⊛ For a $\boxed{\text{metal}}$ the total number of dots in the Lewis symbol is the number of electrons the atom loses to form a cation.

⊛ For a $\boxed{\text{nonmetal}}$, the number of unpaired dots equals

- the number of electrons the atom gains to form an anion

  or,
- the number is shares to form covalent bonds.

⊛ The octet rule:

- when atoms bond, they lose, gain, or share electrons to attain a filled outer level of 8 electron (2 for H and Li)

⇒ Valence electrons are transferred or shared between atoms so that each atom attains 8 valence electrons or the electron configuration of the closest noble gas in the periodic table.

⇒ All of the noble gases, except helium, have 8 valence electrons.

⊛ Ionic Compounds:

- Electrons are transferred forming cations and anions.
- The resulting ions attract each other through electrostatic forces.

⊛ $Na + Cl_2$

Cl_2 → Valence electron - 7

Na → Valence electron - 1

⇒ for 2 Cl, we need 2 Na.

∴ $2Na^+ + 2Cl^- \Rightarrow \cancel{2NaCl_2}$ $Na_2Cl_2 \Rightarrow NaCl$

✲ Three ways to depict electron transfer in the formation of Li^+ and F^-

⇒ Electron configurations:

$$Li\ 1s^2 2s^1 + F\ 1s^2 2s^2 2p^5 \rightarrow Li^+\ 1s^2 + F^-\ 1s^2 2s^2 2p^6$$

⇒ Orbital diagrams:

Li [↑↓] [↑] [| |] ⇒ Li^+ [↑↓] [ ] [| |]
1s 2s 2p → 1s 2s 2p

+ +

F [↑↓] [↑↓] [↑↓|↑↓|↑] F^- [↑↓] [↑↓] [↑↓|↑↓|↑↓]
1s 2s 2p 1s 2s 2p

⇒ Lewis electron-dot symbols:

$$Li\cdot + \cdot\ddot{\underset{..}{F}}: \rightarrow Li^+ + :\ddot{\underset{..}{F}}:^-$$

✲

$Al^{3+} + N^{3-} \Rightarrow AlN$

$2Na^+ + O^{2-} \Rightarrow Na_2O$

More on page 13–16

✲ The Duet Rule:

- Small elements tend towards a helium outer shell, a 'duet' rule.

⇒ He $1s^2$ or [He] helium atom

Li^+ $1s^2$ or [He] lithium cation

Be^{2+} $1s^2$ or [He] beryllium cation

H^- $1s^2$ or [He] hydride anion

Formulas of Ionic Compounds:

- positive and negative charges must balance.
- only one combination carries the correct simplest ratio of whole number.

⇒ $NaCl \Rightarrow Na^+ \; Cl^- \Rightarrow (+1)+(-1)=0$

$CaO \Rightarrow Ca^{2+} \; O^{2-} \Rightarrow (+2)+(-2)=0$

$K_2O \Rightarrow K^+ \; O^{2-} \Rightarrow 2(+1)+(-2)=0$

$Mg_3N_2 \Rightarrow Mg^{2+} \; N^{3-} \Rightarrow 3(+2)+2(-3)=0$

Ions are packed into a 3 dimensional array called a crystal lattice.

⇒ Halite ⇒ $NaCl$ (sodium chloride)

⇒ Wurtzite ⇒ ZnS (zink sulfide)

⇒ Calcite ⇒ $CaCO_3$ (calcium carbonate)

⇒ Muscovite (white mica) ⇒ $KAl_3Si_3O_{10}(OH)_2$

Lattice energy is affected by ionic size and ionic charge.

⇒ ionic size increases, lattice energy decreases. ↓decrease

ionic charge increases, lattice energy increases.

Problem - Slide Page 24

⊛ Properties of Ionic Compounds:

- Ionic compounds tend to be hard, rigid, and brittle, with high melting points.

ions are fixed

- Do not conduct electricity in the solid state.

- Conduct electricity ~~in~~ when melted or dissolved.

ions are free to move

⊛ Interionic attractions are so strong that when an ionic compound is vaporized, ion pairs are formed.

⊛ Crossover Method:

$K^+ \quad S^{2-} \Rightarrow K_2S$

$Ca^{2+} \quad CO_3^{2-} \Rightarrow Ca_2(CO_3)_2$

⊛ ~~What~~ What is the formula for sodium oxide? ($Na + O$)

$Na^+ \quad O^{2-}$

Na_2O

⊛ What is the formula for barium hydroxide?

$Ba \qquad (OH) \Rightarrow ve \Rightarrow 6+1=7$ (1 short)

$Ba^{2+} \qquad (OH)^{1-}$

$\Rightarrow Ba(OH)_2$

⊛ Ionic compound name:

cation name space anion name

$Cu_2SO_3 \Rightarrow$ Cuprous sulfite

$Al(HSO_3)_3 \Rightarrow$ Aluminum (III) hydrogen sulfite

Relation between Ionic compounds & covalent compounds:

Slide Page - 39

Bonding Pairs:

- atom share electrons to achieve a full outer level of electrons. The shared electrons are called a shared pair or bonding pair. H:H or H−H

Lone Pairs:

- an outer-level electron pair that is not involved in bonding is called a lone pair, or unshared pair.

:F̤̈:F̤̈: or :F̤̈−F̤̈:

Properties of a Covalent Bond:

- The $\boxed{\text{bond order}}$ is the number of electron pairs being shared by a given pair of atoms.
  A single bond consists of one bonding pair and has a bond order of 1
- The $\boxed{\text{bond energy}}$ is the energy needed to overcome the attraction between the nuclei and the shared electrons. The stronger the bond the higher the bond energy.
- The $\boxed{\text{bond length}}$ is the distance between the nuclei of the bonded atoms.

⊛ For a given pair of atoms,

- bond order ↑, bond length ↓, bond energy ↑

⊛ A shorter bond is a stronger bond

⊛ Bond length:

decrease →

increase ↓

⊛ Bond energy:

increase →

decrease ↓

⊛ Internuclear distance ⇒ bond length

Covalent radius ⇒ half of the bond length.

Problem, slide- 46

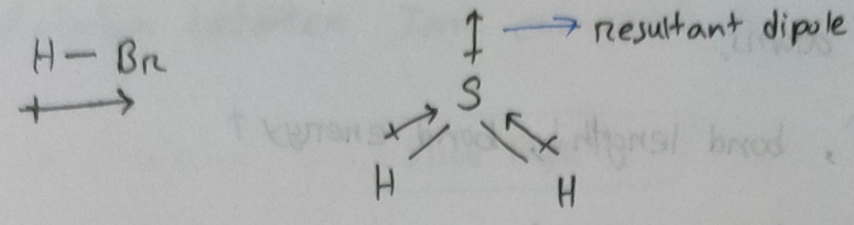
⊛ Three types of forces are known to exist between neutral molecules.

(i) Dispersion force (London Force):
- weak, but increases with increasing molar mass
- present in all atoms and molecule.

(ii) Dipole- Dipole Forces:
- Moderate strength
- present in only polar molecules

(iii) Hydrogen Forces:
- Strong
- present in molecules containing H bonded directly to F, O or N.

Summary on slide - 52



Both are polar molecules.

A species can form hydrogen bonds with water if it contains one of the three electronegative elements (F, O, N) or it has a H atom bonded to one of these three elements.

Relation between bonding
Page - 56

Bond energy and ΔH°

⇒ The heat released or absorbed during a chemical change is due to differentes between the bond energies of reactants and products.

$\Delta H = \Sigma \Delta H$ of reactant bonds broken $-$ $\Sigma \Delta H$ product bonds formed

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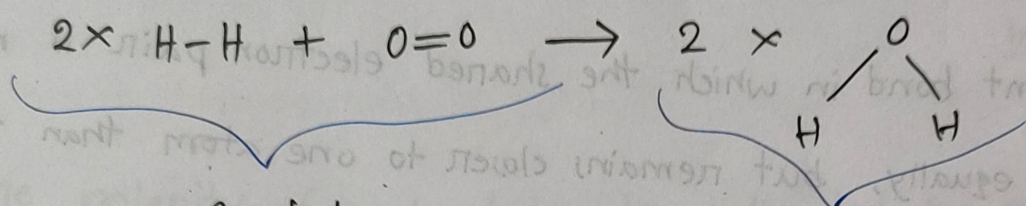
Bond energy:

$$2H_2 + O_2 \rightarrow 2H_2O$$

2 : 1



Reactant

$= 2\times(H-H) + 1\times(O=O)$

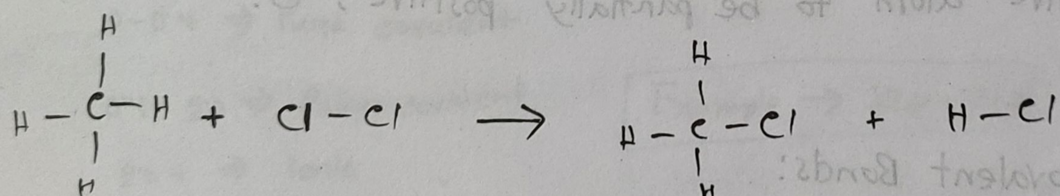
Product $= 2\times 2\times(O-H)$

$\therefore \Delta H^\circ =$ reactant − product

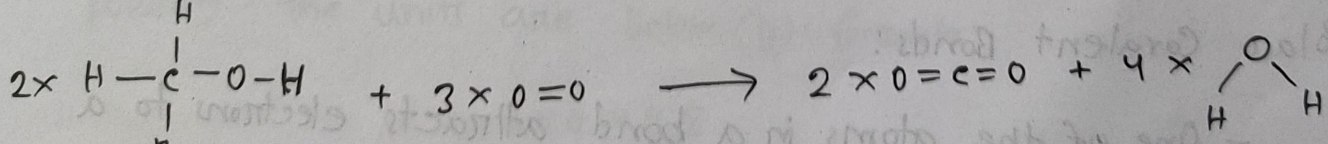
$= \pm \Rightarrow + \Rightarrow$ Endo $\Rightarrow$ absorved

$- \Rightarrow$ Exo $\Rightarrow$ release

* $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$



* $2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$



* Electron equal sharing ⇒ Non-Polar

* μ ⇒ 0-0.5 ⇒ Non Polar

  0.5-2 ⇒ Polar

  >2 ⇒ Ionic

* Electronegativity and Bond Polarity:

⇒ A covalent bond in which the shared electron pair is not shared equally, but remains closer to one atom than the other, is a polar covalent bond.

⇒ The ability of an atom in a covalent bond to attract the shared electron pair is called its electronegativity.

⇒ Unequal sharing of electrons causes the more electronegative atom of the bond to be partially negative, δ^- and the less electronegative atom to be partially positive, δ^+.

* Non-Polar Covalent Bonds:

- Equal sharing of electrons between bonded atoms.

* Polar Covalent Bonds:

- One of the atoms in a bond attracts electrons to a greater extent than the other resulting in an unequal sharing of electron density between the atoms

* Ionic Bonds:
  - No sharing of electrons
  - An electron is transferred from a metal atom to a non-metal atom.

* The polarity of covalent bonds is determined by electronegativity.

increase

increase

* The polarity of a bond can be estimated by calculating the difference in electronegativity between the two atoms.

* In polar covalent bonds the more electronegative atom has a partial negative charge (δ^-)

* The less electronegative atom has a partial positive charge (δ^+)

* Higher electronegativity, increasing ionic character.

  0–0.4 ⇒ Pure covalent

  0.4–2.0 ⇒ Polar covalent $\boxed{\text{Example} \rightarrow \text{Slide-7172}}$

  2.0 + ⇒ ionic

* Dipole moment (μ) is a quantitative measure of molecular polarity - the units are Debye (D) formula:

$$\mu = \delta \times d \times 4.8$$

δ: electron charge

d: in angstroms

⊛ Nonpolar covalent bond ⇒ electron are shared equally.

Polar covalent bond ⇒ electron are shared unequally.

Ionic bonding ⇒ electrons are transferred.

⊛ Diatomic molecules containing atoms of different elements have dipole moments → Polar molecule.

⊛ Diatomic molecules containing atoms of the same elements do not have dipole moments → nonpolar molecules.

$\boxed{\text{Slide - Page - 85}}$

⊛ Two identical polar bonds pointing in opposite directions will cancel. The molecule is nonpolar.

⊛ Three identical polar bonds at 120° from each other will cancel. Nonpolar.

⊛ Three polar bonds in a trigonal pyramidal arrangement (109.5°) will not cancel. Polar Molecule.

⊛ Two polar bond with an angle of less than 180° between them will not cancel. Polar Molecule.

⊛ Four identical with 109.5° will cancel. Nonpolar.

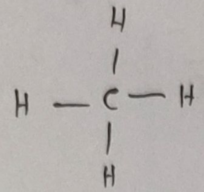
⊛

$\underset{\delta^+}{+\!\!\!\longrightarrow}\ \delta^-$

$\boxed{\text{Slide 90, 91, 89}}$

* CH_4

V.E = 4×1 + 4 = 8

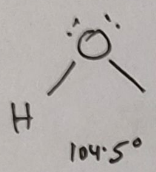


E.S. ⇒ Tetrahedral

M.S. ⇒ Tetrahedral

* H_2O

V.E. = 2×1 + 1×6 = 8

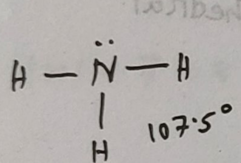


E.S. ⇒ Tetrahedral

M.S. ⇒ Bent (V-shape)

* NH_3

V.E. = 1×5 + 1×3 = 8

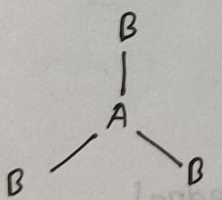


E.S. = Tetrahedral

M.S. = Trigonal pyramidal

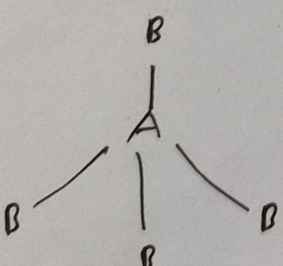
* B — A — B ⇒ Linear ⇒ 180°

*

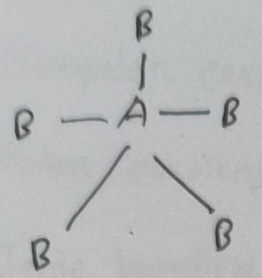


⇒ Trigonal planer ⇒ 120°

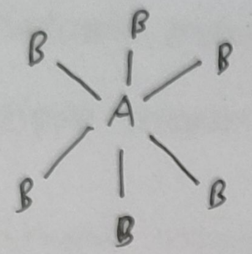
*



⇒ Tetrahedral



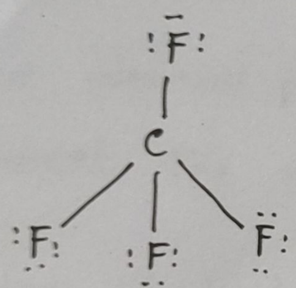
$\Rightarrow$ Trigonal bipyramidal



$\Rightarrow$ Octal

CF_4

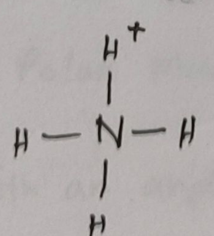
$V.E = 4 + 4 \times 7 = 32$



M.S. = E.S. = Tetrahedral

$NH_4^+ = 1 \times 5 + 4 \times 1 - 1$

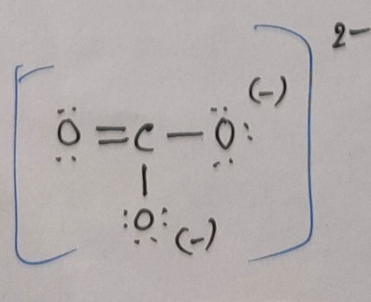
$= 8$



M.S = E.S = Tetrahedral

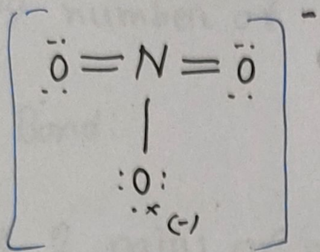
$CO_3^{2-} = 1 \times 4 + 3 \times 6 + 2$

$= 24$

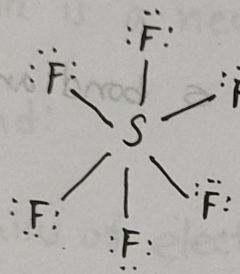


$NO_3^- = 1\times5 + 3\times6 + 1$

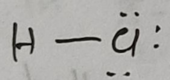
$= 24$



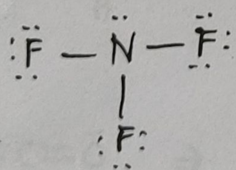
$SF_6 = 6 + 6\times7 = 48$



$HCl = 1\times1 + 7 = 8$



$NF_3 = 5 + 3\times7 = 26$



* Sigma bonds :

$\sigma \Rightarrow$ single bonds are sigma bond

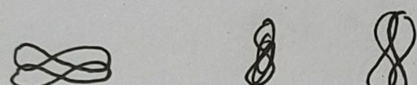
s-s　　s-p　　p-p

overlap

* Pi bonds :

$\pi \Rightarrow$ side ways overlap of p orbitals.

$\Rightarrow$ multiple bonds consist of one σ bond and the additional bond are π bonds.



p-p　　p-p

* Electron associated with π bonds are called π electrons.

* A double bond consist of one σ bond and one π bonds. $\rightarrow$ 2 π electron

* $\equiv$ (Triple bond) $\Rightarrow$ one σ bond
  two π bonds
  $\rightarrow$ 4 π e^-

* Single bond:
  - two atom sharing one pair of electrons.

* Double bond:
  - two atom sharing two pair of electrons.

* Tripple bond:
  - two atom sharing three pair of electrons.

Bond length:

- the distance between the atoms in a bond decneases with increasing number of shaned electron pairs.

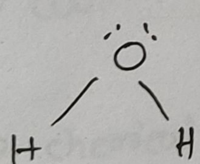
The Double Bond:

- There are 2 pairs of electrons shared
- The first pair is in a σ-bond
- Second pair is a new type of covalent bond (Pi-bond).

The triple bond:

- three pairs of electrons shared
- 1^{st} pair is in σ-bond
- 2^{nd} pair is in π-bond
- 3^{rd} pair is in 2^{nd} π-bond

$:N \equiv N:$ (π_1, σ, π_2)

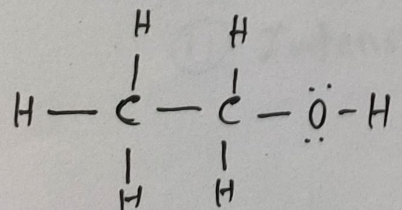
H_2O ⇒  ⇒ σ-bonds = 2

lone pain = 2

CO_2 ⇒ $\ddot{O}=C=\ddot{O}$ ⇒ σ-bonds = 2

π-bonds = 2

lone pain = 4

 ⇒ σ-bonds = 8

lone pairs = 2

⊛ H—C≡N: ⇒ σ- bonds = 2

π- bonds = 2

lone pair = 1

$\boxed{\sigma \text{ Vs } \pi \Rightarrow \text{Slide-113}}$

- A dative covalent bond is formed when an empty orbital of one atom overlaps with an orbital containing a non-bonding pair (lone pair) or electrons of another atom.

- Metallic Bonding:
  - Metallic bond is kind of force that holds atoms together in a metallic substance. In most cases, the outermost electron shell of each of the metal atoms overlaps with a large number of neighboring atoms.

- Alloys:
  - The substances are melted together and mixed then allowed to cool. This is called a alloy.
  - No chemical reaction. Just a mixing of two metals.
  - Two types of alloy:
    1. Substitution alloy ⇒ made from elements of similar chemical properties
       ↳ Rose Gold
    2. Interstitial alloy ⇒ a small proportion of a smaller atoms is added to a metal.
       ↳ Carbon + iron = steel.

* White gold:
  - gold + at least one white metal, silver, nickel ~~or~~ or ~~pla~~ palladium.

* Rose gold:
  - gold + copper
  - used in ~~jw~~ jewellery due to its reddish colour.

* Electrical transmission wires are made from an aluminium alloy.

* Structure of alloy:
  - changes the regular arrangement of the layers of atoms in the metal.
  - slipping of layers of atoms becomes more difficult.
  - Harder and stronger.

* Brass - an alloy of copper and zinc

* Hardening Metals:
  - the way a metal is prepared can have a large impact on how it behaves.
  - The rate at which it cools can have significant effect on the properties of the solid.

❋ There are three ways of treating a metal with heat.

- Annealing
- Quenching
- Tempering

Properties & Relation slide Page-134

Chapter-10

❋ Chemical bonding or Lewis structure:

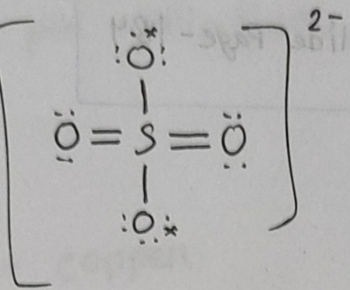
⇒ Molecular formula ⟶ Atom placement ⟶ Sum of valence e^-

- Atom placement ⇒ place atom with lowest EN in center.
- Sum of valence e^-
- Draw single bond
- Remaining valence e^-
- Lewis structure.

❋ Resonance structure ⇒ show ~~avaiabt~~ all possible combination for double bond.

* $\cancel{NF_3}$ $SO_4^{2-} = 6 + 4\times6 + 2$

$= 32$



* $O_2 \Rightarrow V.E = 12$

$\ddot{O}=\ddot{O}$

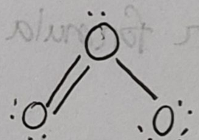
* $F_2 \Rightarrow 14$

$:\ddot{F}-\ddot{F}:$

* $N_2 \Rightarrow 10$

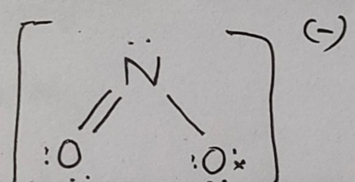
$\ddot{N}\equiv\ddot{N}$

* $O_3 \Rightarrow 18$



* $NO_2^- \Rightarrow 5+12$

$=17$



* bond order $= \frac{1}{2}$ (# electron in bonding MOs − # electron in antibonding MOs)

Molecular Orbital

⊛ N_2

$N = 1s^2\ \underline{2s^2 2p^3}$ → Valence shell

$\overline{\pi^*_{2p}}$ $\overline{\pi^*_{2p}}$

$\overline{\sigma^*_{2p}}$

$\frac{1\!\downarrow}{\sigma^*_{2s}}$

$\frac{1\!\downarrow}{2s}$ $\frac{1\!\downarrow}{2s}$

$2\ \frac{1}{p_x}\ \frac{1}{p_y}\ \frac{1}{p_z}$ $2\ \frac{1}{p_x}\ \frac{1}{p_y}\ \frac{1}{p_z}$

$\frac{1\!\downarrow}{\sigma_{2s}}$

N	N_2	N

$\frac{1\!\downarrow}{\pi_{2p}}$ $\frac{1\!\downarrow}{\pi_{2p}}$

$\frac{1\!\downarrow}{\sigma_{2p}}$

N	N_2	N

(+) (+)

$\therefore$ Bond Order $= \frac{1}{2}(8-2)$

$= \frac{1}{2}\cdot 6$

$= 3$ Ans

⊛ O_2^-

$O(8) = 1s^2\ \underline{2s^2 2p^4}$

$O^-(9) = 1s^2\ \underline{2s^2 2p^5}$

$\frac{1\!\downarrow}{\pi^*_{2p}}$ $\frac{1}{\pi^*_{2p}}$

$\overline{\sigma^*_{2p}}$

$\frac{1\!\downarrow\ 1\ \ 1}{2p}$ $\frac{1\!\downarrow\ 1\!\downarrow\ 1}{2p}$

$\frac{1\!\downarrow}{\pi_{2p}}$ $\frac{1\!\downarrow}{\pi_{2p}}$

$\frac{1\!\downarrow}{\sigma_{2p}}$

$\frac{1\!\downarrow}{\sigma^*_{2s}}$

$\frac{1\!\downarrow}{2s}$ $\frac{1\!\downarrow}{2s}$

$\frac{1\!\downarrow}{\sigma_{2s}}$

O	O_2^-	O^-

$\therefore$ Bond Order $= \frac{1}{2}(8-5) = \frac{1}{2}\cdot 3 = 1.5$

* Bond order > 0

⇒ the molecule is more stable than the separate atoms so it will form.

* Bond order $= 0$

⇒ the molecule is as stable as the separate atoms so it will not form (occurs when equal numbers of electrons occupy bonding and antibonding MOs)

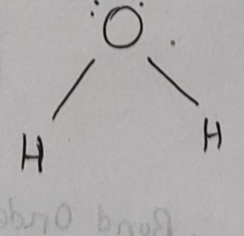
* The higher the bond order, the bond is stronger.

Example, slide-22

* Formal Charge :

FC = # valence electron of central atom − non bonding electrons − $\frac{1}{2}$ · shared electron

* $H_2O = 2 + 6 = 8$



$$FC = 6 - 4 - \frac{1}{2} \cdot 4$$

$$= 6 - 4 - 2$$

$$= 0$$

✱ Formal charge is the charge an atom would have if all electrons were shared equally.

✱ All molecular shape are derived from the five basic geometries :

- Linear ⇒ 180° ⇒ 2 atom + 1
- Trigonal planar ⇒ 120° ⇒ 3 atom + 1
- Tetrahedral ⇒ 109.5° ⇒ 4 atom + 1
- Trigonal bipyramidal ⇒ 120°, 90° ⇒ 5 atom + 1
- Octahedral ⇒ 90° ⇒ 6 atom + 1

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* Molecular shapes with two electron groups

⇒ E.S. ⇒ Linear $\ddot{O}=C=\ddot{O}$ 180°

M.S. ⇒ Linear

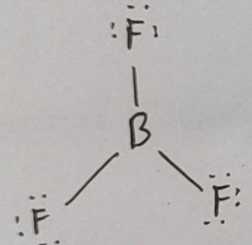
* Molecular shapes with three electron groups:

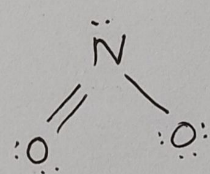
⇒ E.S. ⇒ Trigonal planar

M.S. ⇒ 2 possibilities

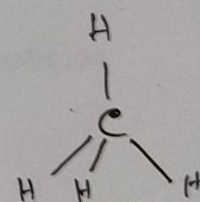
0 Lone Pair ⇒ Trigonal planar ⇒ 120°

1 Lone Pair ⇒ Bent/V-shaped ⇒ <120°





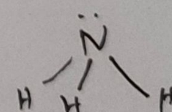
* Molecular shapes with four electron groups:



⇒ E.S. ⇒ Tetrahedral

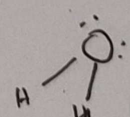
M.S. ⇒ 3 possibilities

0 Lone Pair ⇒ Tetrahedral ⇒ 109.5°



1 Lone Pair ⇒ Trigonal planar ⇒ <109.5°

2 Lone Pair ⇒ Bent ⇒ <109.5°



Moleculan shapes with five electron groups:

⇒ E.S. ⇒ Trigonal ~~p~~ bipyramidal

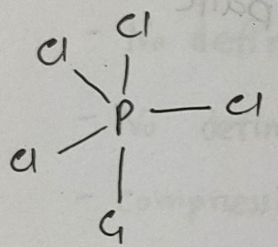
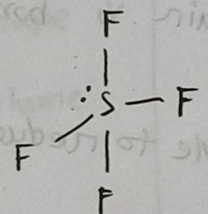
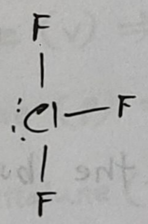
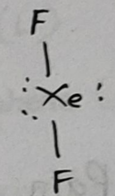
M.S. ⇒ 4 possibilities

0 Lone Pair ⇒ Trigonal bipyramidal ⇒ $90°$ & $120°$

1 Lone Pair ⇒ Seesaw ⇒ $<90°$ & $<120°$

2 Lone Pair ⇒ T-shaped ⇒ $<90°$

3 Lone Pair ⇒ Linear ⇒ $180°$

*Adjust the electron

Moleculan shapes with six electron groups:

⇒ E.S. ⇒ Octahedral

M.S. ⇒ 5 possibilities

0 Lone Pair ⇒ Octahedral ⇒ $90°$

1 Lone Pair ⇒ Square pyramidal ⇒ $<90°$

2 Lone Pair ⇒ Square planar ⇒ $90°$

3 Lone Pair ⇒ T-shaped ⇒ $90°$

4 Lone Pair ⇒ Linear ⇒ $180°$

Example- Slide- 31-33

Problem - 35 - 38

❀ VSEPR ⇒ Valence shell electron pair repulsion theory:

- the shape of a molecule or ion is caused by repulsion between the pairs of electrons, both bond pairs and lone pairs, that ~~so~~ surround the central atom.

⇒ Lone-pair Vs Lone Pair > lone pair Vs bonding pair > bonding pair Vs bonding pair

❀ Lone Pair caused the bonding angle to reduce.

Practice - Slide 41-47

Next Friday - mid
Chapter - 9, 10
11:00 am

L-9/08.01.2024/

Chapter-5

Gases

⊛ State of Matter:

- Solid
- Liquid
- Gas

⊛ Properties of gases

- No definite shape
- No definite volume
- compressible

⊛ Physical Characteristic

- Volume (v) ⇒ liters (L)
- Pressure (P) ⇒ atmosphere (atm)
- Temperature (T) ⇒ kelvin (k)
- Number of atom (n) ⇒ mole/mol

1 mole = $6{\cdot}022 \times 10^{23}$ atoms

⊛ Pressure:

$$P = \frac{\text{force}}{\text{area}}$$

SI units newton/meter^2 = 1 Pa

1 atm = 760 mmHg = 760 ~~torr~~ Torr = 101326 Pa = 1·01325 bar = 14·7 psi

⊛ Manometer

- Device used for measuring the pressure of a gas in a container.

Barometer

- Device used to measure atmospheric pressure.

$\boxed{\text{Problem - Page-6}}$

Temperature Conversion:

$K = C + 273$

$F = 1.80\,C + 32$

Boyle's Law:

- the pressure of a fixed amount of gas at a constant temperature is inversely proportional to the volume of the gas.

$V \propto \frac{1}{P}$

$PV = k$

$\boxed{P_1V_1 = P_2V_2}$ [T, n constant]

Problem - Slide - 10,11

Charles's Law:

- the volume of a fixed amount of gas maintained at constant pressure is directly proportional to its absolute temperature.

$V \propto T$

$\frac{V}{T} = k$

$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$ [P, n constant]

Problem - Slide - 13,14

Avogadro's Law:

- the volume of a sample of gas is directly proportional to the number of moles in the sample at constant temperature and pressure.

$V \propto n$

$$\boxed{\frac{V_1}{n_1} = \frac{V_2}{n_2}} \quad [T, P \text{ constant}]$$

⇒ Avogadro's number $= 6.02 \times 10^{23}$

⇒ At STP, 6.02×10^{23} particles of a gas occupies 22.4 L.

Gay-Lussac Law:

- At constant volume, pressure and absolute temperature are directly related.

$P \propto T$

$P = kT$

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}} \quad [V \text{ constant}]$$

The combined gas law can be used to solve problems where any or all of the variable changes.

$$P_1V_1 = P_2V_2$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\boxed{\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}}$$

$$\Rightarrow \boxed{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}$$

Problem - Slide - 17.

⊛ The ideal gas

$$V \propto \frac{1}{P}, \quad V \propto T, \quad V \propto n$$

$$V \propto \frac{nT}{P}$$

$$V = R\frac{nT}{P}$$

$$\boxed{PV = nRT}$$

gas constant

- $0.0821 \ \frac{L \cdot atm}{mol \cdot K}$
- $8.314 \ \frac{L \cdot kPa}{mol \cdot K}$

Units need to be in m^3, K, and Pa

Problem - Slide - 20 - 24

L-10 / 10.01.2024 /

Molar Mass of a Gas

$$d = \frac{m}{V}$$

$$n = \frac{m}{M}$$

$$PV = nRT = \frac{m}{M} \cdot RT$$

$$M = \frac{m}{V} \frac{RT}{P} = \boxed{d \frac{RT}{P}} \quad g/mol$$

$\boxed{d = \text{density}}$

Composition of Dry Air:

- Nitrogen (N_2) ⇒ 78%
- Oxygen (O_2) ⇒ 21%
- Argon (Ar) ⇒ 0.9%
- Carbon dioxide (CO_2) ⇒ 0.04%

Partial Pressure:

- the preassure of a single gas in a mixture of gases.

Total Pressure:

- the sum of the partial pressure of all the gases in the mixture equals the total pressure.

$$\boxed{P_t = n_t \left(\frac{RT}{V}\right)} \quad t = 1, 2, 3, \cdots$$

Problem- Slide - 29, 30

Kinetic Molecular Theory:

- The particles of the gas are constantly moving.
- The attraction between particles is negligible.
- When the moving gas particles hit another gas particles on the container, they do not stick; but they bounce off and continue moving in another direction.
  - Like billiard balls
- There is a lot of empty space between the gas particles compared to the size of particles.

Diffusion:

- the mixing of gases.

Effusion:

- describes the passage of a gas through a tiny orifice into an evacuated chamber.

Graham's law of effusion:

- the rate of effusion of a gas is inversely proportional to the square root of its molar mass.

Rate of effusion $\propto \sqrt{\frac{1}{M}}$

Real Gas:

- An ideal gas is a hypothetical concept. No gas exactly follows the ideal gas law.
- We must correct for non-ideal gas behavior when,
  - pressure of the gas is high
  - Temperature is low.

$$\boxed{\frac{PV}{nRT} = 1}$$

- Under these condition,
  - Concentration of gas particles is high
  - Attractive forces become important.

Kinetic Molecular Theory model assumed

- no interactions between gas particles
- no volume for the gas particles

1873, Johannes van der Waals

- Correction for attractive forces in gases and liquids
- Correction for volume of the molecules

* The van der Waals equation adjust the ideal gas law to take into account
- the real volume of the gas particles
- the effect of interparticle attraction

$$\underbrace{\left(P + \frac{n^2 a}{V^2}\right)}_{\text{Connected Pressure term}} \underbrace{(V - nb)}_{\text{Connected Volume term}} = nRT$$

$\rightarrow$ factors that influence the attraction between particles

$\rightarrow$ relates to particle volume

Real Gas vs Ideal Gas

Slide - 38 - 40

Quiz - 3

Chapter - 5

15. 01. 2024