**". Obviously, elements from the fourth and higher rows can also exhibit "expanded valence". **EXAMPLE** Phosphorus, electron configuration 2,8,5 can form up to 5 covalent bonds.

e.g. phosphorus pentachloride



## Expanded Octet - Rule of orbitals – exceptions to the rule

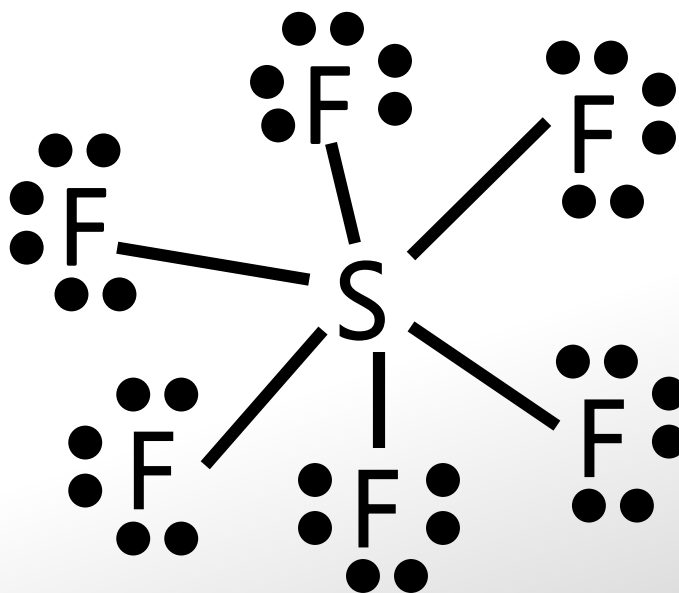


If there are extra non-bonded Pairs of electrons left after all of the periphery atoms are filled in accordance with the *octet rule* then they are placed around the central atom(s) according to the **Rule of Orbitals**. [extension] *The Rule of Orbitals*: the total number of non-bonded pairs and bonded pairs (NBP+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): maximum VSE pairs (NBP+BP) =  $VSO = 1$ ;

$n = 2$  (B, C, N, O, F): maximum VSE pairs (NBP+BP) =  $VSO = 4$  ("octet rule")

$n = 3$  (Al, Si, P, S, Cl): maximum VSE pairs (NBP+BP) =  $VSO = 9$  etc.



SF<sub>6</sub> creates a molecule that has 6 bonded F atoms around the central S atom.

# Lewis Diagrams of ions

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 diagram at the end. *Example carbon dioxide.*

C = 4
O = 6
O = 6
16

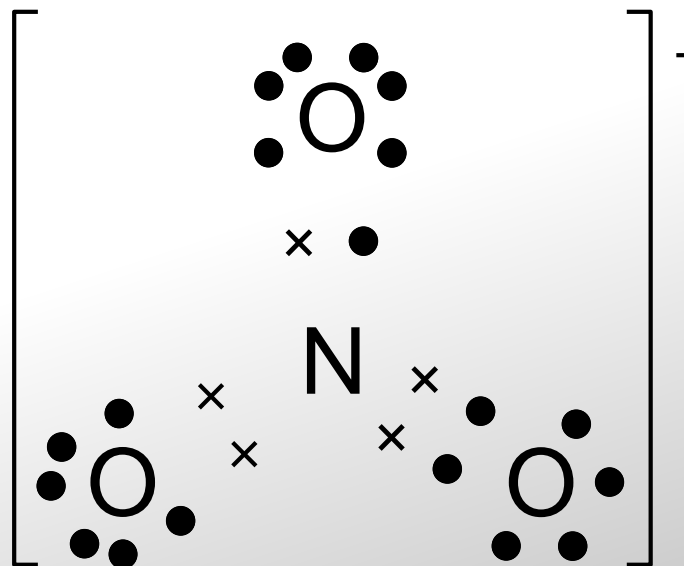


In step 1. For negative ions, add extra e<sup>-</sup> to valence electrons

For positive ions remove e<sup>-</sup> from valence electrons

e.g. for nitrate NO<sub>3</sub><sup>-</sup>

Put bracket around ion and show charge.





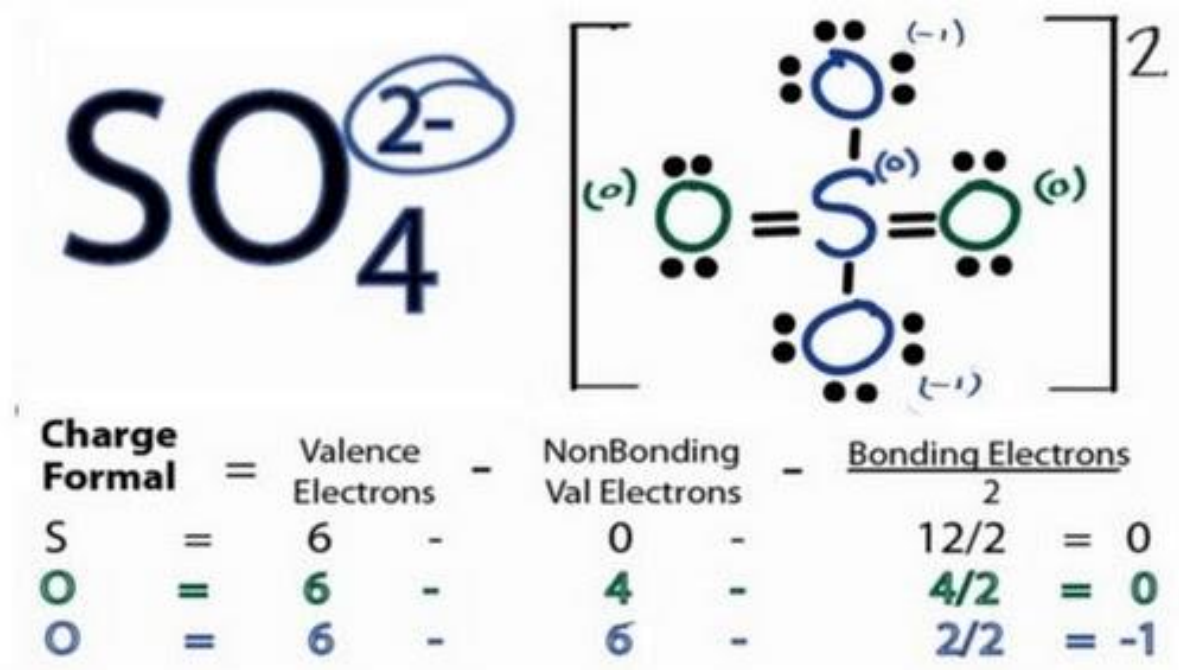
# Lewis Diagrams of ions – Formal Charge



Use formal charge to identify stable canonical form

The Lewis structure for  $\text{SO}_4^{2-}$  includes two double bonds rather than 4 single bonds because this arrangement has the lowest formal charge, FC, where

FC (= no of valence  $e^-$ s - [(no of non-bonded  $e^-$ s) - (1/2 no of bonded  $e^-$ s)]).

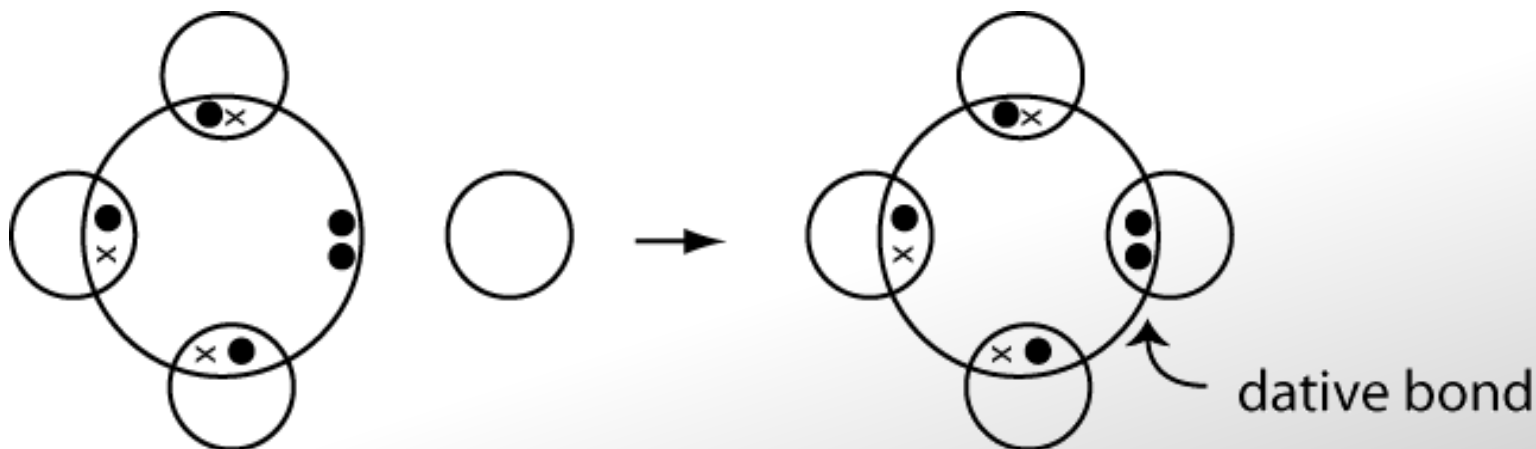


## Dative Bonds

e.g. an ammonium ion is made from an ammonia molecule plus a hydrogen ion.



A dative bond occurs when one atom donates both electrons to form a single covalent bond.



## Bonding the atoms

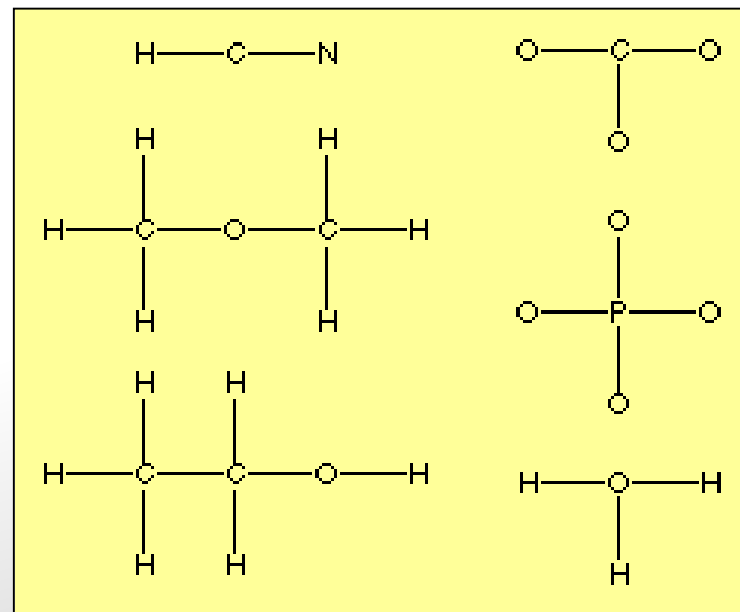
Otherwise, as a general rule, the least electronegative elements are central, the most electronegative elements are peripheral

Often, the formula is written to indicate connectivity.

For example:  $\text{HCN}$  = H bonded to C, C bonded to N, H and N are not bonded.

$\text{CH}_3\text{OCH}_3$  = three H bonded to C1, C1 bonded to O, O bonded to C2, C2 bonded to three H.

$\text{CH}_3\text{CH}_2\text{OH}$  = three H bonded to C1, C1 bonded to C2, C2 bonded to two H and O, O bonded to H.



## Determining Molecular Shapes

**Valence Shell Electron Pair Repulsion Theory** - electrons repel and therefore regions of negative charge move as far apart from each other as possible around a central atom. (used to predict molecular shape.)

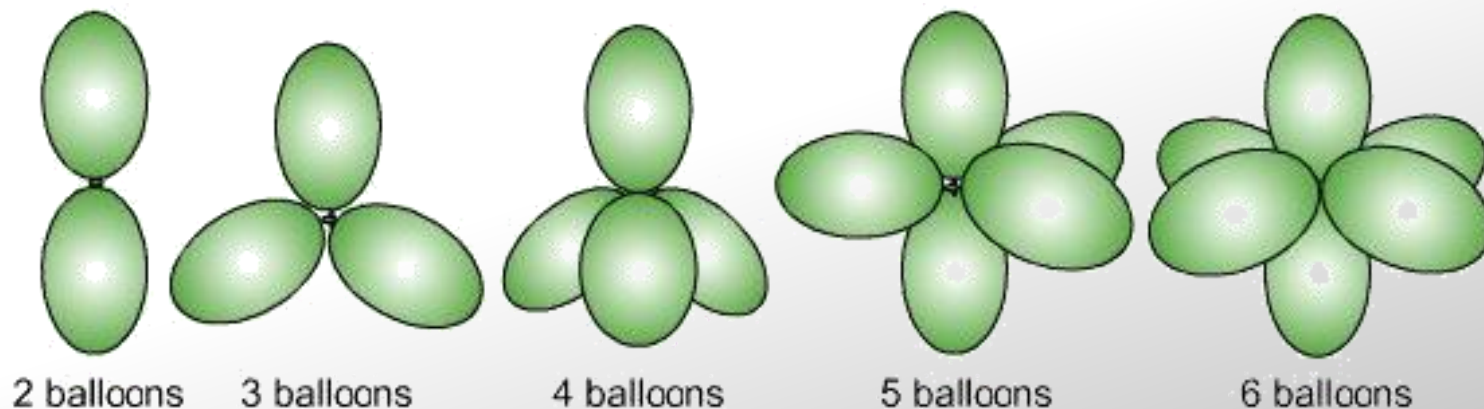
1. Draw the electron dot diagram
2. Count the electron pairs (regions of negative charge) around the central atom.

(Double and triple bonds may be considered as single charge regions)

3. Consider the shape the charge regions repel to.
4. Name the shape formed by the bonded atoms.

Repelling power of charge regions:

triple bond > double bond > lone pair > single bond

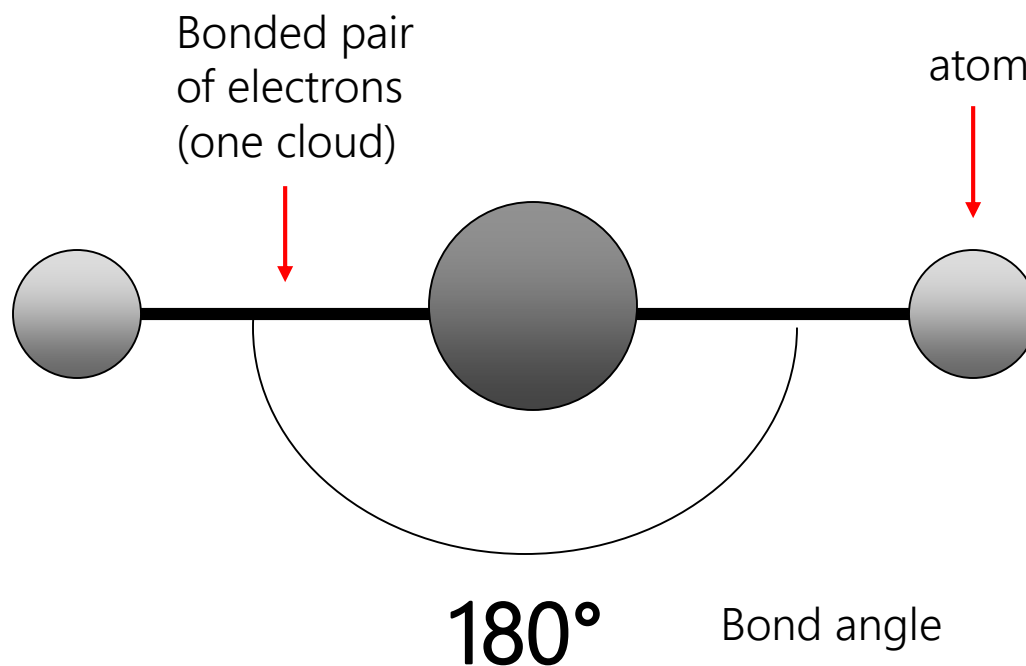


## Molecular Shapes – two regions of charge

Since regions of electrons are negatively charged, they repel each other as far apart as possible (VSEPR) into a **linear geometry / arrangement**. Two negative regions arrange themselves on opposite sides of the central atom.

The bond angle will be  $180^\circ$ .

The shape name is **linear**.

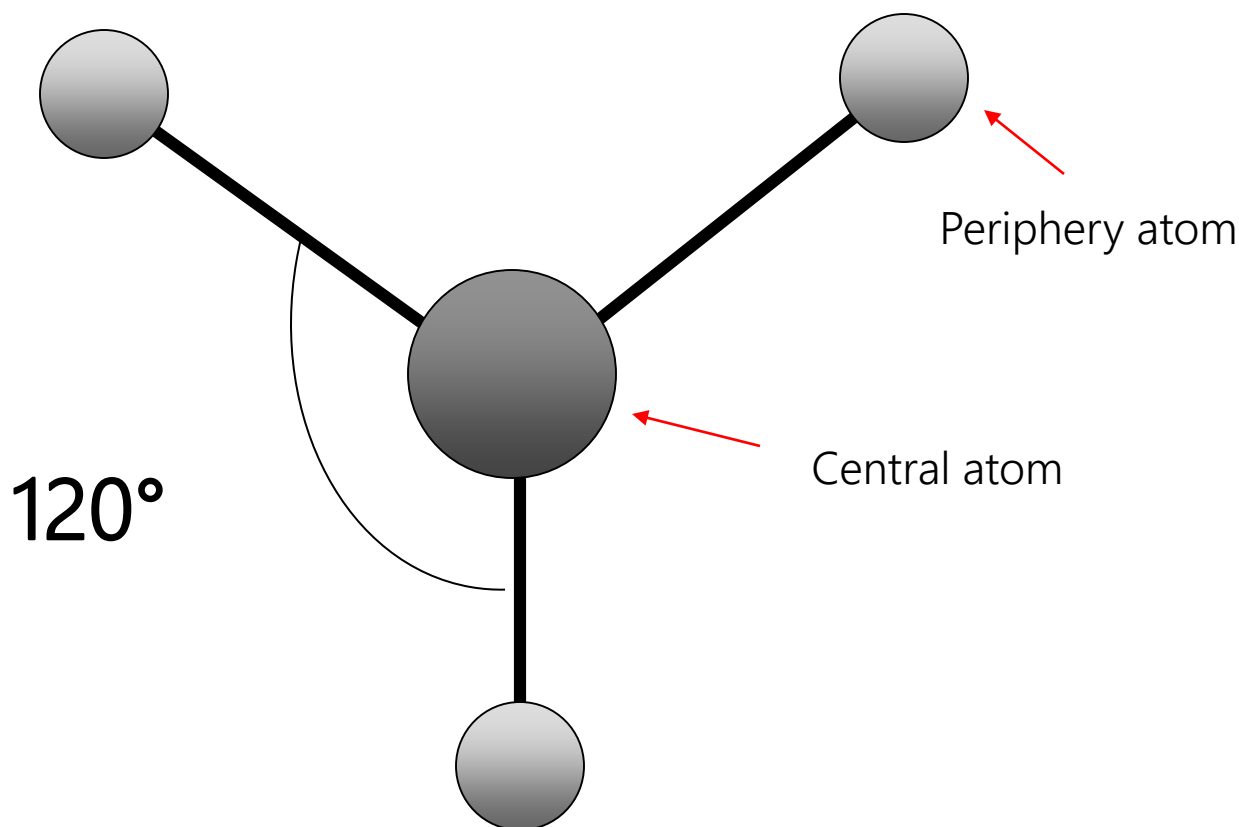


## Molecular Shapes – three regions of charge (0 non-bonding pairs)

Three regions of negative charge will cause a bond angle of  $120^\circ$  as they repel each other into a **trigonal planar geometry / arrangement**.

All the atoms still lie on a flat plane (like a sheet of paper).

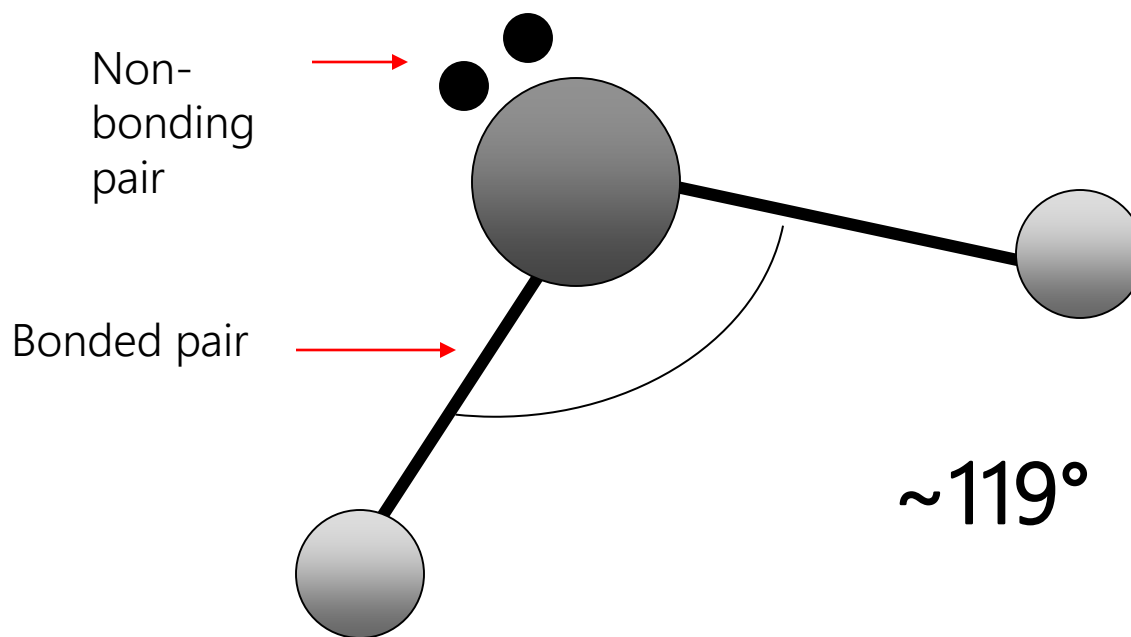
As there are no non-bonded pairs of electrons, the final shape is also a **trigonal planar**.  
(or triangular planar)





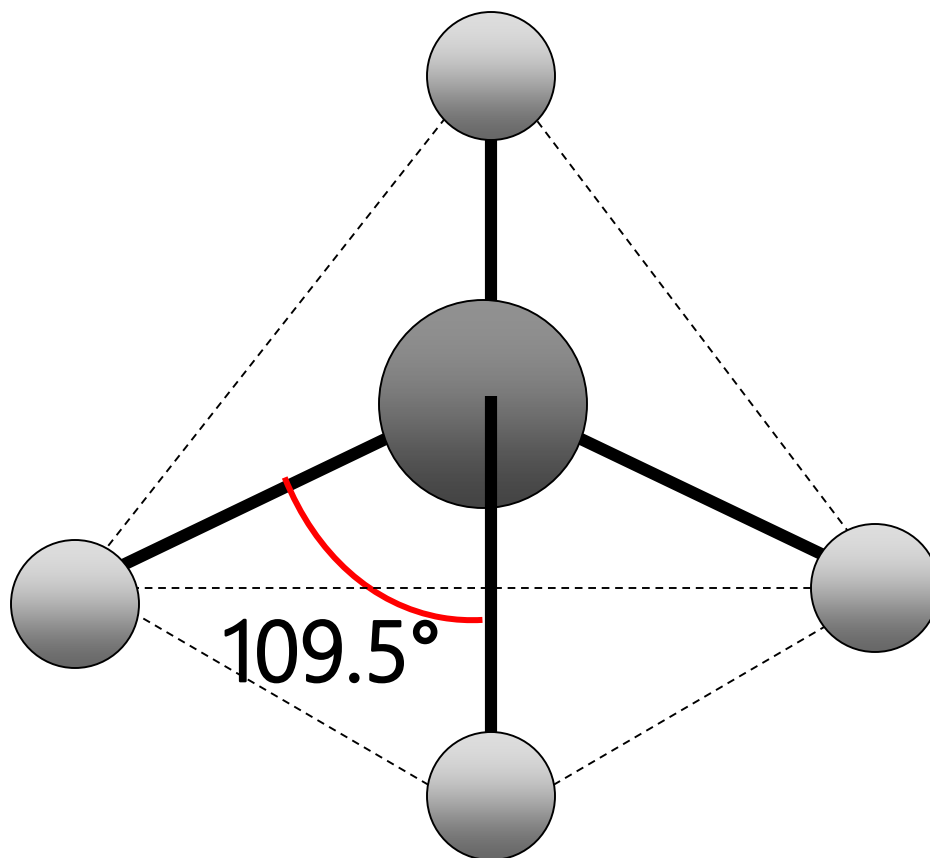
## Molecular Shapes – three regions of charge (1 non-bonding pair)

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 lone pair are only orbiting around one positive nucleus and their negative charge is less 'neutralised' than if they had another nucleus to orbit around. The regions of negative charge repel to a **trigonal planar geometry**. The bond angle between the remaining pairs is approximately  $119^\circ$  to  $120^\circ$ . As there is one non-bonded pair of electrons, the final shape formed by the atoms is called **bent**.



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

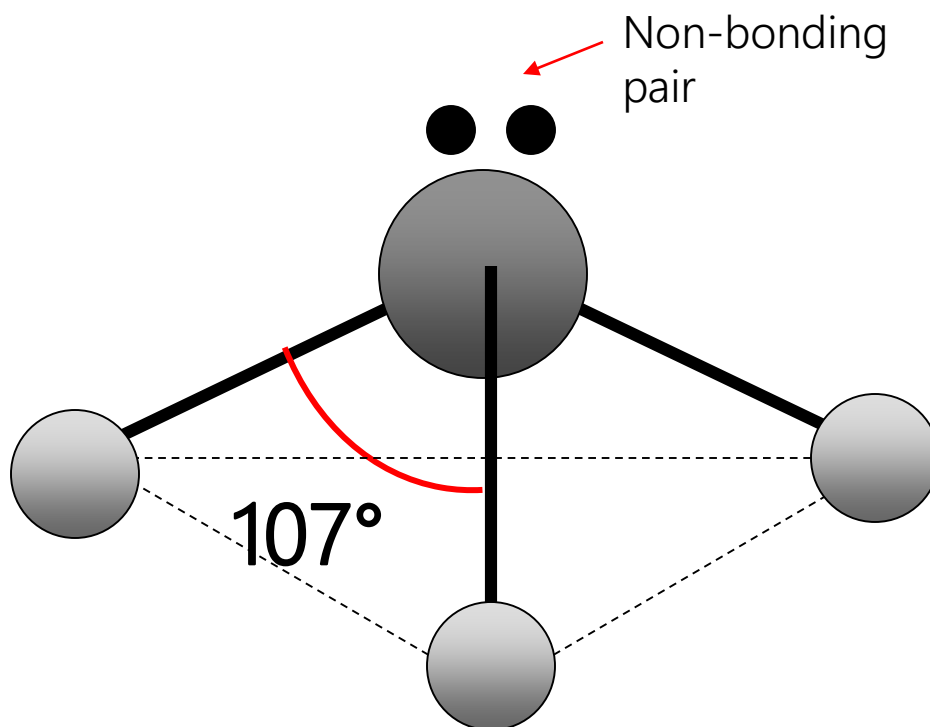
When four regions of negative charge are around a central atom, they repel each other into a 3-dimensional **tetrahedral geometry / arrangement**. The bond angle is now  $109.5^\circ$ . This is because it is a 3-dimensional sphere divided into 4 rather than a circle. This final shape is **tetrahedral**.



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

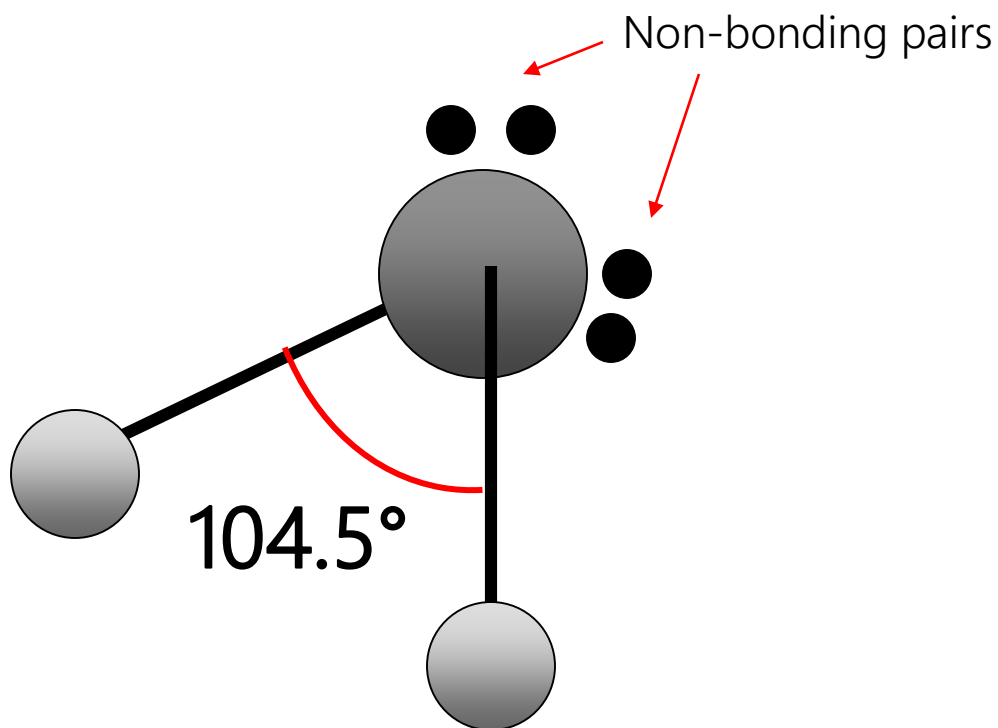
The four regions of negative charge still occupy a 3-dimensional tetrahedral shape. (The non-bonding pair, however, exerts a stronger repulsion to the remaining bonded pairs). The bond angle is  $109.5^\circ$ . ( $107^\circ$  actually because of the stronger non-bonded pair repulsion)

The final shape the bonded atoms form is a **trigonal pyramid** (or a triangular pyramid)



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

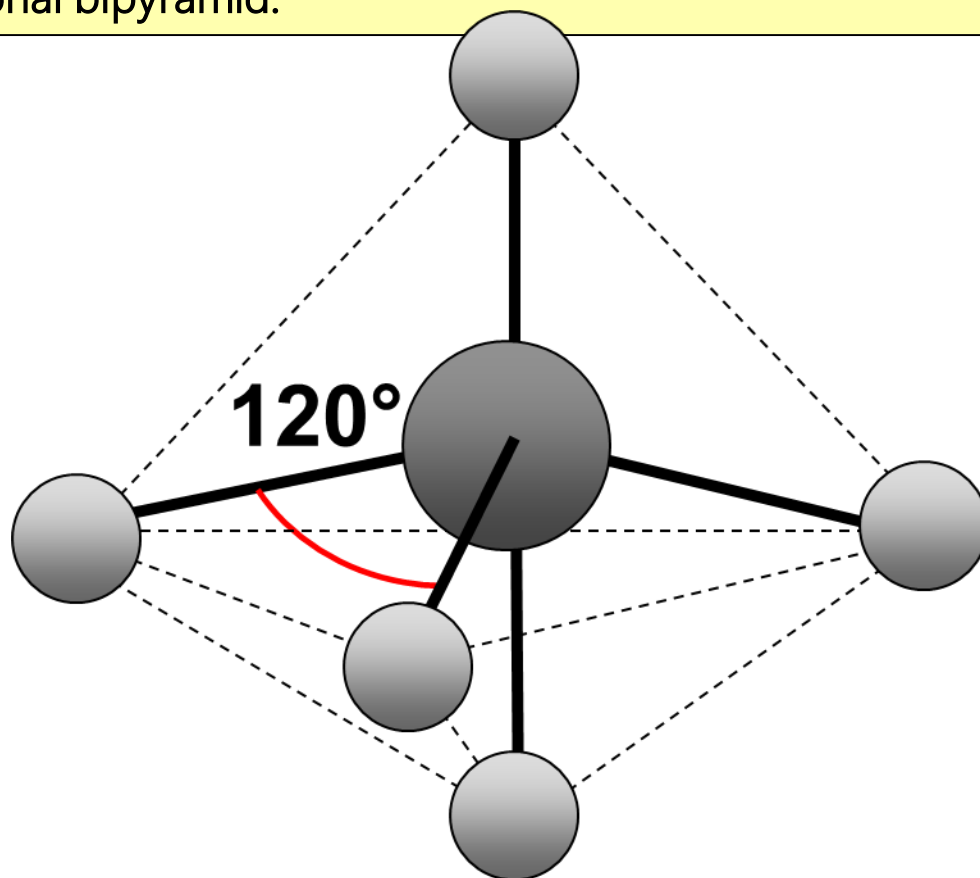
The 4 regions of negative charge repel each other to a (warped) tetrahedral shape. But The two non-bonding pairs create a much stronger repulsion than one lone pair and the bond angle between the remaining bonded pairs is smaller again at approximately  $105^\circ$  (compared to  $119^\circ$  of the bent shape with only 3 regions of negative charge). The final shape the bonded atoms form is called **Bent**.



## Molecular Shapes – five regions of charge (0 non-bonding pairs)

When five regions of negative charge are around a central atom they repel each other into a 3-dimensional shape. The bond angle is now  $120^\circ$ . The central atom has 3 central regions of negative charge that are  $120^\circ$  apart and 1 region above and below,  $180^\circ$  apart

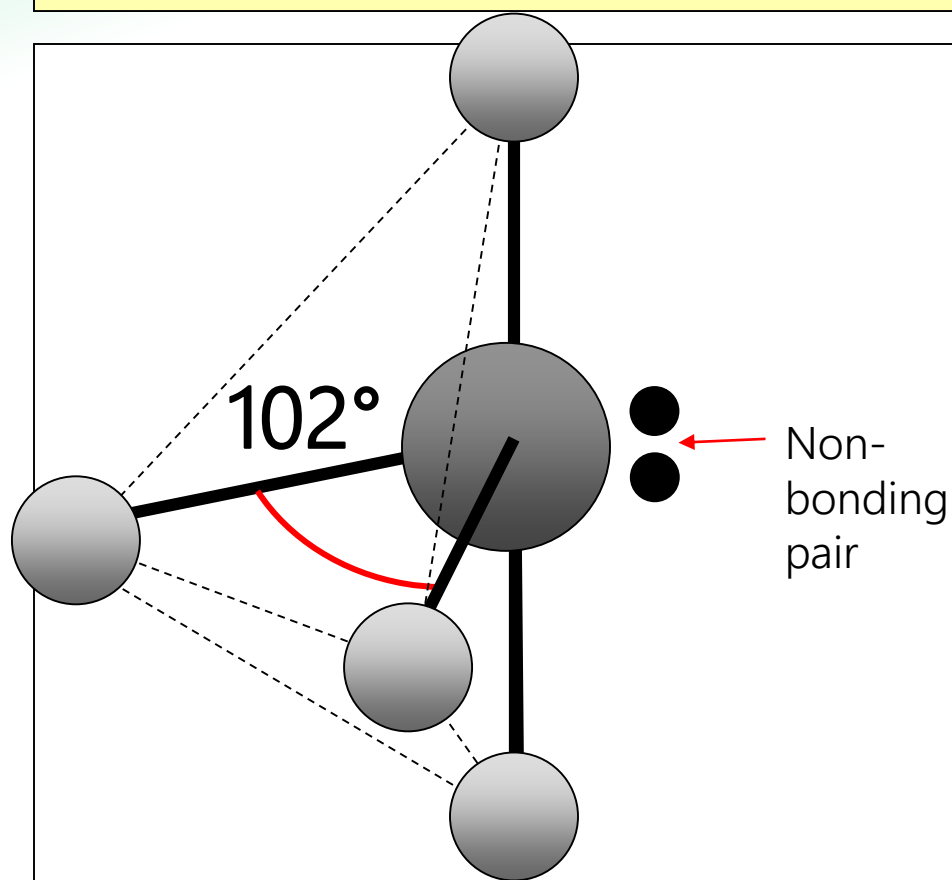
This shape is **trigonal bipyramid**.



## Molecular Shapes – five regions of charge (1 non-bonding pairs)

When five regions of negative charge are around a central atom they repel each other into a trigonal bi-pyramid shape. The bond angle is now  $102^\circ$ . One of the central regions is a non-bonding pair.

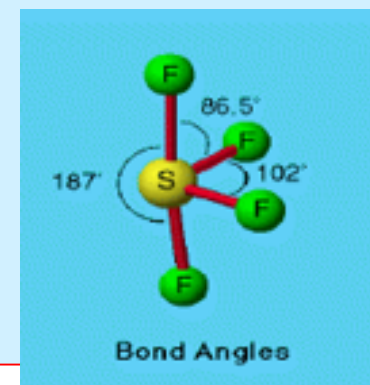
The overall shape formed from bonded regions is a **see-saw**.



### EXAMPLE:

#### Sulfur Tetrafluoride

$\text{SF}_4$ , S at the center with one lone electron pair and four fluoride atoms attached.





## Molecular Shapes – five regions of charge (2 non-bonding pairs)

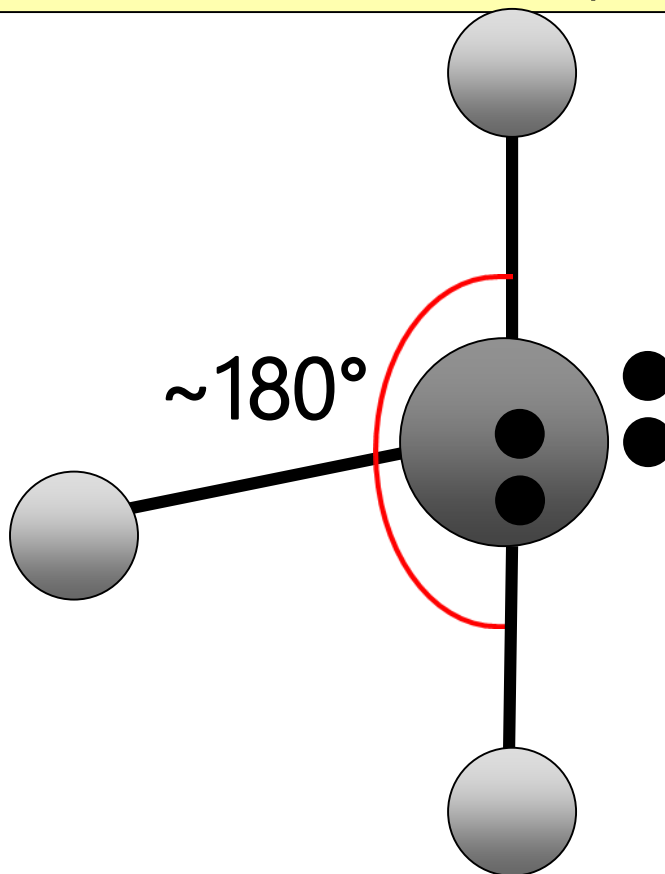
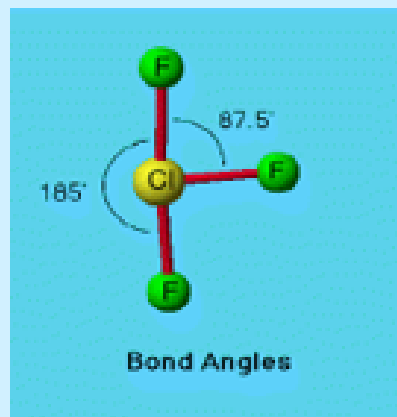
When five regions of negative charge are around a central atom they repel each other into a trigonal bi-pyramid shape. The bond angle is slightly less than  $180^\circ$  between the 2 upright regions of negative charge. Two regions of negative charge in the form of non-bonding pairs exist around the central atom.

The final shape formed by bonded negative charge regions is **T-shaped**.

### EXAMPLE:

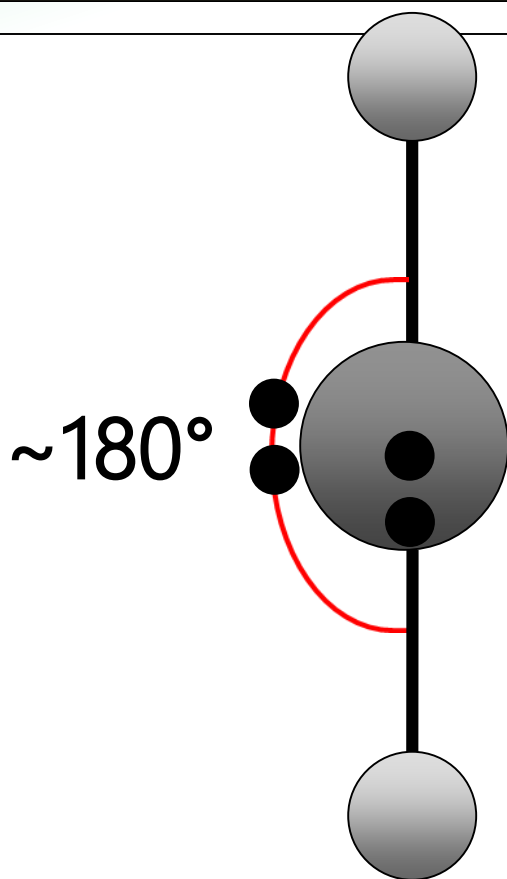
#### Chlorine Trifluoride

$\text{ClF}_3$ , chlorine at the center with three fluorine atoms attached and two lone electron pairs.



## Molecular Shapes – five regions of charge (1 non-bonding pairs)

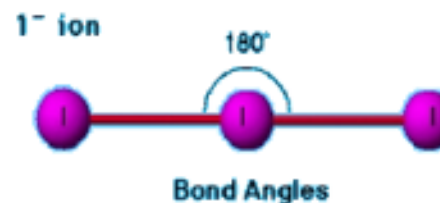
When five regions of negative charge are around a central atom they repel each other into a trigonal bi-pyramid shape. The bond angle is now slightly less than  $180^\circ$ . Three regions around the central atom are in the form of non-bonding pairs. The final shape created by bonded regions of negative charge is **linear**.



### EXAMPLE:

#### Triiodide Ion:

$\text{I}_3^-$ , the I at the center with 3 lone electron pair and two other iodide atoms attached.



## Molecular Shapes – six regions of charge (0 non-bonding pairs)

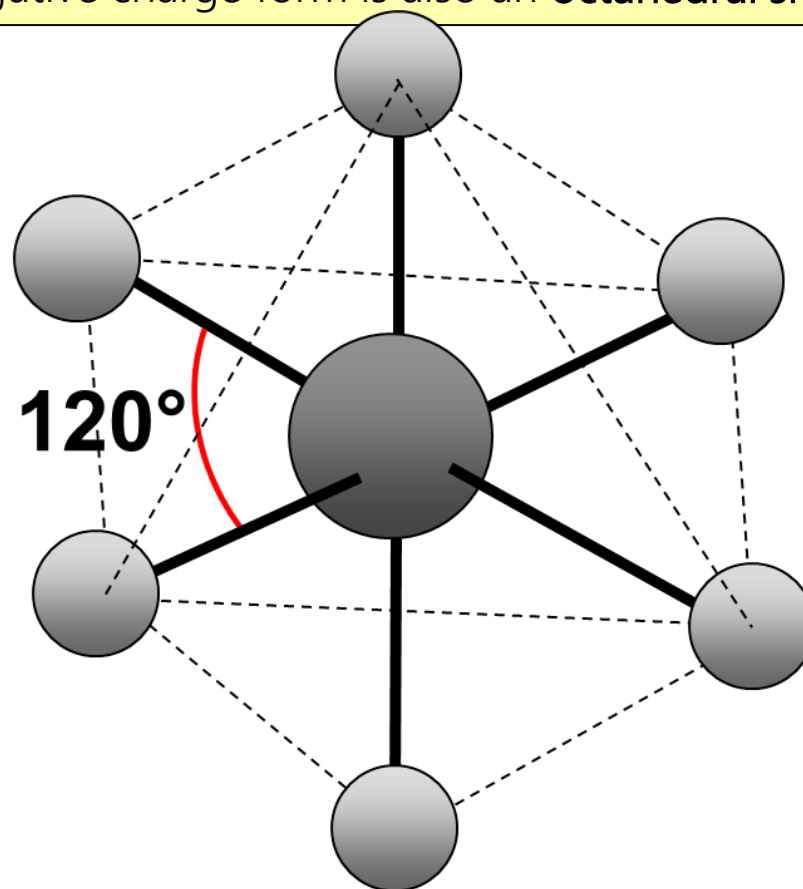
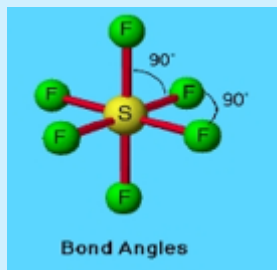
When six regions of negative charge are around a central atom, they repel each other into an octahedral shape. The bond angle is now  $120^\circ$  between the central regions of negative charge. Two regions of negative charge are vertical of the central atom at approximately  $180^\circ$ .

The shape the bonded regions of negative charge form is also an **octahedral shape**.

### EXAMPLE:

#### Sulfur hexafluoride:

An example of octahedral molecular geometry that results from six electron pair geometry is  $\text{SF}_6$ . The sulfur atom has 6 valence electrons. However this is an example where six fluoride atoms are present and the octet is expanded.



## Molecular Shapes – six regions of charge (1 non-bonding pair)

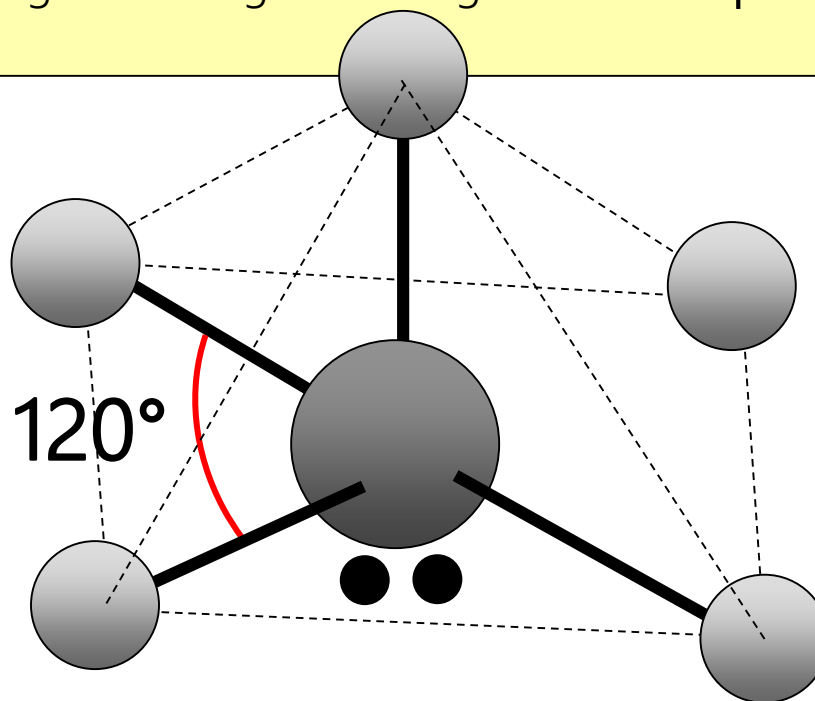
When six regions of negative charge are around a central atom, they repel each other into an octahedral shape. The bond angle is now  $120^\circ$  between the central regions of negative charge. Unlike the 5-region of charge that loses the horizontal atoms first to non-bonding pairs, the 6 region shape loses its vertical atoms first to non-bonding pairs.

The shape the remaining bonded regions of negative charge form is a **square pyramid** shape.

### EXAMPLE:

#### Bromine pentafluoride:

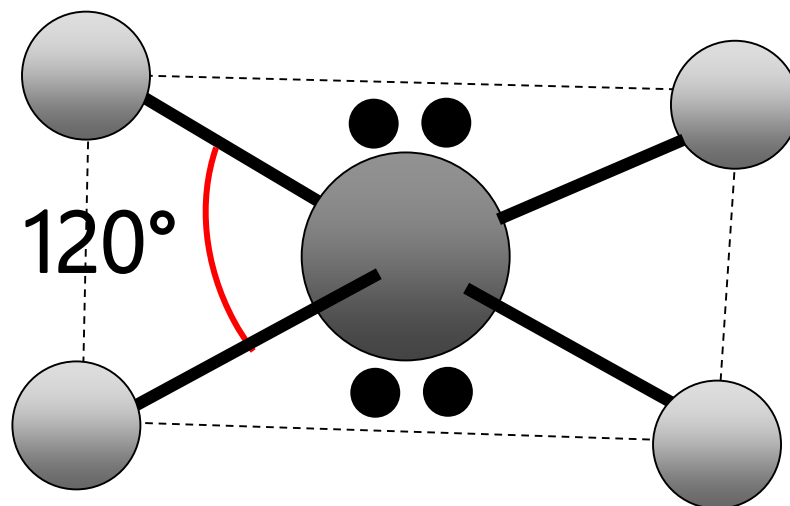
$\text{BrF}_5$ , the Lewis diagram shows Br at the center with one lone electron pair and five fluoride atoms attached.



## Molecular Shapes – six regions of charge (2 non-bonding pairs)

When six regions of negative charge are around a central atom, they repel each other into an octahedral shape. The bond angle is now  $120^\circ$  between the central regions of negative charge. Unlike the 5-region of charge that loses the horizontal atoms first to non-bonding pairs, the 6-region shape loses its two vertical atoms which are replaced by non-bonding pairs.

The shape the remaining bonded regions of negative charge form is a **square planar** shape.

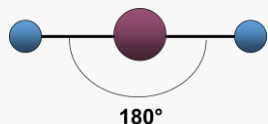


# Determining Molecular Shapes

2 regions

No non-bonding pair

**Linear**

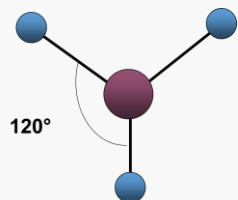


180°

3 regions

No non-bonding pair

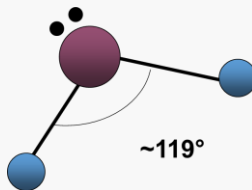
**Trigonal Planar**



120°

1 non-bonding pair

**Bent**

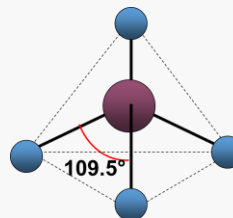


119°

4 regions

No non-bonding pair

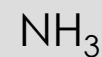
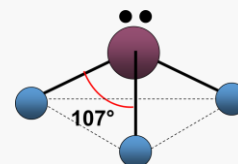
**Tetrahedral**



109.5°

1 non-bonding pair

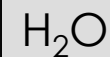
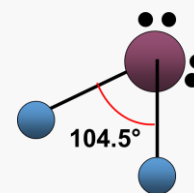
**Trigonal Pyramid**



107°

2 non-bonding pairs

**Bent**

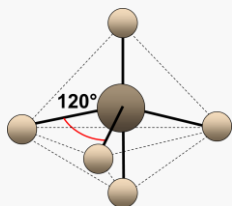


105°

5 regions

no non-bonding pairs

**Trigonal bipyramid**

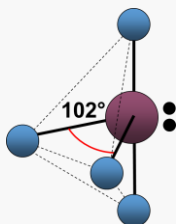


120°

90°

1 non-bonding pair

**See-saw**

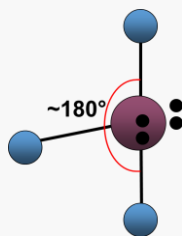


102°

86.5°

2 non-bonding pairs

**T-Shaped**

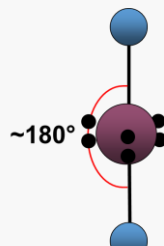


185°

87.5°

3 non-bonding pairs

**Linear**

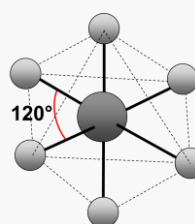


180°

6 regions

no non-bonding pairs

**Octahedral**

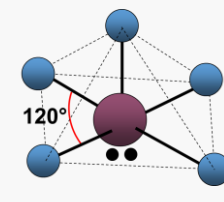


90°

90°

1 non-bonding pair

**Square Pyramid**

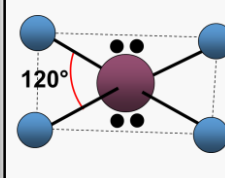


90°

84.8°

2 non-bonding pairs

**Square Planar**



90°

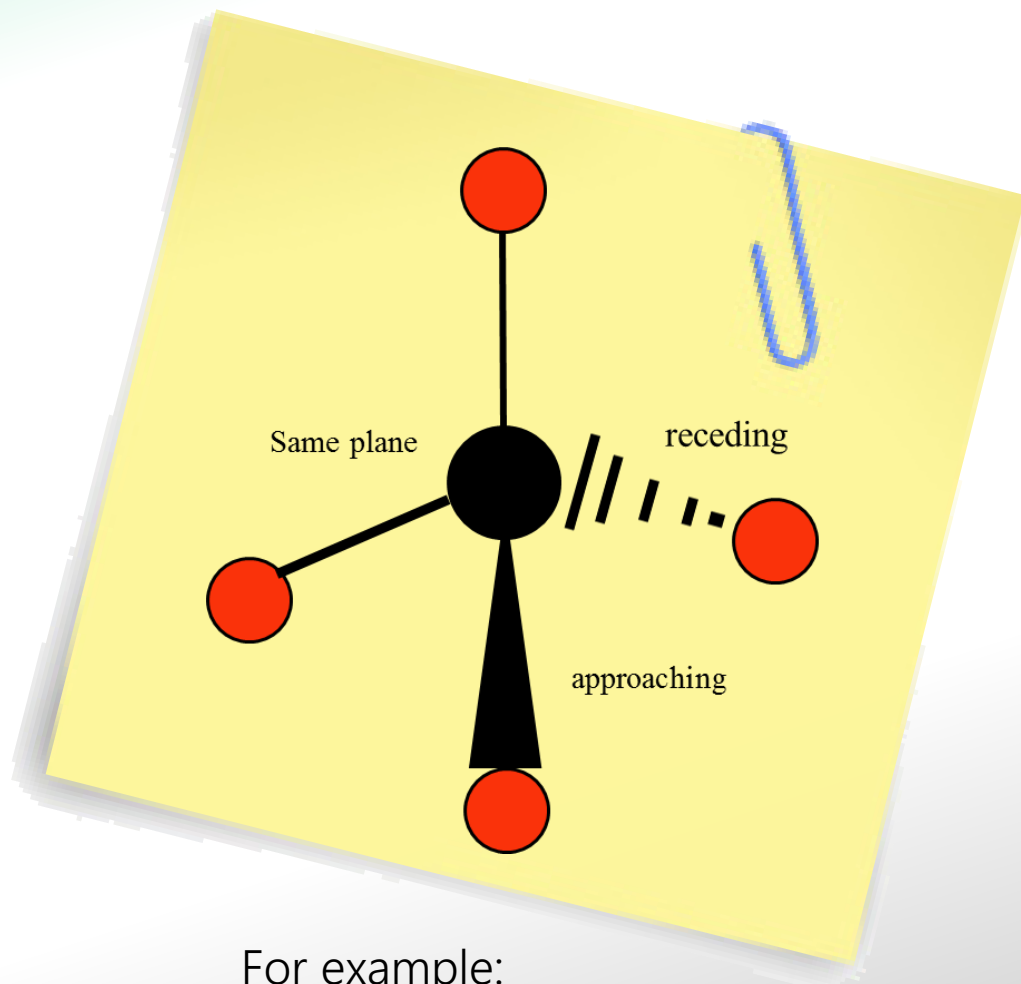
horizontal

vertical



# Drawing 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

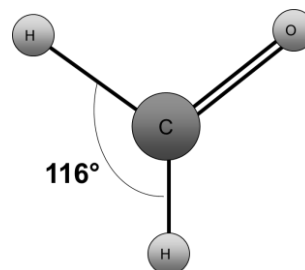


For example:  
Tetrahedral shape

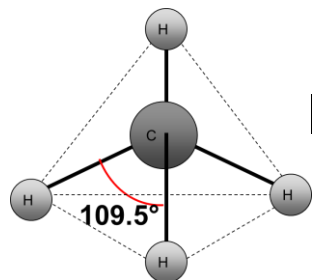
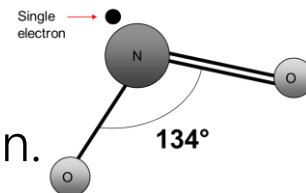
# Minor variations in geometry of shapes



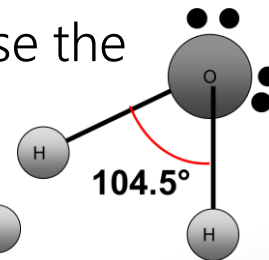
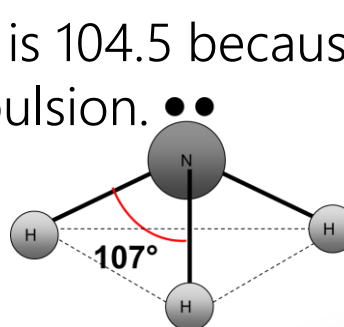
$\text{H}_2\text{C}=\text{O}$  is  $116^\circ$  because of greater repulsion from multiple bonds



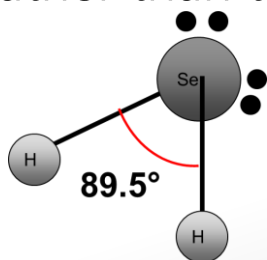
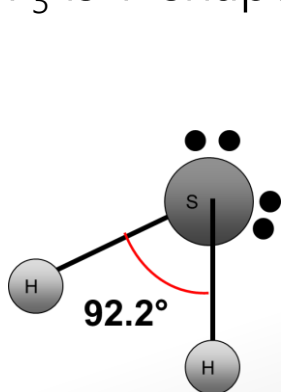
$\text{NO}_2$  is  $134^\circ$  because of smaller repulsion from unpaired electron.



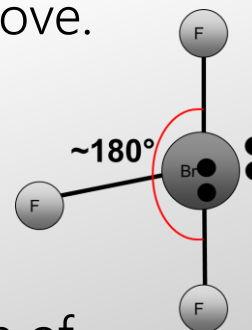
$\text{CH}_4$  is  $109.5^\circ > \text{NH}_3$  is  $107.3^\circ > \text{H}_2\text{O}$  is  $104.5^\circ$  because the lone pair repulsion  $>$  bonded pair repulsion.



$\text{BrF}_3$  is T-shaped rather than trigonal planar for same reason as above.



$\text{H}_2\text{O}$   $104.5^\circ > \text{H}_2\text{S}$   $92.2^\circ > \text{H}_2\text{Se}$   $89.5^\circ$  because of decreasing electronegativity of central atom.



# 2013 Lewis Diagrams and Molecular Shapes

Achieved  
Question

Question 1c: (i) Complete the following table.

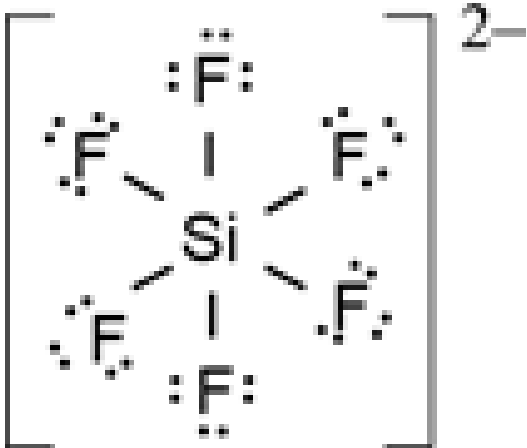
Molecule	$\text{BrF}_3$	$\text{PCl}_6^-$
Lewis diagram	$  \begin{array}{c}  :\ddot{\text{F}}: \\    \\  :\ddot{\text{F}} - \text{Br} : \\    \\  :\ddot{\text{F}}:  \end{array}  $	$  \left[ \begin{array}{ccc}  & :\ddot{\text{Cl}}: & \\  :\ddot{\text{Cl}}: &   & :\ddot{\text{Cl}}: \\  & \text{P} & \\  :\ddot{\text{Cl}}: &   & :\ddot{\text{Cl}}: \\  & :\ddot{\text{Cl}}: &  \end{array} \right]^-  $
Name of shape	T-shaped	Octahedral

# NCEA 2014 Lewis Diagrams and Molecular Shapes

Merit  
Question

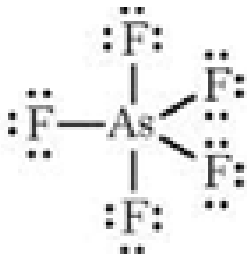
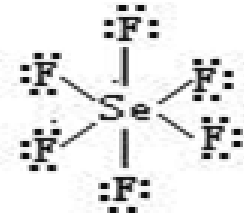
**Question 3a:** In New Zealand, fluoride for water treatment is supplied as sodium fluorosilicate,  $\text{Na}_2\text{SiF}_6$ .

One of the ions formed in the solution from sodium fluorosilicate is  $\text{SiF}_6^{2-}$ .  
Complete the table below.

	$\text{SiF}_6^{2-}$
<b>Lewis diagram</b>	
<b>Name of shape</b>	Octahedral

# NCEA 2015 Lewis Diagrams and Molecular Shapes

Question 3a: Complete the following table

	$\text{AsF}_5$	$\text{SeF}_6$
<b>Lewis diagram</b>		
<b>Name of shape</b>	Trigonal bipyramidal	Octahedral

Achieved  
Question

Achieved  
Question

Only one of each correct required for the mark.

It is highly likely the Lewis structure will be an expanded Octet (more than 4 pairs around the central atom) so "L3" names can be used for shapes.

# NCEA 2016 Lewis Diagrams and Molecular Shapes

Merit  
Question

Question 1c (i) : Complete the following table

	$\text{ICl}_4^-$	$\text{ClF}_3$
Lewis diagram	$\left[ \begin{array}{c} :\ddot{\text{Cl}}: \\   \\ :\ddot{\text{Cl}}-\ddot{\text{I}}-\ddot{\text{Cl}}: \\   \\ :\ddot{\text{Cl}}: \end{array} \right]^-$	$\begin{array}{c} :\ddot{\text{F}}-\ddot{\text{Cl}}-\ddot{\text{F}}: \\   \\ :\ddot{\text{F}}: \end{array}$
Name of shape	$\text{ICl}_4^-$ : Square planar	$\text{ClF}_3$ : T-shaped

Only one name + diagram required for the mark.

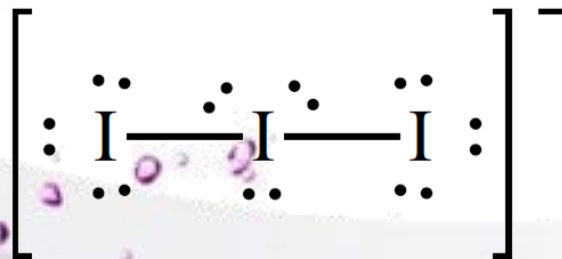
It is highly likely the Lewis structure will be an expanded Octet (more than 4 pairs around the central atom) so "L3" names can be used for shapes.



# NCEA 2017 Lewis Diagrams and Molecular Shapes

MERIT  
Question

**Question 3c:** Iodine forms a linear  $\text{I}_3^-$  ion.  
(i) Draw the Lewis structure for the  $\text{I}_3^-$  ion below.



(ii) Explain why the  $\text{I}_3^-$  ion has a linear shape.

Arrangement of areas of electron density around the central I atom is trigonal bipyramidal due to five regions of negative charge. These areas all repel each other. As there are three non-bonding pairs (in the equatorial area) and two bonded atoms, the shape is linear.

## NCEA 2018 Lewis Diagrams and Molecular Shapes

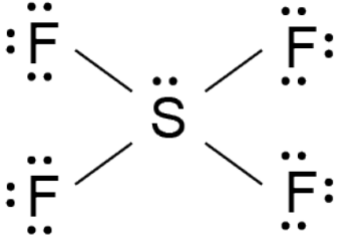
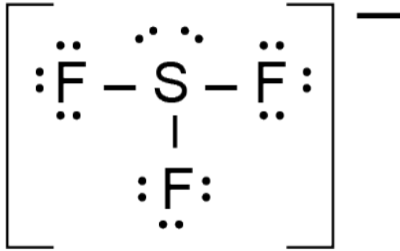
Q 1c (i): Complete the table below.

	$\text{AsF}_5$	$\text{BrF}_5$
<b>Lewis diagram</b>		
<b>Name of shape</b>	trigonal bipyramidal	square-based pyramid

# NCEA 2019 Lewis Diagrams and Molecular Shapes

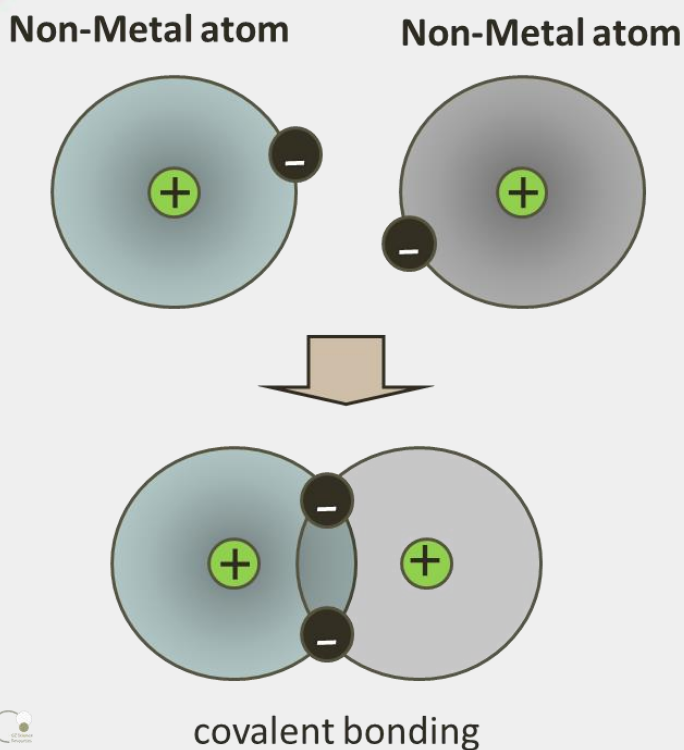
Merit  
Question

Question 1b: Complete the following table.

	$\text{SF}_4$	$\text{SF}_3^-$
<b>Lewis structure</b>		
<b>Name of shape</b>	See-saw	T-shaped

## Covalent Bonding

**Covalent Bonding** occurs when electrons are shared between neighbouring atoms. No ions are formed and there is no transfer of electrons. The compound formed is neutral with no charge.



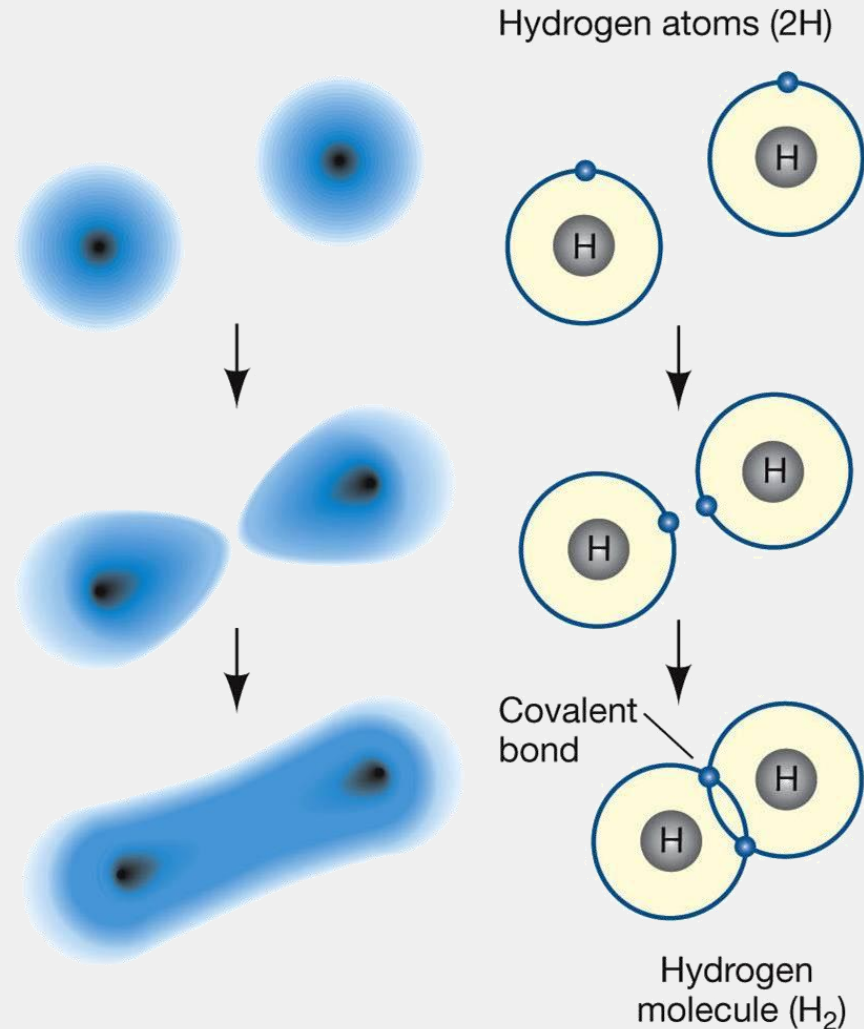
When the electronegativity is similar between atoms then the electrons will be shared evenly creating a **non-polar molecule**. If there is a significant electronegativity difference then dipoles will be created when the electrons orbit for a greater time around the more electronegative atom. This will create a **polar molecule**.

## Background Knowledge

The valance electrons are involved in bonding. These electrons orbit in pairs. The negative charge of the electron pair will attract the positively charged nuclei of other atoms, and this holds the atoms together in a molecule.



## Covalent Bonding

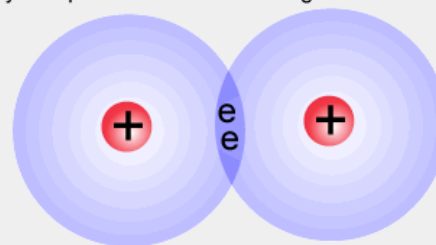


The electron-pair must lie between the nuclei for the attraction to outweigh the repulsion of the two nuclei. This 'sharing' of electrons between atoms creates a covalent bond – giving both atoms the stability of a full outer shell.

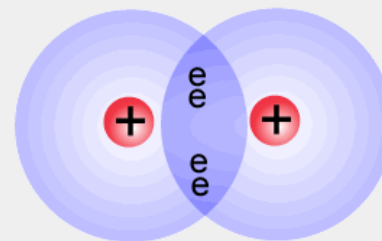
Covalent bonds are normally formed between pairs of non-metallic atoms.

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



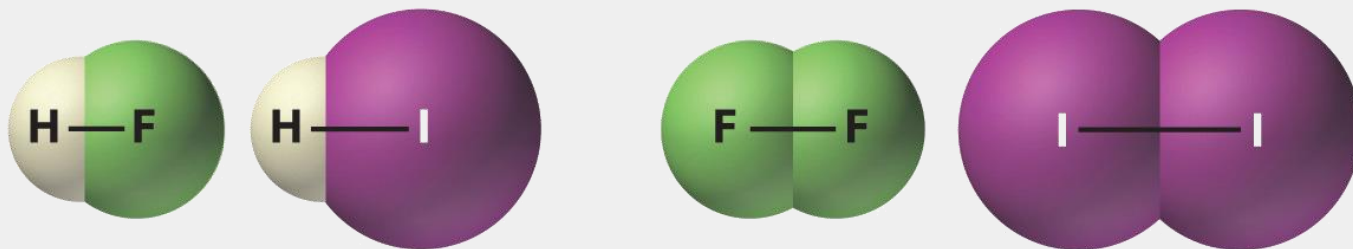
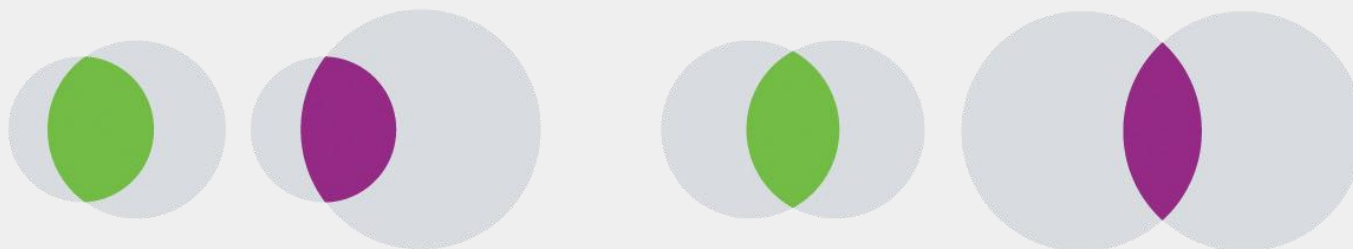
Two pair of electrons hold the nuclei tighter and closer





All covalent bonds are strong. That is it requires 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.

Bonding atoms

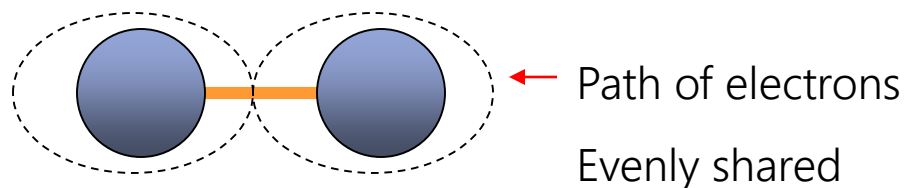
Overlap between  
valence orbitals

(a) H-F versus H-I

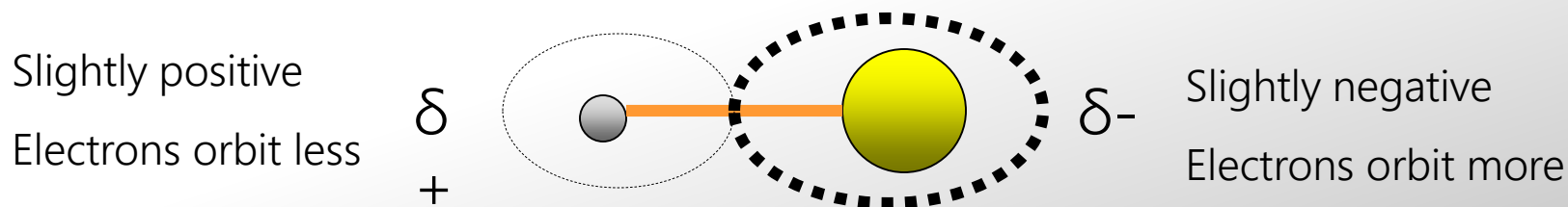
(b) F-F versus I-I

## Polarity and dipoles

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  $I_2$



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 and a **dipole** will form. These bonds become **polar bonds**. Example – hydrochloric acid  $HCl$

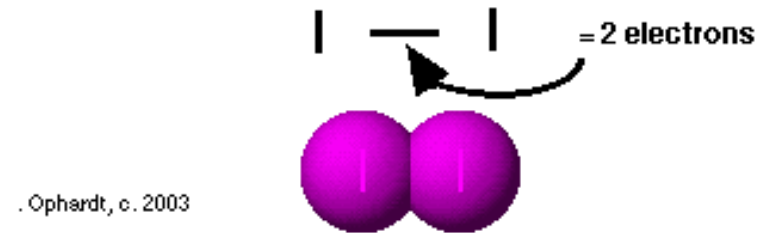


# Polarity

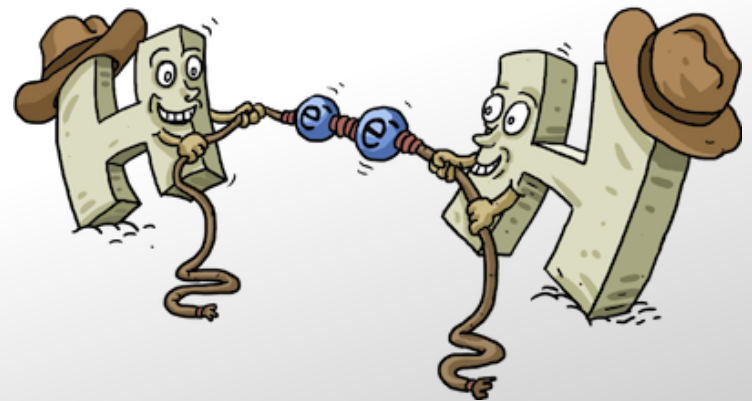
If two bonded atoms are the same, the bond is said to be non-polar. i.e.  $I_2$

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, 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.

**Equal Sharing of electrons between two identical non-metals.**



**Non-Polar Covalent Bond**



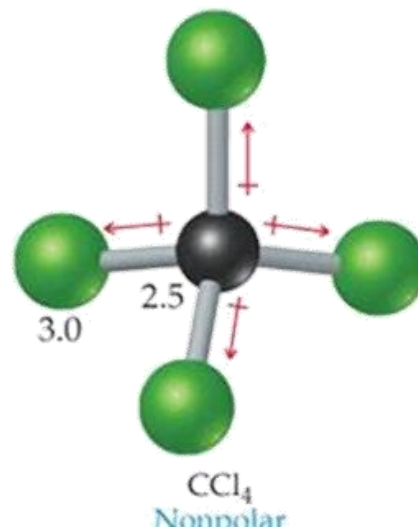
# Symmetry and Polarity

The **overall polarity** of a molecule with polar bonds depends upon whether the molecule is symmetrical or not and if all bonds are the same.

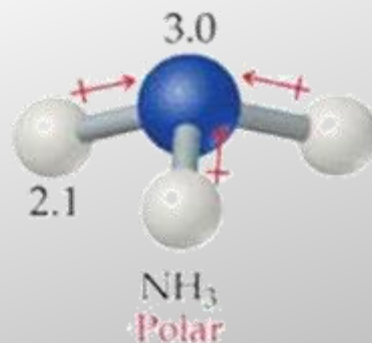
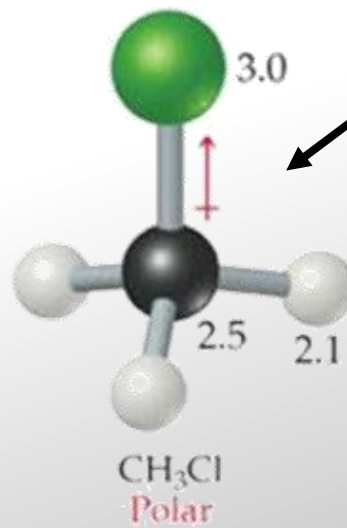
A symmetrical molecule (one where the centres of peripheral atoms coincide) becomes a **non-polar molecule** (only if all bonds are the same) – as the bond dipoles cancel out.

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 molecules

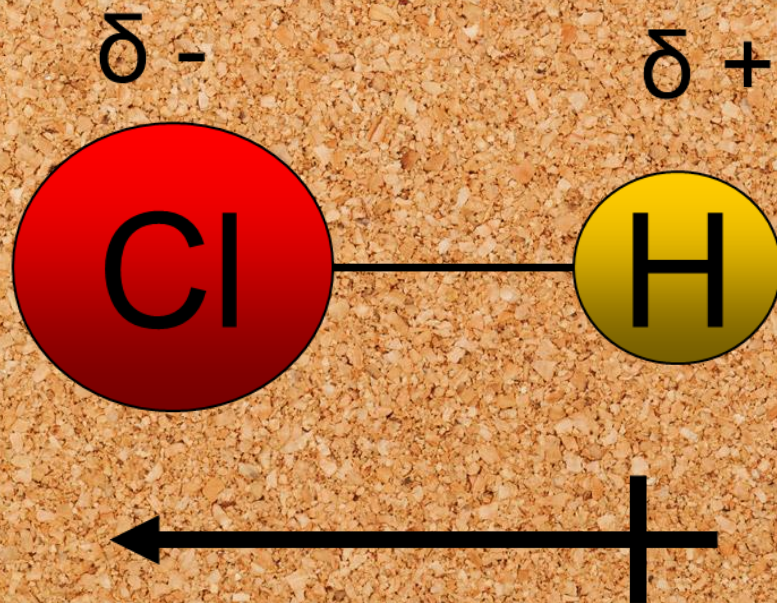


If the bonds (central to peripheral) are different, then even if molecule is symmetrical the dipoles will not cancel out and therefore the molecule is polar overall





# Polarity

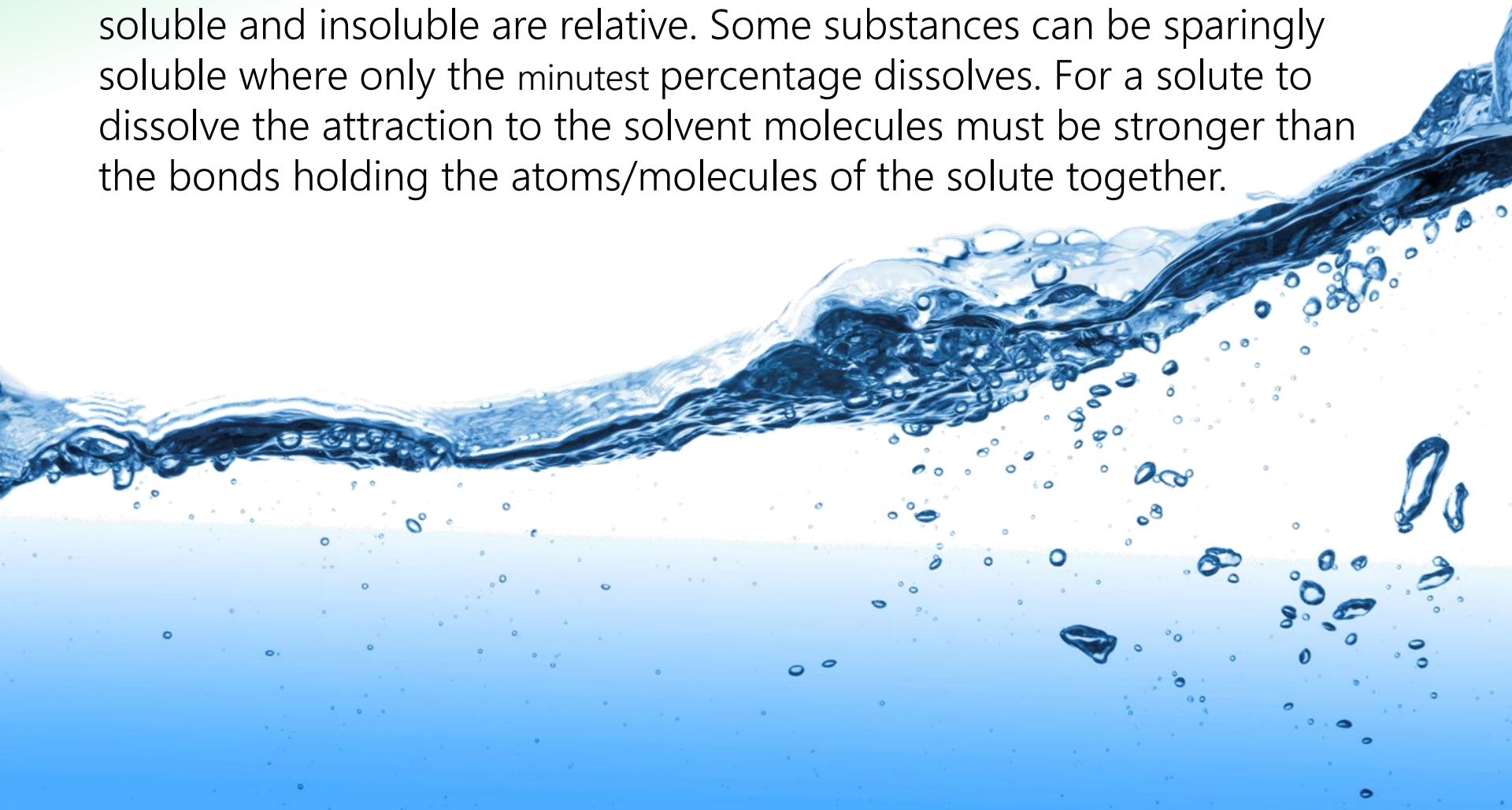


Dipoles may also be shown as an arrow, with a cross, +ve, at the tail. The arrowhead is the -ve end. The more electronegative atom forms the negative end and the less electronegative atom forms the positive end.



# 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 varies 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.



# Dissolving and Polarity

**Polar substances dissolve polar substances.**

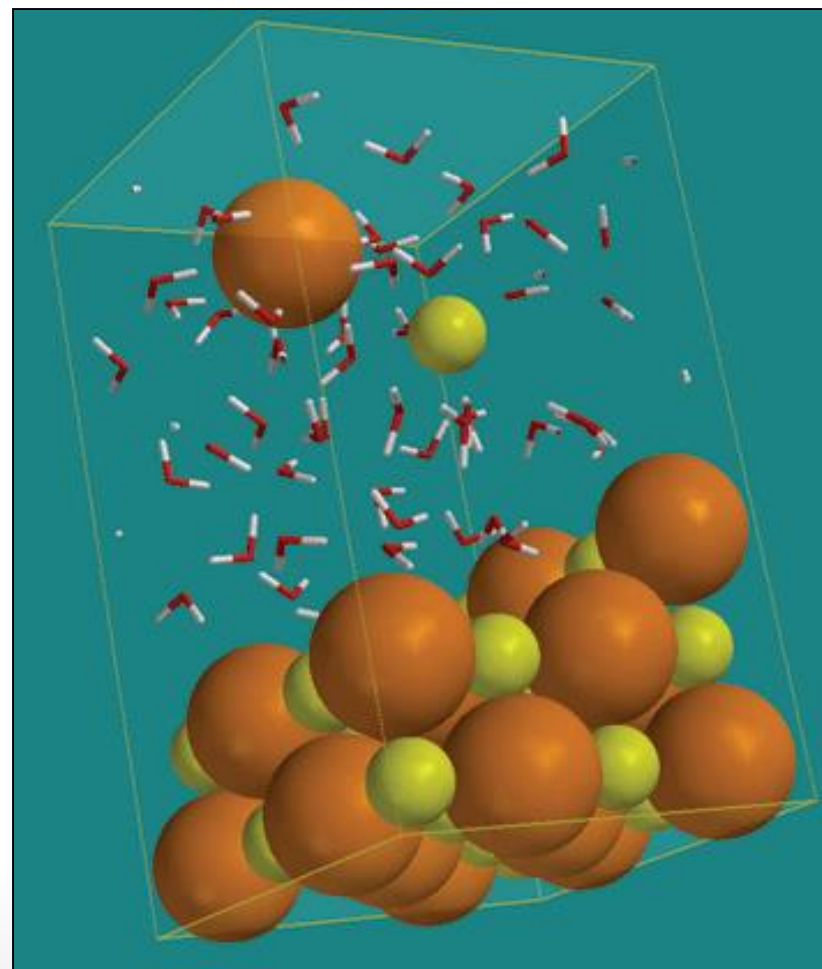
e.g. Water, being polar attracts the molecules of other polar substances (e.g. HCl) and will dissolve them.

**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 naphthalene) by the same weak Van der Waals forces they attract themselves by and so will dissolve non-polar solutes.



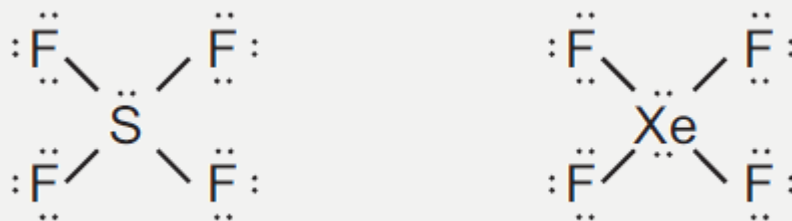
Ionic solid dissolving in water



## NCEA 2013 Polarity

Excellence  
Question

**Question 1c:** (ii) The Lewis diagrams for  $\text{SF}_4$  and  $\text{XeF}_4$  are shown below. Compare and contrast the polarities and shapes of these two molecules.



There is a difference in electronegativity between S and F, so the S-F bonds are polar covalent.  $\text{SF}_4$  has a see-saw shape (distorted tetrahedron) due to the repulsions between four bonding regions and one non-bonding region of charge, which is asymmetric therefore the polarities/dipoles do not cancel. As a result,  $\text{SF}_4$  is a polar molecule.

There is a difference in electronegativity between Xe and F, so the Xe-F bonds are polar covalent.  $\text{XeF}_4$  has a square planar shape, due to the repulsions between four bonding regions and two non-bonding regions of charge; therefore the polarities/dipoles do cancel. As a result,  $\text{XeF}_4$  is a non-polar molecule.

## NCEA 2014 Polarity

Excellence  
Question

**Question 1d:** The halogens make up Group 17 of the periodic table.

(i) The polarity of the HBr molecule is shown below.

$\delta+$   $\delta-$

H — Br

Using this as an example, indicate the polarity of the following bonds by indicating any dipoles present.

$\delta-$   $\delta+$        $\delta+$   $\delta-$   
1. F — Cl      2. At — Cl

(ii) Using your knowledge of trends in the periodic table, circle the atom below that has the greater electronegativity value. Br I

Justify your answer.

**Answer 1d:** Lower electronegativity means less attraction of a bonded atom for a bonding pair of electrons.

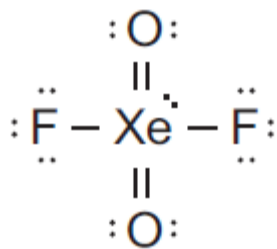
The lower value for iodine indicates that the attraction for the bonding pair in compounds is less than the attraction for bonding pairs in compounds of bromine. As the radii of atoms increase, electronegativity decreases, despite the increased nuclear charge. This is due to more energy levels being added.

**Iodine has a greater number of shells** (5th row) than bromine (4th row). This factor outweighs the increased nuclear charge (53 protons) of the iodine atom, as compared to the bromine atom (35 protons).

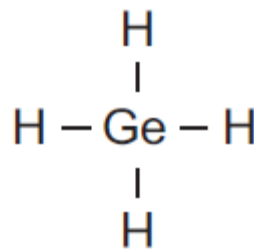
## NCEA 2015 Polarity

Excellence  
Question

**Question 3b:** The Lewis diagrams and shapes for  $\text{XeO}_2\text{F}_2$  and  $\text{GeH}_4$  are shown below. Compare and contrast the polarities and shapes of these two molecules.



see-saw



tetrahedral

**Answer 3b:**  $\text{XeO}_2\text{F}_2$  is polar. It has 5 areas of electron density around the central Xe atom, one of which is a lone pair. Maximum separation for minimum repulsion means that the shape is based on a trigonal bipyramid structure, but the final shape is actually see-saw. The  $\text{Xe}=\text{O}$  bonds are polar, due to the greater electronegativity of O, and the  $\text{Xe}-\text{F}$  bonds even more polar, due to the F atom having the highest electronegativity on the periodic table. The molecule is **not symmetrical**, and so the dipole moments cannot cancel, making the **molecule polar**.

$\text{GeH}_4$  is non-polar. It has 4 areas of electron density around the central Ge atom, all of which are bonded. Maximum separation for minimum repulsion means that the shape is tetrahedral. This is a **symmetrical** structure, thus the bond dipole moments cancel, and therefore the **molecule is non-polar**.

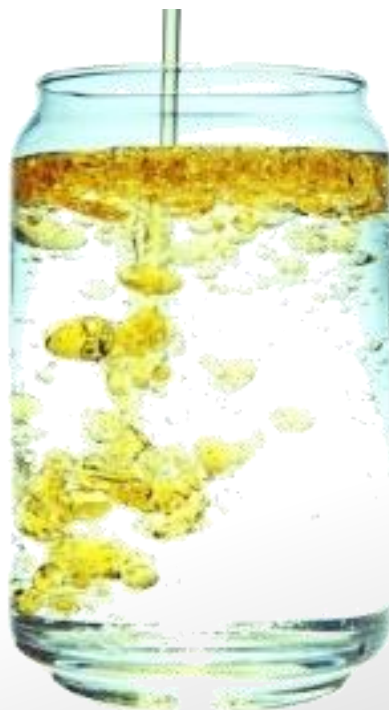
## NCEA 2016 Polarity

Excellence  
Question

**Question 1c (ii):** The Lewis diagram for  $\text{SeF}_6$  is shown beside. Would you expect  $\text{SeF}_6$  to be soluble in water? Explain your answer in terms of the shape and polarity of  $\text{SeF}_6$ .



**Answer 1c (ii): No.** There is an electronegativity difference between Se and F, so the Se-F bonds are polar covalent. The six bond pairs around the central Se atom arrange themselves as far apart as possible to minimise repulsion, so  $\text{SeF}_6$  has an octahedral shape. Since this is a symmetrical shape, the bond dipoles cancel out, so  $\text{SeF}_6$  is a non-polar molecule.



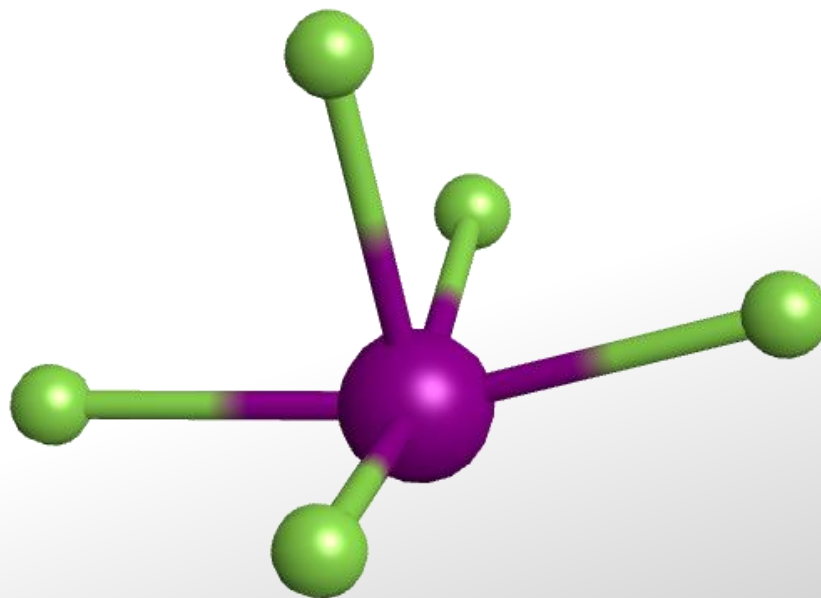
Water is a polar solvent. Non-polar molecules like  $\text{SeF}_6$  are not attracted to polar molecules like water, i.e. the intermolecular attraction between the water molecules and the  $\text{SeF}_6$  molecules is insufficient to overcome the attraction between the water molecules. Therefore,  $\text{SeF}_6$  is insoluble in water.

**Question 3c (iii):**  $\text{IF}_5$  has a square pyramidal shape. Indicate whether the molecule  $\text{IF}_5$  is polar or non-polar. Justify your choice.

**Polar.**

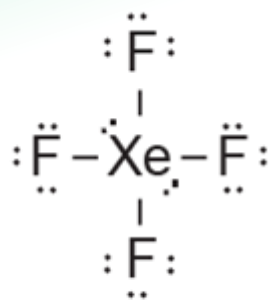
The I-F bond is polar due to a difference in electronegativity.

There are six regions of negative charge giving  $\text{IF}_5$  an octahedral geometry. The five bonded and one lone pair around the central iodine atom gives it the square pyramidal shape. This means the molecule is asymmetric so the bond polarities dipoles don't cancel causing the molecule to be polar.

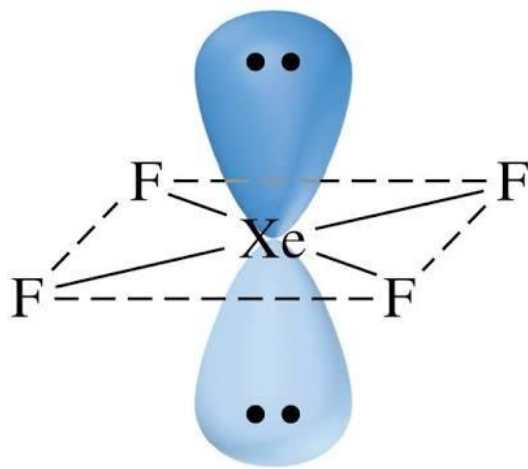


## NCEA 2018 Polarity

Q 1c (ii): The Lewis diagram and shape for  $\text{XeF}_4$  are given below. Elaborate on the shape and polarity of  $\text{XeF}_4$ .



square planar



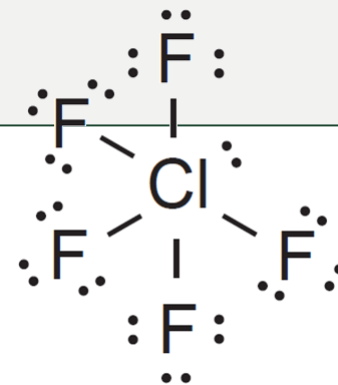
There are six electron clouds about the central atom; four bond pairs and two lone pairs of electrons. The six electron pairs about the central Xe atom are arranged as far apart as possible in an octahedral geometry to minimise repulsion, but due to the two lone pairs,  **$\text{XeF}_4$  has a square planar shape.**

There is an electronegativity difference between Xe and F, so the Xe-F bonds are polar covalent. This molecule is symmetrical due to the position of the two lone pairs around Xe being above and below the plane, so the effect of the bond dipoles cancel, i.e. there is an even spread of charge. **Therefore,  $\text{XeF}_4$  is a non-polar molecule.**

## NCEA 2019 Polarity

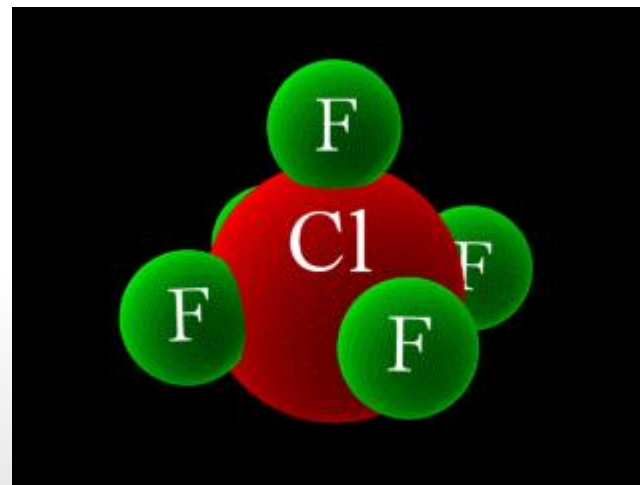
Excellence  
Question

**Question 1d:** The Lewis structure of  $\text{ClF}_5$  is given below. Identify and explain the shape and polarity of  $\text{ClF}_5$ .



$\text{ClF}_5$  has six electron clouds about the central atom, including five bond pairs and one lone pair. The electron clouds repel as far apart as possible; this produces the **square pyramidal shape**.

There is an electronegativity difference between Cl and F, so the Cl–F bonds are polar covalent. The square pyramidal shape arranges these dipoles asymmetrically due to the lone pair on the central atom. The dipoles do not cancel so  $\text{ClF}_5$  is a **polar** molecule.

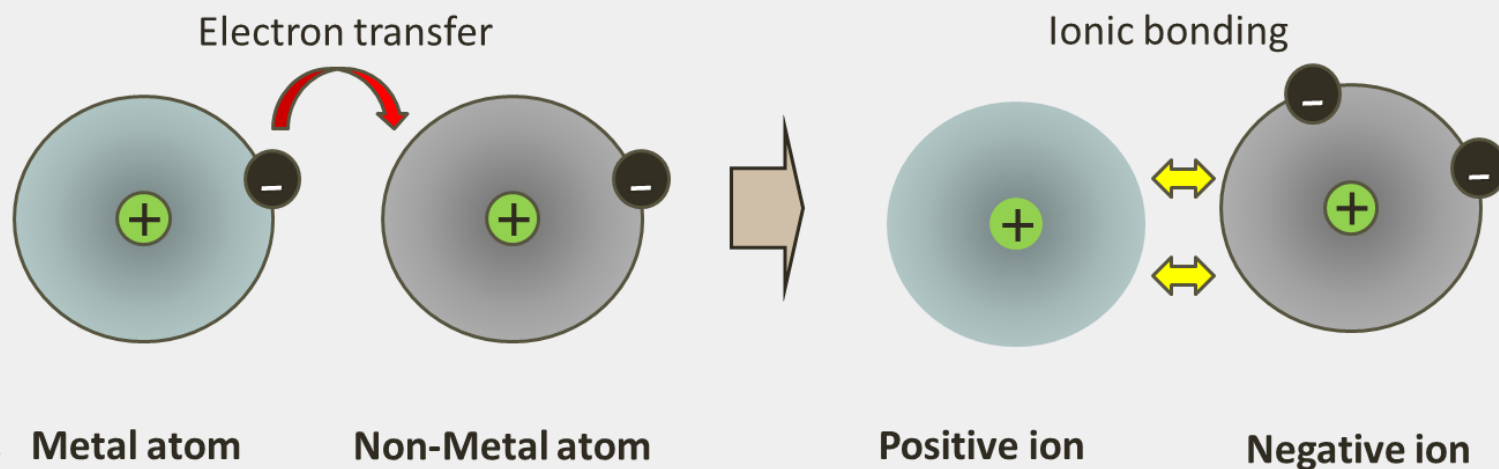




# Ionic Bonding

**Ionic Bonding** is where one atom completely takes valence electrons from another to form ions and the resulting negative and positive ions hold 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.





## Strength of Ionic Bonds

Two factors are involved with determining the strength of ionic bonds.

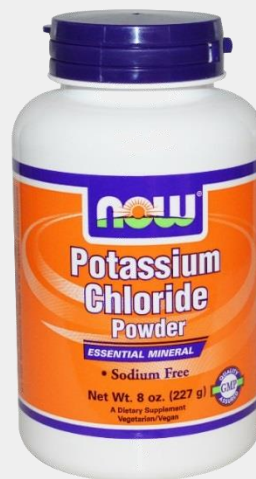
The first is the charge of the ions. A  $2+$  and  $2-$  (e.g.  $\text{MgO}$ ) bond will be much stronger than a  $+$  and  $-$  (e.g.  $\text{KCl}$ ) ionic bond.

Secondarily is the distance between the ions. The closer the ions, the stronger the electrostatic bond.



$\text{MgO}$

Melting point  $2,852\text{ }^{\circ}\text{C}$



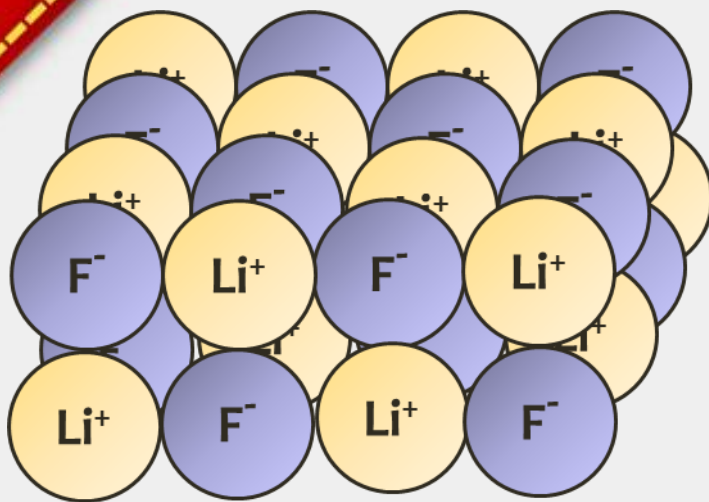
$\text{KCl}$

Melting point  $770\text{ }^{\circ}\text{C}$

# Background Knowledge

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.

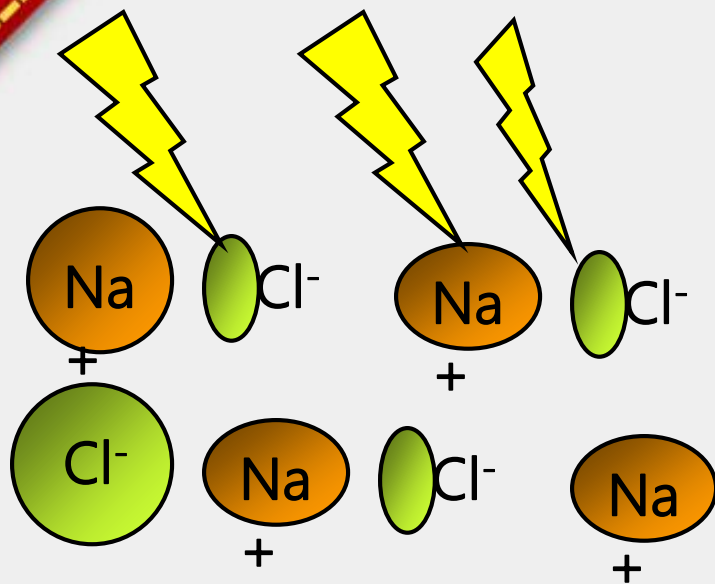


Compounds are neutral substances. For ionic compounds, the charges of the positive ions are balanced by the charges of the negative ions.

The Anion (F) takes the electrons from the Cation (Li) so their outer energy levels have are stable  
Anions and Cations have a strong electrostatic attraction for each other so they bond together as a compound.



An ionic solid is made up of ions held together by strong electrostatic attractions between +ve and -ve ions in a 3-dimensional lattice. Because these strong bonds require a large amount of energy to break the ionic solids have a high melting point.



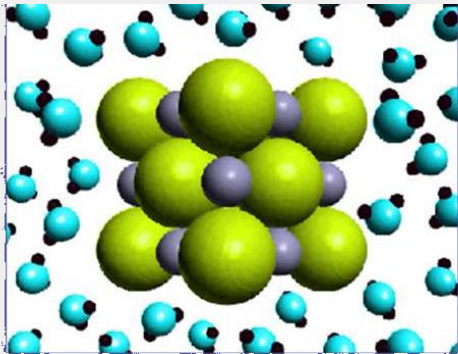
By adding heat energy to the ionic solid this is then transformed into kinetic energy. When the kinetic energy is high enough and the electrostatic bonds are broken then melting point has been reached. Ionic compounds with double charged ions (2+ and 2-) will have a higher melting point as the bonds require more energy to break.

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

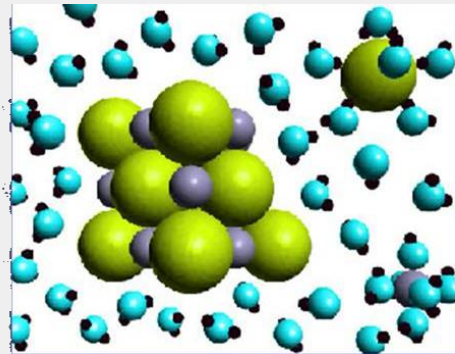
For example:

1. Sodium chloride (NaCl) is an ionic solid
2. Sodium chloride is made up of ions held together by strong 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 however
4. therefore the solid will dissolve and is soluble

**NaCl first place in water**



**Na<sup>+</sup> and Cl<sup>-</sup> ions breaking apart**

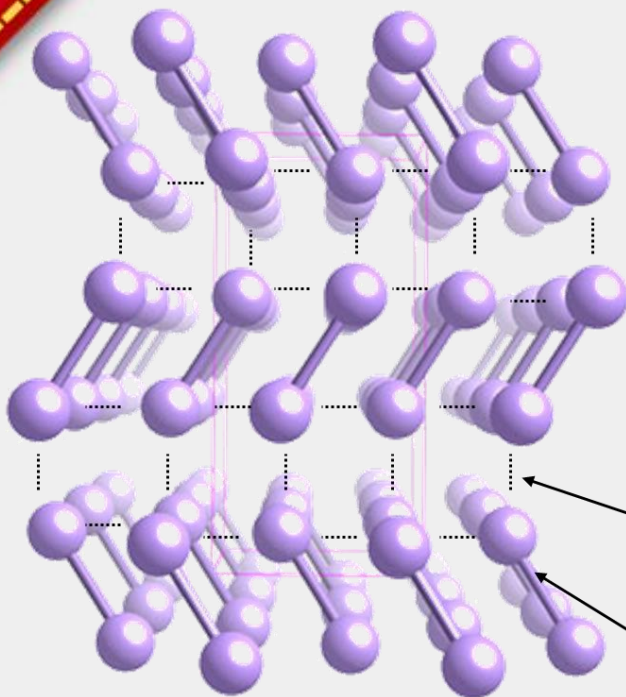


The positive hydrogen end of water is attracted to the anions and the negative oxygen end of water is attracted to the cations



## non-metal + non-metal

Non-polar molecules held together by intermolecular forces caused by temporary dipoles induced by electrons randomly spending more time around one nucleus than the other.

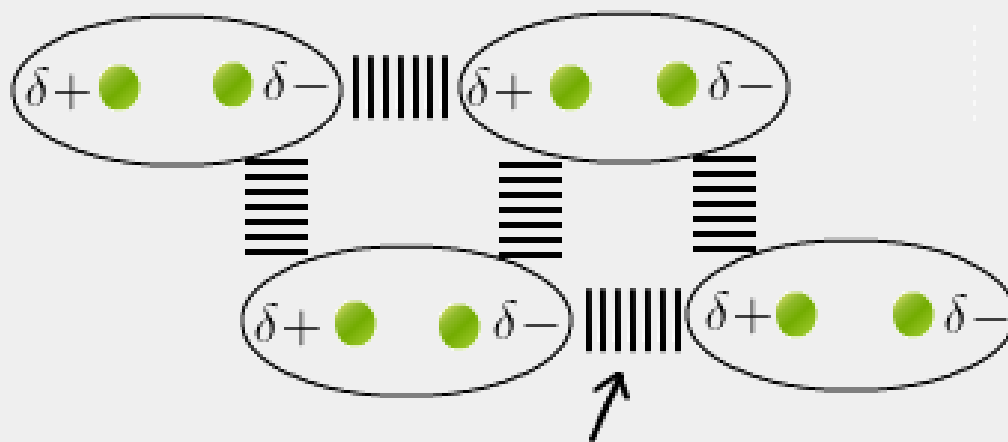


Weak intermolecular bond

Strong covalent bond



Polar molecules held together by **weak intermolecular** forces caused by permanent dipoles induced by electrons spending more time around one nucleus in the molecule that has greater electronegativity than the other. The  $\delta^-$  end of one molecule is attracted to the  $\delta^+$  end of another.



Intermolecular Force between dipoles



# Background Knowledge



## Solids Summary

Name of solid substance	Type of particle in solid	Attractive force broken when solid melts	Attractive force between particle –weak or strong	Relative melting point	solubility
molecular	molecules	Weak intermolecular van der waals	weak	low	Yes if polar No if non-polar
metal	atoms	Metallic bonding	strong	high	no
ionic	ions	Ionic bonding	strong	high	yes
covalent	atoms	Covalent bonding	strong	high	no

# Weak Intermolecular Force

Weak intermolecular forces of attraction occur between molecular solids

## 3 types of forces:

Instantaneous dipole (ID-ID) – temporary dipole – all molecules have this type

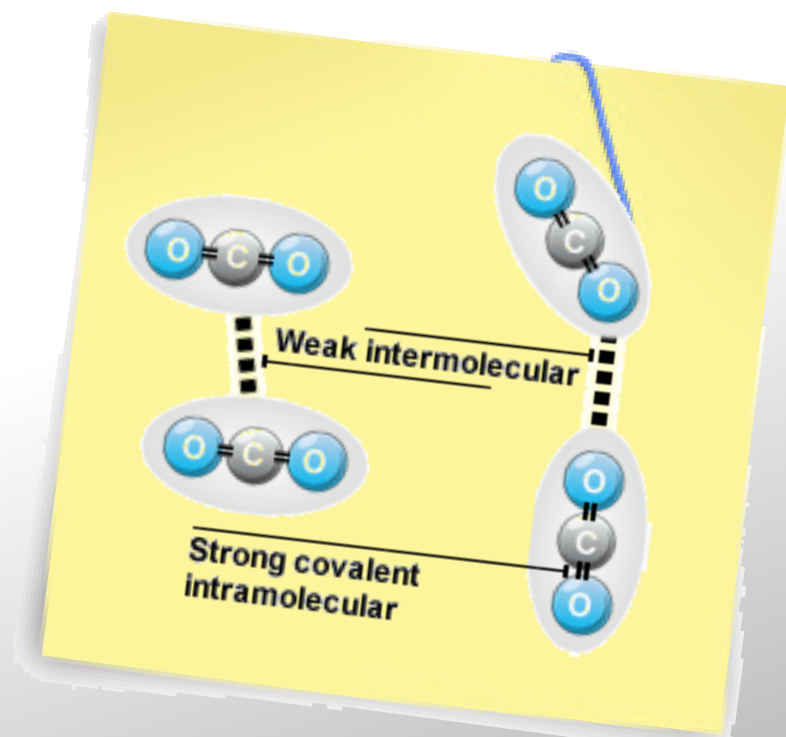
Permanent dipole (PD-PD) – permanent dipole – only polar molecules have this type **in addition** to ID-ID

Hydrogen bonding (HB) – permanent dipole – molecules with H bonded to O, N or F have this type of bonding

### 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.



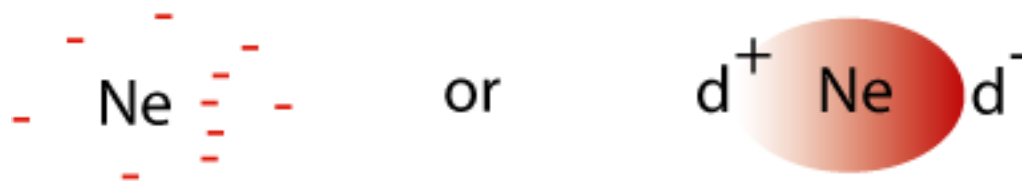
## Instantaneous dipole (ID – ID) Attractions

The inert gases (group 18) are monatomic, and must be non-polar, yet the atoms attract together to form liquids and freeze to form solids. Likewise, non-polar molecules like cyclohexane condense and freeze. This suggests that some kind of bonding force operates between **non-polar** monatomic elements or **non-polar** molecules.

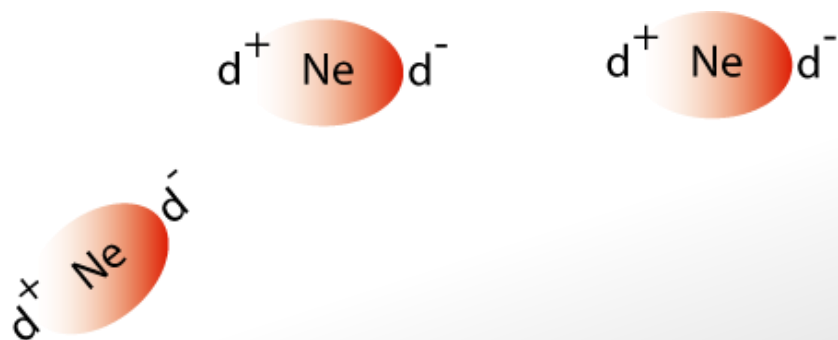
Gp 18	M.P. /°C		Gp 17	M.P. /°C		Misc.	M.P. /°C
He	-272		F <sub>2</sub>	-220		O <sub>2</sub>	-219
Ne	-249		Cl <sub>2</sub>	-101		N <sub>2</sub>	-210
Ar	-189		Br <sub>2</sub>	-7		S <sub>8</sub>	113
Kr	-157		I <sub>2</sub>	114		C <sub>6</sub> H <sub>12</sub>	6.6
Xe	-108					polythene	85 - 110

## Instantaneous/temporary Dipole

In any monatomic element or molecule, the electrons are moving rapidly. At any one instant, the arrangement of electrons about the nucleus will not be symmetric. This creates an instantaneous dipole (ID).



An atom with an ID will attract electrons in a neighbouring atom. This causes an induced dipole in this atom. The instantaneous dipole atom then attracts the induced dipole atom.

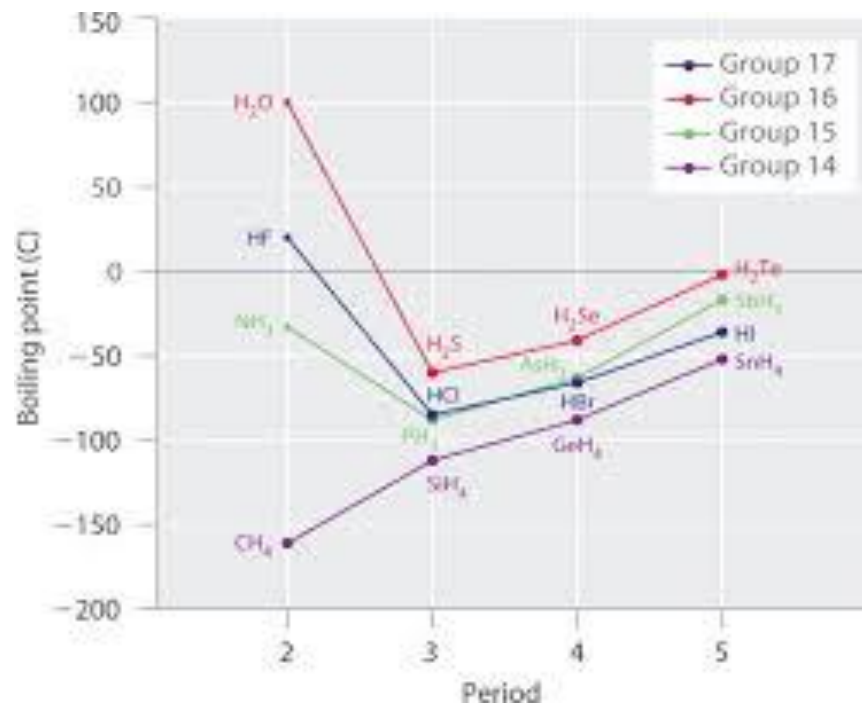


A instantaneous dipole (ID) in one atom or molecule will induce an ID in a neighbouring atom/molecule. The slightly negative end of one atom/molecule will repel electrons in a neighbouring atom/molecule and 'induce' that region to be more positive - and therefore ID-ID bonding occurs as the slightly negative end of one atom/molecule is attracted to the other slightly positive end of the other atom/molecule

## Increasing molar mass, increases the strength of the ID-ID bond

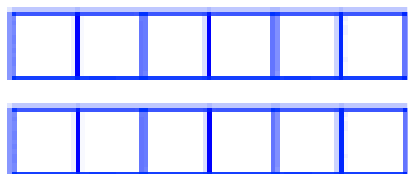
As molecules increase in molar mass, and therefore the total number of electrons, then the strength of the ID-ID bond increases as well. This is seen as an increase in melting point / boiling point of the larger molecules.

This is caused because the greater number of electrons are able to create a larger possible dipole (difference in negative and positive ends of the dipole) and strengthen the bonding between 2 molecules.

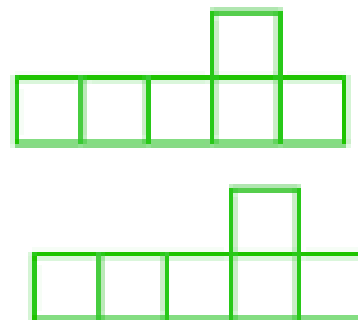


## Straight chained vs branching chained molecules

Molecules that have side chains do not “pack” together so well as single chain molecules so cannot get as close to other molecules – all things being equal **straight chained** versions of molecules tend to have **higher MP and BP**



melting point:  $-95^\circ\text{C}$



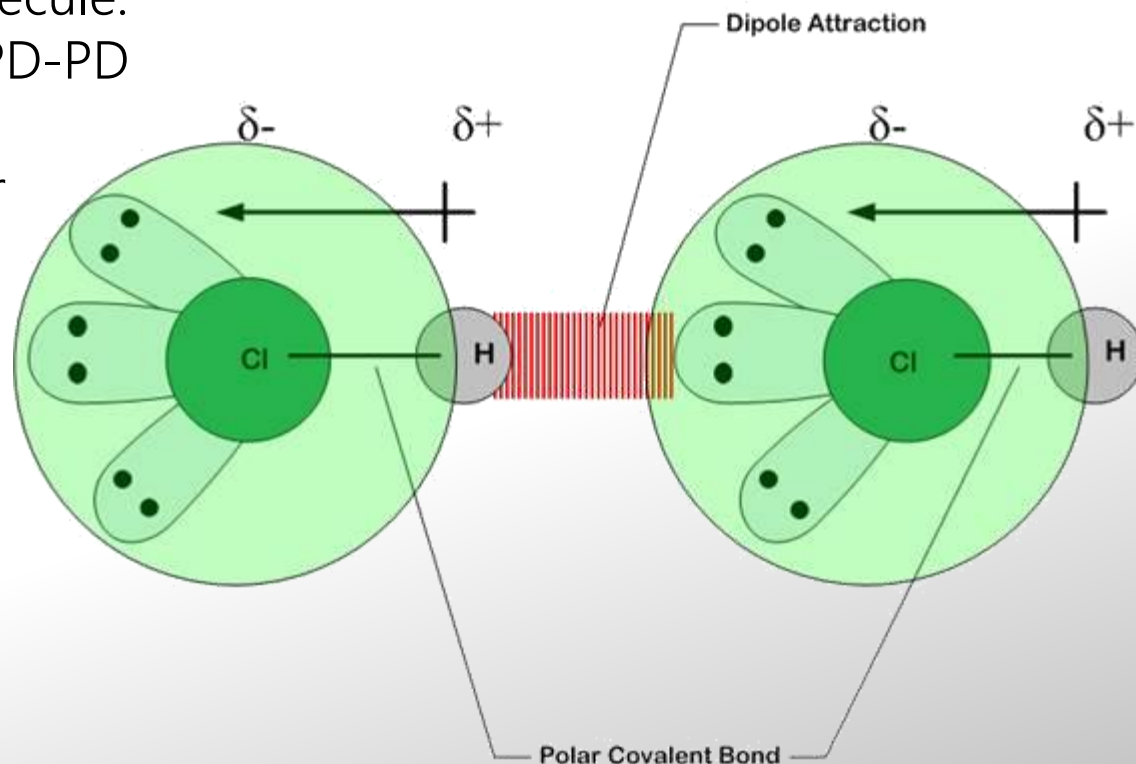
melting point:  $-154^\circ\text{C}$

# Permanent Dipole

## Permanent dipole (PD – PD) Attractions

Polar molecules have an imbalance of charge caused by an electronegativity difference between its constituent atoms. The weak molecular force always involves instantaneous dipole but in small molecules permanent dipole will play a larger part in the bonding. The slightly negative end of one molecule will be attracted to the slightly positive end of another molecule.

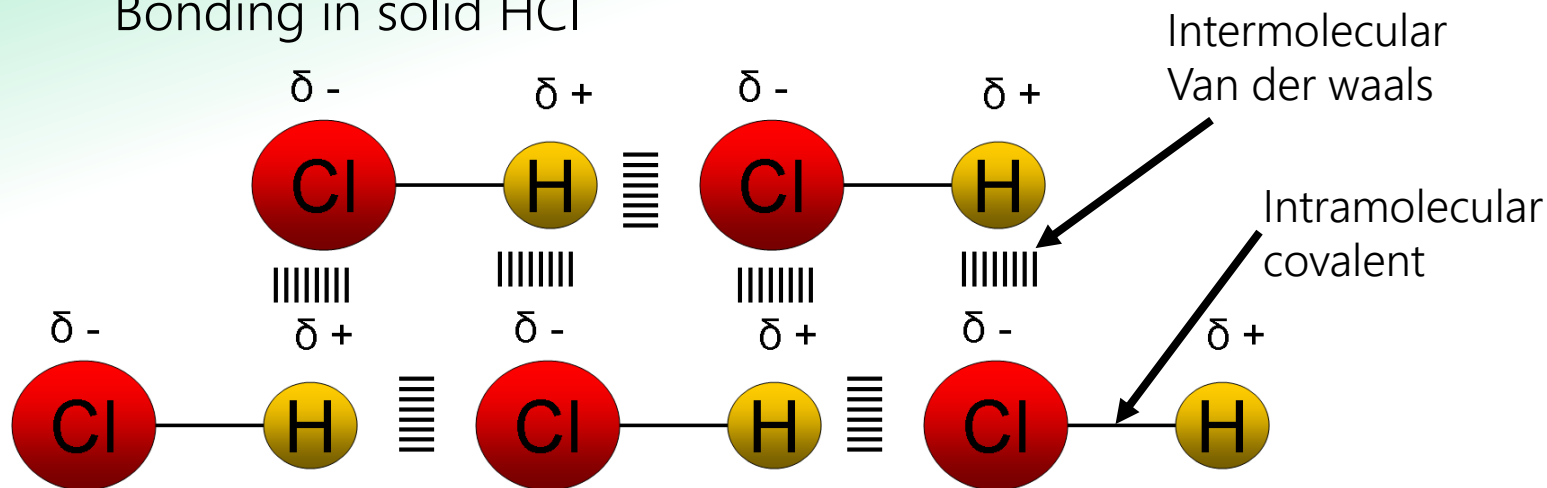
The higher strength of the PD-PD bonding means that polar molecules will have a higher boiling/melting point to non-polar molecules (with only ID-ID bonding) of similar molar mass.





## Permanent Dipole (PD – PD) Attractions

e.g. Bonding in solid HCl



The weakness of PD – PD bonding is revealed by the low melting points of PD molecular solids, by comparison with other types of solids.

e.g.	hydrogen chloride (PD – PD)	-114 °C
	acetone (PD – PD)	-95 °C
	NaCl (ionic)	801 °C
	copper (metallic)	1085 °C

Molecules that have side chains do not “pack” together so well as single chain molecules so cannot get as close to other molecules – all things being equal **straight chained** versions of molecules tend to have **higher MP and BP**

## Strength of ID-ID attractions compared with PD-PD

ID – ID attractions increase in strength as the number of electrons in the atom or molecule increase. See trend He – Xe. Increasing MPs show increasing force of intermolecular bonds.

Small molecules that have 'ID-ID only' have lower MPs than small molecules with both PD-PD and ID-ID, indicating that 'ID-ID only' is a weaker bond than 'PD-PD + ID-ID'. e.g.

	Type	Num e-	MP (°C)
F <sub>2</sub>	ID-ID	18	-220
HCl	ID-ID + PD-PD	18	-114

## Strength of ID-ID attractions compared with PD-PD

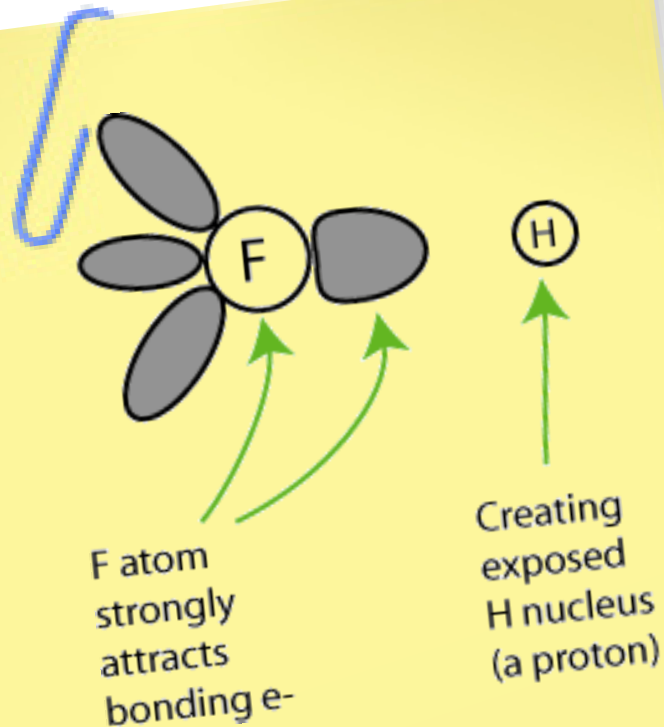
When both ID-ID and PD-PD are present in a series of compounds, which has the greater influence on MP?

	Ne	MP (°C)
HCl	18	-114
HBr	36	-87
HI	54	-51

ID-ID increase down this series due to increasing number of e<sup>-</sup> per molecule. But PD-PD increase going up series due to increasing electro-negativity difference between the atoms in the molecules. As MPs increase going down the series it seems that increasing ID-ID is making the stronger contribution to inter-molecular forces. For polar molecules with higher numbers of electrons such as HI, ID-ID makes a stronger contribution to intermolecular forces than PD-PD.

# Hydrogen Bonding

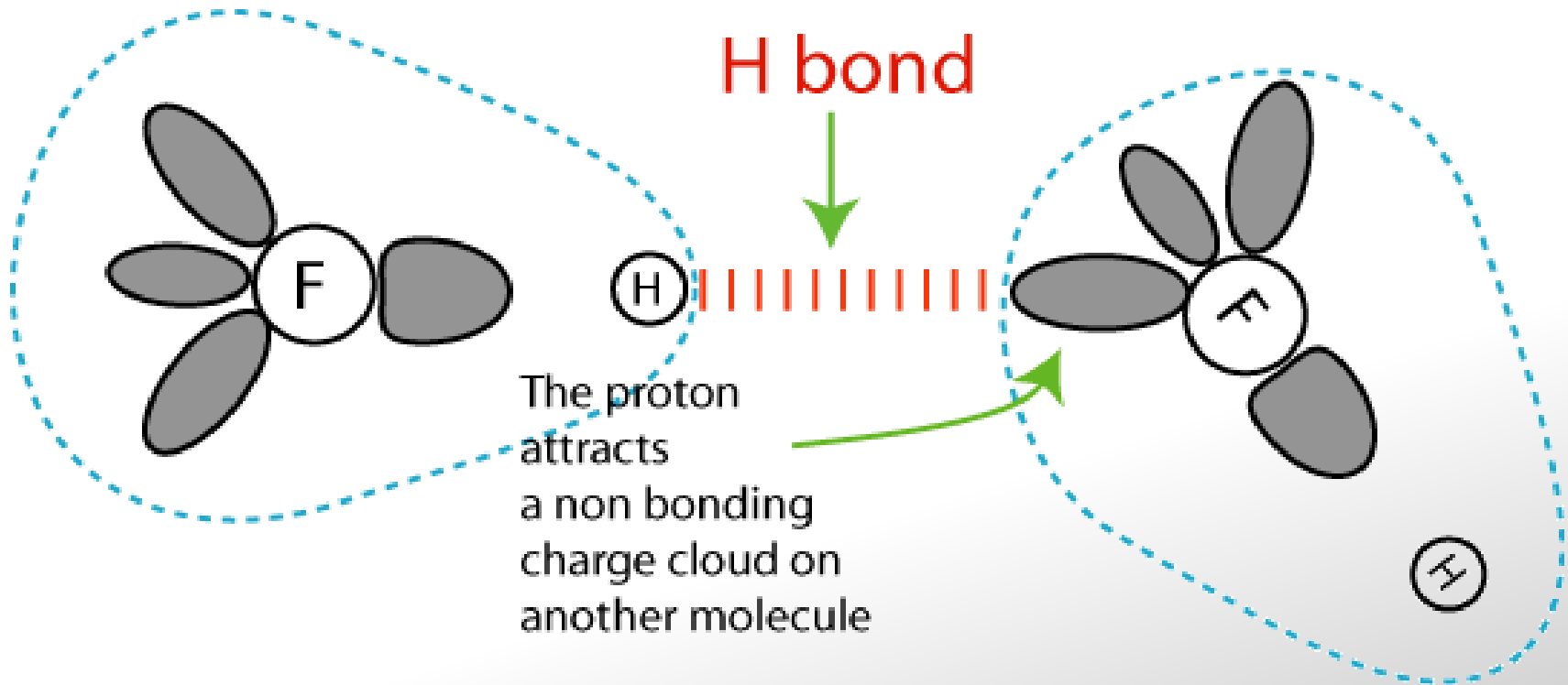
This is a special type of PD – PD bonding. It occurs when **hydrogen is bonded to N, O or F**. Since these are the three most electronegative elements and since they have very small atoms, (which concentrates their charge), they have a strong polarising effect on the covalent bond with hydrogen. Hydrogen, having no inner electron shells, is left as an exposed proton.



This proton, being very small indeed, exerts a strong attraction to a non-bonding pair of e<sup>-</sup> on a neighbouring molecule.

# Hydrogen Bonding

The hydrogen bond is around 10% of the strength of a covalent bond, and is the strongest of the three types of intermolecular attraction.



## Hydrogen Bonding anomalies



Boiling points of ammonia, water and hydrogen fluoride.

On average:

Ammonia can form two H-bonds per molecule.

(1 for the three H and 1 for the non-bonding pair of e-)

Water can form four H-bonds per molecule.

(2 per H and 2 per non-bonding pair of e-).

Hydrogen fluoride can form two H-bonds per molecule.

(1 for the three non-bonding pair of e- and 1 for the H)

	Ne	MP (°C)		Ne	MP (°C)			Ne	MP (°C)
NH <sub>3</sub>	17	-33	H <sub>2</sub> O	10	0		HF	10	-83
			H <sub>2</sub> S	18	-86		HCl	18	-114
			H <sub>2</sub> Se	36	-64		HBr	36	-87
			H <sub>2</sub> Te	54	-49		HI	54	-51

This accounts for the anomalously high M.P.s for NH<sub>3</sub>, H<sub>2</sub>O and HF when considered alongside the other hydrogen compounds of their groups.



# Hydrogen Bonding of Water

The high surface tension of water is explained by the presence of a hydrogen-bonded network of water molecules at the surface. This network is sufficiently strong to enable a needle to be floated on the surface of the water.

H bonding gives ice a network tetrahedral structure (like diamond). When ice melts, this partially collapses, and molecules move closer together so water is denser.

(density-water =  $1.0 \text{ g cm}^{-3}$       density-ice =  $0.92 \text{ g cm}^{-3}$ )

In contrast most molecular solids expand on melting.



## Answering Intermolecular Questions

<b>Non-Polar</b> Temporary (Instantaneous) dipole ID - ID	<b>Polar</b> Permanent dipole +Instantaneous dipole PD – PD + ID - ID	<b>Polar (H-O, H-N, H-F)</b> Hydrogen Bonding + Instantaneous dipole HB – HB + ID - ID
<p>The (.....) molecular solid is non-polar and so the only intermolecular forces would be due to temporary dipole interactions.</p> <p>This type of intermolecular bonding is the weakest of the three so these molecular solids will have the lowest boiling point.</p> <p>Generally as the molar mass increases there are more electrons and more instantaneous dipole-dipole interactions so all things being equal, the boiling points would increase.</p>	<p>The (.....) molecular solid is polar and contains both Permanent dipole and instantaneous dipole interactions.</p> <p>The permanent dipole is generally the stronger force and will be the dominant force found in the molecular solid and therefore these substances tend to have a higher boiling point than the non-polar substances.</p> <p>The instantaneous dipole forces will become more significant if the molar mass of the molecule becomes greater.</p>	<p>The (.....) molecular solid has the highest boiling point of the three. This is because it has hydrogen bonds which are very strong intermolecular forces.</p> <p>The electronegativity difference between the H-.... Bond creates a very polar molecule which takes a lot of energy to break the intermolecular bond between molecules hence a high boiling point.</p> <p>These molecular substances also contain instantaneous dipoles which only become significant when the molar mass is large.</p>

## NCEA 2013 Intermolecular Forces

Excellence  
Question

**Question 3a:** Use the information in the table above to compare and contrast the boiling points of hydrazine, fluoromethane, and decane in terms of the relative strengths of the attractive forces between the particles involved.

Molecule	Boiling point/ °C
Hydrazine, $\text{N}_2\text{H}_4$	114
Fluoromethane, $\text{CH}_3\text{F}$	-78.4
Decane, $\text{C}_{10}\text{H}_{22}$	174

$\text{N}_2\text{H}_4$  is a polar molecule. (Due to the presence of the highly polar N-H bonds), there is hydrogen bonding between  $\text{N}_2\text{H}_4$  molecules.

$\text{CH}_3\text{F}$  is also a polar molecule. (Due to the presence of the C-F bond), there are permanent dipole attractions between the  $\text{CH}_3\text{F}$  molecules.

The attractive forces due to permanent dipoles in  $\text{CH}_3\text{F}$  must be weaker than the attractive forces due to hydrogen bonding in  $\text{N}_2\text{H}_4$ , because  $\text{CH}_3\text{F}$  boils at a lower temperature and they are similar masses so temporary dipole attractions are similar.

$\text{C}_{10}\text{H}_{22}$  is a non-polar molecule. The only attractive forces between the  $\text{C}_{10}\text{H}_{22}$  molecules are due to temporary dipoles. However, since  $\text{C}_{10}\text{H}_{22}$  is a significantly larger molecule than  $\text{N}_2\text{H}_4$ , and  $\text{CH}_3\text{F}$ , it is more polarisable / has more electrons / greater molar mass, so its temporary dipole attractions are even stronger than the hydrogen bonds in  $\text{N}_2\text{H}_4$ . As a result,  $\text{C}_{10}\text{H}_{22}$  requires the most heat energy to break its intermolecular forces and therefore has the highest boiling point.

## NCEA 2014 Intermolecular Forces

Achieved  
Question

**Question 2a:** The boiling points of ammonia,  $\text{NH}_3$ , fluorine,  $\text{F}_2$ , and hydrogen chloride,  $\text{HCl}$ , are given in the table below.

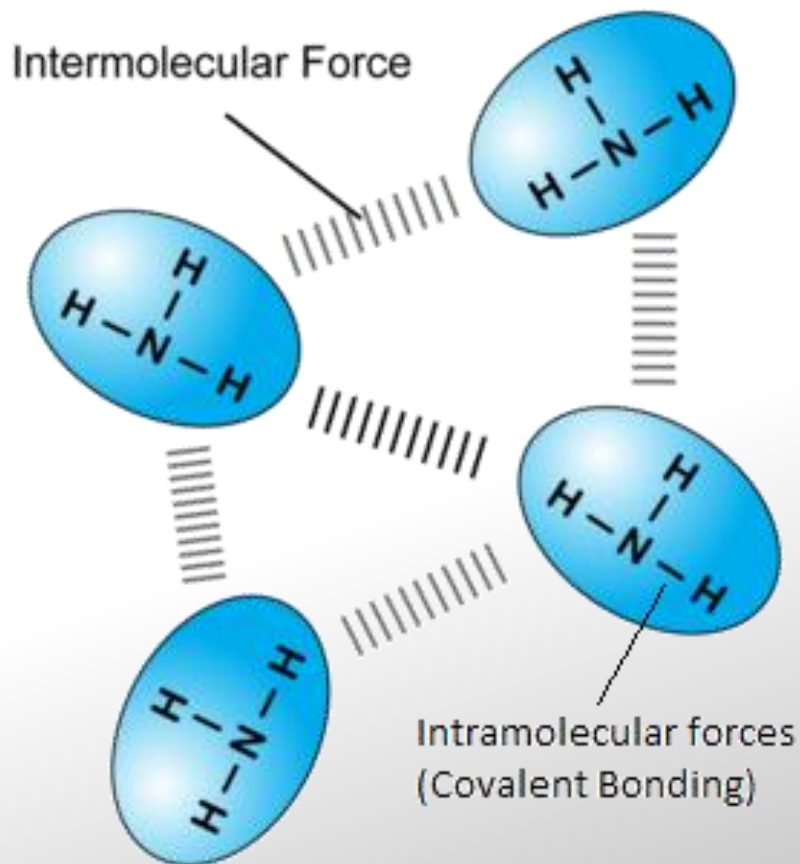
Complete the table to identify the attractive forces between the molecules in their liquid state.

Molecule	Boiling point/ $^{\circ}\text{C}$	Attractive forces
Ammonia, $\text{NH}_3$	-33	Hydrogen bonds, instantaneous dipoles
Fluorine, $\text{F}_2$	-188	Instantaneous dipoles
Hydrogen chloride, $\text{HCl}$	-85	Permanent dipoles, instantaneous dipoles

In answering the question don't forget that all molecules have instantaneous dipoles, plus PD if they are polar or HD if they have H-O, H-N, or H-F bonds

**Question 2b:** Discuss the differences between the boiling points of  $\text{NH}_3$  and  $\text{HCl}$ , in terms of the strength of the attractive forces between the particles involved. Then describe why  $\text{F}_2$  has the lowest boiling point.

**Answer 2b:**  $\text{NH}_3$  and  $\text{HCl}$  both have temporary and permanent dipoles, as they are polar molecules. However,  $\text{NH}_3$  has H-bonding, which means the boiling point is higher due to these stronger forces of attraction.  $\text{HCl}$  has a permanent dipole, but not H-bonding.  $\text{F}_2$  has the lowest boiling point, due to having only temporary dipoles. This type of intermolecular bonding is the weakest of the three so these molecular solids will have the lowest boiling point.



## NCEA 2015 Intermolecular Forces

Achieved  
Question

**Question 3c:** The two molecules below have the same molecular formula ( $C_5H_{12}O$ ) but have different boiling points.

- List all the forces of attraction between these molecules in each of their liquid states.
- Use the information above to explain the difference in the boiling points of pentan-1-ol and dimethylpropan-1-ol by comparing and contrasting the relative strengths of the attractive forces between the molecules involved.

Name	Pentan-1-ol	Dimethylpropan-1-ol
Structure	$CH_3 - CH_2 - CH_2 - CH_2 - CH_2 - OH$	$  \begin{array}{c}  CH_3 \\    \\  CH_3 - C - CH_2 - OH \\    \\  CH_3  \end{array}  $
Boiling point	138°C	113°C

In answering the question cross off all of those factors that are the same (Molar mass, HB and PD bonding) and look for differences – side chains and no side chains – How does this effect bonding [the closer, the stronger the bond]



# NCEA 2015 Intermolecular Forces

Excellence  
Question

Name	Pentan-1-ol	Dimethylpropan-1-ol
Structure	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{OH}$	$  \begin{array}{c}  \text{CH}_3 \\    \\  \text{CH}_3 - \text{C} - \text{CH}_2 - \text{OH} \\    \\  \text{CH}_3  \end{array}  $
Boiling point	138°C	113°C

Answer 3c: (i) FORCES

- hydrogen bonding
- permanent dipoles
- instantaneous dipoles.

(ii) The attractive forces due to the **hydrogen bonding and permanent dipoles are similar** between the molecules in both liquids, as they both have one OH group, which causes the molecule to be polar and take part in hydrogen bonding.

The **two molecules have the same mass**, and so the same number of electrons involved in the weak instantaneous dipoles.

However, the **pentan-1-ol molecule has no side chains** and so the main chains can get closer to each other (less steric hindrance, greater surface area), thus the instantaneous dipoles are stronger / greater in pentan-1-ol, and therefore the boiling point is higher.

## NCEA 2016 Intermolecular Forces

Achieved  
Question

**Question 2a:** Identify all the attractive forces between particles of the following compounds in their liquid state.

Compound	$\Delta_{\text{vap}}H^\circ / \text{kJ mol}^{-1}$	Attractive forces
NaCl	194	NaCl: <b>ionic bonds</b> .
HCl	16.0	HCl: Permanent dipole-dipole attractions, temporary dipole-dipole attractions.
CH <sub>3</sub> Cl	22.0	CH <sub>3</sub> Cl: Permanent dipole-dipole attractions, temporary dipole-dipole attractions.

This year they also included an example of a molecule with ionic bonds –  
**Watch for this**

## NCEA 2016 Intermolecular Forces

**Question 2b (i) :** The standard enthalpy of vaporisation,  $\Delta_{\text{vap}}H^\circ$ , of sodium chloride, NaCl, hydrogen chloride, HCl, and chloromethane, CH<sub>3</sub>Cl, are given in the table below. Explain why  $\Delta_{\text{vap}}H^\circ(\text{NaCl})$  is significantly higher than both  $\Delta_{\text{vap}}H^\circ(\text{HCl})$  and  $\Delta_{\text{vap}}H^\circ(\text{CH}_3\text{Cl})$ .

Compound	$\Delta_{\text{vap}}H^\circ / \text{kJ mol}^{-1}$
NaCl	194
HCl	16.0
CH <sub>3</sub> Cl	22.0

**Answer 2b: (i)** Much more heat energy is required to overcome the attraction between its particles and convert NaCl from a liquid to a gas than HCl and CH<sub>3</sub>Cl, because NaCl has strong ionic bonding between its ions compared to weak intermolecular bonding between the HCl and CH<sub>3</sub>Cl molecules.

**Question 2b:(ii)** Explain why  $\Delta_{\text{vap}}H^\circ(\text{CH}_3\text{Cl})$  is greater than  $\Delta_{\text{vap}}H^\circ(\text{HCl})$ .

**Answer 2b: (ii)** Both HCl and CH<sub>3</sub>Cl are polar molecules and therefore have permanent dipole-dipole attractions and temporary dipole-dipole attractions between their molecules. However, CH<sub>3</sub>Cl has a larger molar mass and therefore more electrons, so its temporary dipole-dipole attractions are stronger than between the HCl molecules. This means more heat energy is required to overcome the attractions between liquid CH<sub>3</sub>Cl molecules, so it has a higher  $\Delta_{\text{vap}}H^\circ$ .

## NCEA 2017 Intermolecular Forces

Excellence  
Question

Molecule	Boiling Point / °C	$M / \text{g mol}^{-1}$
Hydrazine, $\text{N}_2\text{H}_4$	114	32
Iodomethane, $\text{CH}_3\text{I}$	42.4	142
Decane, $\text{C}_{10}\text{H}_{22}$	174	142

**Question 2a:** Use the information in the table above to compare and contrast the boiling points of the substances below.

In your answers, you should:

- list the types of intermolecular forces present for each substance
- explain the relative strength between the particles involved.

(a) (i) Hydrazine and iodomethane.

**Hydrazine** ( $\text{N}_2\text{H}_4$ ) hydrogen bonding, (permanent dipole attractions) temporary dipole attractions.

For **iodomethane** ( $\text{CH}_3\text{I}$ ) permanent and temporary dipole attractions.

The hydrogen bonds between  $\text{N}_2\text{H}_4$  molecules are stronger than the permanent dipole forces between  $\text{CH}_3\text{I}$  molecules therefore require more energy to break resulting in a higher boiling point.

The presence of hydrogen bonding outweighs the expected higher temporary dipole in  $\text{CH}_3\text{I}$  due to the greater molar mass.

## NCEA 2017 Intermolecular Forces

Excellence  
Question

Molecule	Boiling Point / °C	$M$ / g mol <sup>-1</sup>
Hydrazine, N <sub>2</sub> H <sub>4</sub>	114	32
Iodomethane, CH <sub>3</sub> I	42.4	142
Decane, C <sub>10</sub> H <sub>22</sub>	174	142

**Question 2a:** Use the information in the table above to compare and contrast the boiling points of the substances below.

(a)(ii) Iodomethane and decane.

**Iodomethane** (CH<sub>3</sub>I) has permanent and temporary dipole attractions.

**Decane** (C<sub>10</sub>H<sub>22</sub>) has temporary dipole attractions.

Despite the molecules having the same molar mass, decane C<sub>10</sub>H<sub>22</sub> has stronger intermolecular attractions. Decane is a longer molecule, when compared to the spherical shape of iodomethane CH<sub>3</sub>I, so it has a greater surface area / electron cloud / number of electrons meaning stronger temporary dipole attractions.

Therefore, more energy is needed to break these attractions, resulting in a higher boiling point.

## NCEA 2017 Intermolecular Forces

MERIT  
Question

Molecule	Boiling Point / °C	$M$ / g mol <sup>-1</sup>
Hydrazine, N <sub>2</sub> H <sub>4</sub>	114	32
Iodomethane, CH <sub>3</sub> I	42.4	142
Decane, C <sub>10</sub> H <sub>22</sub>	174	142

**Question 2b:** Explain why the solubility of hydrazine in water is greater than that of decane in water.

Hydrazine is a polar molecule. Decane is non-polar. As water is a polar solvent, the hydrazine will be more soluble than the decane.

The attractive forces between the molecules of hydrazine are less than the attractive forces between the hydrazine and water molecules, and therefore it is

more soluble than decane, where the attractive forces between the decane molecules are greater than the attractive forces between the decane and water molecules.

## NCEA 2017 Intermolecular Forces

MERIT  
Question

**Question 3a:** Chlorine,  $\text{Cl}_2$ , bromine,  $\text{Br}_2$ , and iodine,  $\text{I}_2$ , are all halogens. Bromine is a liquid at room temperature.

(i) In the box below, tick the type(s) of intermolecular attractions in **liquid** bromine.

Intermolecular attraction	Tick (✓)
Temporary dipole-dipole attractions	✓
Permanent dipole-dipole attractions	
Hydrogen bonding	

(ii) Explain why bromine is a liquid at room temperature, whereas chlorine is a gas.

Bromine is a larger molecule than chlorine so the temporary dipole intermolecular attractions are greater. Thus, more energy is required to separate the molecules and so  $\text{Br}_2$  has a higher boiling point and is a liquid at room temperature

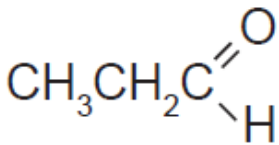




## NCEA 2018 Intermolecular Forces

**Q 2a (i):** The standard enthalpy of vaporisation,  $\Delta_{\text{vap}}H^\circ$ , of methanol, propan-1-ol, and propanal, are given in the table below.

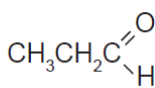
List all the forces of attraction between the molecules in their liquid state.

Molecule	$\Delta_{\text{vap}}H^\circ$ /kJ mol <sup>-1</sup>	$M$ /g mol <sup>-1</sup>	Attractive forces
Methanol CH <sub>3</sub> -OH	38	32	Methanol: hydrogen bonding, permanent dipole attractions, temporary dipole attractions.
Propan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> -OH	47	60	Propan-1-ol: hydrogen bonding, permanent dipole attractions, temporary dipole attractions.
Propanal 	30	58	Propanol: permanent dipole attractions, temporary dipole attractions.

## NCEA 2018 Intermolecular Forces

**Q 2a (ii):** The standard enthalpy of vaporisation,  $\Delta_{\text{vap}}H^\circ$ , of methanol, propan-1-ol, and propanal, are given in the table below.

Compare and contrast the enthalpy of vaporisation of methanol, propan-1-ol, and propanal. Your answer should include an explanation of the relative strength of the attractive forces between the molecules.

Molecule	$\Delta_{\text{vap}}H^\circ$ /kJ mol <sup>-1</sup>	<i>M</i> /g mol <sup>-1</sup>
Methanol CH <sub>3</sub> -OH	38	32
Propan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> -OH	47	60
Propanal 	30	58

Both methanol and propan-1-ol have temporary dipoles, permanent dipoles, and hydrogen bonding between the molecules. Since propan-1-ol has a larger electron cloud than methanol, it has stronger temporary dipole attractions. As a result, propan-1-ol has a higher enthalpy of vaporisation/ requires more energy to separate the molecules than methanol. Propanal has temporary dipoles and permanent dipoles between the molecules.

Permanent dipole attractions are weaker than hydrogen bonding, so propanal has a lower enthalpy of vaporisation/requires less energy to separate the molecules than methanol and propan-1-ol.

Even though propanal and propan-1-ol have electron clouds of similar size and would therefore have temporary dipole attractions of similar strength, the hydrogen bonding in propan-1-ol has more influence on the enthalpy of vaporisation of/energy required to separate than the permanent dipole attractions.

## NCEA 2019 Intermolecular Forces

Achieved  
Question

**Question 3a:** List all the forces of attraction between the following molecules in their liquid state.

Molecule	Boiling point/ °C	Attractive forces
Ammonia, $\text{NH}_3(\ell)$	-33.3	$\text{NH}_3$ Hydrogen bonding, (permanent dipoles), temporary dipoles.
Ethane, $\text{C}_2\text{H}_6(\ell)$	-88.6	$\text{C}_2\text{H}_6$ Temporary dipoles / instantaneous dipoles.
Methanamine, $\text{CH}_3\text{NH}_2(\ell)$	-6.3	$\text{CH}_3\text{NH}_2$ Hydrogen bonding, permanent dipoles, temporary dipoles..

## NCEA 2019 Intermolecular Forces

Excellence  
Question

**Question 3b:(i)** Using the data in the below table, identify the molecule that has the strongest forces of attraction between its molecules.

Molecule	Boiling point/ °C	Attractive forces
Ammonia, $\text{NH}_3(\ell)$	-33.3	$\text{NH}_3$ Hydrogen bonding, (permanent dipoles), temporary dipoles.
Ethane, $\text{C}_2\text{H}_6(\ell)$	-88.6	$\text{C}_2\text{H}_6$ Temporary dipoles / instantaneous dipoles.
Methanamine, $\text{CH}_3\text{NH}_2(\ell)$	-6.3	$\text{CH}_3\text{NH}_2$ Hydrogen bonding, permanent dipoles, temporary dipoles..

Methanamine and ethane have electron clouds of similar size / similar molar mass and would therefore have intermolecular temporary dipole attractions of similar strength. However, methanamine also has hydrogen bonding due to the N–H bond. As this is the strongest intermolecular force, it requires a larger amount of heat energy to break. Therefore, **methanamine** has a higher boiling point than ethane.

**Question 3b:** (ii) Justify why methanamine has a higher boiling point than ethane.

**Question 3b:(iii)** Justify why methanamine has a higher boiling point than ammonia.

Both methanamine and ammonia have intermolecular hydrogen bonding due to the N–H bond, which causes strongest type of intermolecular force. However, methanamine has a significantly larger electron cloud / larger molar mass. This means the temporary dipole attractions between methanamine molecules will be stronger, and will therefore require more heat energy to break. So, methanamine has a higher boiling point than ammonia.