# Chapter 9
# Bonding – 2
# Polar Covalent Bond, Electronegativity, Formal Charge, Resonance

Dr. Sapna Gupta

# Writing Lewis Structures

1. Draw the skeleton structure of the molecule or ion by placing the lowest electronegative element in the center.
2. Add the total number of valence electrons. Subtract electron(s) if is a cation and add electron(s) if anion
3. Share one pair of electrons between each atom, and subtract those from the total number of electrons.
4. Distribute electrons to the atoms surrounding the central atom or atoms to satisfy the octet rule.
5. Distribute the remaining electrons as pairs to the central atom or atoms.
6. Add multiple bonds if atoms don't have the octet.

**Hint**:

H and halogens have single bonds (unless halogen is in the center)

O and S has two bonds (two single or one double)

N and P has three bonds (three single, one double one single or triple)

C has four bonds (different combination)

# Steps for Writing Lewis Structure

**Steps for Drawing Lewis Structures**

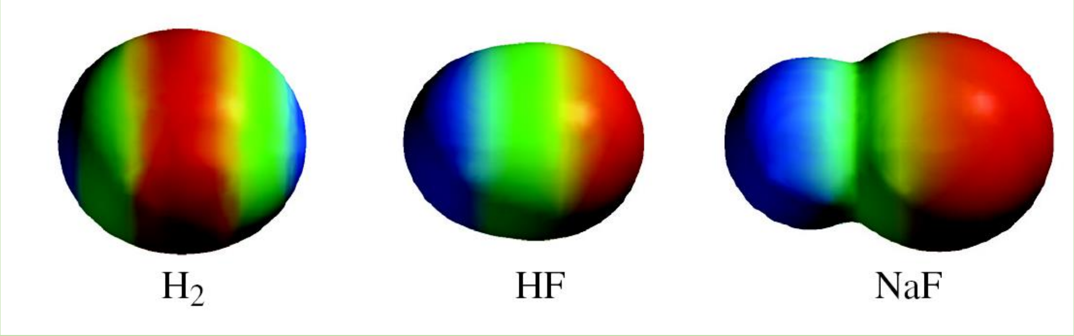
| Step | $CH_4$ | $CCl_4$ | $H_2O$ | $O_2$ | $CN^-$ |
|---|---|---|---|---|---|
| 1 | H<br>H—C—H<br>H | Cl<br>Cl—C—Cl<br>Cl | H—O—H | O—O | C—N |
| 2 | 8 | 32 | 8 | 12 | 10 |
| 3 | 8 – 8 = 0 | 32 – 8 = 24 | 8 – 4 = 4 | 12 – 2 = 10 | 10 – 2 = 8 |
| 4 | H<br>H—C—H<br>H | :C̤̈l:<br>:C̤̈l—C—C̤̈l:<br>:C̤̈l: | H—O—H | :Ö̤—Ö̤: | :C—N̤̈: |
| 5 | — | — | H—Ö̤—H | — | — |
| 6 | — | — | — | :O=O: | $[:C\equiv N:]^-$ |

# Polar Covalent Bond

- ***Nonpolar covalent bond*** = electrons are shared *equally* by two bonded atoms
- ***Polar covalent bond*** = electrons are shared *unequally* by two bonded atoms

| $M{:}X$ | $M^{\delta+}X^{\delta-}$ | $M^{+}X^{-}$ |
|---|---|---|
| **Pure covalent bond** | **Polar covalent bond** | **Ionic bond** |
| Neutral atoms held together by *equally* shared electrons | Partially charged atoms held together by *unequally* shared electrons | Oppositely charged ions held together by electrostatic attraction |

- Electron density distributions




red ⇒ high electron density
green ⇒ intermediate electron density
blue ⇒ low electron density

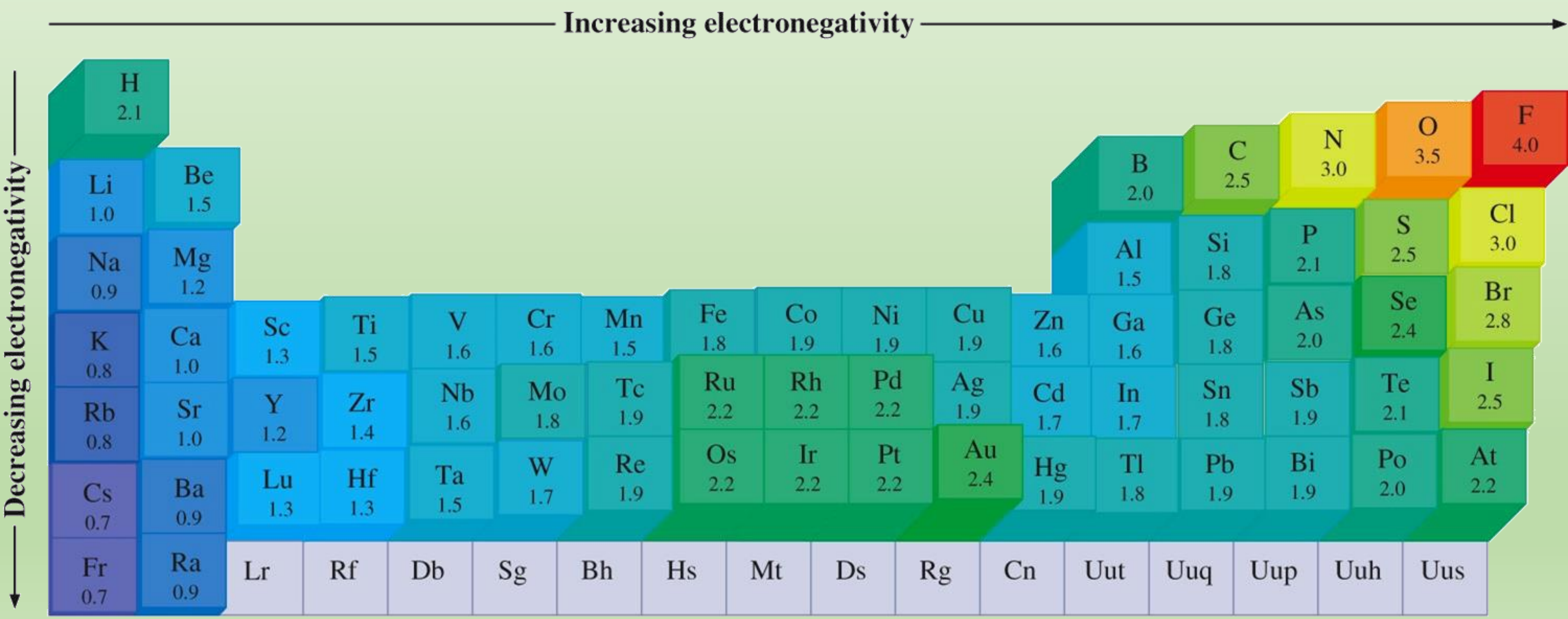
$\delta+$ $\delta-$
H – F

$\longmapsto$
H – F

alternate representations

# Electronegativity

- **Electronegativity,** *X*, is a measure of the ability of an atom in a molecule to attract bonding electrons to itself. Electronegativity increases across the group and decreases down a group.



# Polar Bond, Contd…

The difference in electronegativity between the two atoms in a bond is a rough measure of bond polarity.

When the difference is very large, an ionic bond forms. When the difference is large, the bond is polar. When the difference is small, the bond is nonpolar.

**Solved Problem:**

Using electronegativities, arrange the following bonds in order by increasing polarity: C—N, Na—F, O—H.

For C—N, the difference is 3.0 (N) – 2.5 (C) = 0.5.

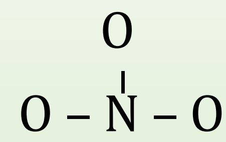
For Na—F, the difference is 4.0 (F) – 0.9 (Na) = 3.1.

For O—H, the difference is 3.5 (O) – 2.1 (H) = 1.4.

Bond Polarities: C-N < O-H < Na - F

## Solved Problem:
## What is the Lewis structure of $NO_3^-$?

1) Draw skeletal structure with central atom being the least electronegative.



2) Add valence electrons. Add 1 for each negative charge and subtract 1 for each positive charge.

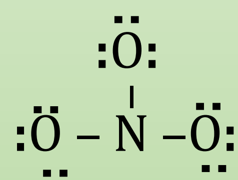
$$NO_3^- \Rightarrow (1 \times 5) + (3 \times 6) + 1 = 24 \text{ valence } e^-$$

$24\,e^-$

3) Subtract 2 for each bond in the skeletal structure.

$-\,6\,e^-$

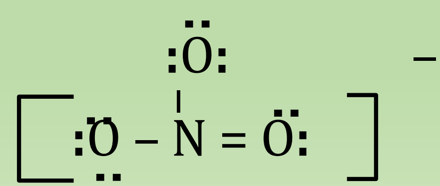
4) Complete electron octets for atoms bonded to the central atom except for hydrogen.



$18\,e^-$

5) Place extra electrons on the central atom.

6) Add multiple bonds if atoms lack an octet.



$24\,e^-$

# Lewis Structures and Formal Charge

The electron surplus or deficit, relative to the free atom, that is assigned to an atom in a Lewis structure.

$$\text{Formal Charge} = \boxed{\text{Total valence electrons}} - \boxed{\text{Total non-bonding electrons}} - \frac{1}{2}\boxed{\left(\text{Total bonding electrons}\right)}$$

Example: $H_2O$ = $\mathrm{H}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{O}}}:\mathrm{H}$

| H: | | O: | |
|---|---|---|---|
| orig. valence $e^-$ | = 1 | orig. valence $e^-$ | = 6 |
| – non-bonding $e^-$ | = –0 | – non-bonding $e^-$ | = –4 |
| – 1/2 bonding $e^-$ | = –1 | – 1/2 bonding $e^-$ | = –2 |
| formal charge | = 0 | formal charge | = 0 |

Formal charges are not “real” charges.

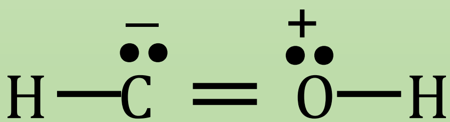
# Formal Charge, contd..

- A Lewis structure with *no* formal charges is generally better than one with formal charges.
- Small formal charges are generally better than large formal charges.
- Negative formal charges should be on the more electronegative atom(s).

Example: $H_2CO$

H C O H  or  H C O ?  (with H above and H below the second C)

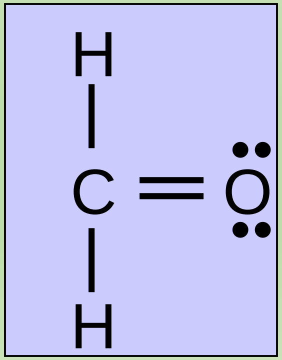
Answer:



For carbon
4-2-3= **-1**

For oxygen
6-2-3=**+1**

For carbon
4-0-4= **0**

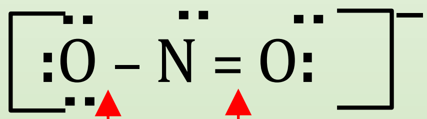


For oxygen
6-4-2=**0**

# Resonance

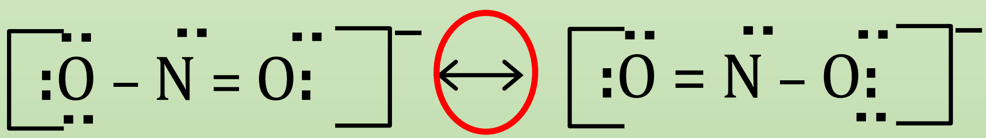
***Resonance*** is the movement of electrons over two or more bonds. In these cases two or more equally valid Lewis structures can be written.

Example: $NO_2^-$



These two bonds are known to be identical.

<u>Solution:</u>



Two ***resonance structures***, their average or the ***resonance hybrid***, best describes the nitrite ion.

The double-headed arrow indicates resonance.

# Key Points

- Lewis structures
- Bond polarity
- Electronegativity
- Formal charge
- Resonance structures