

[A First Year blog on Melting and Boiling Points: Part 3, 16th February 2019](#)

(This is a **revision** blog. It is merely a *summary* of some of the things that you need to know. *Everything in this blog can be found in the First Year, Foundation Book.*)

This four-part blog has so far consisted of

Part 1 An examination of what happens when things Melt and Boil.

Part 2 An explanation of the forces that hold (i) solid molecular substances and (ii) liquid molecular substances together viz. Hydrogen bonds, and the three forces of attraction known collectively as van der Waals'/London/Debye forces of attraction. In general these are

- A Hydrogen bonds, and
- B1 Permanent dipole to Permanent dipole forces of attraction
- B2 Permanent Dipole to Induced dipole forces of attraction, and
- B3 Temporary/Momentary/Instantaneous dipole to Induced dipole forces of attraction.

In this week's blog I shall concentrate on

(Part 3) A detailed examination of 'Hydrogen Bonds', and a bit more about 'dipoles'.

In next week's blog I shall talk to you about

(Part 4) The sort of questions that are asked in the 'A' Level exams requiring candidates to explain the Melting and Boiling temperatures of different substances.

However, there is a problem in that many schools have stopped submitting their students for the First Year 'A' Level exams. In the State Grammar School in which I taught last Autumn, the school saved **£70,000** by not entering pupils for the First Year 'A' Level exams (and instead, it set internal exams for the Lower Sixth students to take). The sort of questions that are asked in the Second Year exams are not the same as those in the First Year. In a sense therefore this four-part blog on Melting and Boiling Points is a bit more advanced than First Year stuff, but the stuff on Hydrogen Bonds this week will be helpful to those students wanting to get into Medicine.

In last week's blog I told you about

- the permanent total separation of charge (as in *ionic substances* such as NaCl)
 - the permanent partial separation of charge (as in *polarised molecules* such as Water, H₂O), and
 - the temporary partial separation of charge (as in *non-polarised molecules* such as CCl₄),
- and we discussed the forces that underlie **inter**-molecular forces of attraction¹ i.e. the forces of attraction between molecules that hold molecular substances together viz.

- A) Hydrogen bonds, and
- B) the three forces of attraction known collectively as Van der Waals/London forces of attraction viz.
 - B1) Permanent dipole to Permanent dipole forces of attraction²
 - B2) Permanent dipole to Induced dipole forces of attraction, and
 - B3) Temporary/Momentary/Instantaneous dipole to Induced dipole forces of attraction.

B3 is always present as an inter-molecular force of attraction, but B1 and B2 are much more powerful and can overwhelm it. However, I did not actually tell you much about the forces themselves, therefore in this week's blog, let me do so.

¹ **Intra**-molecular forces of attraction hold the atoms in a molecule together, while **Inter**-molecular forces of attraction are the forces of attraction between molecules that hold the molecules together.

² A dipole exists whenever there is a separation of charge (**no matter how small that separation of charge may be**). A total separation of charge would result in **ions** in an ionic substance rather than in a molecular substance.

A) Hydrogen Bonds

I have already said elsewhere that in a future blog I will tell you how the front bit of one molecule can grab hold of and not let go of the end bit of another molecule – and indeed the principles that I will outline will form the basis of something called “polymers” in “Organic Chemistry” (where, to use a simile, the head of one snake bites the tail of another snake, which then bites the tail of another snake, which in turn bites the tail of another snake and so on). However, for those of you who want to go on and do medicine of some sort (i.e. become nurses/pharmacists/doctors/dentists/or whatever) then it is imperative that you understand DNA/genes/chromosomes/proteins/nucleic acids/peptides, etc – so please do assimilate the principles of [Hydrogen Bonding](#) that are in this blog because it is one of the building blocks of many of the above structures.

Hydrogen bonds and van der Waals' forces of attraction³ are the two simplest **inter**-molecular forces of attraction that we now need to study.

Could I remind you that, by definition, for a bond to form **there must be a pair of bonding electrons**, and if no pairs of bonding electrons are involved, then no bond has been formed.

In an “Hydrogen Bond” a pair of electrons is involved, but (as we shall see) the two electrons are only partially involved. A true “bond” is thus not formed, therefore “Hydrogen Bonds” are not **bonds** in the strictest sense of the word – but they have now been called Hydrogen “*Bonds*” for so long that chemists just accept this name/this nomenclature for this force of attraction.

However, no *pairs of electrons* are directly involved in van der Waals' forces of attraction, therefore van der Waals' forces of attraction should NOT be called “bonds” at all, and instead they should be called just plain and simple “forces of attraction”. **vdW forces are not bonds. They are just forces of attraction!**

A Hydrogen bond is a quasi-bond⁴ that is created between two molecules each of which has a **permanent dipole (that is caused by a permanent *partial* separation of charge in each molecule)** – and the distinguishing features of the Hydrogen bond are that

- i) **Hydrogen bonds can be formed only by compounds that contain at least one Hydrogen atom bonded to either a [Fluorine](#)/or an [Oxygen](#)/or a [Nitrogen](#) atom, (and H bonds will NOT form with any other elements other than F/or O/or N), and**
- ii) **Hydrogen bonds are formed between an H^{δ+} species and a Lone Pair of electrons on the F/or O/ or N atom.**⁵

If you were to consult a data book, you would see that the three most electronegative elements in the Periodic Table are Fluorine (F)/ Oxygen (O)/and Nitrogen (N), and if an atom of any of these three elements is bonded to an Hydrogen atom, then such an atom will pull or *heavily* distort the electron cloud/the electron density distribution (edd) of the lone electron in the Hydrogen atom towards itself.⁶

³ I shall not attempt to differentiate between the work done by Johannes Diderik van der Waals and Fritz London and Peter Debye.

⁴ If this is the first time that you have encountered the Latin word “quasi”, then in English it means “as if”, but it has come to have the usage “sort of”/“similar to”, therefore in this context a Hydrogen Bond is a “sort of bond”/it is similar to/it resembles a proper or a real “bond”.

⁵ A “lone pair” of electrons is one where the two electrons are NOT already involved in a covalent bond. They are therefore often also referred to as “an unbonded pair of electrons”.

⁶ Both Oxygen and Chlorine are ‘next to’ fluorine in the Periodic Table, but Chlorine is not as electronegative as Oxygen because its outermost electrons are in a ‘3p’ orbital which is farther away from the nucleus than the electrons in Oxygen’s ‘2p’ orbital. A Chlorine atom’s ability to attract electrons (even though it has 9 more protons) is thus not as great as that of an Oxygen atom because of the **distance** and the **shielding** effects.

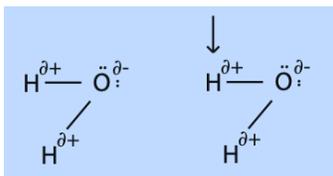
In fact, so distorted does the edd of the Hydrogen electron become that the Hydrogen atom's electron is almost pulled off the Hydrogen atom altogether; and, even though it becomes an $H^{\delta+}$ Hydrogen *atom*, it is **very nearly** transformed into an H^+ *ion*!

I have elsewhere compared the condition " δ^- " to the situation where a greedy bloke in bed pulls the duvet off his partner and leaves her as " δ^+ ". Please now imagine that you are the partner of that greedy bloke! If the bed in one of the spare bedrooms in your house were made up, then would you not be tempted to say to yourself "Blow this for a lark. I'm going off to sleep in the spare bed!" – and why therefore should the $H^{\delta+}$ atom not be tempted to do exactly the same thing?

As it happens, F/O/and N atoms are so greedy/are so electronegative that the poor little $H^{\delta+}$ atom has almost lost her electron, therefore she is desperate to get her electron back – and indeed she is so desperate for her electron that she would be almost tempted to marry someone else if the other person were prepared to give her back her 'lost' electron!⁷

As soon as an $H^{\delta+}$ atom that is bonded to an F/O/N atom sees a lone pair of electrons on a neighbouring F–H / O–H /or N–H molecule (or indeed if it encountered any area of high electron density), then it would be so attracted to that lone pair of electrons that **it will form a quasi-bond with that neighbouring molecule through the lone pair of electrons on that neighbouring molecule.**

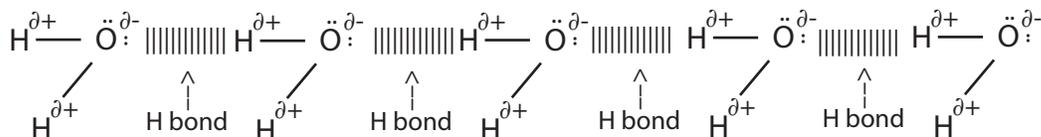
As an example let us look at two neighbouring Water molecules in a beaker of Water. The highly electronegative O atoms in each of the two water molecules below have almost totally removed the electron density from their two H atoms and have thus made them into $H^{\delta+}$ species, and it is clear from the way that I have drawn the diagram that the $H^{\delta+}$ atom that I have identified by the arrow will be highly attracted to the lone pair of electrons on the $O^{\delta-}$ atom to the left of it (*this $O^{\delta-}$ atom being rich in electron density*)



and the $H^{\delta+}$ atom will be sorely tempted to form a bond with that $O^{\delta-}$ atom through the lone pair of electrons on the $O^{\delta-}$ atom!⁸ **NB At 'A' Level, Hydrogen bonding is always executed through a lone pair of electrons!**

If a real/a proper covalent bond (*and in this case it would be a **dativ**e bond*) were formed between the $H^{\delta+}$ atom and the $O^{\delta-}$ atom, then clearly scientists would not have had to invent a new name for the sort of bond that actually forms between the $H^{\delta+}$ atom and the $O^{\delta-}$ atom. In fact, a real/a proper bond is **not** formed between the $H^{\delta+}$ atom and the $O^{\delta-}$ atom, and it is only a *quasi-bond* i.e. a "**hydrogen bond**" is formed.

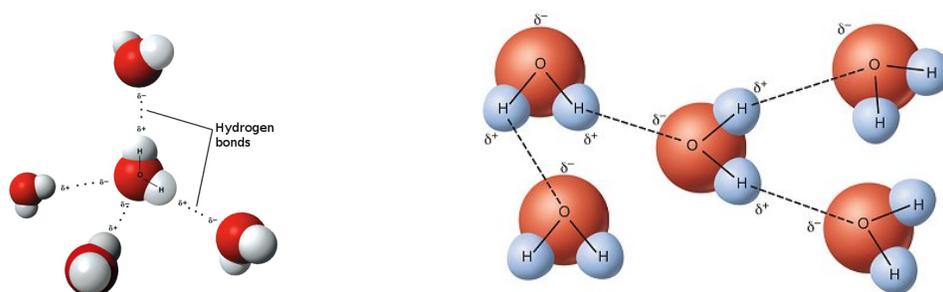
In the diagram below I have shown five water molecules with Hydrogen bonds between them viz.



⁷ The process of attributing a human emotion to a non-human object is called "anthropomorphism".

⁸ As it happens, EVERY $H^{\delta+}$ atom in water is attracted to every $O^{\delta-}$ atom in the water, but each $H^{\delta+}$ atom will be most attracted to its *closest* $O^{\delta-}$ atom.

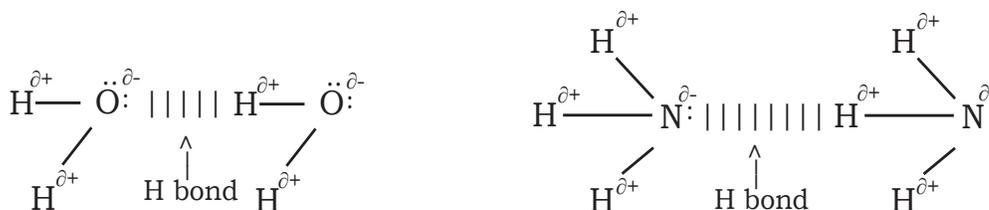
and the diagram below shows how they may be arranged tetrahedrally in space



Source: Wikipedia

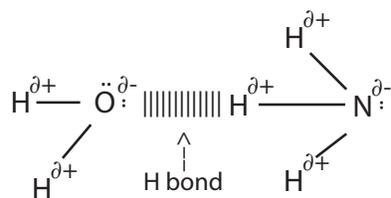
In the exam, if you are asked to draw Hydrogen bonding in Water, or Hydrogen bonding in liquid Ammonia then please draw your diagrams **EXACTLY** as I have drawn them below (with all the “ δ^+ ”s and all the “ δ^- ”s in the diagrams!

Please do remember that at ‘A’ Level **the H bond must always be linked through one of the lone pairs of electrons on the F/O/or N atom.**



An $H^{\delta+}$ atom can bond (via a lone pair of electrons) with the atoms of any of the three elements F/O/and N, therefore **Water molecules can form Hydrogen bonds not just with other Water molecules but also with say Ammonia molecules** (and this it **DOES** do). In the exam if you are asked to show this, then please draw your diagram **EXACTLY** as I have drawn it below. *[Please note that if molecules of Water and Ammonia come into contact with each other, then Hydrogen bonds will be present not just between Water molecules and other Water molecules, and between Ammonia molecules and other Ammonia molecules, but also between Water molecules and Ammonia molecules.]*

Below is an example of how a molecule of Water and a molecule of Ammonia might be hydrogen bonded.



In the exam, you will almost certainly lose a mark

- if you do not show the Hydrogen bond going through a lone pair of electrons, and
- if you do not show all your delta pluses and all your delta minuses.

The Hydrogen bonds could have been formed by **any** of the $H^{\delta+}$ atoms on the Ammonia molecule and through either of the lone pairs of electrons on the $O^{\delta-}$ atom, but in the exam you do not have a great deal of time to start thinking “Now, how shall I draw this?”. Instead you need to do bread-and-butter stuff like this very quickly so that you have enough time to think about the *non-standard* stuff – therefore please draw the diagrams exactly as I have drawn them for you on page 4.

Please note that any substance that contains an H atom bonded to an F/O/or N atom will be so denuded of its electron density that it will form Hydrogen bonds between the molecules of the substance viz.

Molecule 1 The H bond Molecule 2



The “:” above indicates a lone or an unbonded pair of electrons. In fact, each F atom will have three unbonded or lone pairs of electrons, and each O atom will have two of them, and each N atom will have one (but, to keep my diagram simple, I have not shown every single lone pair of electrons).

F/O/and N are the three most electronegative elements that there are, and it is only F/O/and N atoms that have sufficient electronegativity to drag the electron cloud almost completely off an H atom and thus cause $H^{\delta+}$ atoms to be attracted by a lone pair of electrons on a neighbouring F/O/or N atom. No other element, I repeat “No other element” is sufficiently electronegative to be able to do what F/O/and N are able to do.

If you look at the Pauling values for electronegativity (below), you will see that Chlorine has the same electronegative value as Nitrogen – but I promise you that **Chlorine cannot form Hydrogen bonds!** There is no such thing as an absolute value for electronegativity, and (the massively talented) Linus Pauling in a sense “made up” his values based on experiments – but Pauling’s electronegativity value is **not** a sufficient guide as to Chlorine’s ability to form Hydrogen bonds.

Chlorine atoms bonded to Hydrogen atoms (e.g. as in HCl) are **not** able to form Hydrogen bonds. Could you please now look at a Hill & Holman in your school library, and go to the Index and look up “Hydrogen Bonds” – and you will then read the evidence for the material that we have discussed in this blog.

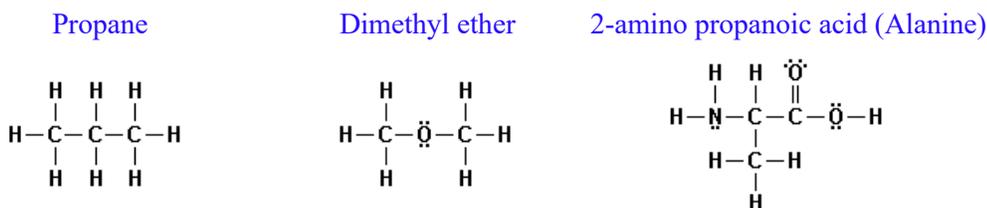
Postscript : I do hope that you are now getting the hang of using reference books for different purposes.

Pauling’s Values of Electronegativities

1	2	3	4	5	6	7	8	9	10																							
1 H 2.1									2 He																							
3 Li 1.0	4 Be 1.5																															
11 Na 0.9	12 Mg 1.2																															
19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr															
37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.8	49 In 1.8	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe															
55 Cs 0.7	56 Ba 0.9	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn	
87 Fr 0.7	88 Ra 0.9	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112																					

Source: Oklahoma State University

Now could you tell me which of these three molecules below has Hydrogen bonding in it? (*Hint: please remember that something must denude the H atom of its electron density.*)



I have told you in quite a lot of detail about Hydrogen bonding, so let me now turn my attention to van der Waals'/London/Debye forces of attraction (which I talked about in last week's blog).

B) van der Waals/London/Debye forces of attraction,

NOT bonds, just forces of attraction between molecules

(and please remember that we are talking about *molecular* substances, and NOT ionic substances.)

I should like you to appreciate very clearly that *if two or more entities remain bonded together, then some sort of force of attraction must be binding them together*. My wife died 10 months short of our 50th anniversary, therefore clearly there must have been something that held us together for such a long period of time – and, since I absolutely adored her, the force that held us together was *love!*

The earth is roughly 4.5 billion years old, and thus many atoms and molecules have stayed together for just a *tiny* bit longer than my wife and I were married – but *love* does not explain why atoms and molecules stay together. Last week we looked at metallic and ionic bonding, so [let now look at what holds molecules together](#).

I do hope that you have the difference between **INTRA-molecular covalent bonds** and **INTER-molecular forces of attraction** firmly in your head viz.

- *The strong bonds **INSIDE** or within a molecule are called “**INTRA-molecular**” covalent bonds, and they hold together the **atoms** inside a **molecule** (e.g. the bond that holds two Hydrogen atoms together in a molecule of Hydrogen is an intra-molecular covalent bond), and*
- *The weak forces of attraction **BETWEEN** molecules are called “**INTER-molecular**” forces of attraction, and in this blog we are going to talk about inter-molecular forces of attraction.⁹*

van der Waals/London/Debye forces of attraction

Having talked about Hydrogen bonds earlier in this blog, we are now going to talk about three different types of *forces of attraction*, that are often collectively called “van der Waals’ forces of attraction” because the Nobel Laureate Johannes Diderik van der Waals (which I shall shorten to “vdW”) did the initial work on these forces of attraction.¹⁰ They are

⁹ A good way of remembering what an **INTER**-molecular bond is, is to remember that an Inter-City train is the train that runs **BETWEEN** cities, and an **INTER**-molecular force of attraction is thus the force of attraction **BETWEEN** molecules!

¹⁰ Other scientists also did work in this field – but it is the concepts that I want to discuss, and not the names that are given to the different types of force involved. The reason that it is important for you to understand the nature of bonds and different types of forces of attraction is that when you understand the nature of the binding forces involved, then you will understand lots of other things such as the different boiling points of different substances/the amount of energy involved in different reactions/etc.

- B1) the **Permanent dipole to Permanent dipole**¹¹ (**PD to PD**) force of attraction
 B2) the **Permanent dipole to Induced dipole (PD to ID)** force of attraction, and
 B3) the **Momentary or Temporary or Instantaneous dipole to Induced dipole (MD to ID)** force of attraction (often also called “dispersion forces” or “London forces” after Fritz London, the Nobel Laureate who did some of the work on them).

The names of each of these forces indicates the nature of the force involved, but before we start to discuss vdW forces of attraction, I should like to remind you that we have talked about a number of different types of bonds so far viz.

- Metallic bonds** i.e. bonds between positively charged metal ions and a sea of negatively charged delocalised electrons in a **piece of metal**
- Ionic/electrovalent bonds** i.e. bonds between positively charged metal cations and negatively charged non-metal anions in an **ionic substance**
- Covalent intra-molecular bonds** i.e. bonds between the positively charged protons in the nuclei of the constituent atoms of a **molecular substance** and the negatively charged shared pairs of electrons (and we will talk about these bonds in much greater detail in later blogs), and
- Inter-molecular bonds.**

Let us now talk about the three vdWaals/London/Debye forces of attraction, but before we go any further, we need to understand clearly the terms that we shall be using, and these are

- “permanent dipole”
- “induced dipole”
- “momentary” or “temporary” or “instantaneous” dipole

and please note that we chemists are extraordinarily lazy/imprecise/unscientific in our choice of terms here, because the word “dipole” implies two poles i.e. a complete separation of charge, and **there is no complete separation of charge here** – otherwise we would be talking about ionic substances and not molecular substances!

What we are talking about here are MOLECULAR substances, and we are trying to answer the question “What is it that holds together the molecules of a molecular substance?”.

The separation of charge with which we are concerned here is either

- a **permanent** partial separation of charge that is caused in molecules by the fact that some elements are more electronegative than others (as in H₂O), or
- a **temporary/momentary/instantaneous** partial separation of charge that is caused in molecules either by induction from a neighbouring molecule or spontaneously because the edd in an atom and a molecule is random and can therefore be either uniform or non-uniform.

If you need to remind yourself of this aspect of inter-molecular forces of attraction, then please go back to last week’s blog and re-read it.

¹¹ Strictly speaking it should be “entity with a permanent separation of charge to entity with a permanent separation of charge” force of attraction, but we shorten it to “dipole to dipole” because that description is much less cumbersome.

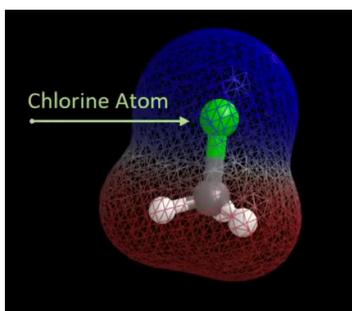
a) Permanent Dipole (PD)

When we want to emphasise the fact a substance is **ionic** e.g. NaCl, then we can write it as Na^+Cl^- or $[\text{Na}]^+[\text{Cl}]^-$, but if we want to emphasise that the substance is **molecular**, then we use a different convention viz. to convey the polarity involved in HCl we write it as “ $\text{H}^{\delta+}-\text{Cl}^{\delta-}$ ” where the arrow head points at the more electronegative element e.g.

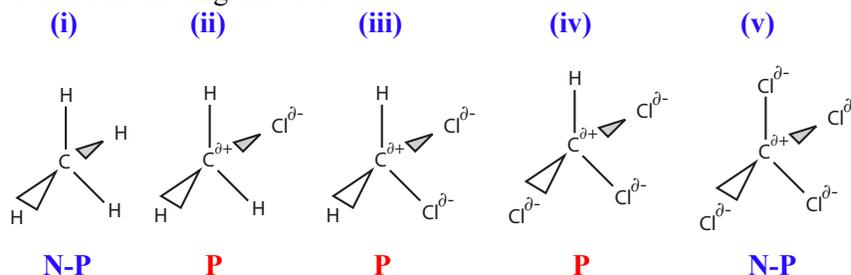
- for Oxygen and Hydrogen “ $\text{H}^{\delta+}-\text{O}^{\delta-}$ ” or “ $\text{H}^{\delta+}-\text{O}^{\delta-}$ ”
 - for Chlorine and Hydrogen “ $\text{H}^{\delta+}-\text{Cl}^{\delta-}$ ” or “ $\text{H}^{\delta+}-\text{Cl}^{\delta-}$ ”
 - for Nitrogen and Hydrogen “ $\text{H}^{\delta+}-\text{N}^{\delta-}$ ” or “ $\text{H}^{\delta+}-\text{N}^{\delta-}$ ”
 - for Sulphur and Hydrogen “ $\text{H}^{\delta+}-\text{S}^{\delta-}$ ” or “ $\text{H}^{\delta+}-\text{S}^{\delta-}$ ”
 - for Oxygen and Sulphur “ $\text{S}^{\delta+}-\text{O}^{\delta-}$ ” or “ $\text{S}^{\delta+}-\text{O}^{\delta-}$ ”
- and so on.

This arrow “ $\text{H}^{\delta+}-\text{Cl}^{\delta-}$ ” is thus very important when it comes to highlighting the difference in electronegativity between two elements e.g. the slide below comes from Dr Wayne Breslyn’s excellent body of work that he has put on the web, and it shows the $\text{Cl}^{\delta-}$ atom in Chloromethane (CH_3Cl) pulling the electron density on the molecule towards itself (as shown in blue below). This would leave the central Carbon atom denuded of electron density i.e. the Carbon atom becomes $\text{C}^{\delta+}$, and the C–Cl bond can thus be shown as “ $\text{C}^{\delta+}-\text{Cl}^{\delta-}$ ”. In Organic Chemistry, this difference in electronegativity is very often *vitaly* important and it is symbols such as these that are used to draw attention to this fact.

In the slide below, the green atom is the Chlorine atom/the white atoms are the Hydrogen atoms/and the Grey atom in the middle is the Carbon atom.



There are five molecules in the diagram below.



P = The **WHOLE** molecule is polarised.

N-P = There may be polarised bonds in the molecule, but the molecule itself is **NOT** polarised.

In turn from Left to Right the substances are called

- i) Methane
- ii) Chloromethane
- iii) Dichloromethane
- iv) Trichloromethane, and
- v) Tetrachloromethane (or Carbon Tetrachloride)

a non-polar molecule
a polarised molecule
a polarised molecule
a polarised molecule
a non-polar molecule

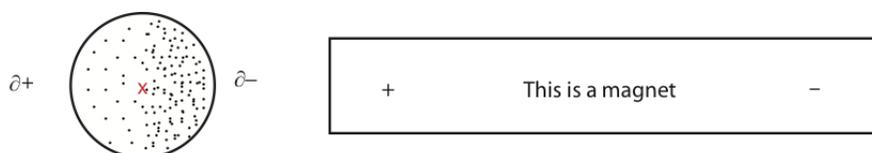
It is sometimes said that there is no separation of charge in a C—H bond (because the difference in electronegativity between Carbon and Hydrogen is quite small viz. 0.4 units) – but there is unquestionably a sufficient difference in electronegativity between Carbon and Chlorine for there to be a separation of charge in every C—Cl bond viz. $C^{\delta+}—Cl^{\delta-}$, therefore *Chloromethane/Dichloromethane/ and Trichloromethane (but **not** Tetrachloromethane)* are all partially polarised **molecules** i.e. **the molecules THEMSELVES** (and not just the C—Cl bonds) have a “ $\delta+$ ” aspect and a “ $\delta-$ ” aspect to them.

Right at the end of this blog I will do for you the Trigonometry to show you why molecules such as BCl_3 and CCl_4 are not polarised molecules. (I will leave it right to the end because **not** all of you will be doing ‘A’ Level Maths.)

b) Induced Dipole (ID)

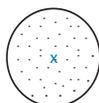
The verb “to induce” means to persuade or to cause someone to do something, and in this context, an “Induced Dipole” consists of a molecule or an atom where there is normally no separation of charge but it has polarity that has been induced into it by the polarity of a nearby molecule.

Last week I told you that a magnet would cause a separation of charge to occur in an atom or a molecule in which there is normally no separation of charge and you would get the following,

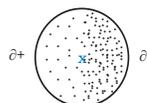


and if you were to place any molecule that has a permanent separation of charge in it (e.g. CH_3Cl) next to a molecule that has no separation of charge, then the CH_3Cl molecule will have the same effect as a magnet and it will thus **INDUCE POLARITY** into the molecule that until then had no polarity.

Molecule (A) with uniform electron density distribution (edd)



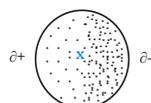
Molecule (B) such as CH_3Cl where there is a permanent partial separation of charge caused by the fact that Cl is more electronegative than C.



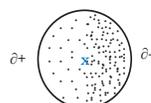
Cl would be on the Right / C in the middle / and the three H atoms are on the Left. The overall shape of the molecule would be a “tetrahedron”.

Put the CH_3Cl molecule next to molecule A, and the $\delta-$ bit on the CH_3Cl will push the edd on molecule A away from it, and thus A will no longer have uniform density distribution, but it will instead have **POLARITY INDUCED** into it.

Molecule B (CH_3Cl)



Molecule A



The edd on A has been repelled to the right thus causing A to have **INDUCED POLARITY**.

c) Momentary/Temporary/Instantaneous Dipole

Could you please note carefully that Dispersion forces of attraction (or London forces after Fritz London) occur millions of times per second (literally **MILLIONS** of times per second), and every single molecule in the world has them, and has them continuously.

NB Fritz London used the term “Dispersion” with reference to the dispersion of electrons in an electron cloud (e.g. uniform or non-uniform) – and **please remember that in the whole of this blog we are trying to understand what it is that holds together the molecules of a molecular substance.**

What I am about to say is **not** how Fritz London arrived at his theory of Dispersion Forces. He used Quantum Mechanics to arrive at his conclusion (and thus got his Nobel Prize). I am using nothing but common sense to arrive at the same conclusion – but it is only because Fritz London’s apple fell on my head that I was able to see what he saw. (*If you do not understand the allusion to Isaac Newton, then just press on and do not worry about apples falling on my head.*)

In molecules that have no separation of charge, their electron density distribution will be random and uniform because there is nothing to place the edd on one side rather than on another side.

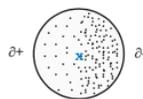
However, if the edd is random, then there are no rules that govern the distribution and therefore they **could** be on one side rather than another.

Normal edd of a given atom/molecule at any ONE point in time.



A randomly and evenly distributed edd

Edd of the same atom/molecule when for one brief moment in time there is **MOMENTARILY or **INSTANTANEOUSLY** more electron density on one side of it than there is on the other side of it.**



*An **unevenly** distributed edd*

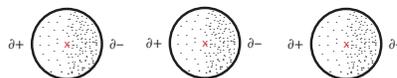
Last week I told you that for the brief space of time in which the atom or the molecule above has a partial separation of charge or a dipole, then for that brief moment in time there is no difference between it and an entity that has a permanent dipole.

For that brief space of time in which an entity possessing an Instantaneous or Momentary Dipole has such a dipole, then it is able to do all the things that a Permanent Dipole can do. For example, it can temporarily induce polarity into another entity. In the diagram below, the middle molecule briefly takes on an Instantaneous dipole, and thus **INDUCES** polarity into its immediate neighbours.

The middle molecule in that brief moment of time in which it takes on a **Momentary or an **Instantaneous** dipole**



The same molecule in the next brief moment of time in which it has induced a dipole into its immediate neighbours



However, as soon as the momentary or instantaneous polarity of the first atom or molecule is lost, then it would cease to induce polarity into its neighbours.

That is exactly what happens! At one moment of time the edd will be uniform and at the next moment of time it will not be uniform, and then at the next moment of time it will be uniform and then at the following moment in time it will not be uniform – and this switching in the edd will go on millions of time per second for as long as the molecule exists.

Please do **NOT** read any further until you have read and re-read that point and you have understood it fully.

The following table sets the context for the strengths of these forces. All the numbers below will be affected by the number of electrons in the molecule (but let us not make things too complicated). In general therefore

Type of bond/force of attraction	Energy required to break the bond/the force of attraction (but the amount of energy can vary enormously according to the exact bond involved)
Covalent and electrovalent bonds	Many hundreds and often even thousands of kJ mol ⁻¹
Hydrogen “bonds”	about 10-50 kJ mol ⁻¹
PD to PD	about 5 kJ mol ⁻¹
PD to ID	about 3 kJ mol ⁻¹
MD to ID	about 1 kJ mol ⁻¹

If you are stumbling over any of what I am saying, then please do go back and re-read last week’s blog and the early part of this week’s blog.

A) Permanent dipole to Permanent dipole force of attraction

This force of attraction binds together the *molecules* of a substance that has **permanent dipoles** in its molecules.¹² **Please note that we are not talking about metallic or ionic substances.** Metallic and ionic substances DO have permanent dipoles in them, but the force of attraction that holds a metallic substance together is omni-directional **metallic** bonding, and an ionic substance is held together by omni-directional **electrovalent** bonding and NOT by permanent dipole to permanent dipole forces of attraction! There is a distinction here (albeit a subtle one, and it involves the magnitudes of the forces of attraction) of which you need to be aware.

A dipole (as the name implies) has two poles, a positive pole and a negative pole, and when the positive pole of a *molecule* that has a permanent dipole is attracted to the negative pole of a neighbouring *molecule* that has a permanent dipole, then a **permanent dipole to permanent dipole force of attraction** is formed.

Of course, while the *positive* pole of a molecule is forming a force of attraction with the negative pole of the molecule next to it, then its *negative* pole is forming a force of attraction with the positive pole of the molecule that is on the other side of it, and so on.

¹² Please remember that there is a HUGE number of “polar” or “polarised” substances (i.e. substances that have permanent dipoles), and examples of such substances would be HCl/H₂SO₄/H₃PO₄/CH₃Cl/CH₂Cl₂/CHCl₃/etc. However, molecules that have polarised bonds will not themselves be polarised if the dipoles inside the molecule cancel each other out – therefore even though the *molecules* of CCl₄/BF₃/BCl₃/etc have polarised bonds, **the MOLECULES themselves do NOT have permanent dipoles**, therefore the force that binds them together CANNOT be a permanent dipole to permanent dipole force of attraction.

B) Permanent dipole to Induced dipole forces of attraction

Let us now ask the question “what would happen if we were to put even just one molecule of a permanently polarised substance into a container full of a non-polarised molecular substance?” – and the answer (I hope) is obvious.

The permanently polarised molecule will induce polarity into its non-polarised neighbours who will immediately induce polarity into their neighbours and so on, therefore the force of attraction that holds the two sorts of molecules together is a **Permanent dipole to INDUCED dipole** force of attraction. The molecules will have Induced Dipoles, but the species that is starting off the induction is a permanent dipole. If we were now to remove the permanently polarised molecule, then we would be left with a container of molecules that have no permanent separation of charge, therefore the force of attraction that would be holding the molecules together would revert to being the Momentary dipole to an Induced dipole force of attraction.

C) Momentary (or Instantaneous) dipole to Induced dipole force of attraction

Let us consider a humble 1 litre (i.e. 1 cubic decimetre, or 1dm³) beaker of any liquid substance that is non-polar (i.e. the substance does not have molecules that are permanently polarised). In a future blog we will talk about the mathematical aspects of ‘A’ Level Chemistry, and we will see how easy it is to calculate how many molecules there are in 1 litre of any substance – but for the moment let me tell you that there is a very large number of molecules in 1 litre of any liquid.

We have already seen that the probability of a non-polar molecule taking on a Momentary or Instantaneous dipole at any instant is very large, and if there are grillions of molecules in a 1 litre beaker, then **there must at any one time be a VERY large number of molecules in the beaker that have a Momentary or Instantaneous dipole!** If this is so, (and this is the lovely thing), then even though the substance may be non-polar, it is a fact that **EVERY SINGLE MOLECULE** (i.e. *every single molecule*) *in the substance will at any one moment in time ALWAYS have either a Momentary dipole in it or else it will have an Induced dipole in it!* **The substance is completely NON-POLAR and yet every single molecule in it ALWAYS has a dipole!** Staggering isn’t it?

If you were not surprised by what I said just now, then please do take another moment to think about what I have just said. Why do you think that the molecules were staying together? Did they love each other? Had they signed a contract to stay together? Had they appeared before a priest/rabbi/mullah/ponghgyi¹³/whatever to swear that they would stick by each other through thick and thin? Please ask yourself what it was that **YOU** thought was holding all the molecules of the **non-polar** substance together, and now think about the consequence of every single non-polar molecule of the substance actually having a dipole and **that being the force that holds the substance together!**

Actually, surprising though this thought may be, we should **not** be at all surprised by this, because our common sense tells us that for the molecules of a non-polar substance to stay together *there must be some sort of force of attraction holding all the molecules together* – and, as it happens, that force of attraction is not love but it is indeed the force of attraction of Momentary or Instantaneous dipoles to Induced dipoles!

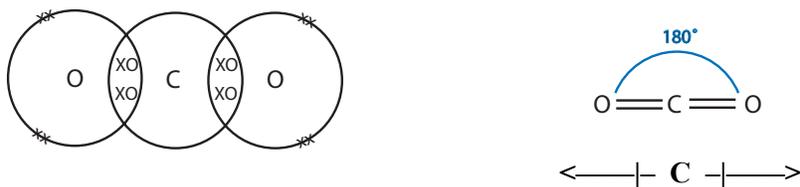
Even in a non-polar substance, every molecule always, ALWAYS has a dipole in it (either Momentary or Induced) and it is the force of attraction between the oppositely charged parts of these dipoles that holds the molecules of a non-polar (or a non-polarised) substance together.

¹³ A Buddhist Burmese priest.

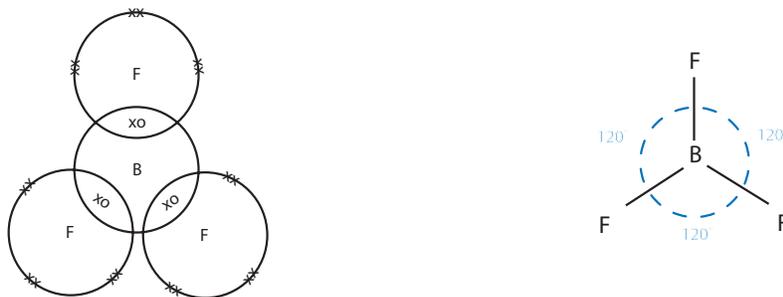
OK, now for some Trigonometry.

The cancellation of dipole moments in a symmetrical molecule such as CO₂ / BF₃ / CCl₄

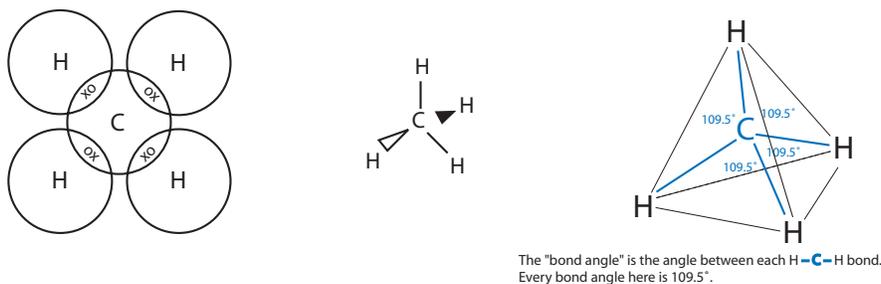
The **size** of a dipole is measured by its dipole moment (μ), but all that I want to show you here is that dipole moments can cancel each other out, and if they do so, then although there may be polarised bonds inside a molecule, the **molecule** itself will not be polarised. For example, the dipole moments in a molecule of CO₂ are exactly equal and opposite to each other therefore CO₂ molecules are **not** polarised molecules.



In molecules such as BF₃ / BCl₃ / BI₃ you have to do a bit of trigonometry to see that the dipole moments cancel each other out. The bit of Trig that you need is that the dipole moment of the bottom right hand BF dipole vertically downwards is given by $BF \cos 60 = 0.5BF$, therefore the bottom right hand BF dipole vertically downwards plus the bottom left hand BF dipole vertically downwards = $0.5BF + 0.5BF = 1.0BF$. The dipole moment of BF vertically downwards is thus exactly equal and opposite to BF vertically upwards, therefore the two dipole moments cancel each other out and the molecule itself (although it has individual dipoles in it) is **not** polarised.



The Trig in a symmetrical tetrahedron such as CCl₄ is determined by $\cos 109.5 = -0.3333$, and 3 lots of $-0.3333 = -1.0$, therefore a symmetrical tetrahedral molecule such as CCl₄ will not be polarised. The diagram that follows is that of CH₄ but the Trig will be just the same for CCl₄ as for CH₄.



In the final part of this four-part blog (next week), I will tie together everything that we have talked about in the first three parts of the blog.