

# Electronegativity analysis essay



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This page explains what electronegative is, and how and why it varies around the Periodic Table. It looks at the way that electronegative differences affect bond type and explains what is meant by polar bonds and polar molecules. If you are interested in electronegative in an organic chemistry context, you will find a link at the bottom of this page.

**What is electronegative Definition** Electronegative is a measure of the tendency of an atom to attract a bonding pair of electrons. The Pauling scale is the most commonly used.

Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7.

**What happens if two atoms of equal electronegative bond together?**

Consider a bond between two atoms, A and B. Each atom may be forming other bonds as well as the one shown - but these are irrelevant to the argument. [P] If the atoms are equally electronegative, both have the same tendency to attract the bonding pair of electrons, and so it will be found on average half way between the two atoms. To get a bond like this, A and B would usually have to be the same atom.

You will find this sort of bond in, for example, H<sub>2</sub>O or CO molecules. [pick]

Note: It's important to realize that this is an average picture. The electrons are actually in a molecular orbital, and are moving around all the time within that orbital. [pick] This sort of bond could be thought of as being a "pure" covalent bond - where the electrons are shared evenly between the two atoms.

What happens if B is slightly more electronegative than A? B will attract the electron pair rather more than A does. [pick] That means that the B end of the bond has more than its fair share of electron density and so becomes slightly negative.

At the same time, the A end (rather short of electrons) becomes slightly positive. In the diagram, "[pick]" (read as "delta") means "slightly" - so [pick]+ means "slightly positive".

Defining polar bonds This is described as a polar bond. A polar bond is a covalent bond in which there is a separation of charge between one end and the other - in other words in which one end is slightly positive and the other slightly negative. Examples include most covalent bonds. The hydrogen-chlorine bond in HCl or the hydrogen-oxygen bonds in water are typical.

What happens if B is a lot more electronegative than A? In this case, the electron pair is dragged right over to B's end of the bond.

To all intents and purposes, A has lost control of its electron, and B has complete control over both electrons. Ions have been formed. [pick] A "spectrum" of bonds The implication of all this is that there is no clear-cut division between covalent and ionic bonds. In a pure covalent bond, the electrons are held on average exactly half towards one end. How far does this dragging have to go before the bond counts as ionic? There is no real answer to that.

You normally think of sodium chloride as being happily ionic solid, but even here the sodium hasn't completely lost control of its electron. Because of the

properties of sodium chloride, however, we tend to count it as if it were purely ionic. [pick] Note: Don't worry too much about the exact cut-off point between polar covalent bonds and ionic bonds. At Eleven, examples will tend to avoid the grey areas - they will be obviously covalent or obviously ionic. You will, however, be expected to be realistic that those grey areas exist. [pick] Lithium iodide, on the other hand, would be described as being "ionic with some covalent character".

In this case, the pair of electrons hasn't moved entirely over to the iodine end of the bond. Lithium iodide, for example, dissolves in organic solvents like ethanol - not something which ionic substances normally do. Summary ; No electronegative difference between two atoms leads to a pure non-polar covalent bond. ; A small electronegative difference leads to a polar covalent bond.

; A large electronegative difference leads to an ionic bond. Polar bonds and polar molecules In a simple molecule like HCl, if the bond is polar, so also is the whole molecule. What about more complicated molecules?

In CH<sub>4</sub>, each bond is polar. [pick] Note: Ordinary lines represent bonds in the plane of the screen or paper. Dotted lines represent bonds going away from you into the screen or paper.

Wedged lines represent bonds coming out of the screen or paper towards you. [pick] The molecule as a whole, however, isn't polar - in the sense that it doesn't have an end (or a side) which is slightly negative and one which is slightly positive. The whole of the outside of the molecule is somewhat negative, but there is no overall separation of charge from top to bottom, or

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from left to right. By contrast, CHECK is polar. E hydrogen at the top of the molecule is less electronegative than carbon and so is slightly positive. This means that the molecule now has a slightly positive “ top” and a slightly negative “ bottom”, and so is overall a polar molecule.

A polar molecule will need to be “ lop-sided” in some way. Patterns of electronegative in the Periodic Table ere most electronegative element is fluorine. If you remember that fact, everything becomes easy, because electronegative must always increase towards fluorine in the Periodic Table. [pick] Note: This simplification ignores the noble gases. Historically this is because they

Nerve believed not to form bonds - and if they don't form bonds, they can't have an electronegative value. Even now that we know that some of them do form bonds, data sources still don't quote electronegative values for them.

[pick] Trends in electronegative across a period As you go across a period the electronegative increases. The chart shows electromagnetisms from sodium to chlorine - you have to ignore argon. It doesn't have an electronegative, because it doesn't form bonds. [pick] Trends in electronegative down a group As you go down a group, electronegative decreases. If it increases up to fluorine, it just decrease as you go down.

) The chart shows the patterns of electronegative in Groups 1 and 7. [pick] Explaining the patterns in electronegative The attraction that a bonding pair of electrons feels for a particular nucleus depends ; the number of protons in the nucleus; ; the distance from the nucleus; ; the amount of screening by inner electrons. Note: If you aren't happy about the concept of screening or <https://assignbuster.com/electronegativity-analysis-essay/>

shielding, it would pay you to read the page on nomination energies before you go on.

The factors influencing nomination energies are just the same as those influencing electromagnetisms. Use the BACK button on your browser to return to this page. Why does electronegativity increase across a period? Consider sodium at the beginning of period 3 and chlorine at the end (ignoring the noble gas, argon).

Think of sodium chloride as if it were covalently bonded. [pick] Both sodium and chlorine have their bonding electrons in the 3-level. The electron pair is screened from both nuclei by the 1s, 2s and 2p electrons, but the chlorine nucleus has 6 more protons in it.

It is no wonder the electron pair gets dragged so far towards the chlorine that ions are formed. Electronegativity increases across a period because the number of charges on the nucleus increases. That attracts the bonding pair of electrons more strongly.

Why does electronegativity fall as you go down a group? Think of hydrogen fluoride and hydrogen chloride. The bonding pair is shielded from the fluorine's nucleus only by the 2s electrons. In the chlorine case it is shielded by all the 1s2s2p6 electrons.

In each case there is a net pull from the centre of the fluorine or chlorine of +7. But fluorine has the bonding pair in the 2-level rather than the 3-level as it is in chlorine.

If it is closer to the nucleus, the attraction is greater. As you go down a group, electronegativity decreases because the bonding pair of electrons is increasingly distant from the attraction of the nucleus. [pick] Warning! As far as I am aware, none of the UK-based A level (or equivalent) syllabuses any longer want the next bit. It used to be on the AKA syllabus, but has been removed from their new syllabus.

At the time of writing, it does, however, still appear on at least one overseas A level syllabus (Malta, but there may be others that page. It is an alternative (and, to my mind, more awkward) way of looking at the formation of a polar bond.

Reading it unnecessarily just risks confusing you. [pick] Here plainsong ability of positive ions. What do we mean by "plainsong ability"? In the discussion so far, we've looked at the formation of polar bonds from the point of view of the distortions which occur in a covalent bond if one atom is more electronegative than the other.

But you can also look at the formation of polar covalent bonds by imagining that you start from ions. Solid aluminum chloride is covalent. Imagine instead that it was ionic. It would contain  $\text{Al}^{3+}$  and  $\text{Cl}^-$  ions.

The aluminum ion is very small and is packed with three positive charges - the "charge density" is therefore very high. That will have a considerable effect on any nearby electrons. [pick] Note that the aluminum ions polarise the chloride ions. In the case of aluminum chloride, the electron pairs are dragged back towards the aluminum to such an extent that the bonds become covalent.

But because the chlorine is more electronegative than aluminum, the electron pairs won't be pulled half way between the two atoms, and so the bond formed will be polar. Factors affecting polarizability Positive ions can have the effect of polarizing (electrically distorting) nearby negative ions. The polarizability depends on the charge density in the positive ion. Polarizability increases as the positive ion gets smaller and the number of charges gets larger. As a negative ion gets bigger, it becomes easier to polarize.

For example, in an iodide ion,  $I^-$ , the outer electrons are in the 5-level - relatively distant from the nucleus. A positive ion would be more effective in attracting a pair of electrons from an iodide ion than the corresponding electrons in, say, a fluoride ion " here they are much closer to the nucleus. Aluminum iodide is covalent because the electron pair is easily dragged away from the iodide ion. On the other hand, aluminum fluoride is ionic because the aluminum ion can't polarize the small fluoride ion sufficiently to form a covalent bond.

ELECTRONEGATIVE This page deals with electronegativity in an organic chemistry context. If you want a broader view of electronegativity, there is a link at the bottom of the page. What is electronegativity? Electronegativity (element) is given a value of 4.0, and values range down to cesium and francium which are the least electronegative at 0.7.

What happens if two atoms of equal electronegativity bond together? The most obvious example of this is the bond between two carbon atoms. Both atoms will attract the bonding pair to exactly the same extent.



That means that on average the electron pair will be found half way between the two nuclei, and you could draw a picture of the bond like this: [pick] It is important to realism that this is an average picture. The electrons are actually in a sigma orbital, and are moving constantly within that orbital. [pick] Help! A sigma orbital is a molecular orbital formed by end-to-end overlap between the bonding in methane and ethane. [pick] ere carbon-fluorine bond Fluorine is much more electronegative than carbon.

The actual values on the Palling scale are Carbon I fluorine 12. 5 14. That means that fluorine attracts the bonding pair much more strongly than carbon does. The bond - on average - will look like this: [pick] NH is fluorine more electronegative than carbon? A simple dots-and-crosses diagram of a C-F bond is perfectly adequate to explain it. (pick] ere bonding pair is in the second energy level of both carbon and fluorine, so in the absence of any other effect, the distance of the pair from both nuclei would be the same.

The electron pair is shielded from the full force of both nuclei by the 1s electrons - again there is nothing to pull it closer to one atom than the other.

BUT, the fluorine nucleus has 9 protons whereas the carbon nucleus has only 6. Allowing for the shielding effect of the 1s electrons, the bonding pair feels a net pull of about 4+ from the carbon, but about 7+ from the fluorine. It is this extra nuclear charge which pulls the bonding pair (on average) closer to the fluorine than the carbon.

[pick] Help! You have to imagine what the bonding pair " sees" if it looks in towards the nucleus. In the carbon case, it sees 6 positive protons, and 2

negative electrons. That means that there will be a net pull from the carbon of about 4+. The shielding

Nullity actually be quite as high as 2-, because the 1s electrons spend some of their time on the far side of the carbon nucleus - and so aren't always between the bonding pair and the nucleus.

Incidentally, thinking about electrons looking towards the nucleus may be helpful in picturing what is going on, but avoid using terms like this in exams.

[pick] ere carbon-chlorine bond ere electromagnetisms are: | carbon | chlorine 13. 0 ere bonding pair of electrons will be dragged towards the chlorine but not as much as in the fluorine case. Chlorine isn't as electronegative as fluorine.

Why isn't chlorine s electronegative as fluorine? Chlorine is a bigger atom than fluorine. Fluorine:  $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^1$  chlorine:

$1s^2 3s^2 3p_x^2 3p_y^2 3p_z^1$  Help! If you aren't happy about this, read the article on orbital.

Use the BACK button In the chlorine case, the bonding pair will be shielded by all the 1-level and 2-level electrons. The 17 protons on the nucleus will be shielded by a total of 10 electrons, giving a net pull from the chlorine of about 7+. That is the same as the pull from the fluorine, but with chlorine the bonding pair starts off further away from the nucleus because it is in the 3-level.

Since it is further away, it feels the pull from the nucleus less strongly.

Bond polarity and inductive effects Polar bonds Rhine about the carbon-fluorine bond again. Because the bonding pair is pulled towards the fluorine end of the bond, that end is left rather more negative than it Mould otherwise be. The carbon end is left rather short of electrons and so becomes slightly positive. [pick] ere symbols [pick]+ and [pick]- mean “slightly positive” and “slightly negative”. You read [pick]+ as “delta plus” or “delta positive”. We describe a bond having one end slightly positive and the other end slightly negative as being polar.