

11/4/05

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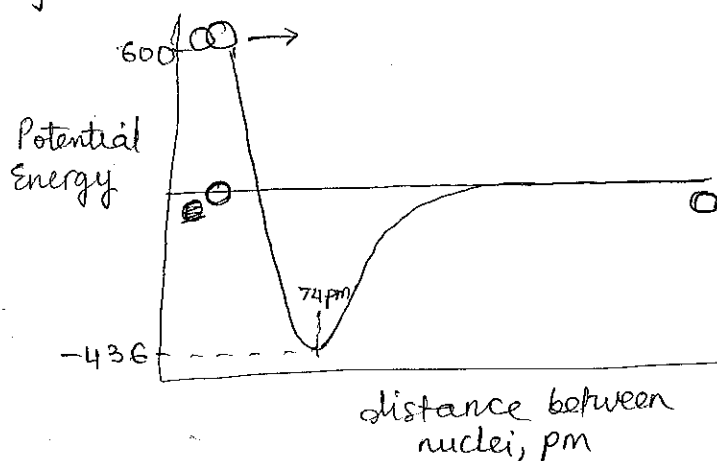
CHM 123 - Lecture (Friday 10:30am)

CHAPTER 8 - Covalent Bonding

- Covalent Bonds - shared electron pair that connect atoms in molecular compounds.
- Covalent bond typically occur between non metal atoms (CO , H_2O , HCl , organic molecules).

Covalent Bond Lengths

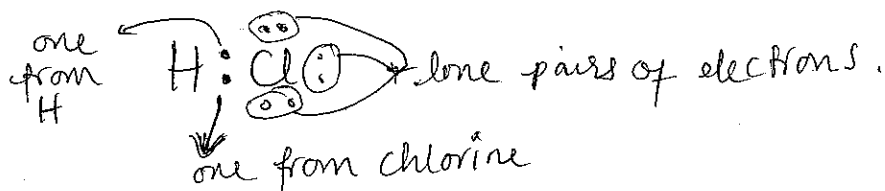
e.g. H_2



- At zero potential energy, H atoms are far apart.
- At 600, atoms are too close and overlap to an extent that repulsion takes place. Energy is high.
- At -436, atoms overlap just enough to form a stable bond ($\text{H}-\text{H}$), bond distance is 74 pm and energy is negative.

Single Bonds and Lewis Structures

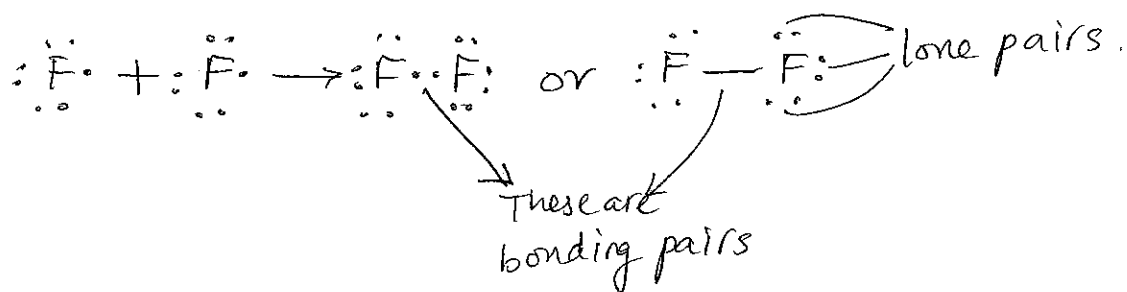
- Octet Rule - whenever possible, valence electrons in covalent compounds distribute so that each main group element is surrounded by 8 electrons (except hydrogen which wants 2).
- Atoms that form covalent bonds do this by sharing electrons.
- Hydrogen: one valence electron $\text{H}^\bullet$ (Lewis-dot structure)
- Chlorine: seven valence electrons $\cdot\ddot{\text{Cl}}:$



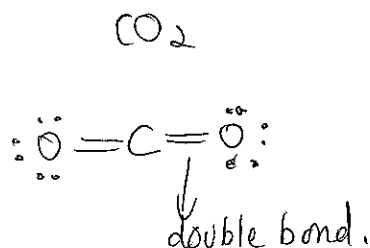
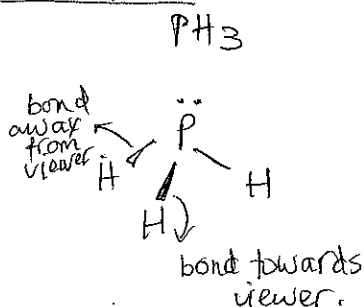
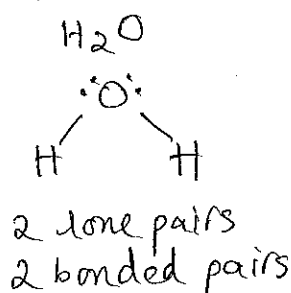
e^- around Cl = 8 } octet rule satisfied for Cl
 # e^- around H = 2 } + 2 electrons for H by sharing of electrons.

Lewis Structures

- the # of covalent bonds determined by # of electrons atoms must share to complete octet.
- single covalent bond - 2 atoms share one electron pair.
- shared electrons are counted with each of the atoms in the bond.



Lewis Structure Examples



Guidelines for Drawing Lewis Structures

- Count up the total # of valence electrons available. If the species is an ion, either add the negative charge (anion) to the total or subtract the positive charge (cation).

e.g. SOCl_2 (typically first atom is in the center, in this case S is in the center).

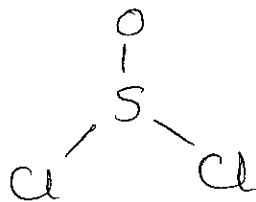
$$\text{S: } 6 \text{ valence } e^- \times 1 = 6$$

$$\text{O: } 6 \text{ valence } e^- \times 1 = 6$$

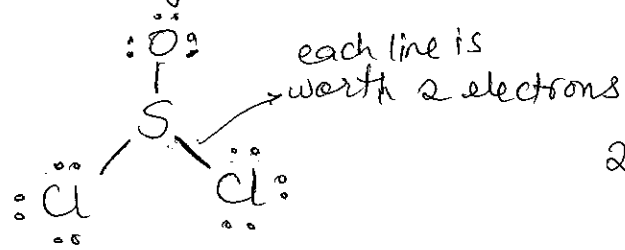
$$\text{Cl: } 7 \text{ " " } \times 2 = 14$$

$$\text{Total valence } e^- = 26$$

- ② Draw a skeleton structure by joining with a single line (shared pairs of electrons/bonds).



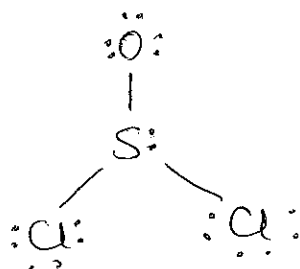
- ③ Place lone pairs around each non-H atom to satisfy octet rule, starting with terminal atoms



24 valence electrons in current structure.

We have 26 valence electrons.

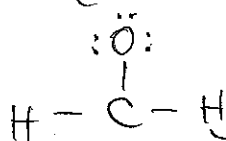
- ④ Place any leftover electron pairs on the central atom—even if it will give it more than an octet.



26 valence electrons

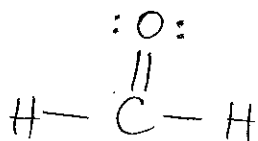
- ⑤ If number of electrons around central atom is less than eight, change single bonds around the central atom to multiple bonds. Be sure to alter the lone pairs on terminal atoms.

e.g. CH_2O (C in center)



- 12 valence e^-
- carbon only has 6 electrons around it, not an octet.
Hydrogen will NEVER have lone pairs. Already has 2 electrons. ~~to make~~

- ⑥ Make double bond between Carbon and ~~atom~~ oxygen.



12 valence electrons and each atom has an octet, Hydrogen has 2 electrons.

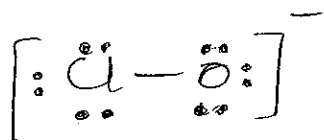
Example:

(a) ClO^-

$$Cl: 1 \times 7 \text{ val } e^- = 7$$

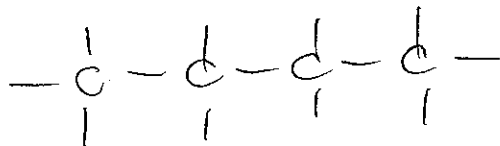
$$O: 1 \times 6 \text{ val } e^- = 6$$

$$\begin{array}{r} \text{(neg. charge)} \quad (-) : 1 \times 1 \text{ " " } = 1 \\ \hline \text{Total valence} = 14 e^- \end{array}$$



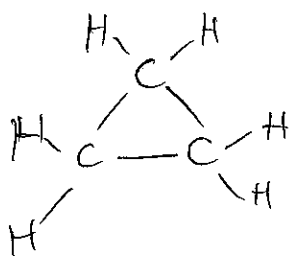
Single Bonds and Hydrocarbons

- Carbon unique - forms strong bonds with carbon as well as other atoms like H, O, N, S and halogens.

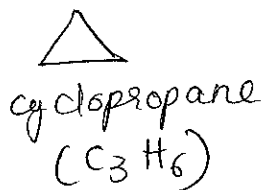


- Each carbon can form 4 covalent bonds.
- hydrocarbons contain only C and H.
- Alkanes (contain only C-C and C-H bonds). Known as saturated hydrocarbons. general formula: $C_n H_{2n+2}$ ($n = \# \text{ of carbons}$).
- Alkanes can be straight chained or branched.

- Cycloalkanes can also exist (C_nH_{2n})



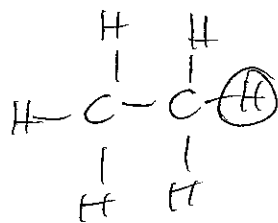
or



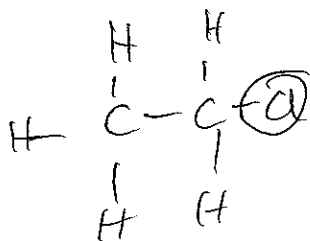
cyclobutane

Substitutions in Hydrocarbons

- one or many C-H bonds can be replaced with bonds between C and other atoms.



ethane



chloroethane