Guidelines for Writing Lewis Structures

We illustrate using PCl_3.

Step 1. Count the total number of valence electrons in the molecule or ion.

PCl_3 = neutral molecule

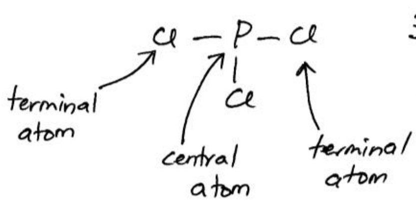
P (Group 5A) : 2.8.5 : has valence electron = $5e^-$

Cl (Group 7A) : 2.8.6 : has valence electron = e^-

∴ Total no. of valence electron = 5+3(7) = 5 + 21

= $26e^-$

Step 2. Use atomic symbol to draw a skeleton structure by joining the atoms with shared pairs of electrons (a single line) and substract two valence electrons for each bond.

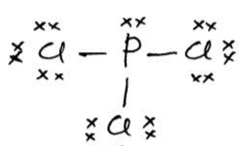


3 P–Cl bonds × $2e^-$ = $6e^-$

So 26 − 6 = $20e^-$ remaining.

Step 3. Place the remaining electrons in pairs so that each atom ends up with eight electrons (except H) to satify the octet rule, starting with the terminal atom

- First, place lone pairs on the terminal atom (surrounding atom) which is <u>more electronegative</u> to give each an octet.
- If any electrons remain, place them around the central atom.



Total no. of electron = $6e^-$ (bonding e^-) + $20e^-$ (non-bonding electron)

(10 lone pairs)

Step 4: Place any leftover electrons on the central atom, even if it give the central atom more than 8 electrons/octet.

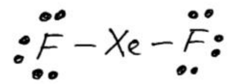
eg. XeF_2 or ClF_3

Xe : $8e^-$

$2F$: $2 \times 7e^- = 14e^-$

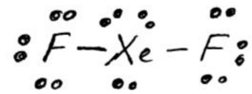
No. of valence electrons $= 22e^-$

F − Xe − F — $4e^-$ used up



$12e^-$ used up

$16e^-$



(Balance $= (22-6)e^-$
$= 6e^-$

place these e^- at the central atom

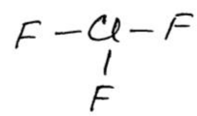
3 lone pairs (6 electrons) at the central atom)

eg. ClF_3

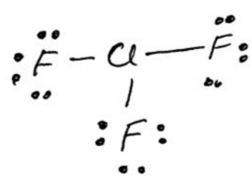
Cl : $7e^-$

$3F$: $3(7e^-) = 21e^-$

No. of valence electron $= 28e^-$



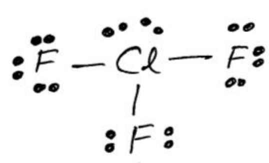
$6e^-$ used up (3 single bonds)



$18e^-$ used up. (9 lone pairs)

Used up electrons $= (6+18)e^-$
$= 24e^-$

Balance $= (28-24) = 4e^-$



place these electron at the central atom Cl.

Step 5 If the number of electrons around the central atom is less than eight, change single bonds to the central atom to multiple bonds.

- central atom can form double bond (2 shared pairs) or triple bond (3 shared pairs) - known as multiple bonds.

Remember !!

1. In nearly all compounds :
   - Hydrogen atoms form one bond (single bond)
   - Carbon atoms form four bonds [$-\overset{|}{\underset{|}{C}}-$, $-\overset{|}{C}=$ $=C=$ $-C\equiv$]
   - Nitrogen atoms form three bonds [$-\underset{|}{N}-$ / $=N-$ / $N\equiv$]
   - Oxygen atoms form two bond : [$-O-$ / $O=$]
   - Halogen form one bond when they are terminal atoms (surrounding central atom).
   - Fluorine always a surrounding atom.

2. Although Lewis structures can predict the number of covalent bonds an atom will form - they do not give an accurate representation of where electrons are located in a molecule.

3. Lewis structures are not meant to convey the shapes of molecules. Lewis structures are used to predict geometries by a method based on the repulsions between valence-shell electron pairs. (VSEPR model). ie repulsion among pairs of bonding and lone-pair electrons at the central atom.

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