	Patterns	Nuclear Charge	Distance	Shielding	Overall effect
1	Increase across period	Increase	Constant	Constant	Increase across period
2	Last element in period To first element in next period Decrease	Increase	Increase	Increase	Shielding and distance outweigh increased nuclear charge
3	Be → B Decrease	Increase	Increase	Increase	Shielding and distance outweigh increased nuclear charge
4	Down a group Decrease	Increase	Increase	Increase	Shielding and distance outweigh increased nuclear charge
5	N → O Decrease				Spin pair repulsion

KCl

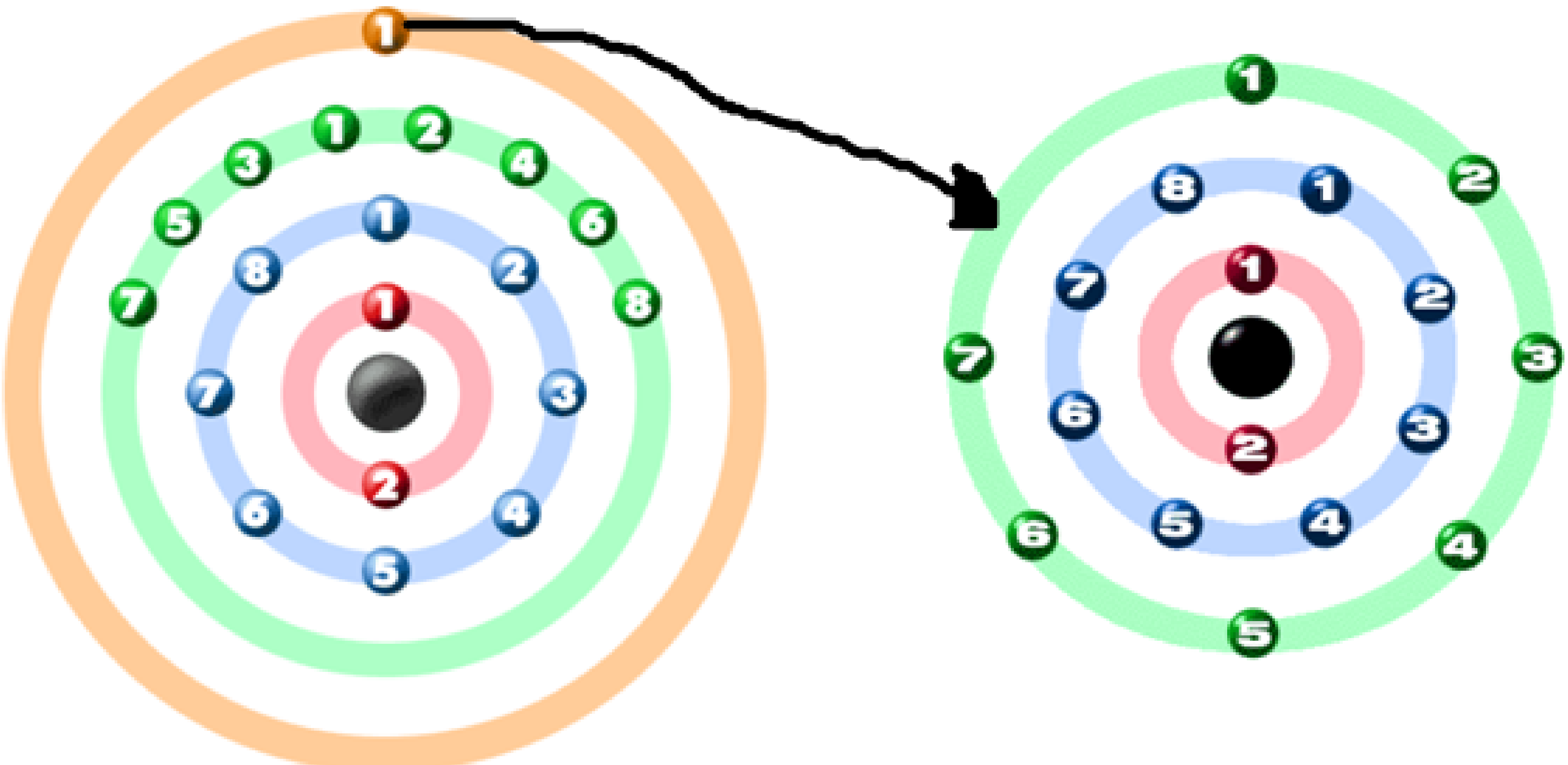
Ionic bonds (aka electrovalent)

K = 19

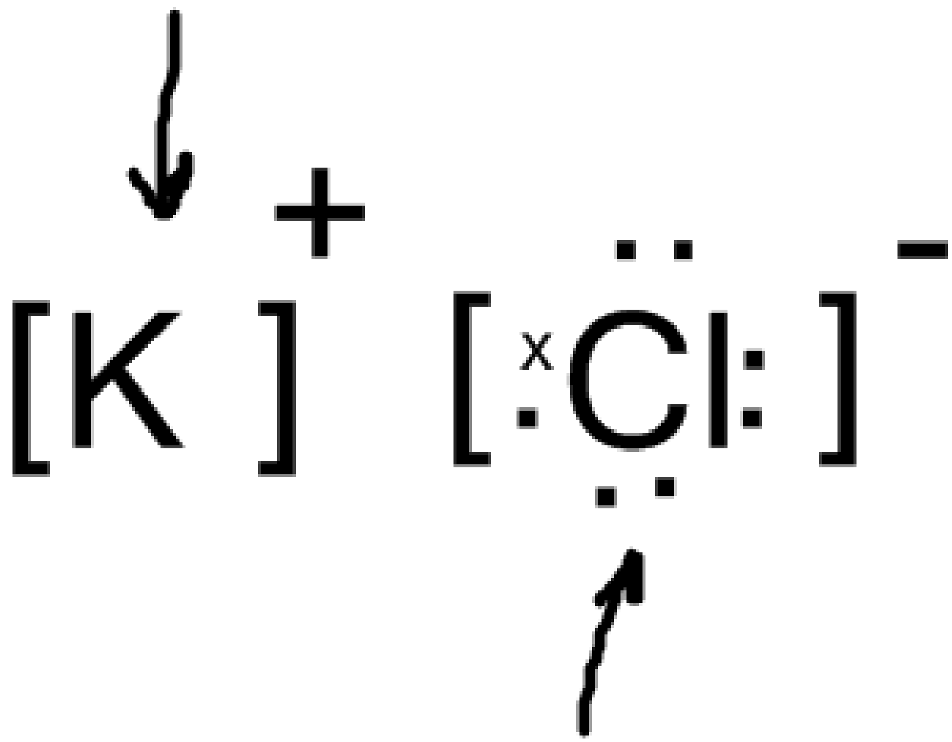
$1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 4s^1$

Cl = 17

$1s^2\ 2s^2 2p^6\ 3s^2 3p^5$



The lost electron stops shielding the other electrons from the nucleus so there is a greater pull from the nucleus and the atom becomes tightly bond decreasing the size.



The gained electron increases the shielding of the other electrons from the nucleus and the atom increases in size.

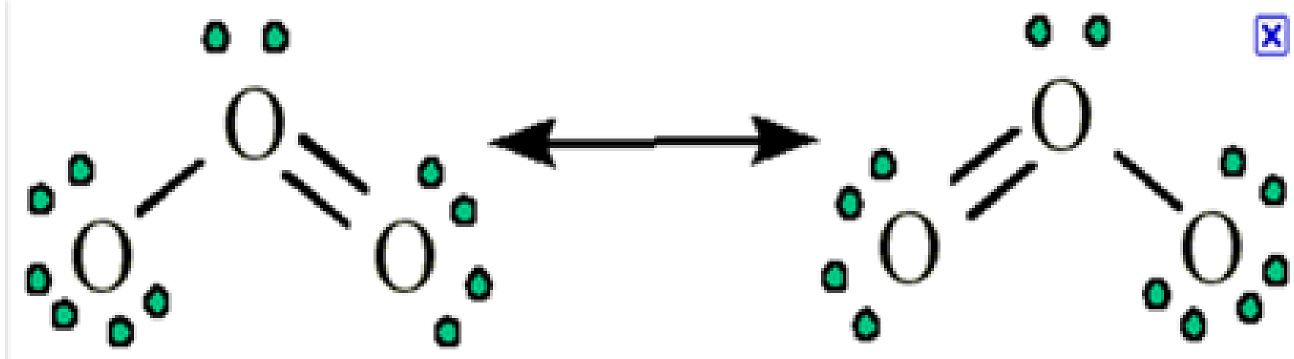
dot and cross diagrams show:

1. the outer electrons only
2. that the charge of the ion is spread evenly, by using square brackets
3. the charge on each ion, written in the top righthand corner of the square brackets.

1) draw out dot structures
2) calculate # of possible electrons
3) arrange elements so you have a center. If Carbon is present it goes in the center, otherwise the element with the lowest electronegativity does.
4) put the first 8 electrons around the center element, then add the rest to each nonmetal (not hydrogen), so that every element is surrounded by 8 electrons.

If a molecule has more than one structure, it is called a resonance structure.

Ozone, O_3



Draw out:

1. Na_2O
2. CaO
3. $MgCl_2$

The Octet Rule:

Main-group elements tend to react so they have eight valence electrons, by sharing (covalent) their outermost s and p orbitals.

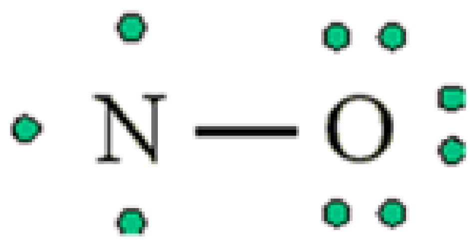
Hydrogen has a single electron and therefore has the following configuration. H $1s^1$

The configuration for chlorine is: Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Most main-group elements tend to form covalent bonds according to the octet rule but there are exceptions:

3 ways to tell:

1. odd # of electons - NO

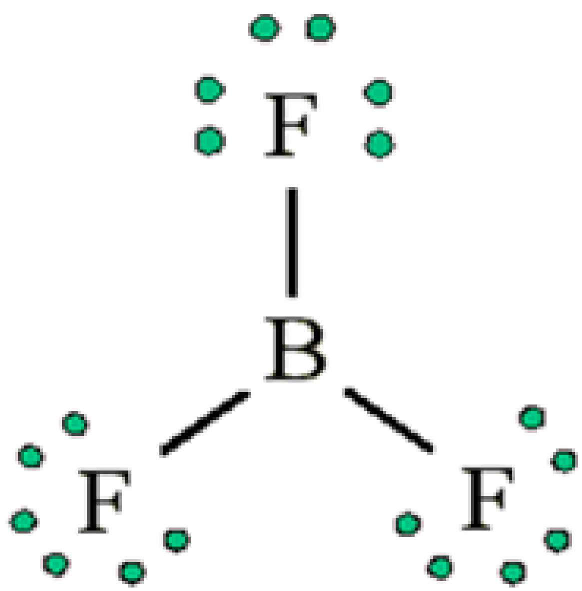


Nitrogen has 3 and Oxygen has 6, total = 9. It will never have 8 in the outer shell.

Electron deficient

2. Less than possible: B, Be -

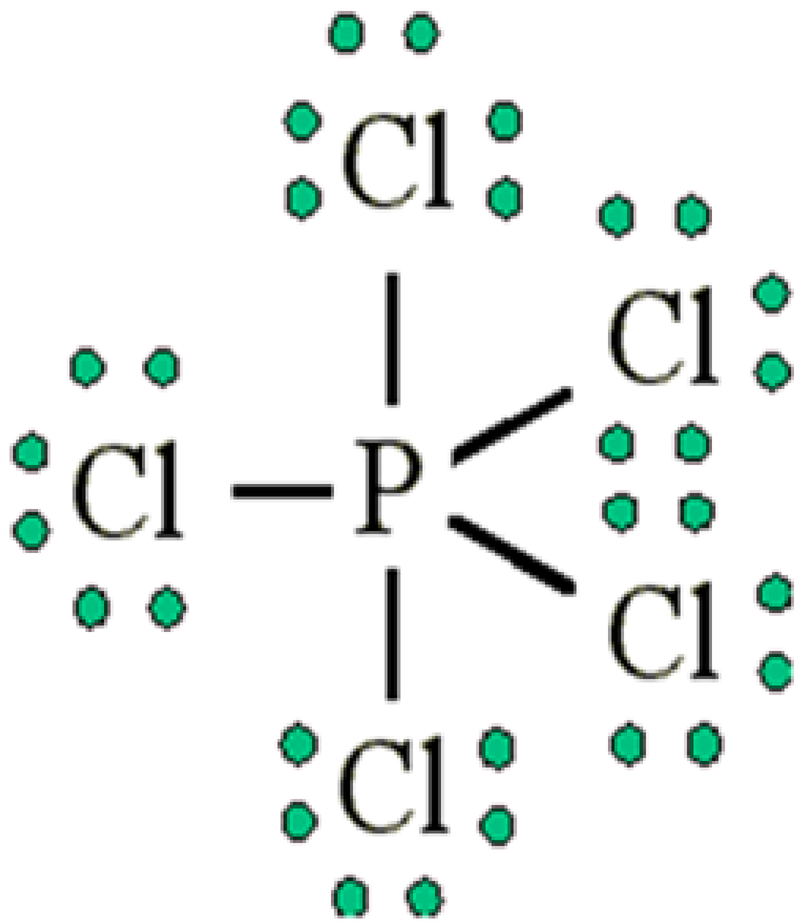
BF_3



Boron has 3, F has 7 this gives 24 electrons, not enough to have 8 surrounding each.

3. More than possible: PCl_5, PF_5, SF_6, XeF_4, ClO_4^-

Expanded Octet



-If negative charge add
indicated # of electrons.
-If positive subtract indicated #
of electrons.

Covalent bonds

Shared pair of electrons

sharing 2 electrons = single bond, increased bond length, decreased bond energy (weaker bond)

sharing 4 electrons = double bond

sharing 6 electrons = triple bond, shortest bond length, requires a lot of energy to break the bonds.

bond breaking = endothermic +
bond making = exothermic -

covalent bonds

type	strength	length	energy
single	weakest	longest	lowest
triple	strongest	shortest	highest

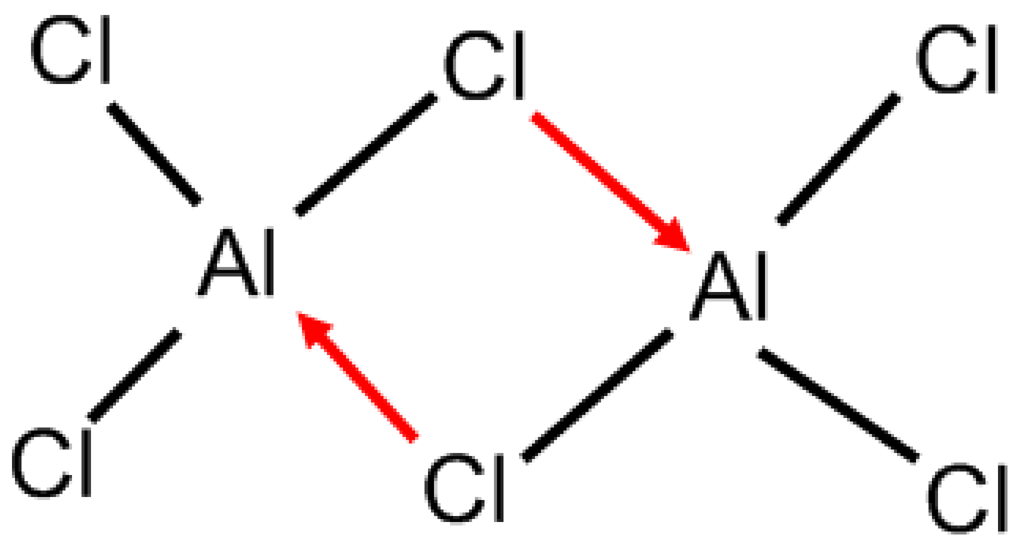
bond breaking = endothermic +
bond making = exothermic -

Bond energy (enthalpy) - energy required to break one mole of a given bond in a gaseous molecule

Co-ordinate (dative covalent bond)

when one atom provides both the electrons needed for a covalent bond and a second atom that has an unfilled orbital to accept the electrons.

head of arrow points toward unfilled orbital



Ionic	Covalent
Metal & Nonmetal	Nonmetal & Nonmetal
Solid at room temp.	Soft and squishy
High melting/boiling point	Low melting/boiling point
Conducts electricity when melted or when dissolved in water	Does not conduct electricity in water.
	Usually does not dissolve in water.
Hard - able to resist force	If it contains carbon or hydrogen it will burn.
Brittle - force will cause fracture but won't form a powder.	
arranged in repeating patterns	

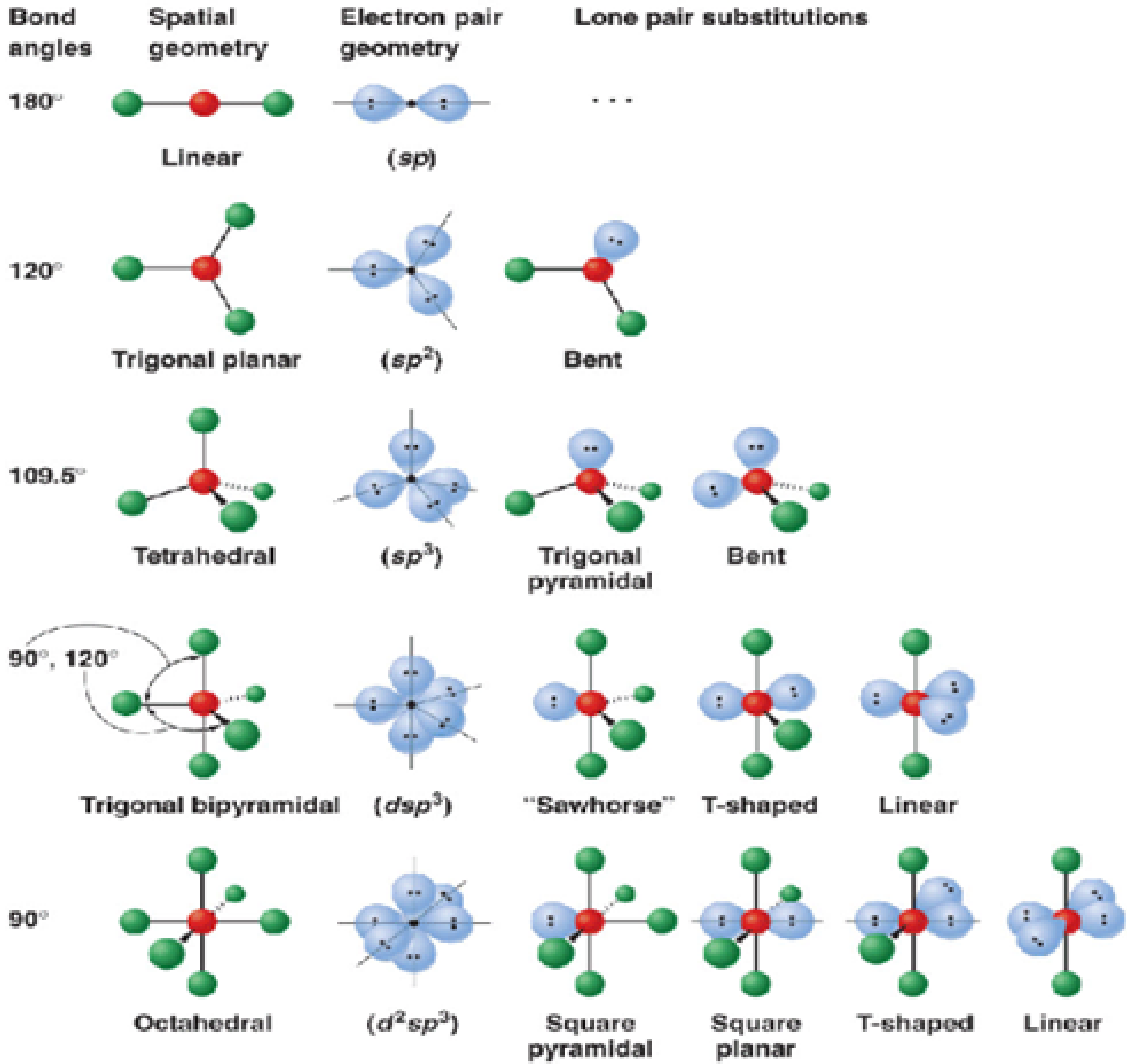
Pairs of electrons not used in bonding are called Lone pairs

The lone pairs of electrons have the greatest repulsion on each other.

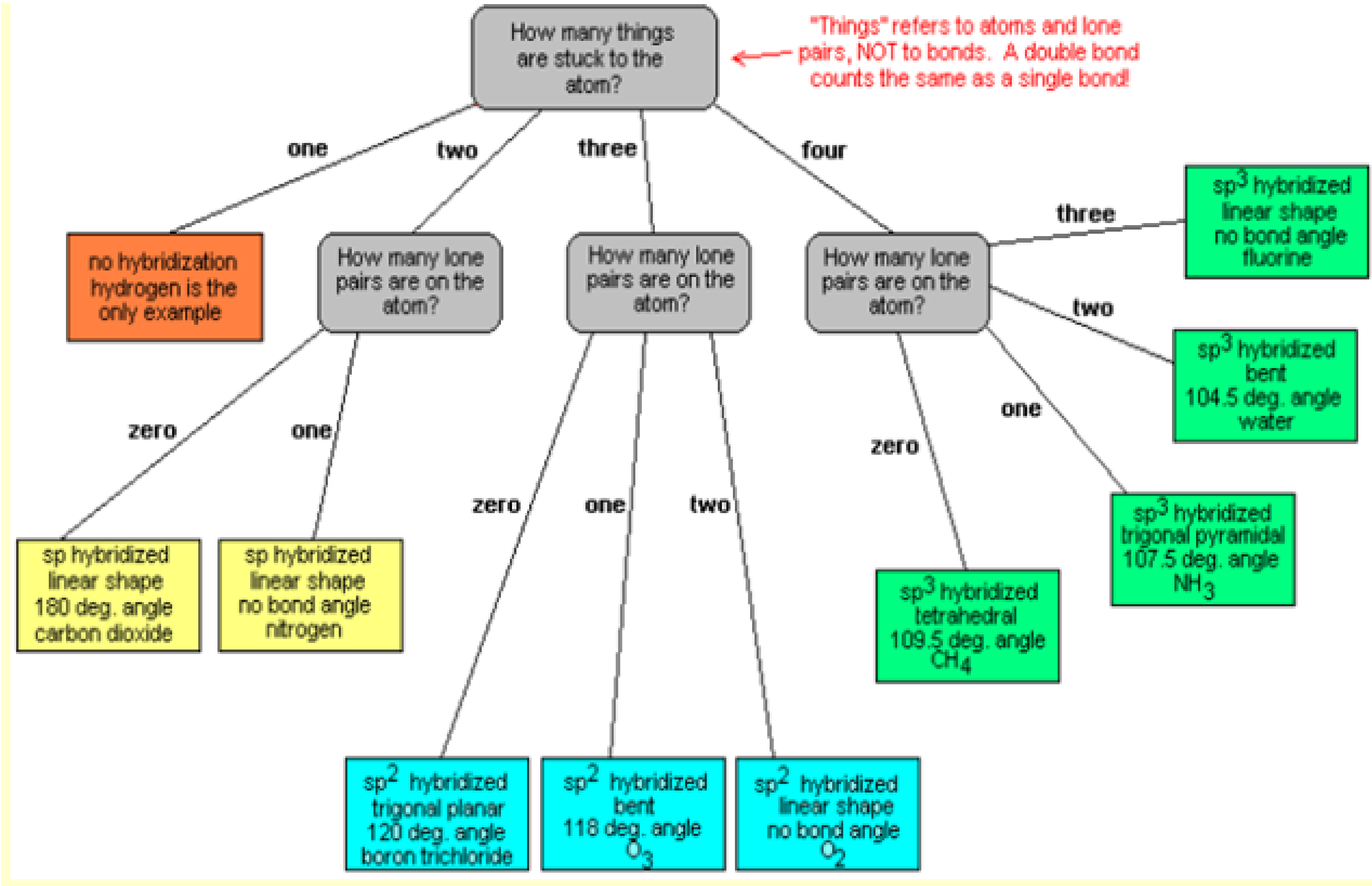
This is what gives water its 104.5 angle
the lone pairs of electrons on the oxygen repel each other

lone pair:lone pair > lone pair:bond pair > bond pair:bond pair

The differences in the electron pair repulsions determine the shape and bond angles a molecule.


Bond angles
Spatial geometry
Electron pair geometry
Lone pair substitutions
180°
Linear
(sp)
120°
Trigonal planar
(sp²)
Bent
109.5°
Tetrahedral
(sp³)
Trigonal pyramidal
Bent
90°, 120°
Trigonal bipyramidal
(dsp³)
"Sawhorse"
T-shaped
Linear
90°
Octahedral
(d²sp³)
Square pyramidal
Square planar
T-shaped
Linear

Shapes of the molecules



try these:

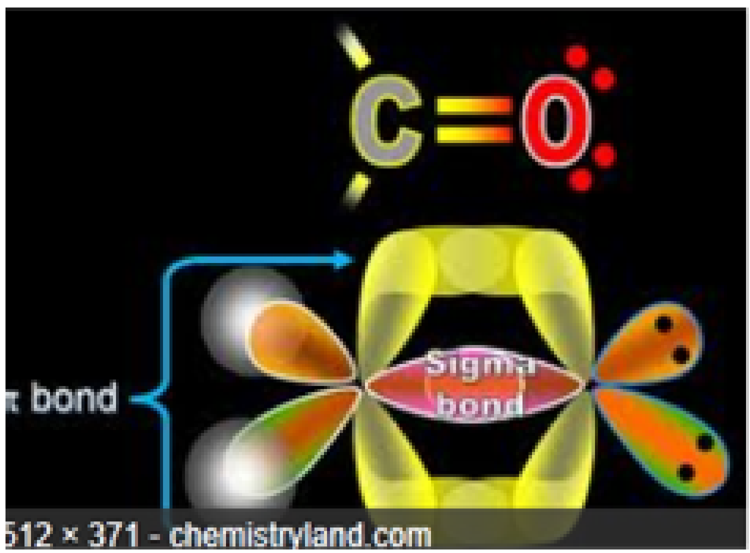
1) carbon tetrabromide
2) phosphorus trichloride
3) oxygen
4) the chlorine atom in hydrochloric acid (HCl)
5) boron trichloride
6) CH_2O
7) sulfur difluoride
8) either carbon atom in C_2H_2

1) sp^3, tetrahedral, 109.5 degrees.
2) sp^3, trigonal pyramidal, 107.5 degrees.
3) sp^2, linear, no bond angle
4) sp^3, linear, no bond angle
5) sp^2, trigonal planar, 120 degrees
6) sp^2, trigonal planar, 120 degrees
7) sp^3, bent, 104.5 degrees
8) sp, linear, 180 degrees

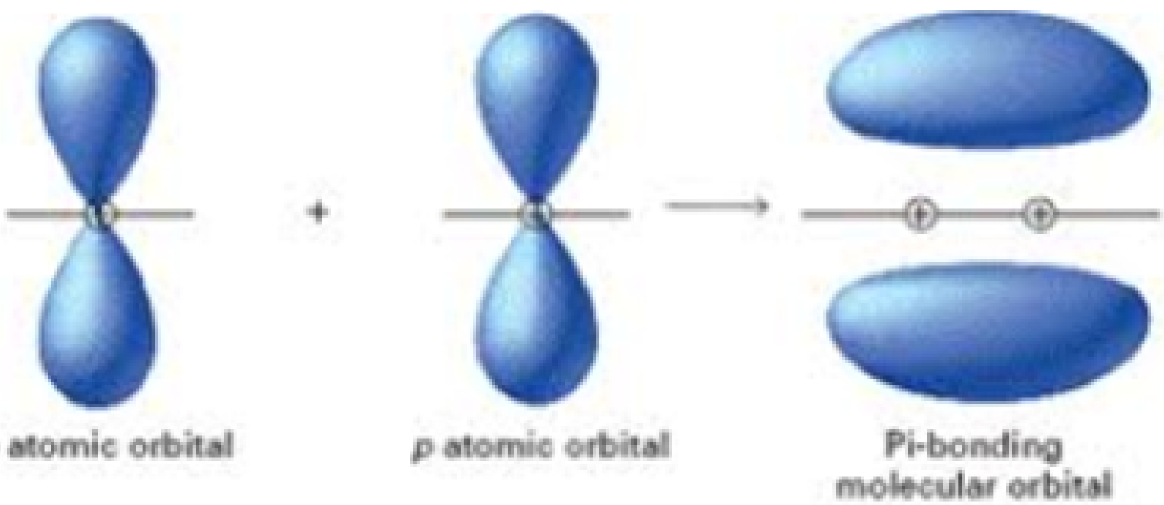
Molecular Shapes: VSEPR Summary

Electron Groups			Lone Pairs	Bonds	Geometry	Examples	
180	2	sp	0	2	Linear	$BeCl_2$	AB_2
120	3		0	3	Trigonal planar	BF_3	AB_3
118	3	sp^2	1	2	Bent	SO_2	AB_2E
109.5	4		0	4	Tetrahedral	CH_4	AB_4
107.5	4	sp^3	1	3	Trigonal pyramidal	NH_3	AB_3E
104.5	4		2	2	Bent	H_2O	AB_2E_2
90 A 120 E	5		0	5	Trigonal bipyramidal	PCl_5	AB_5
	5	dsp^3	1	4	See-saw	SF_4	AB_4 E
	5		2	3	T-Shaped	ClF_3	AB_3 E_2
	5		3	2	linear	I_3^-	AB_2 E_3
90	6		0	6	Octahedral	SF_6	AB_6
	6	d_2sp^3	1	5	Square pyramidal	$SbCl_5^{2-}$	
	6		2	4	Square planar	XeF_4	

sigma bonds
form when s and s orbitals overlap
when s and p orbitals overlap
when p and p orbitals overlap



pi bonds
form when 2 p orbitals overlap an cover an s orbital



These bonds determine
the strength of the bond.
the greater the overlap,
the stronger the bond.

remember as length increases strength and energy decreases

Representative bond lengths:

single	sp^3–sp^3	154 pm (1.54 Å)
	sp^3–sp^2	150 pm (1.50 Å)
	sp^3–sp	146 pm (1.46 Å)
double	sp^2–sp^2	147 pm (1.47 Å)
	sp^2–sp	143 pm (1.43 Å)
triple	sp–sp	137 pm (1.37 Å)

Metallic bonding

Metal and a metal, in which electrons are delocalized (move freely around in a sea of electrons)

-high electrical and thermal conductivity
-strong absorbers and reflectors of light, this reflection of light is responsible for the luster (shiny appearance)
-Malleability - ability to be beaten into sheets
-ductility - ability to be drawn into wire

planes of the atoms slide past each other without breaking bonds.

sea of delocalized electrons

Intermolecular forces

within covalent compounds

1) VDW - london dispersion, induced dipole

Argon, I_2, Cl_2, F_2 NONPOLAR

2) Dipole - dipole

H_2S, CH_3Cl, I-Cl POLAR

3) Hydrogen bonding H-FON & LP-FON

VDW- nonpolar (equal sharing of electrons)

weak forces of attraction between all molecules caused by the formation of temporary dipoles

fairly unreactive

dispersion

forces increase with increased electrons which increases contact points in the molecule and the boiling point

I>Br>Cl>F

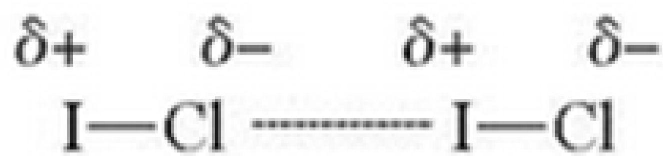
the longer the molecule the higher the boiling point

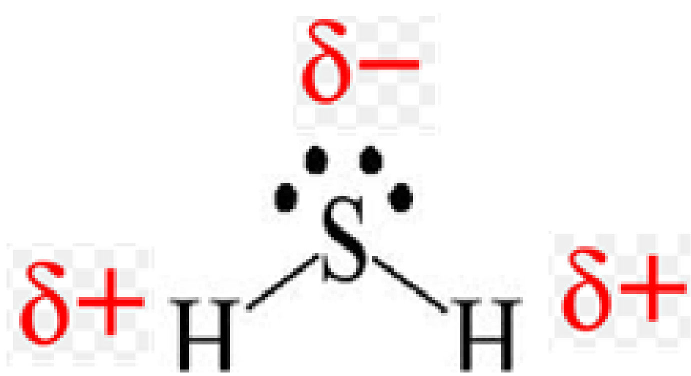
Butane>Propane>Ethane>Methane

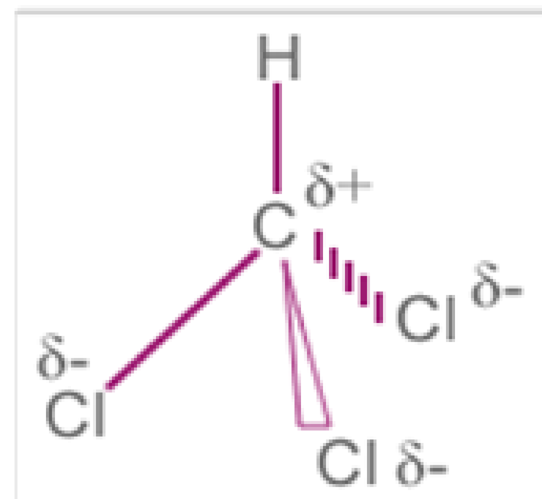
Dipole - Dipole (polar molecules)

electrons are not shared equally, one end ...partial positive...partial negative

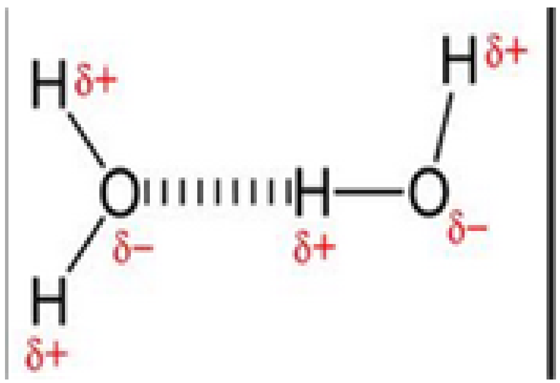
reactive due to carbon being deficient in electrons so it will be attacked by a nucleophile (nucleophilic substitution pg.233)

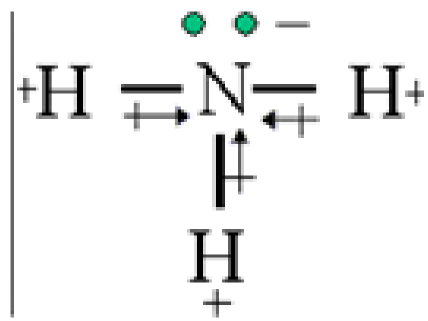






Hydrogen bonds H-FON & LP-FON

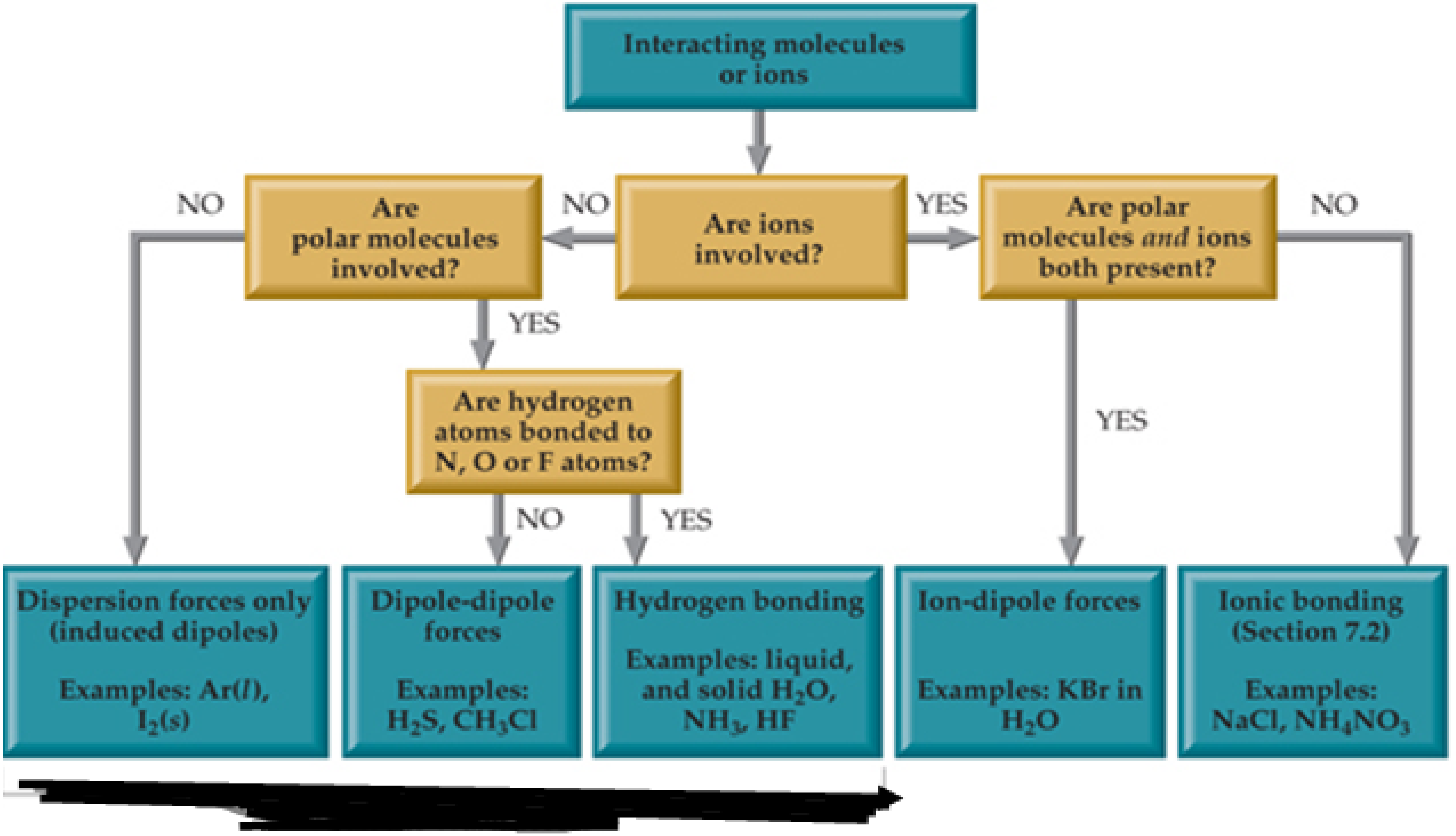




Boiling points
HB > DD > VDW (dispersion)

Solubility - "like dissolves like"
polar & polar, nonpolar & nonpolar

Conductivity
M>I (molten) >C


Interacting molecules or ions
Are ions involved?
NO
YES
Are polar molecules involved?
NO
YES
Are polar molecules *and* ions both present?
NO
YES
Are hydrogen atoms bonded to N, O or F atoms?
NO
YES
Dispersion forces only (induced dipoles)
Examples: Ar(*l*), $I_2(s)$
Dipole-dipole forces
Examples: H_2S, CH_3Cl
Hydrogen bonding
Examples: liquid, and solid H_2O, NH_3, HF
Ion-dipole forces
Examples: KBr in H_2O
Ionic bonding (Section 7.2)
Examples: NaCl, NH_4NO_3

Substance #1	Predominant Intermolecular Force	Substance #2	Predominant Intermolecular Force	Substance with Higher Boiling Point
(a) $HCl(g)$		I_2		
(b) CH_3F		CH_3OH		
(c) H_2O		H_2S		
(d) SiO_2		SO_2		
(e) Fe		Kr		
(f) CH_3OH		CuO		
(g) NH_3		CH_4		
(h) $HCl(g)$		NaCl		
(i) SiC		Cu		

Boiling points

HB > DD > Dispersion (VDW)

Substance #1	Predominant Intermolecular Force	Substance #2	Predominant Intermolecular Force	Substance with Higher Boiling Point
(a) HCl(g)	dipole-dipole	I_2	dispersion	I_2
(b) CH_3F	dipole-dipole	CH_3OH	hydrogen bonding	CH_3OH
(c) H_2O	hydrogen bonding	H_2S	dipole-dipole	H_2O
(d) SiO_2	Covalent	SO_2	Covalent	SiO_2
(e) Fe	metallic bonding	Kr	dispersion	Fe
(f) CH_3OH	hydrogen bonding	CuO	Ionic bonding	CuO
(g) NH_3	hydrogen bonding	CH_4	dipole-dipole	NH_3
(h) HCl(g)	dipole-dipole	NaCl	Ionic bonding	NaCl
(i) SiC	Covalent	Cu	metallic bonding	SiC

note: Silicon Dioxide is a giant covalent (molecular) compound see page 85

Warning! There's a bit of a problem here with modern syllabuses. The majority of the syllabuses talk as if dipole-dipole interactions were quite distinct from van der Waals forces. Such a syllabus will talk about van der Waals forces (meaning dispersion forces) and, separately, dipole-dipole interactions.

All intermolecular attractions are known collectively as van der Waals forces. The various different types were first explained by different people at different times. Dispersion forces, for example, were described by London in 1930; dipole-dipole interactions by Keesom in 1912.

Cambridge wants you to treat them separately