Structures, Shapes and Polarity of Molecules

Level 2 recap:

- Polar and non polar bonds
- Lewis diagrams
- Lone pairs
- Shapes
- Polarity

Do now: Brainstorm what you know/remember about these L2 concepts…

Bond polarity

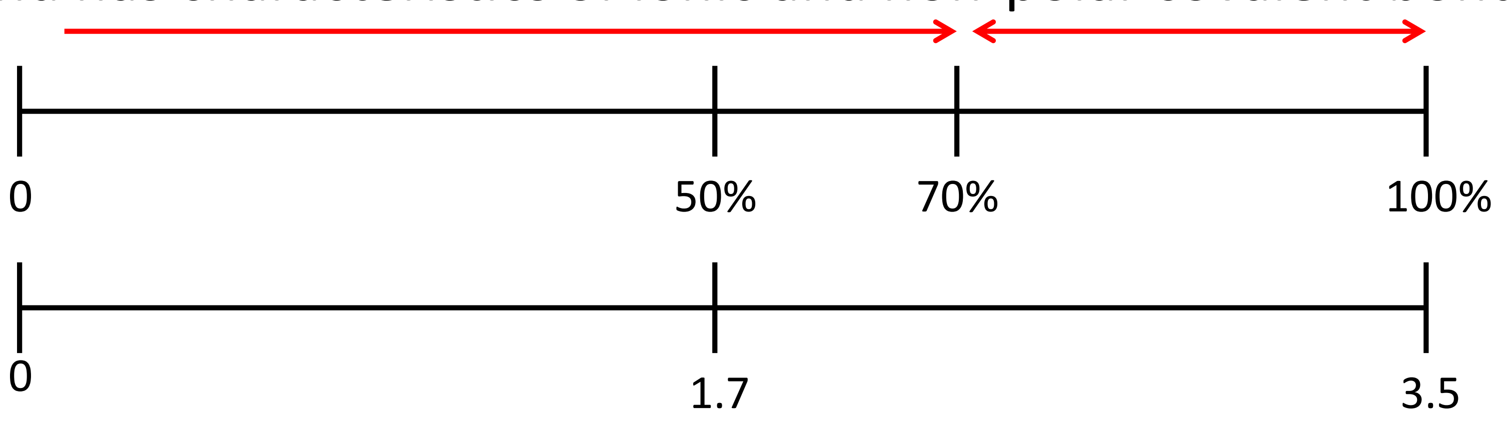
Differences in electronegativity between atoms tell us about the type of bonding between atoms

workbook
pg 23

There are 3 types of bonding between atoms

Ionic	Polar covalent	Non-polar covalent
Difference in electronegativity > 2.1	Difference in electronegativity 0.5 – 2.0	Difference in electronegativity < 0.5

Bonding between atoms is a continuum - each polar covalent bond has characteristics of ionic and non-polar covalent bonding



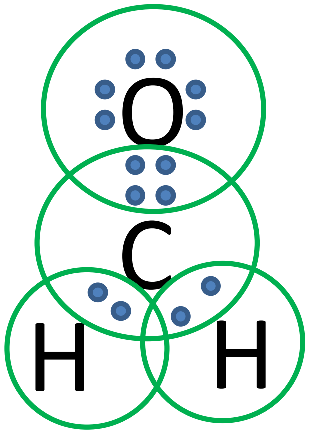
Lewis Structures

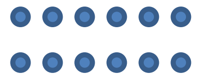
We are only interested in valence electrons

Atoms share a pair of electrons to form a covalent bond

Single bonds – one pair, double bonds – two pairs, triple bonds – three pairs

1. Place atoms around the central atom
2. Count the **total** number of valence electrons
3. Place 2 electrons between each pair of atoms
4. Place remaining electrons around outside atoms
   so they have a full valence shell
5. Place remaining electrons around central atom
   so it has a full valence shell
6. Check each atom has a full valence shell
7. If the central atom does not have a full valence shell move
   pairs of electrons to form double and triple bonds

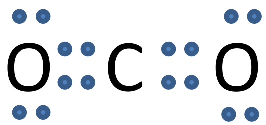




Do now:

Draw Lewis diagrams for the following compounds

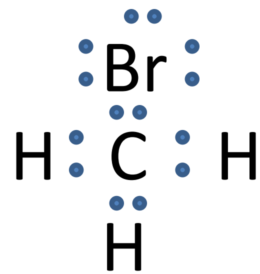
CO_2



linear

non-polar

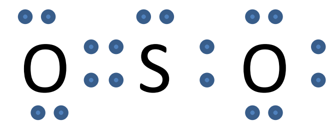
CH_3Br



tetrahedral

polar

SO_2



bent

polar

What are the shapes of these compounds?

Are these compounds polar or non-polar?

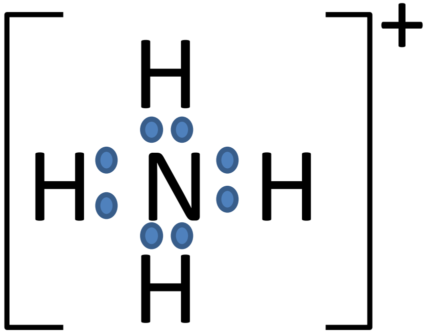
Lewis Structures

We can draw Lewis structures for ions

A + charge means we have 1 less valence electron in the Lewis structure

A – charge means we have 1 more valence electron in the Lewis structure

We need to draw the structure with brackets and indicate the charge eg NH_4^+



workbook
pg 27 Q3

Lewis Structures

Some atoms do not obey the octet rule – Be, B
Be requires only 4 electrons in its valence shell to be stable
B requires only 6 electrons in its valence shell to be stable
eg $BeCl_2$, BH_3

This year we learn that some atoms can accommodate more than 8 electrons in their valence shells and still be stable. Follow the same rules for drawing Lewis diagrams, make sure you count the number of electrons you have to work with.

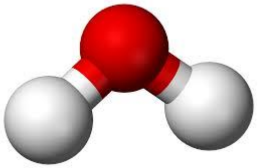
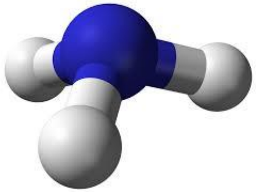
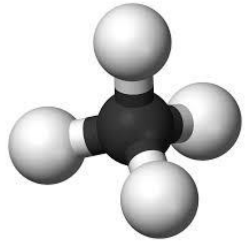
Lewis Structures

P, S, As, Cl, I, Br, Xe can all accommodate more than 8 electrons in their outer shell. This expanded valence shell can be 10 or 12 electrons.

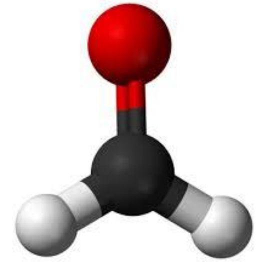
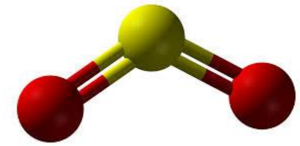
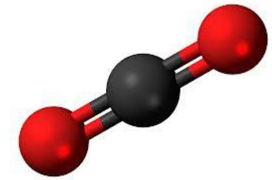
It is important to count how many electrons you need to include in your valence shell other wise you will miss lone pairs of electrons on the central atom.

eg PCl_5 BrF_5 I_3^- XeF_4 SF_6

SF_4 ClF_3

Shapes

From last year:

Bonding regions	Non-bonding regions	Shape	Bond angle	Example
4	0	tetrahedral	109°	methane (CH_4)
3	1	trigonal pyramid	< 109° (107°)	ammonia (NH_3)
2	2	bent	< 109° (105°)	water (H_2O)
3	0	trigonal planar	120°	BCl_3
2	1	bent	< 120°	SO_2
2	0	linear	< 180°	CO_2

Do now:

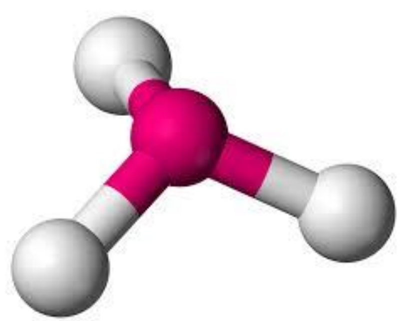
Draw Lewis diagrams and state the shape for the following compounds

SCl_2

4 regions of charge
2 bonding,
2 non-bonding
bent

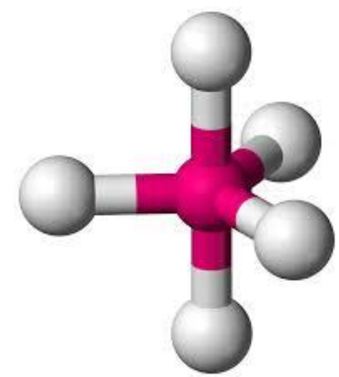
PCl_3

4 regions of charge
3 bonding,
1 non-bonding
trigonal pyramid



PCl_5

5 regions of charge
5 bonding,
0 non-bonding
trigonal bipyramid

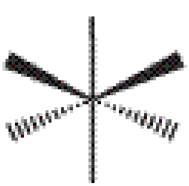


Shapes

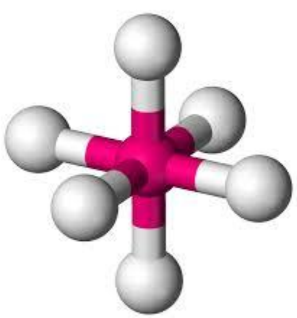
New this year:

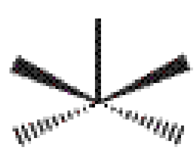
6 regions of charge around the central atom

Bonding regions	Non-bonding regions	Shape	Bond angle	Example
6	0	octahedral	90°	SF_6, PCl_6^-, SiF_6^-
5	1	square pyramidal	90°	BrF_5
4	2	square planar	90°	XeF_4, BrF_4^-, ICl_4^-

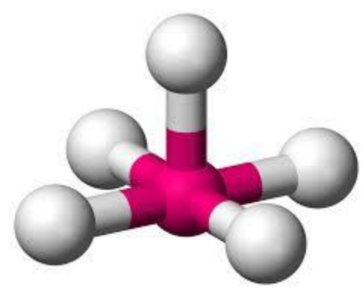


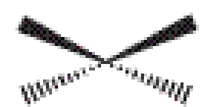
Octahedral



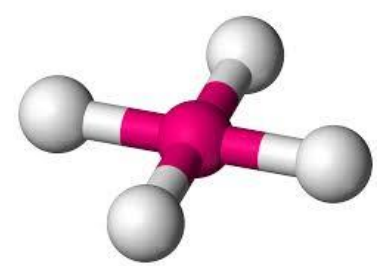


Square Pyramidal





Square Planar

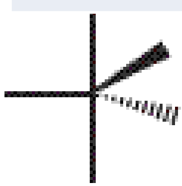


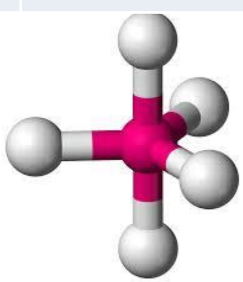
Shapes

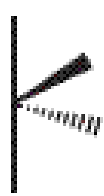
workbook
pg 34 and 35

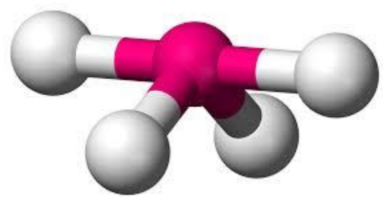
New this year:
5 regions of charge around the central atom

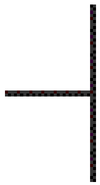
Bonding regions	Non-bonding regions	Shape	Bond angle	Example
5	0	trigonal bipyramid	120°, 90°	PCl_5, AsF_5
4	1	seesaw	180°, 120°, 90°	SF_4
3	2	T-shaped	90°	BrF_3, ClF_3
2	3	linear	180°	XeF_2, I_3^-

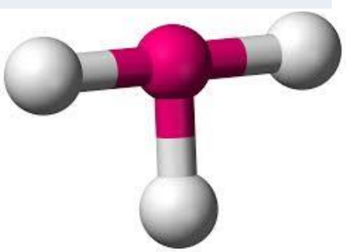

Triagonal Bipyramidal




Seesaw



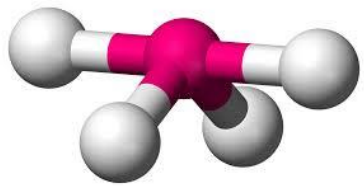

T-Shaped



Shapes

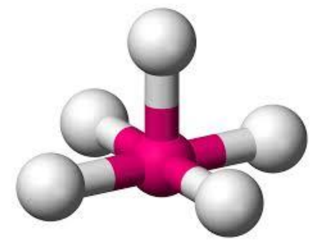
Draw Lewis diagrams and state the shape for the following compounds

SF_4



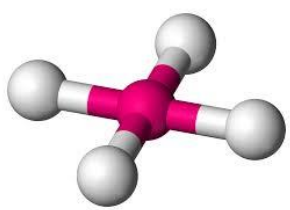
5 regions of charge
4 bonding,
1 non-bonding
see-saw

BrF_5



6 regions of charge
5 bonding,
1 non-bonding
square pyramid

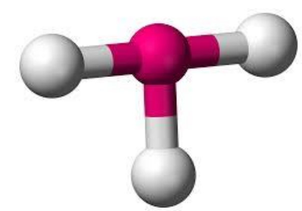
ICl_4^-



6 regions of charge
4 bonding,
2 non-bonding
square planar

XeF_2

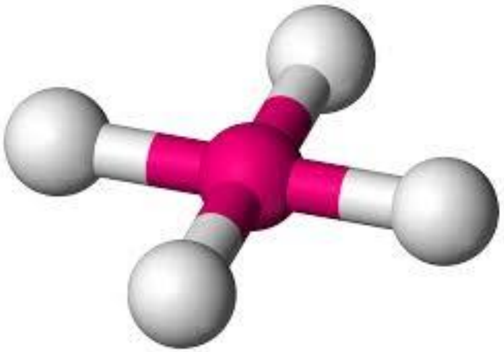
5 regions of charge
2 bonding,
3 non-bonding
linear



ClF_3

5 regions of charge
3 bonding,
2 non-bonding
T-shaped

Polarity

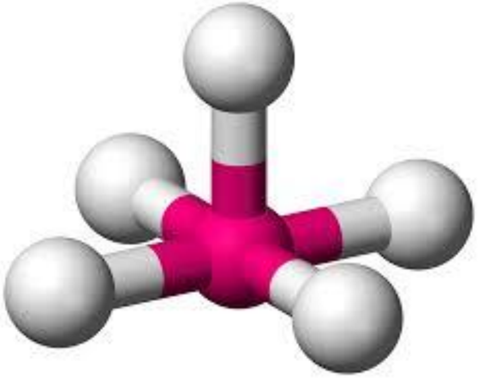


Non-polar molecules

- Same atoms bonded together
- Symmetrical around each bond (bonds can be polar)
- No lone pair (square planar molecules are an exception)
- **No net dipole** (bond dipoles cancel out)

Polar molecules

- Different atoms bonded together (polar bonds)
- Unsymmetrical around each bond
- Lone pair(s)
- **Net dipole** (bond dipoles do not cancel out)



Polarity

Decide if these shapes will form compounds that are polar or non-polar and why

Octahedral

non- polar

Square pyramidal

polar

Square planar

non- polar

Trigonal bipyramid

non- polar

Seesaw

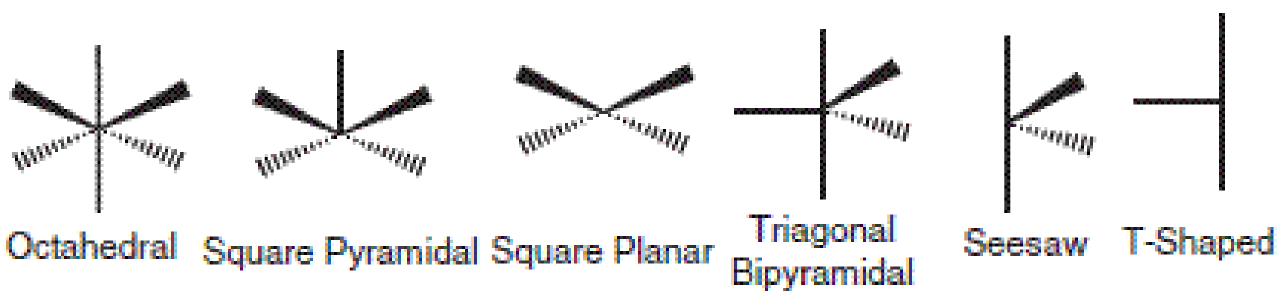
polar

T-shaped

polar

Linear

non- polar



2013 Exam Q 1 (c) (i)

(c) (i) Complete the following table.

Molecule	BrF_3	PCl_6^-
Lewis diagram		
Name of shape		

2013 Exam Q1 (c) (ii)

(ii) The Lewis diagrams for SF_4 and XeF_4 are shown below.

Compare and contrast the polarities and shapes of these two molecules.

N SF_4 is the see saw shape, whereas XeF_4 is the square planar shape. N

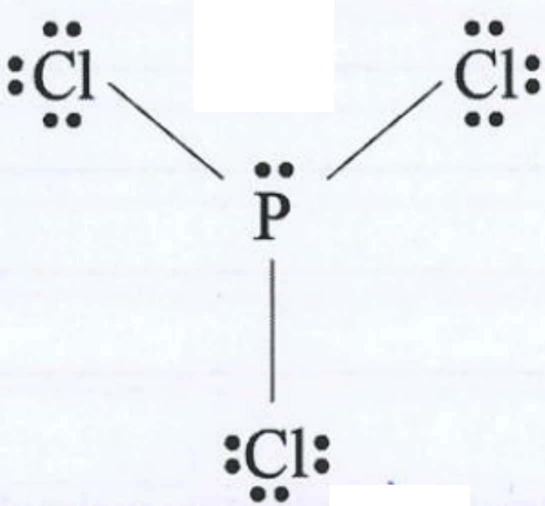
On your worksheet you have an A, M and E exemplar. Compare the exemplars and see if you can write bullet points on what needs to be covered in an Excellence response.

Do now:

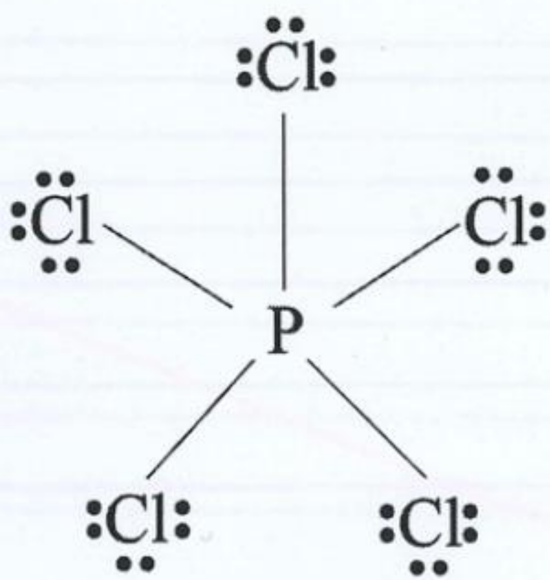
Answer the following exam question

(ii) The Lewis structures for the two molecules PCl_3 and PCl_5 are shown below. Compare and contrast the shapes and the polarities of these two molecules.

PCl_3



PCl_5



2013 Practise exam Q 1 (c) (ii)

Key points in your answer:

For Achieved: (2 of these)

- Both shapes OR
- Both polarities correct OR
- States electronegativity of P is greater than Cl OR
- States P – Cl bond is polar OR
- States symmetry of both molecules

For Merit:

- Makes links between TWO of: electronegativity, dipole moment, symmetry

For Excellence:

- Makes links between THREE of: electronegativity, dipole moment, symmetry

Both PCl_3 and PCl_5 contain polar Cl bonds, because Cl is more electronegative than P therefore Cl attracts the bonding electrons closer toward itself, leaving the Cl end of the bond slightly negative and the P end slightly positive. However, PCl_3 is a polar molecule whilst PCl_5 is a non polar molecule. In PCl_3, the central P atom has 4 areas of electron density around it which repel each other as far apart as possible due to VSEPR but as only 3 are bonded to Cl atoms, a trigonal pyramid shape is observed. This shape is asymmetrical so the centre of positive charge is not in the same place as the centre of negative charge, the dipoles do not cancel each other out and the PCl_3 molecule is polar.

PCl_5 is non-polar because of its different shape. In PCl_5, the central P atom has 5 areas of electron density around it which repel each as far apart as possible and as all 5 are bonded to Cl atoms, a trigonal bipyramid shape is observed. This is a symmetrical shape so the centre of positive charge is in the same place as the centre of negative charge, the dipoles cancel each other out and the overall PCl_5 molecule is non-polar //