Electronegativity and Polarity

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Electronegativity

Definition: Electronegativity is the relative tendency of an atom in a molecule to attract electrons in a covalent bond to itself.

- Electronegativity is measured on the **Pauling scale** (ranges from 0 to 4).
- The **most** electronegative element is **fluorine** and it is given a value of 4.0.
- The ***higher*** the electronegativity value of an atom, the ***greater*** is the ability of an atom of that element to attract electrons to itself.

KEY

Electronegativities of elements

INCREASING ELECTRONEGATIVITY →

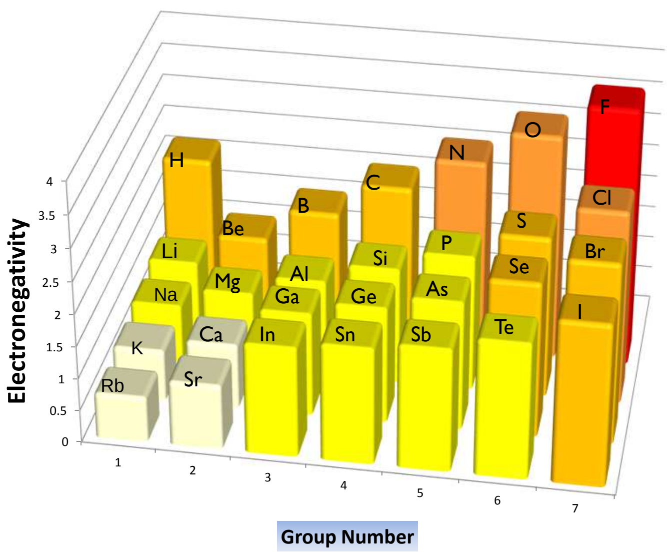
↓ *DECREASING ELECTRONEGATIVITY*

						H 2.1										
Li 1.0	Be 1.5											B 2	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.7	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1														

Numbers refer to the **Pauling Scale** of Electronegativity

KEY

Electronegativity



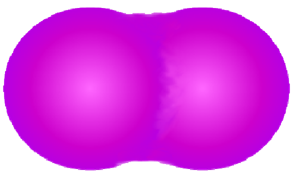
Ionic or covalent bonds?

- Not all bonds are *purely* ionic or *purely* covalent.
- Some covalent bonds have a partial ionic character: '**polar bonds**'.
- To deduce the degree of **bond polarisation**, we consider the *difference* in electronegativity between the two atoms in the bond.

KEY

Pure covalent bond (non-polar)

- Difference in electronegativity is **zero or very small** (< 0.3).
- Each atom pulls on the electrons to the **same extent** and they are **equally shared**.
- E.g. Cl_2
  - Electronegativity (Cl) = 3.0
  - **Electronegativity difference** (Cl – Cl) = 3.0 – 3.0 = **0**



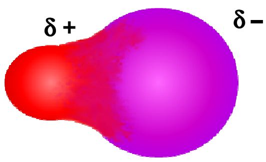
Covalent bond
Electrons shared equally

N.B. Bonds between **identical non-metal atoms** will always be **non-polar** as the electronegativity difference will be zero.

KEY

Covalent with ionic character (Polar bond)

- Difference in electronegativity between atoms is **around** 0.3 to 1.7.
- **One atom** will pull electron pair **closer** to its end.
- E.g. HCl
  - Electronegativity (H) = 2.1 Electronegativity (Cl) = 3.0
  - **Electronegativity difference** (H – Cl) = 3.0 – 2.1 = **0.9**



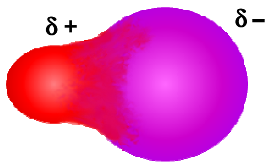
Polar Covalent bond
Electrons shared unequally

Polar covalent bond or **polar bond** is a covalent bond with greater electron density around one of the two atoms.

KEY

Representation of polar bonds

$$\overset{\delta+}{\mathrm{H}} - \overset{\delta-}{\mathrm{Cl}}$$



Polar Covalent bond
Electrons shared unequally

δ indicates a **slight deviation** from being neutral.

The separation of charge **induces** a **dipole**: the bond is polar.

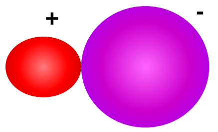
The ***greater*** the electronegativity difference, the ***greater*** the bond polarity.

δ – resides on the **more** electronegative atom; **δ +** on the **less** electronegative atom.

KEY

Pure ionic

- Difference in electronegativity between atoms is **large** (> 1.7).
- An electron is completely transferred from one atom to another.
- E.g. NaCl
  - Electronegativity (Na) = 0.9 Electronegativity (Cl) = 3.0
  - **Electronegativity difference** (Na–Cl) = 3.0 – 0.9 = **2.1**



Ionic bond
Electrons are transferred

Ionic bonds are generally formed between **reactive metals** and **non-metals** where the difference in electronegativity is **large**.

KEY

Non-polar, polar or ionic?

In general if electronegativity difference between two bonded atoms is:

< 0.3 , usually between identical nonmetal atoms, called nonpolar covalent

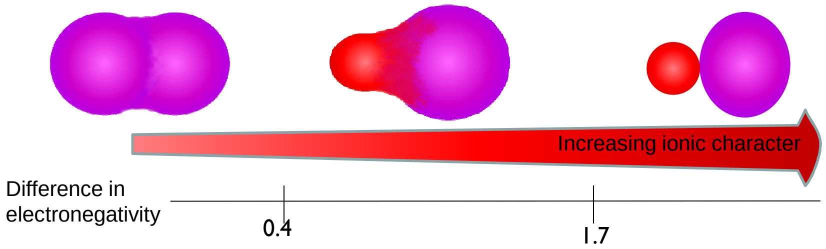
0.3 to 1.7, 2 different nonmetals called polar covalent

1.7 or greater, usually nonmetals and reactive metals, is ionic

Non polar Covalent bond
Electrons shared equally

Polar Covalent bond
Electrons shared unequally

Ionic bond
Electrons are transferred



Polar bonds and polar molecules

- For a **molecule** to be polar it must have a **net dipole**.
- To work out if a molecule is polar we consider:
  - the **polarity** of each of the **individual bonds**
  - the **shape** of the molecule.

EXTRA

Non-polar molecules

- A symmetric molecule (**all bonds identical *and* no lone pairs**) will **not** be polar ***even if individual bonds within the molecular ARE polar.***

e.g. CO_2

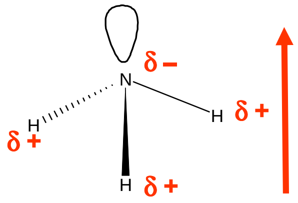
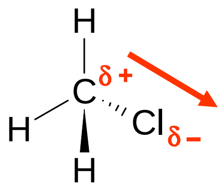
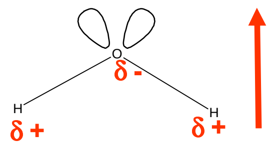
$$\overset{\delta -}{O} = \overset{\delta +}{C} = \overset{\delta -}{O}$$

The individual dipoles on the bonds '**cancel out**' due to the symmetrical shape of the molecule.
There is no **NET** dipole moment: the **molecule** is **NON POLAR**

EXTRA

Polar molecules

- A non-symmetrical molecule (**different bonds *OR* the same bonds *and* having lone pairs**) will be **polar** if **individual bond(s)** are themselves **polar**.

		
Polar N-H bonds Molecule non-symmetrical (**lone pair**)	**Polar** C-Cl bond Molecule non-symmetrical (**different bonds**)	**Polar** O-H bonds Molecule non-symmetrical (**lone pairs**)

In each case, the dipoles **do not** cancel out.
There is a **net dipole moment** and the **molecule** is **polar**.

EXTRA