

LECTURE NOTES

ON

CHM 211: Basic Inorganic Chemistry

DEPARTMENT OF CHEMISTRY,
UNIVERSITY OF AGRICULTURE, ABEOKUTA(UNAAB)

LECTURERS IN CHARGE

Dr.(Mrs) C.A. Akinremi and Dr. S.A. Amolegbe

COURSE DESCRIPTION

CHM 211: Basic Inorganic Chemistry

2units

Wave mechanical treatment of atomic structure, Periodicity and periodic table.

Chemical bonding. Bonding theories; Valency bond theory; Molecular orbital theories.

Inorganic stereochemistry, Nomenclature of Inorganic compounds. Chemistry of group

IIIA (boron group), VA (Nitrogen group), noble gas. Introduction to first row

transition metal chemistry.

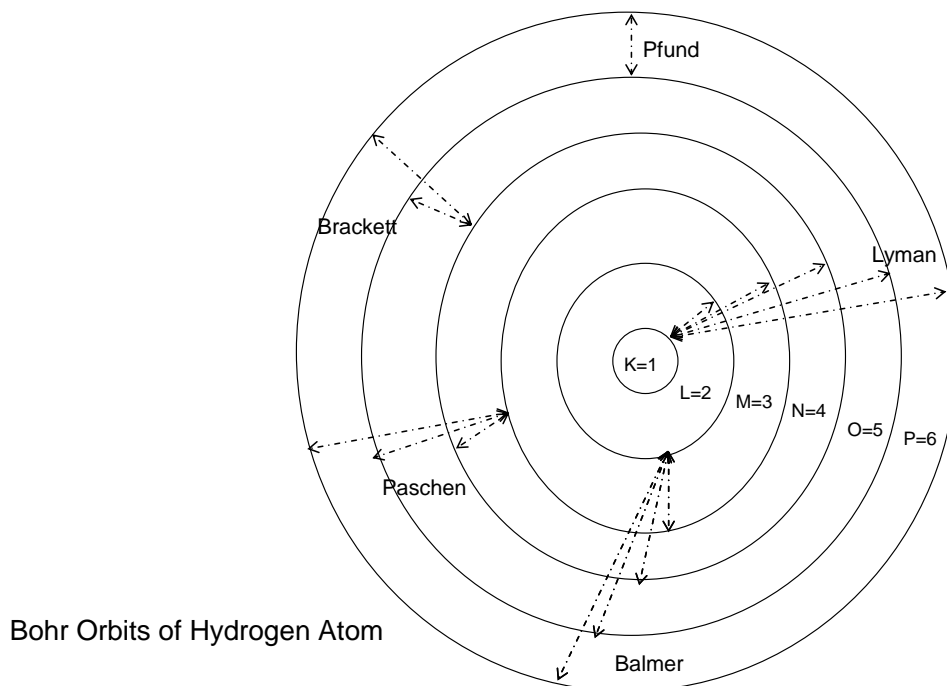
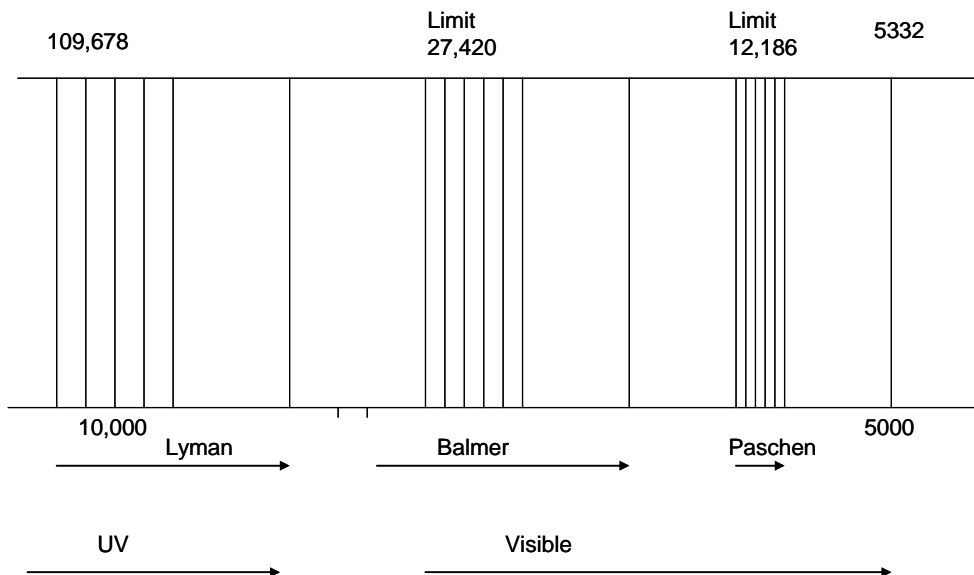
Wave Mechanical Treatment of Atomic Structure

Atomic Structure

Atomic structure which includes the number and arrangement of the electrons governs its chemical and physical properties that is the possible distribution of the electrons about the nucleus, the energies of the electrons and their magnetic properties etc.

Simplest Atom- Hydrogen Atom

When an electrical discharge is passed through hydrogen at low pressure, a pink-coloured glow is observed. That is, the excited hydrogen atom emits light of definite frequencies. If this is examined by a simple spectroscope, a few sharp coloured lines are observed. The lines of spectrum are named after the men who discovered them.



The spectrum of atomic hydrogen

Lines	n_i	n_f	Region
Lyman series	$n_i=1$	$n_f=2,3,4$ etc	In the Ultraviolet region
Balmer series	$n_i=2$	$n_f=3,4,5$ etc	In the Visibl region
Paschen series	$n_i=3$	$n_f=4,5,6$ etc	In the Infrared region
Brackett series	$n_i=4$	$n_f=5,6,7$ etc	In the far Infrared region
Pfund series	$n_i=5$	$n_f=6,7,8$ etc	In the far Infrared region

When the lines corresponding to a particular wavelength are examined, it is seen that they all fit the equation:

$$\frac{1}{\lambda} = R \left[\frac{1}{n_i^2} - \frac{1}{n_h^2} \right]$$

Where λ = wavelength of the particular line

R is Rydberg's constant = $1.097 \times 10^7 \text{ m}^{-1} = 1.097 \times 10^5 \text{ cm}^{-1}$ or $109,677 \text{ cm}^{-1}$,

n_h = quantum number describing the higher energy state

n_i = quantum number describing the lower energy state

(i.e. n_i and n_h are integers where n_h is greater than n_i)

$$\frac{1}{\lambda} = \bar{\nu} \text{ - Wave number}$$

$$\bar{\nu} = R \left[\frac{1}{n_i^2} - \frac{1}{n_h^2} \right]$$

1. Lyman series

$$\bar{\nu} = R \left[\frac{1}{1^2} - \frac{1}{n_h^2} \right] \text{ where } n_h = 2,3,4,5$$

2. Balmer series

$$\bar{\nu} = R \left[\frac{1}{2^2} - \frac{1}{n_h^2} \right] \text{ where } n_h = 3,4,5$$

3. Paschen series

$$\bar{\nu} = R \left[\frac{1}{3^2} - \frac{1}{n_h^2} \right] \text{ where } n_h = 4,5,6$$

4. Brackett series

$$\bar{\nu} = R \left[\frac{1}{4^2} - \frac{1}{n_h^2} \right] \text{ where } n_h = 5,6,7$$

5. Pfund

$$\bar{\nu} = R \left[\frac{1}{5^2} - \frac{1}{n_h^2} \right] \text{ where } n_h = 6, 7, 8$$

Bohr's Theory of atomic spectrum of hydrogen

In 1911, Rutherford provided the idea of an atom consisting of a positively charged nucleus surrounded by electrons and Bohr used this model of an atom to explain the anomaly of the discrete internal energy.

In the year 1913, Bohr made the following assumptions;

1. Electrons rotated around the nucleus of an atom in definite circular orbit and that a definite orbit corresponded to a definite electronic energy
2. An electron in a circular orbit of a small radius would have a lower (that is more negative) energy than an electron in an orbit of large radius.
3. Only certain orbits were possible for electrons [explanation=I can be shown from classical mechanics(physics) that the energy of an electron in a circular orbit around a proton is solely mainly determined by the radius of the orbit.

Bohr imposed one condition so that he could derive the Rydberg formula; He stated that in each orbit, the angular momentum of the electron, mvr must be quantized that is, half a value given by;

$$mvr = \frac{nh}{2\pi} \text{ where } m, v, r \text{ and } h \text{ are mass, velocity of electron, radius of the orbit and Planck's}$$

constant respectively and n is called the quantum number of the orbit.

By using this assumption, it was determined that for each orbit, the energy and radius were given by

$$\Delta E = h\nu = \frac{hc}{\lambda} = hc\bar{\nu}$$

Where ΔE = difference in energy in the two levels

By knowing the mass and the charge of electrons, Bohr calculated the radii of the possible orbits for the hydrogen atom from this equation;

$$r_n = \frac{\epsilon_0 h^2 n^2}{\pi m_e e^2} \text{ (this expression lies on the centrifugal force acting on the electron)}$$

Where ϵ_0 = permittivity of a vacuum = $8.854 \times 10^{-12} \text{ Fm}^{-1}$

h = Planck's constant = $6.626 \times 10^{-34} \text{ Js}$

n = 1, 2, 3... describing a given

m_e = electron rest mass = $9.109 \times 10^{-31} \text{ kg}$

e = charge on an electron = $1.602 \times 10^{-19} \text{ C}$

NOTE

A general equation from all the spectral frequency can be calculated

$$E_n = -\frac{Rch}{n^2}$$

If electron is in a level designated n_H and falls into a lower level n_L , then energy emitted is

$$\Delta E = E_{n_L} - E_{n_H} = -\frac{Rch}{n_L^2} + \frac{Rch}{n_H^2}$$

$$= Rch \left[\frac{1}{n_L^2} - \frac{1}{n_H^2} \right]$$

$$\Delta E = h\nu = Rch \left[\frac{1}{n_L^2} - \frac{1}{n_H^2} \right]$$

$$\lambda = c/\nu, \therefore 1/\lambda = R \left[\frac{1}{n_L^2} - \frac{1}{n_H^2} \right]$$

This equation is precisely same as the first one arrived at empirically from the study of the spectral lines of atomic hydrogen.

The Concept of Duality Behaviour of Electrons

Light radiation can be considered either as a stream of photons or as a wave motion.

In 1924, de Broglie pointed out that a similar duality exists in the case of electrons. De Broglie showed by arrangement, based on the theory of relativity that if an electron of momentum P could be described by a wave theory, in terms of a wavelength λ then

$$\lambda = h/P \text{ or } h/mv$$

Thus λ can be expressed in terms of electron energy rather than momentum. A kinetic energy ($\frac{1}{2}mv^2$) of an electron mass m and velocity v is given by

$\frac{1}{2}mv^2 = E - V$ where E is total energy $V =$ potential energy thus

$$v^2 = 2 \left(\frac{E - V}{m} \right) \quad \text{and} \quad \lambda = \frac{h}{mv} = \frac{h}{m \sqrt{2 \left(\frac{E - V}{m} \right)}}$$

BONDING THEORIES; VALENCE BOND THEORY, MOLECULAR BOND THEORIES.

Lewis Structure

Lewis proposed that a covalent bond is formed when two neighboring atoms share one or more electron pair. Also, some molecules have non-bonding pairs of electrons on the atom, which are called lone pair. These electrons contribute to the shape and reactivity of the molecules (chemical properties) but do not directly bond the atoms together.

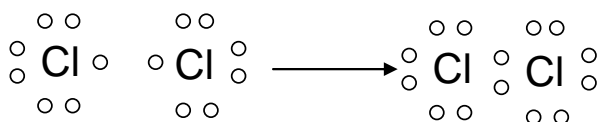
Octet Rule

For many light atoms following a Lewis structure, a stable arrangement is attained when the atom is surrounded by eight electrons (valence electrons corresponding to the s and p electrons outside the noble gas core –ie s^2p^6 configuration).

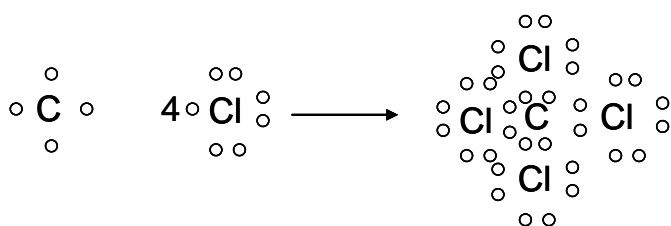
Lewis explained the existence of some molecules by proposing the Octet Rule which states that. "Each atom shares electrons with neighboring atoms to achieve a total of eight valence electrons (an octet)". This explains the observed valencies in a large number of cases. The atoms continue to form bonds until they have made up an octet of electrons.

Example:

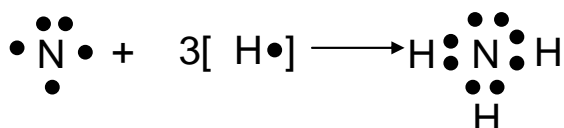
1. Chlorine: This atom has 7 electrons in its outer shell, so by sharing one electron with another Cl atom both atoms attain an octet and form a chlorine molecule (Cl_2).



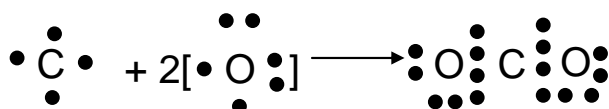
2. CCl_4 :



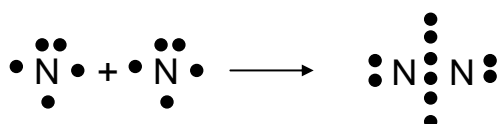
3. NH_3 :



4. CO_2 :



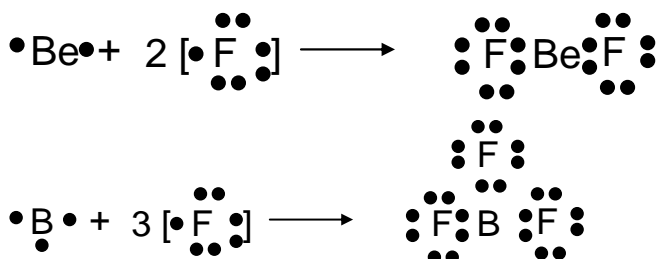
5. N_2 :



From these, it can be seen that four electrons are shared between two atoms giving rise to a double bond. 6 electrons shared gives rise to triple bonds.

Exceptions

1. Hydrogen is an exception which is stable with two valence electrons
2. Beryllium(2,2) and B(2,3)



Be and B have less than 4 outer electrons and even if all the outer electrons are used to form bonds, an octet cannot be attained.

3. When atoms have an extra energy level which is close in energy to the p level and may not accept electrons and can be used for bonding. Eg PF_3 obeys octet rule but PF_5 doesnot. [P=2,8,5]. This is because PF_5 has 10 outer electrons: one 3s, three 3p and one 3d orbitals.
4. For molecules with odd numbers of electrons, octet rule doesnot apply eg $\text{NO}=11$ electrons, $\text{ClO}_2=19$
5. Also compounds with more than four covalent bonds must break the octet rule. This becomes more obvious after the first two periods of 8 elements.

Valence Theory.

This theory was proposed by Linus Pauling. This theory considers the interaction of separate atoms as they are brought together to form a molecule. It is a way of expressing the lewis theory in terms of wave concepts. In VB theory, a bond between atoms A and B is formed when two atomic orbitals, one from each atom, merge with one another, and the electrons they contain pair up (so that their spins are $\downarrow\uparrow$). The merging of orbitals gives rise to constructive interference—*i.e.*, an enhancement of amplitude—between the wavefunctions in the areas where they overlap, and hence an enhanced amplitude results in the internuclear region. As a consequence of the formation of this region of heightened amplitude, there is an increased probability of finding the electrons in the internuclear region and, by implication, a lowering of the energy of the molecule.

In summary, a single bond in a Lewis structure corresponds to a σ bond of VB theory. A double bond corresponds to a σ bond plus a π bond, and a triple bond corresponds to a σ bond plus two π bonds.

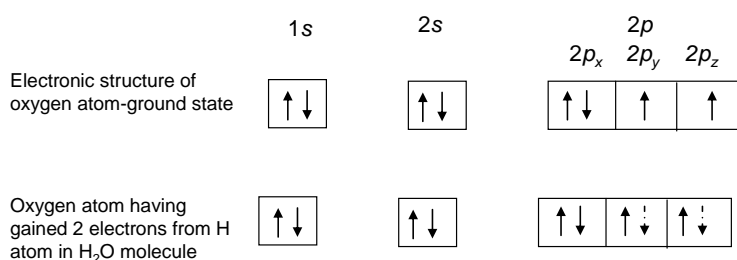
Illustration of Valence Bond theory

1. H₂



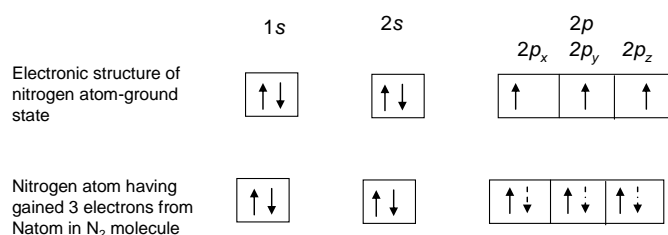
Here, the VB wavefunction is formed by spin pairing of the electrons in the two contributing atomic orbitals. This electron distribution is called a **σ bond**. This has a cylindrical symmetry around the internuclear axis, and the electrons in it have zero orbital angular momentum about that axis.

2. H₂O



First, note that the valence-shell electron configuration of an oxygen atom is $2s^2 2p_x^2 2p_y^1 2p_z^1$, with an unpaired electron in each of two $2p$ orbitals, and is the Lewis diagram for the atom. Each hydrogen atom has an unpaired $1s$ electron ($H\cdot$) that can pair with one of the unpaired oxygen $2p$ electrons. Hence, a bond can form by the pairing of each hydrogen electron with an oxygen electron and the overlap of the orbitals they occupy. The electron distribution arising from each overlap is cylindrically symmetrical around the respective O–H axis and is called a **σ bond**. The VB description of H₂O is therefore that each hydrogen atom is linked to the oxygen atom by a σ bond formed by pairing of a hydrogen $1s$ electron and an oxygen $2p$ electron.

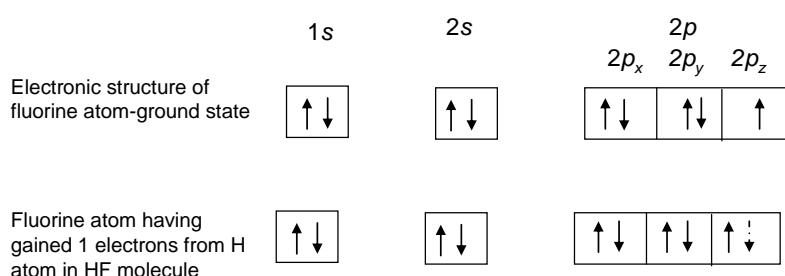
3. N₂



π bond

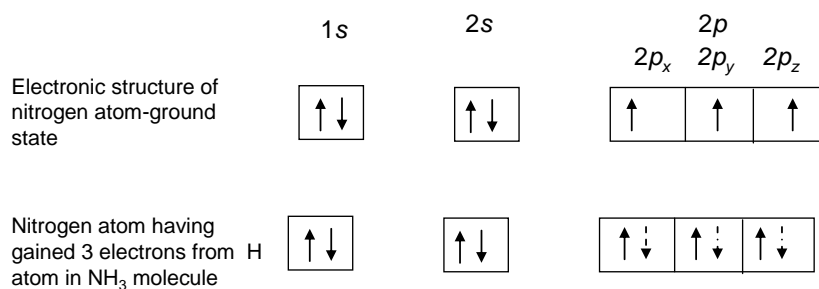
Each nitrogen atom has the valence-shell electron configuration $2s^2 2p_x^1 2p_y^1 2p_z^1$. If the z direction is taken to lie along the internuclear axis of the molecule, then the electrons in the two $2p_z$ orbitals can pair and overlap to form a σ bond. However, the $2p_x$ orbitals now lie in the wrong orientation for head-to-head overlap, and they overlap side-to-side instead. The resulting electron distribution is called a π bond. A **π bond** also helps to hold the two atoms together, but, because the region of maximum electron density produced by the overlap is off the line of the internuclear axis, it does not do so with the same strength as a σ bond. The $2p_y$ electrons can pair and overlap in the same way and give rise to a second π bond. Therefore, the structure of an N_2 molecule consists of one σ bond and two π bonds. Note how this corresponds to and refines the Lewis description of the $:N\equiv N:$ molecule.

4. HF



H has a singly occupied s orbital that overlaps with a singly filled $2p$ orbital on F

5. NH₃



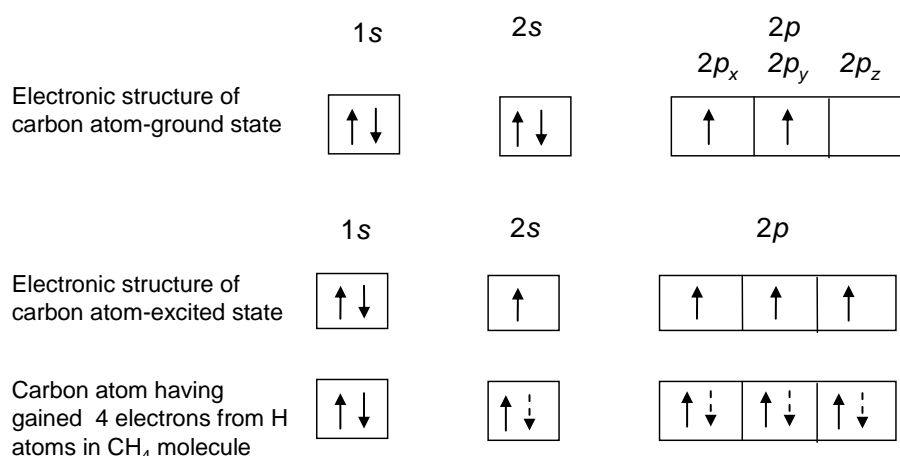
N has three singly filled $2p$ orbital each of which overlaps with a singly occupied s -orbital.

Promotion of electrons

Valence bond theory runs into an apparent difficulty with CH₄. The valence-shell electron configuration of carbon is $2s^2 2p_x^1 2p_y^1 2p_z^1$, which suggests that it can form only two bonds to hydrogen atoms, in which case carbon would have a valence of 2. The normal valence of carbon is 4, however. This difficulty is resolved by noting that only the overall energy of a molecule is important, and, as long as a process leads to a lowering of energy, it

can contribute even if an initial investment of energy is required. In this case, VB theory allows promotion to occur, in which an electron is elevated to a higher orbital. Thus, a carbon atom is envisaged as undergoing promotion to the valence configuration $2s^1 2p_x^1 2p_y^1 2p_z^1$ as a CH_4 molecule is formed. Although promotion requires energy, it enables the formation of four bonds, and overall there is a lowering of energy. Carbon is particularly suited to this promotion because the energy involved is not very great; hence the formation of tetravalent carbon compounds is the rule rather than the exception.

Promotion is the excitation of an electron to an orbital of higher energy in the course of the bond formation. Promotion of electrons may occur if the outcome is to achieve more or stronger bonds and a lower overall energy



Hybridization

Hybrid orbitals are formed when atomic orbitals on the same atom interfere.

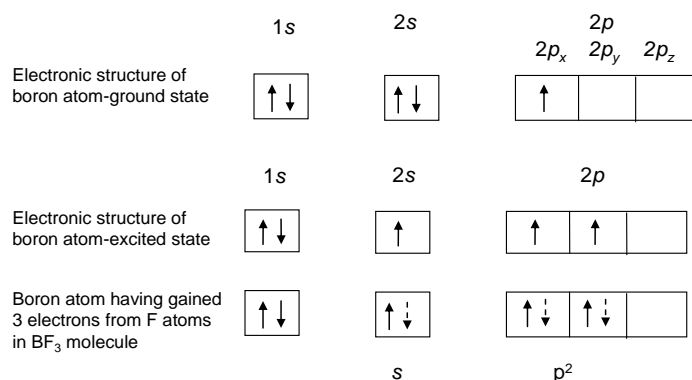
The discussion is not yet complete, however. From physical and chemical evidence, the 4 bonds in methane are equivalent. But from VB theory, 3 bonds are of one type (2p) and the fourth bond is of distinctly different character (2s.)

Quantum mechanical considerations resolve this dilemma by invoking hybridization. Hybridization is the mixing of atomic orbitals on the same atom or when atomic orbitals on the same atom interfere, giving rise to hybrid orbitals. For effective mixing, the energy of the orbitals must be the same.

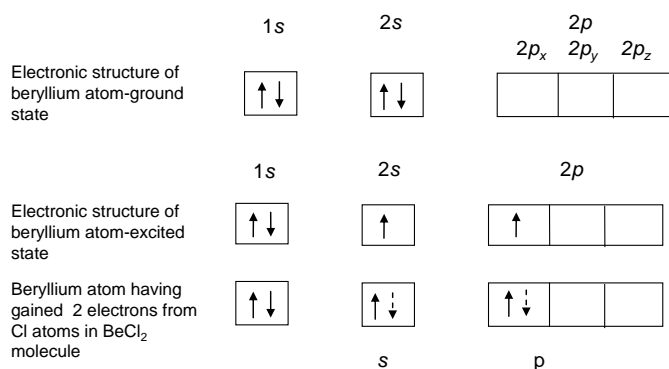
sp³ hybridisation: Each is built from one 2s and three 2p orbitals. It is called an sp^3 hybrid orbital. Since each sp^3 hybrid orbital has the same composition, all four bonds are identical apart from their orientation in space. Here, the electron density distribution in the promoted atom is equivalent to the electron density in which each electron occupies a hybrid orbital formed by the interference between 2s and 2p orbitals. Each hybrid orbital contains an unpaired electron and can form a σ bond by pairing with a 1s electron of a hydrogen atom. Hence, the VB structure of methane is described as consisting of four equivalent σ bonds

formed by overlap of the s orbitals of the hydrogen atoms with sp^3 hybrid orbitals of the carbon atom.

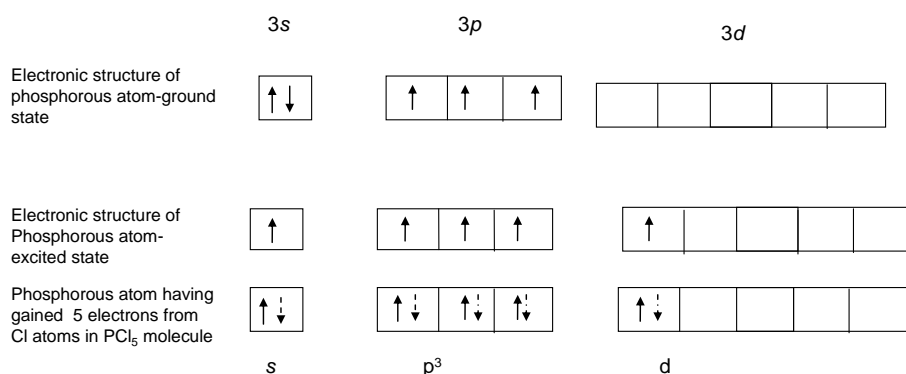
sp^2 hybridisation: The structure of ethylene can be examined in VB terms to illustrate the use of hybridization. To reproduce the Lewis structure given earlier, it is necessary to contrive a double bond (*i.e.*, a σ bond plus a π bond) between the two carbon atoms. Such a bonding pattern can be achieved by selecting the carbon 2s orbital, from which an electron has been promoted, and two of its 2p orbitals for hybridization, leaving one 2p orbital unhybridized and ready for forming a π bond. When one 2s and two 2p orbitals are hybridized, they form sp^2 hybrid orbitals, which have lobelike boundary surfaces that point to the corners of an equilateral triangle; the unhybridized 2p orbital lies perpendicular to the plane of the triangle. Each of the orbitals contains a single electron. Two of the hybrids can form σ bonds to two hydrogen atoms, and one of the hybrids can form a σ bond to the other carbon atom (which has undergone similar hybridization). The unhybridized 2p orbitals are now side-by-side and can overlap to form a π bond. This description conforms to the Lewis description. It also explains naturally why ethylene is a planar molecule, because twisting one end of the molecule relative to the other reduces the overlap between the 2p orbitals and hence weakens the π bond. All double bonds confer a torsional rigidity (a resistance to twisting) to the parts of molecules where they lie. BF_3 is another example:



sp hybridisation: This means one s atomic orbital and one p atomic orbital mixes to form a set of 2 hybrid orbitals with different directional properties. Eg $BeCl_2$ where $be = [He]2s^2$



Mixing d orbitals:



However, d orbitals are in general too large and have high energy to mix completely with s and p orbitals. The presence of electronegative atoms cause contraction of the d orbitals and lower their energy, so mixing is possible. Also, the number of d orbitals occupied by electrons affects the mixing, the more the number the shorter the distance. So d-orbitals probably do participate in bonding in cases where d-orbital contraction occurs.

Number of Outer orbitals	Type of Hybridisation	Arrangement
2	sp	linear
3	sp^2	Trigonal planar
4	sp^3	Tetrahedral
5	a. sp^3dz^2 b. $sp^3d(x^2-y^2)$	a. Trigonal bipyramid b. square based pyramid
6	sp^3d^2	Octahedron
7	sp^3d^3	Pentagonal bipyramid
[4	dsp^2	Square planar

Molecular Orbital Theory

This considers the ways in which electrons are influenced by the presence of two or more nuclei, the electrons in bonded atoms being said to occupy molecular orbitals.

Example H_2 molecule:

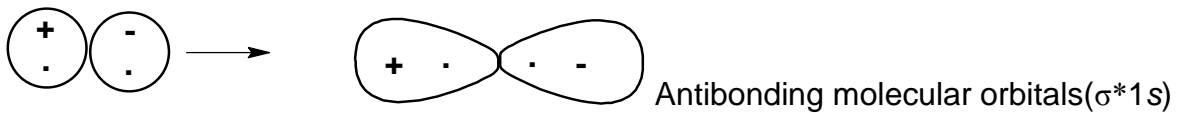
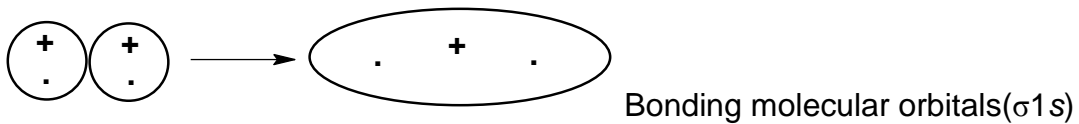
If 2 hydrogen atoms are first considered to be widely separated, then each electron is influenced by its own nucleus, each electron occupying a 1s atomic orbitals. As each atom come close to one another, they also come under the influence of two H nuclei and form a chemical bond. Under these conditions, the separate atomic orbitals can be said to merge and form molecular orbitals.

The number of possible molecular orbitals is equal to the number of the separate atomic orbitals in the separated atom. Hydrogen molecule has two atomic orbitals, hence two possible molecular orbitals. The 2 molecular orbitals are obtained by adding and subtracting the two atomic orbitals. This is called Linear combination of atomic orbitals abbreviated as **LCAO method**. The two possible molecular orbitals are represented by

$$\varphi_{\text{bonding}} = \varphi_{A.1s} + \varphi_{B.1s}$$

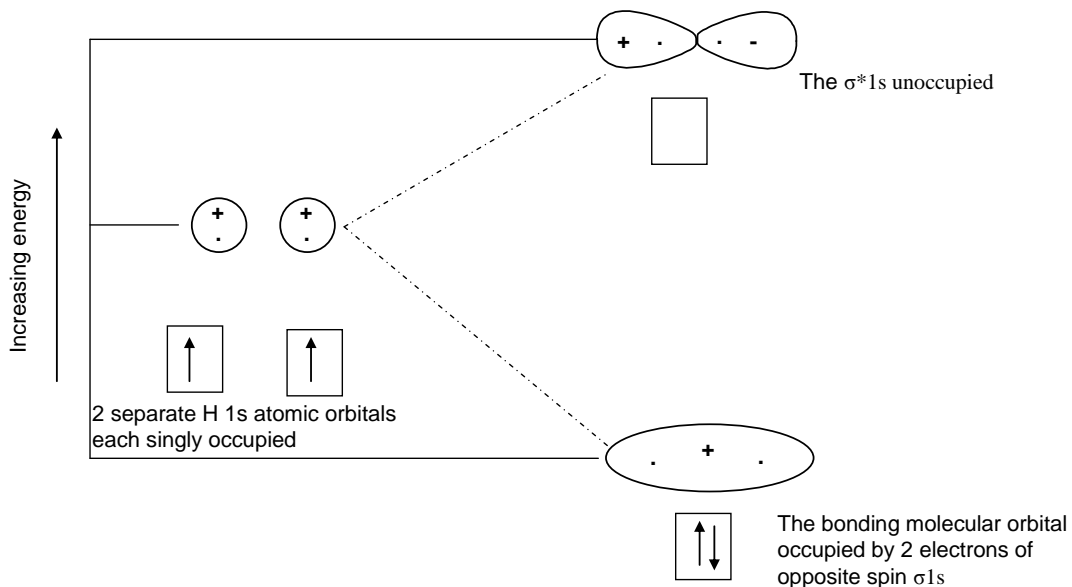
$$\varphi_{\text{antibonding}} = \varphi_{A.1s} - \varphi_{B.1s} \text{ where A and B are the 2 H atoms.}$$

A bonding molecular orbital is obtained by adding together the separate atomic orbitals and this results in the build-up of negative charge between the two nuclei. An antibonding molecular orbital is obtained by subtracting two atomic orbitals and this leads to a draining away of negative charge between the two nuclei.



A bonding molecular orbital is of lower energy than two separate atomic orbitals, while, antibonding molecular orbitals has a higher energy than the two separate atomic orbitals. Also, two electrons with opposite spins can occupy each molecular orbital-Pauli exclusion principle.

Hence for the hydrogen molecule, both electrons occupy the bonding molecular orbital, a chemical bond being established because this bonding molecular orbital is of lower energy than two singly atomic orbital of the H atoms.



Conditions for atomic orbitals to give molecular orbitals.

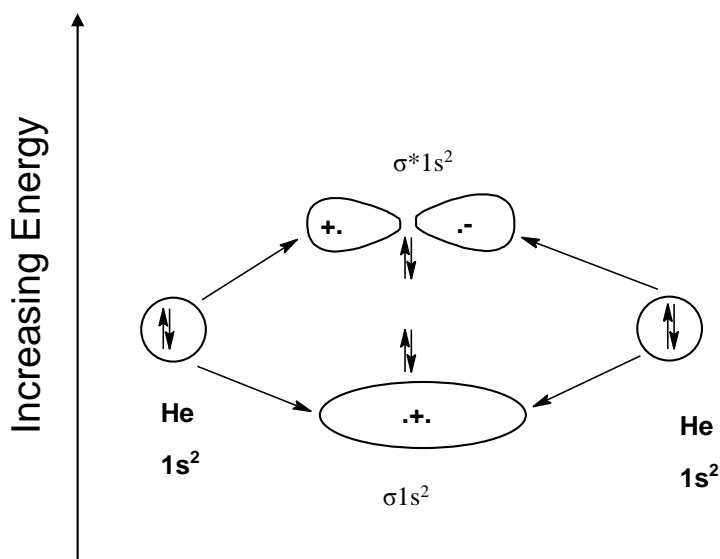
1. The atomic orbital must overlap to an appreciable extent otherwise there will be no significant build up of charge between the nuclei and therefore no chemical bond will be established.
2. In the region of overlap, the wave functions to the separate atomic orbitals must be of the same sign, otherwise there will be no build up of charge between the two nuclei.
3. The energies of the separate atomic orbitals must be of comparable magnitude.

Application of Molecular orbital Theory For s atomic orbitals

Note that Bond order = $\frac{1}{2}$ [Number of bonding electrons - Number of antibonding electrons]

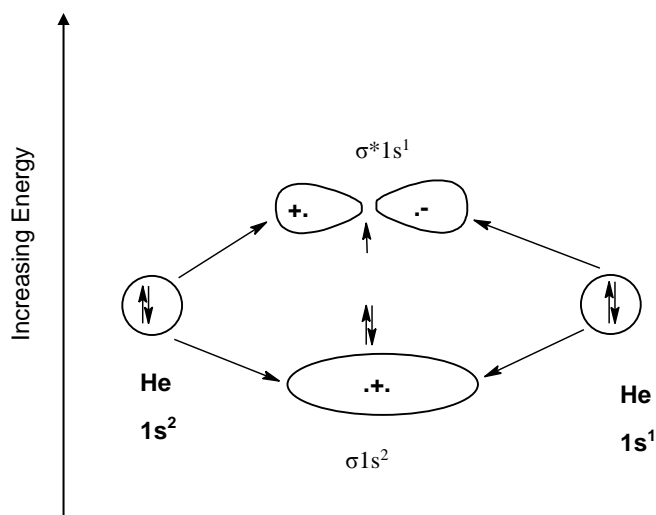
He₂ and He₂⁺

He diatomic molecule



The lower energy of the bonding molecular orbital is effectively cancelled out by the higher energy of antibonding molecular orbital so no chemical bond is established, hence no He₂ molecule is formed.

He diatomic ion.



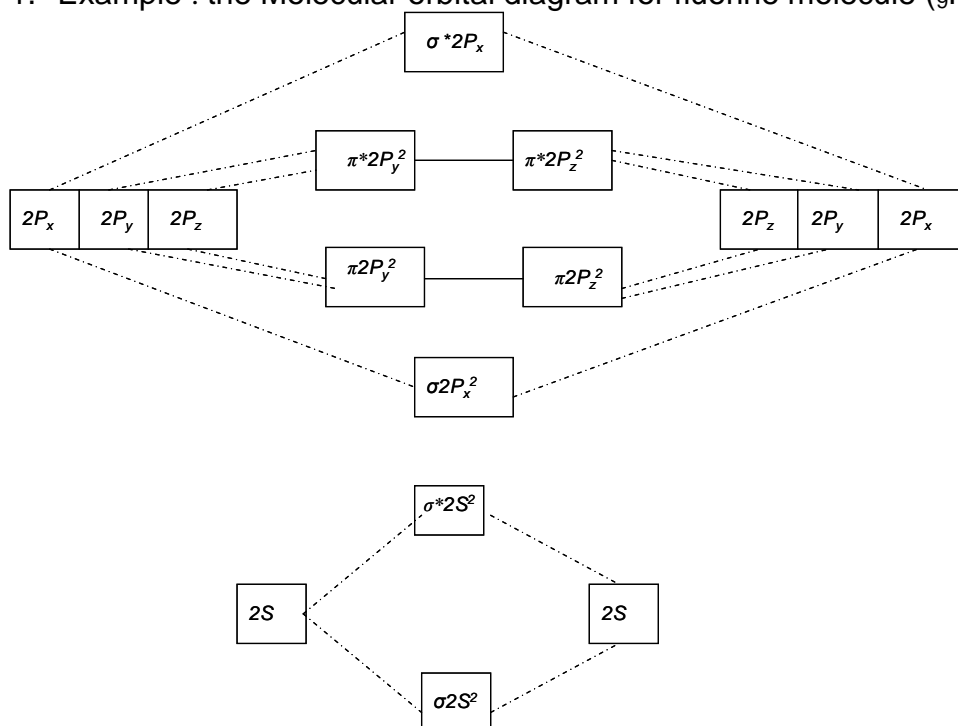
He_2^+ contains 3 electrons and 2 are in a bonding molecular orbital while the third occupy the antibonding molecular orbital. There is some overall bonding and this specie can exist. It can be obtained by subjecting He gas at a low pressure to a high energy discharge.

Application of Molecular orbital Theory For p atomic orbitals

P atomic orbitals can overlap in two distinct ways:

1. Two $2p_x$ can overlap along x-axis and give rise to σ bonding and antibonding molecular orbital.
2. Two $2p_x$ can also overlap laterally and again give rise to π bonding and antibonding molecular orbital.

1. Example . the Molecular orbital diagram for fluorine molecule (${}_9\text{F}$).



Therefore the molecular orbitals of fluorine molecule is ;

$$\text{F}_2 = \sigma_{1s}^2 \sigma_{1s}^{*2} \sigma_{2s}^2 \sigma_{2s}^{*2} \sigma_{2p_x}^2 \prod_{2p_y}^2 \prod_{2p_z}^2 \prod_{2p_y}^{*2} \prod_{2p_z}^{*2}$$

Chemistry of Group IIIA (Boron Group)

This group marks the beginning of the p-block elements. Members Include: Boron ($^{10}_5\text{B}$), Aluminium ($^{27}_{13}\text{Al}$), Gallium ($^{70}_{31}\text{Ga}$), Indium ($^{115}_{49}\text{In}$), Thallium ($^{204}_{81}\text{Tl}$). Boron is the only non-metallic in the group and the group is also electron deficient. This calls for its separate studies.

Boron ($^{10}_5\text{B}$): $[\text{He}] 3s^2 3p^1$

Aluminium ($^{27}_{13}\text{Al}$): $[\text{Ne}] 3s^2 3p^1$

Gallium ($^{70}_{31}\text{Ga}$): $[\text{Ar}] 3d^{10} 4s^2 4p^1$

Indium ($^{115}_{49}\text{In}$): $[\text{Kr}] 4d^{10} 5s^2 5p^1$

Thallium ($^{204}_{81}\text{Tl}$): $[\text{Xe}] 4f^{14} 5d^{10} 6s^2 6p^1$

Elemental boron has properties that places it on the borderline between metal and non metals. It is a semiconductor, not a metallic conductor and chemically it must be classed as a non metal. In general, B chemistry resembles that of Si more closely than that of Al, Ga, In and Tl.

Isolation of the Element

Boron is not an abundant element, but it occurs in some concentrated deposits in arid parts of California, and elsewhere, as borax $\text{Na}_2\text{B}_4\text{O}_5(\text{OH})_4 \cdot 8\text{H}_2\text{O}$ and kernite, $\text{Na}_2\text{B}_4\text{O}_5(\text{OH})_4 \cdot 2\text{H}_2\text{O}$.

Compounds of Group IIIA Elements

1) Oxygen compounds of boron

- (i) Boron oxides
- (ii) Boric acid
- (iii) Borates

2) The Boron Trihalides

- (i) Boron Trifluorides
- (ii) Boron Trichlorides

3) The Boron Hydrides

- (i) Boranes
- (ii) Diboranes
- (iii) Higher boranes

4) Complexes and organometallic compounds of Boron

Aluminum is the commonest metallic element in the earth's crust and occurs in rocks such as feldspars and micas. More accessible deposits are hydrous oxides such as bauxite,

$\text{Al}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ and cryolite, Na_3AlF_6 . Gallium and indium occur only in traces in aluminum and zinc ores.

Compounds of aluminum includes alumina Al_2O_3 (α and γ), $\text{Al}(\text{OH})_3$ thallium oxide (Tl_2O) as well as binary compounds: Al_4C_3 .

Chemistry of Group VA (Nitrogen Group)

Elements in Group VA are referred to as nitrogen group. They include: Nitrogen ($^{14}_7\text{N}$), Phosphorous ($^{31}_{15}\text{P}$), Arsenic ($^{74.9}_{33}\text{As}$), Antimony ($^{121.8}_{51}\text{Sb}$), Bismuth ($^{209}_{83}\text{Bi}$).

Nitrogen ($^{14}_7\text{N}$): $[\text{He}] 2s^2 2p^3$

Phosphorous ($^{31}_{15}\text{P}$): $[\text{Ne}] 3s^2 3p^3$

Arsenic ($^{74.9}_{33}\text{As}$): $[\text{Ar}] 3d^{10} 4s^2 4p^3$

Antimony ($^{121.8}_{51}\text{Sb}$): $[\text{Kr}] 4d^{10} 5s^2 5p^3$

Bismuth ($^{209}_{83}\text{Bi}$): $[\text{Xe}] 4f^{14} 5d^{10} 6s^2 6p^3$

Nitrogen occurs in the nature mainly as dinitrogen N_2 which comprises 78% by volume of the earth's atmosphere. The great strength of $\text{N}=\text{N}$ bond is principally responsible for the chemical inertness of N_2 and for the fact that most simple nitrogen compounds are endothermic even though they may contain strong bonds. N is obtained by liquefaction and fractional distillation of liquid air.

Nitrogen Compounds

Anionic compounds

(i) Nitrides- N^{3-} , (ii) Nitrites- N^{2-} (iii) Azide - N_3^-

Nitrogen hydrides

(i) NH_3 (ii) NH_4^+ salts (iii) N_2H_4

Oxides of Nitrogen

(i) N_2O (ii) NO (iii) NO_2 (iv) N_2O_4 (v) N_2O_5

Compounds of other Group elements

(i) PH_3 (ii) PX_3 (where $\text{X} = \text{F}, \text{Cl}, \text{Br}$ or I)

Chemistry of Noble Gases VIIIA /0 (Noble gases)

The last main group consist of individual atoms too "noble" to interact with others. The group 8A/0 elements display regular trends in physical properties and very low, if any, reactivity.

The group consists of helium(He), the second most abundant element in the universe, neon (Ne) and argon(Ar); krypton (Kr) and Xenon(Xe), the only members for which compounds are

known; and finally radioactive radon (Rn). The elements of this group form no compounds when considered early member (He, Ne & Ar) for that reason were called the “inert gases”. However, this is no longer true. Compounds are now known which contain the elements krypton, xenon and radon. The reason why it is easier to form compounds of the latter is that the outer electrons of Kr and Xe are so much further from the nucleus and are therefore less tightly held.

Electronic Configuration:

He : $1s^2$

Ne: $1s^2 2s^2 2p^6$

Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$

Kr: $[Ar] 3d^{10} 4s^2 4p^6$

Xe: $[Kr] 4d^{10} 5s^2 5p^6$

Rn: $[Xe] 4f^{14} 5d^{10} 6s^2 6p^6$

Uses of Noble Gases

Helium and neon are used in cryogenic devices to attain and maintain low temperatures. The inert character of helium and argon makes them useful in arc welding to provide a gas shielding around the metal electrodes. Neon, helium, argon and xenon are used in various types of lamps, flash lamps, glow lamps and stroboscopic lights. They are also used in electronic devices ranging from lasers to cathode-ray tubes.

Compounds of Noble Gases

Compounds of the inert gases are characterized by bonds with highly electronegative atoms such as O/F. Typical of these simple compounds are the xenon fluorides including XeF_6 , XeF_4 , XeF_2 , $XeOF_4$, XeO_3 , XeO_4 . The first rare gas compound is $[XePtF_6]$. Formation of clathrate compounds of the noble gases are also known.

Introduction to first row transition metal chemistry

Transition elements are metals of partially and completely filled d- orbital (typical and non-typical). They are found in the 4th, 5th and 6th periods of the periodic table forming 1st, 2nd and 3rd rows. They share properties similar to metals but with certain peculiar properties which includes:

- 1) Variable oxidation state
- 2) Formation of coloured compounds.
- 3) Catalytic behaviour.
- 4) Formation of complex ions.
- 5) Magnetic effect etc.