

SCH4U – Electronegativity & Polar Molecules

Describe the difference between a polar covalent bond and a non-polar covalent bond:

Polar Covalent Bond: When ΔEN is between 0.5 and 1.7 the bond is considered polar covalent. In this scenario the 2 atoms are sharing the electrons unequally and the atom with the higher EN will pull the electrons closer to itself, and therefore have a higher electron density. Example: O-H (an oxygen atom bound to a hydrogen atom)

Non-Polar Covalent Bond: When ΔEN is less than 0.5 the bond is considered a non-polar covalent bond. For example an H-H bond (2 hydrogen atoms bound to each other)

When do polar covalent bonds form?

Define electronegativity:

Electronegativity is the ability of an atom in a molecule to attract shared electrons. Some atoms have very high electronegativity values, such as Fluorine, and they tend to hog all of the shared electrons in a molecule. Others, such as potassium, have very low electronegativity values, and they tend to give up their shared electrons rather easily

Atoms with high electronegativity values tend to pull shared electrons towards them.

Atoms with very low electronegativity values give up their shared electrons easily.

Using arrows indicating increasing electronegativity values, show the trend for electronegativity on the periodic table:

Group (vertical)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period (horizontal)																		
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		

Periodic table of electronegativity using the Pauling scale

To determine the type of bond that forms between 2 atoms (ionic, polar covalent, non-polar covalent) you can calculate the Difference in electronegativity

Large ΔEN indicates what type of bonds: ionic or Polar covalent

Small ΔEN indicates what type of bond: non-polar covalent

Complete the following table:

ΔEN	Bond Type
< 0.5	non-polar covalent
0.5 - 1.7	Polar covalent
> 1.7	ionic

Determine the ΔEN and classify the type of bond that forms between potassium (K) and bromine (Br):

$$K = 0.82$$

$$Br = 2.96$$

$$2.96 - 0.82 = 2.14 \text{ - Ionic Bond}$$

Determine the ΔEN and classify the type of bond that forms between nitrogen (N) and hydrogen (H):

$$N = 3.04$$

$$H = 2.20$$

$$3.04 - 2.20 = 0.84 \text{ = Polar Covalent}$$

In bonds that have a polar character, the more electronegative atom becomes slightly Negative and the atom with the lower electronegativity becomes slightly positive.

Define dipole:

a separation of positive and negative charges in space

Write the symbols for the negative end of a dipole and the positive end of a dipole:

We can represent the dipole using some common symbols: δ^+ is used to represent the positive end of a dipole and δ^- is used to represent the negative end of a dipole, we also use an arrow with a line at the back of it to show the difference in charge across a bond, the head of the arrow points towards the negative end of the dipole

When using an arrow to describe a dipole, the arrow points towards the negative end.

Molecular Polarity

Bonds can be polar or non-polar, entire molecules can also either be polar or non-polar.

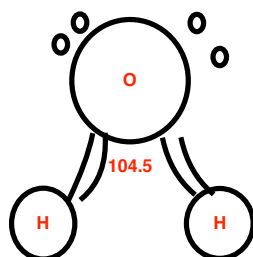
Describe why water from a tap would be attracted to a balloon that was rubbed on someone's head.

When you rub the balloon you touch the balloon to your hairs. when this happens a few electrons are transferred from the hair to the balloon. Rubbing repeats this millions of times, leaving the balloon charged.

The water starts off neutral, with the same amount of positive and negative charges. When you bring the charged balloon near to the water it pushes electrons away, leaving a positively charged area which is attracted to the balloon

Draw the VSEPR predicted shape for H_2O and indicate the bond dipoles and net polarity of the molecule:

See below **



Non-polar molecules

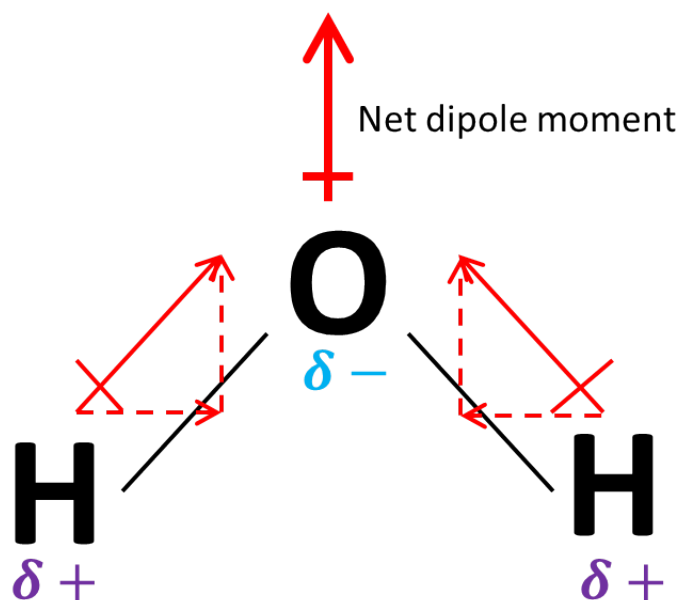
Non-polar molecules have no Separation of charge

What affects the polarity of a molecule:

1. Shape
2. Electronegativity

Describe how some molecules with bond dipoles can be non-polar:

A molecule can possess polar bonds and still be nonpolar. If the polar bonds are evenly (or symmetrically) distributed, the bond dipoles cancel and do not create a molecular dipole.



*Dipole moment is a vector quantity. Hence it can be broken down into a vertical and horizontal component as shown above (see dotted lines).

The horizontal components are pointing in opposite directions, hence will cancel out each other. But the vertical components are pointing in the same direction.

This gives rise to a net dipole moment, leaving oxygen with a permanent partial negative charge.

Complete the following table:

Non-polar molecular structures with bond dipoles

Molecule Description	General Example	Specific Example
Bent with 2 identical bonds	B-A-B	H ₂ O
Trigonal Planar with 3 identical bonds		SO ₃
Tetrahedral with 4 identical bonds		CH ₄

Determining the net polarity of a molecule:

1. Draw the Lewis structure.
2. Determine the predicted shape of the molecule using VSEPR.
3. Identify the electronegativity of each atom in the molecule and determine the partial charges in the molecule.
4. Draw bond dipoles and determine whether the molecule has a net dipole.

Predict the polarity of NH₃

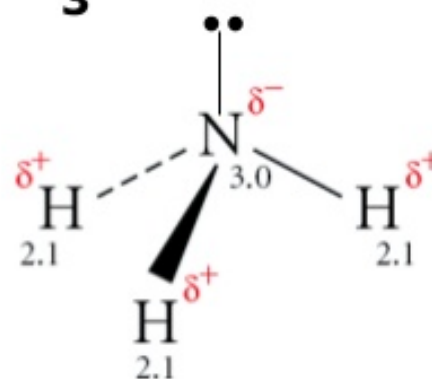
Predict the polarity of CF₄

Predict the polarity of SF₂

MOLECULAR SHAPE AND POLARITY

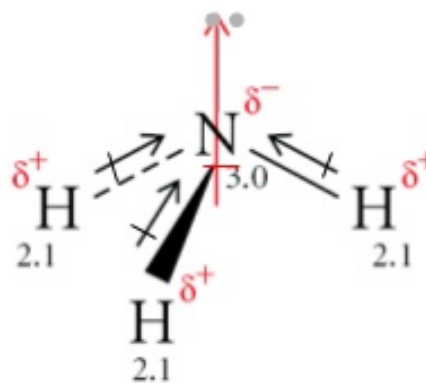
Example: **NH₃**

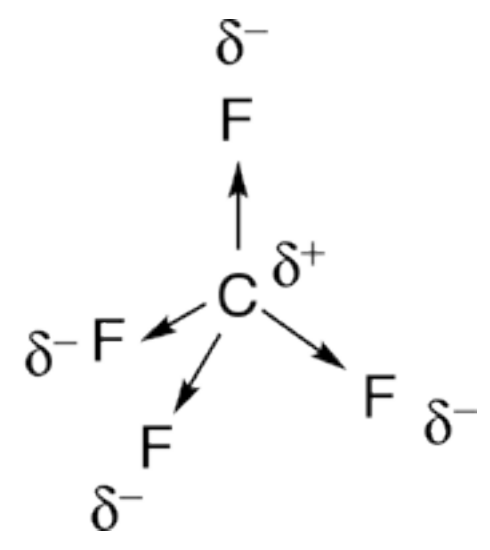
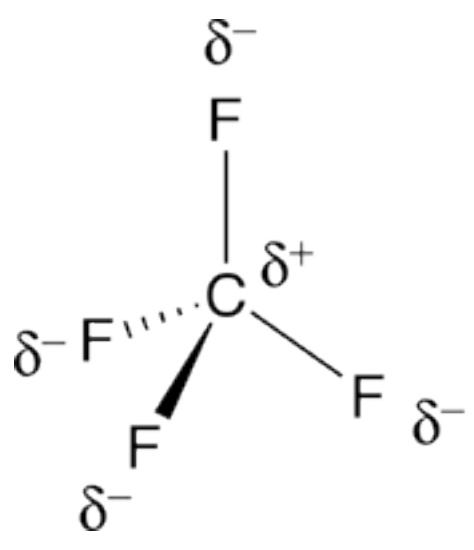
3) Add the electronegativity of the atoms and assign δ^+ and δ^- to the bonds



4) Draw in the bond dipoles

\therefore NH₃ is polar because it has polar bonds that do not cancel to zero.

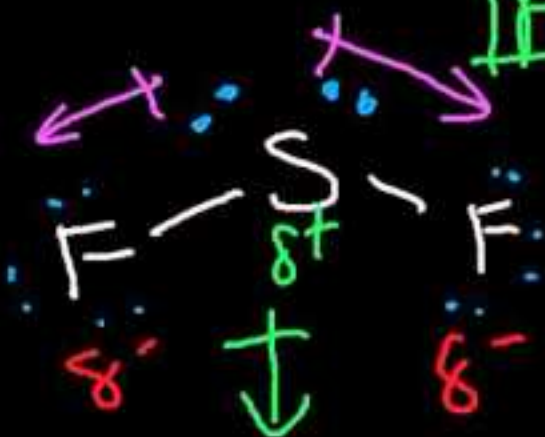




SF₂

$$6 + 2(7) = 20$$

$$\boxed{\text{polar}} - \frac{16}{4}$$



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