Chapter 4 : Chemical bonding
Lesson 15 : Types of bonds

* Introduction :

As we Know ; Chemical Compounds are formed by chemical Combination "bonding" between their atoms , but HOW ??

* Types of Chemical bonds :

Bond	Ionic	Covalent	Metallic
1. Element	Metal + Nonmetal	Two nonmetals	Metals
2. Bonding	e^- - transfer from metal (low IE) to nonmetal (high E.A) Thus, electrostatic attraction between Cation and anion	sharing of e^-'s between the two nonmetals (high EA), Thus attraction through shared electrons.	Valence electrons are shared as Pool between the metal atoms. "Nuclei in sea of electrons" "Electron Sea Model"
3. Example	$Na^+ Cl^-$ Cation Anion	H_2O water	$Na_{(s)}$ Sodium metal

* Classification of Covalent bonds :

They are three different types of Covalent bonding based on the no. of shared electrons.

Covalent bond.	Single	Double	Triple
1. No. of Shared e⁻'s	Two	four	Six
2. No. of Shared e⁻ pairs	One	Two	Three
3. No. of Lines	Single Line	Double Line	Triple Line
4. Examples	H—H ; C—C C—H ; H—Cl	O=O ; C=O C=C ; C=N	N≡N ; C≡N C≡C

* **In general;**

The type of bonding used is the suitable way for the atom to reach "octet rule" (or the noble gas e^- configuration).

e.g.

$^{8}O: 1s^2\ 2s^2\ 2p^4$ So, No. of valence electrons = 6 (Group No. 6A)

So, Oxygen atom makes bond in such way that it gains two electrons to reach the "octet rule"

$$^{8}O: 1s^2\ 2s^2\ 2p^4 \xrightarrow[\text{sharing of } 2e^- \text{ (Covalent bond)}]{\text{gains } 2e^- \text{ (ionic bond)}} O^{2-}: 1s^2\ \underbrace{2s^2\ 2p^6}_{\text{octet}}$$

* Lewis Representations:

Lewis structures allow us to easily see the number of valence electrons in an atom. Lewis notation uses two dots (duet) to represent a pair of electrons.

e.g.

$^{8}O : 1s^2 2s^2 2p^4$ Valence electrons = 6

Thus;

Lewis structure for O-atom is : $:\ddot{\underset{\cdot}{O}}\cdot$

N.B. Dots represent the electrons that the element could lose and empty spaces indicate the electrons that are needed to satisfy the octet rule.

Examples:

$\dot{Li}$ $\underset{\cdot}{\dot{Be}}$ $\cdot\dot{B}\cdot$ $\cdot\underset{\cdot}{\dot{C}}\cdot$ $\cdot\underset{\cdot}{\ddot{N}}\cdot$ $:\underset{\cdot\cdot}{\ddot{F}}\cdot$ $:\underset{\cdot\cdot}{\ddot{Ne}}:$

(Put the electrons firstly single then make pairing)

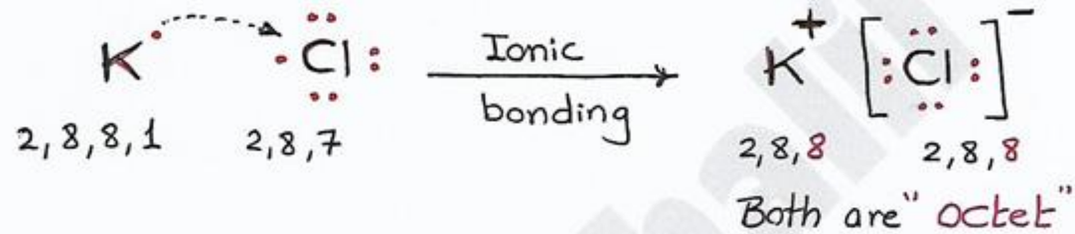
N.B.

Two elements not subject to Octet rule : H & He because they have only one E-level of max e^--capacity = 2

$\dot{H}$: hydrogen can only lose or gain one electron.

* Lewis structures for ionic Compounds:

Lewis structure represents ionic bonding by moving electron dots from metal to nonmetal, where the lewis structure of anion is usually written in brackets with charge in its upper right corner:



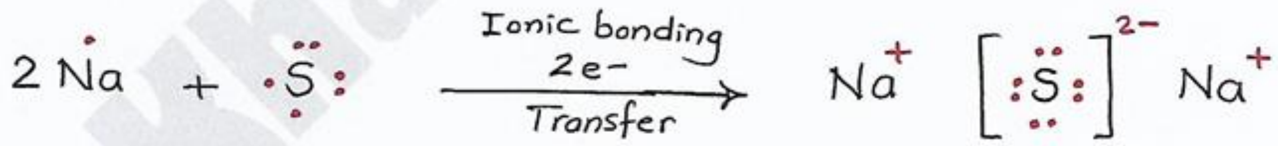
Another example;

Ionic Compound of sodium and sulfur needs two Na-atoms and one S-atom, Explain Why ??

Na-atom must lose one valence electron to obtain "octet", and.

S-atom " gain two electrons to obtain "octet"

Thus: Two Na-atoms are needed to combine with one S-atom.



* Lattice Energy:

Ions are arranged in a pattern "Crystal lattice" where the crystal lattice maximizes attractions between cations and anions leading to more stable arrangement.

extra stability of crystal lattice is measured as "Lattice Energy".

N.B.
⇒ Lattice Energy increases with :

* Increasing the charge on the ion.
* Decreasing the size of ion.

* Lattice energy: the energy needed to separate a mole of a solid ionic compound into its gaseous ions.

Q. What are the properties of ionic compounds?

A.

- Physical properties: a- high melting points > 300°C.
  b. high boiling points.
  c. Hard and brittle solids.
- Solid ionic compounds:
  a. All are crystalline at room temperature.
  b. don't conduct electricity, but in liquid state they are strong electrolyte "Conductors"
  c. Liquid states are thermal insulators.
- Solubility of ionic compounds:
  Many solid ionic compounds are soluble in water.

Chapter 4: Chemical Bonding
Lesson 16: Factors Affecting Bond Type

* Introduction:

Oxygen and sulfur (Group 6A) form covalent compounds with hydrogen in atomic ratio 1:2 (H_2O & H_2S). However, their reactivity is quite different (H_2S is more reactive).

Reactivity depends on:

1. Type of bonds within the compound.
2. polarity.
3. Bond energy (energy required to break the bonds).
4. Energy required to separate the molecules.

* Covalent Bonding & Lewis Structures:

Molecular compounds contain covalent bonds (of shared electrons).
Lewis theory represents the covalent bonds by drawing the neighboring atoms to share e⁻-pairs to obtain octet rule (or duet for hydrogen).

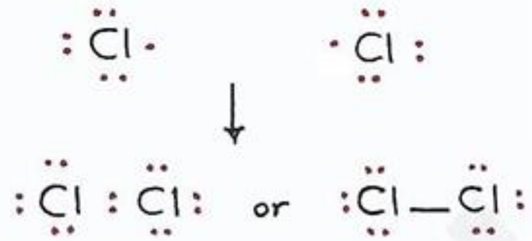
Example; water H_2O

Lewis structures for atoms

Lewis structures for molecule

Example; Diatomic Halogen Molecules

Lewis structure :



Both are Octet

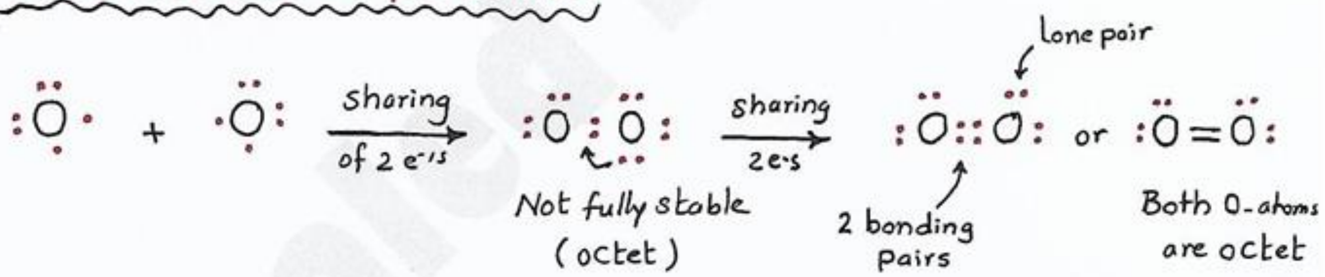
Q. Draw Lewis structure for H_2 molecule?

$\dot{H} + \dot{H} \longrightarrow H : H$ or $H—H$

H obeys duet rule

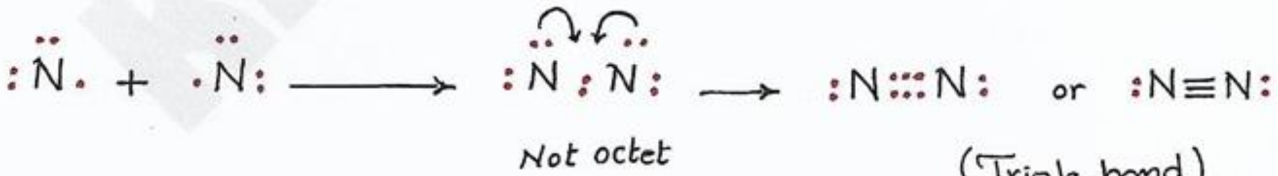
* Lewis structures for Multiple bonds :

Oxygen :



(Double bond)

Nitrogen :



(Triple bond)

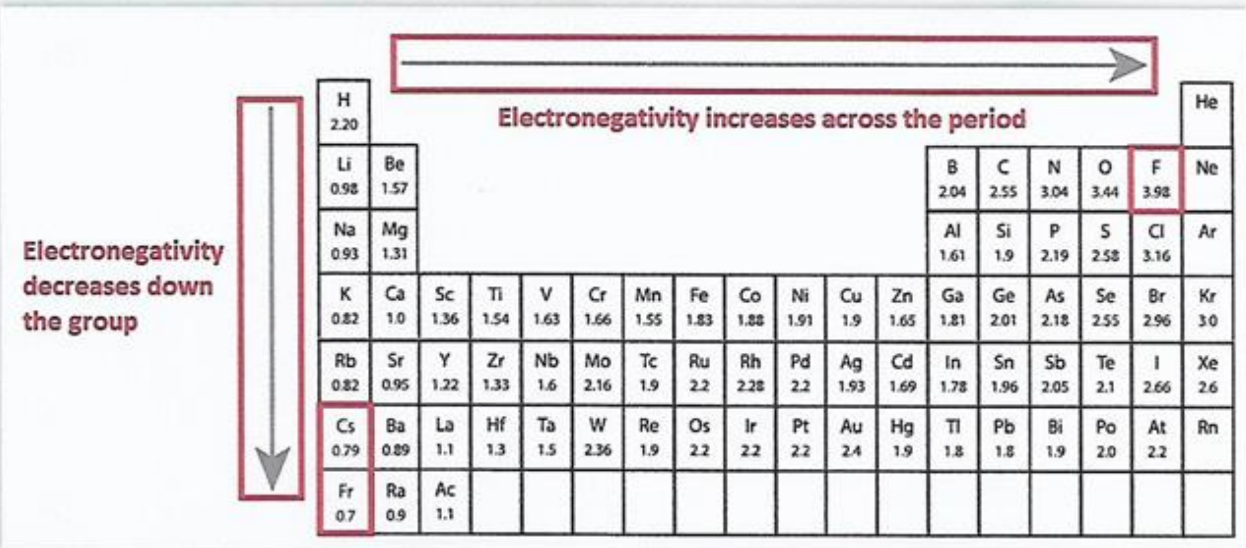
* Electronegativity : (E.N.)

It is the relative ability of atoms to attract shared electrons.

N.B.

1. EN (nonmetal) > EN (metal)
2. F-atom has the highest EN-value (4.0) while Cs and Fr atoms have the lowest EN-value (0.7).
3. As EN difference increases as the bond polarity increases.

. Electronegativity values and Trends in periodic table :



H 2.20																	He
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
Na 0.93	Mg 1.31											Al 1.61	Si 1.9	P 2.19	S 2.58	Cl 3.16	Ar
K 0.82	Ca 1.0	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.9	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.0
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.2	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
Cs 0.79	Ba 0.89	La 1.1	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9	Ac 1.1															

* Electronegativity & Bond Polarity :

* The bond polarity depends on the elements' electronegativities.

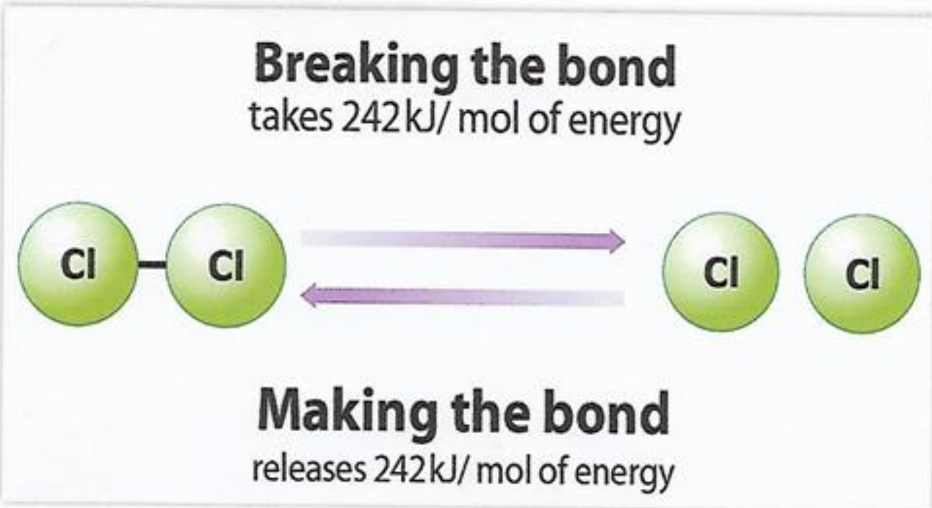
Type of bond

$\Delta EN = 0$	$\Delta EN = 0.1-0.4$	$\Delta EN = 0.5-1.9$	$\Delta EN \geq 2.0$
∴ Pure Covalent (Equal sharing of e⁻'s)	Nonpolar Covalent	Polar Covalent	Ionic (Transfer of e⁻'s)
e.g. H — H	Cl — Br	$\overset{\delta+}{H}-\overset{\delta-}{Cl}$	$Na^{+}\ Cl^{-}$
$\Delta EN = 0$	$\Delta EN = 0.2$	$\Delta EN = 0.96$	$\Delta EN = 2.23$

* Bond Energies :

The amount of energy, in gaseous state, that it takes to break 1 mol of a bond in a compound.

In general, Chemical reactions involve breaking of bonds in reactants and making new bonds in products.



In general, the more electrons two atoms share, the stronger the covalent bond.

C≡C (837 kJ) > C=C (611 kJ) > C—C (347 kJ)
C≡N (891 kJ) > C = N (615 kJ) > C—N (305 kJ)

* Trends in bond strength:

* Bond Lengths :

Bond Length is the distance between the nuclei of bonded atoms.

In general,

lewis theory predicts that the more electrons two atom share, the shorter the bond should be when comparing bonds of like atoms.

Bond length : (pm)

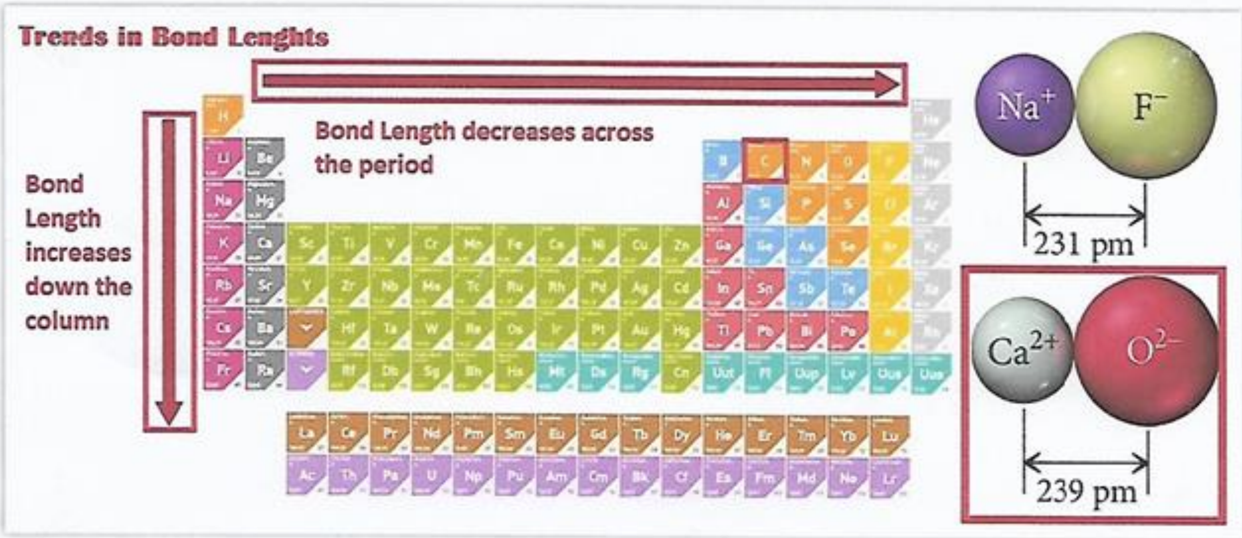
C≡C < C=C < C—C
(120) (134) (154)

Bond Length for halogens:

F–F < Cl—Cl < Br—Br < I—I
(143) (199) (228) (266)

* Trends in bond length :

By Comparing the average bond length of a bond between two particular atoms in a variety of Compounds.



* Bond length decreases across the period:

$$C-C > C-N > C-O$$

154 147 143 pm

* Bond length increases down the group :

$$F-F < Cl-Cl < Br-Br$$

144 198 228 pm

* In general, as bonds get longer, they also get weaker.

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