

BP202T. PHARMACEUTICAL ORGANIC
CHEMISTRY –I (Theory)



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UNIT-II

UNIT-II 10 hours

Alkanes*, Alkenes* and Conjugated dienes*

- SP³ hybridization in alkanes, Halogenation of alkanes, uses of paraffins.
Stabilities of alkenes, SP² hybridization in alkenes

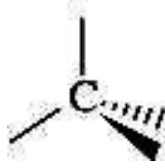


PART3

Sp³ Hybridization in Alkanes, Halogenation of Alkanes, Uses of Paraffin's

Hybridization:

- Hybridization and general principles explain how covalent bonding in organic chemistry is possible.
- Hybridization happens when atomic orbitals mix to form a new atomic orbital. The new orbital can hold the same total number of electrons as the old ones. The properties and energy of the new, hybridized orbital are an 'average' of the original unhybridized orbitals.

Types of Hybridization

Type of hybrid	<i>sp³</i> <i>hybridization</i>	<i>sp²</i> <i>hybridization</i>	<i>sp</i> <i>hybridization</i>
Diagram			
Atomic orbitals used	s, p, p, p	s, p, p	s, p
Orbitals Combined	s-orbital + 3 p-orbitals	s-orbital + 2 p-orbitals	s-orbital + 1 p-orbital
Resulting Orbitals	4 sp ³ orbitals (no p-orbitals)	3 sp ² orbitals + 1 p-orbital	2 sp orbitals + 2 p-orbitals
Number of hybrid orbitals formed	4	3	2
Number of atoms bonded to the C	4	3	2
Geometry	tetrahedral	flat triangular	Linear
Ideal angle	109.5 ^o	120 ^o	180 ^o
Bonds	single bonds	double bonds	triple bonds

- **Bond Lengths:** mostly dependent on atomic size, bond order, and hybridization

o Multiple Bonding: Bond length depends strongly on bond order (length: single > double > triple)

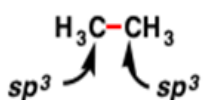
Bond Lengths (Å)

sp^3-sp^3	C—C	1.54	sp^3-sp^3	C—O	1.42
sp^2-sp^2	C=C	1.34	sp^2-sp^2	C=O	1.22
$sp-sp$	C≡C	1.20			

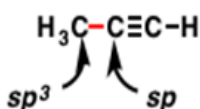
Effect of hybridization on length of single bonds: C—H and C—C bonds shorten slightly with increased s character on carbon.

The six types of carbon-carbon σ -bonds

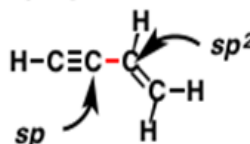
sp^3-sp^3 σ -bond



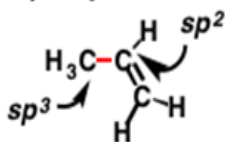
sp^3-sp σ -bond



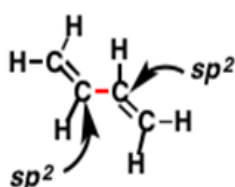
$sp-sp^2$ σ -bond



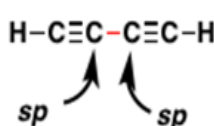
sp^3-sp^2 σ -bond



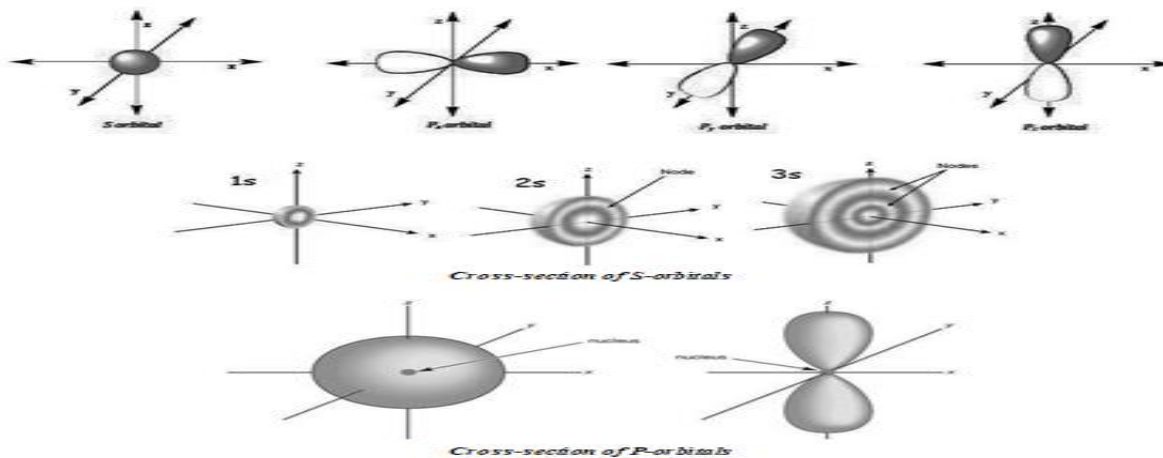
sp^2-sp^2 σ -bond



$sp-sp$ σ -bond

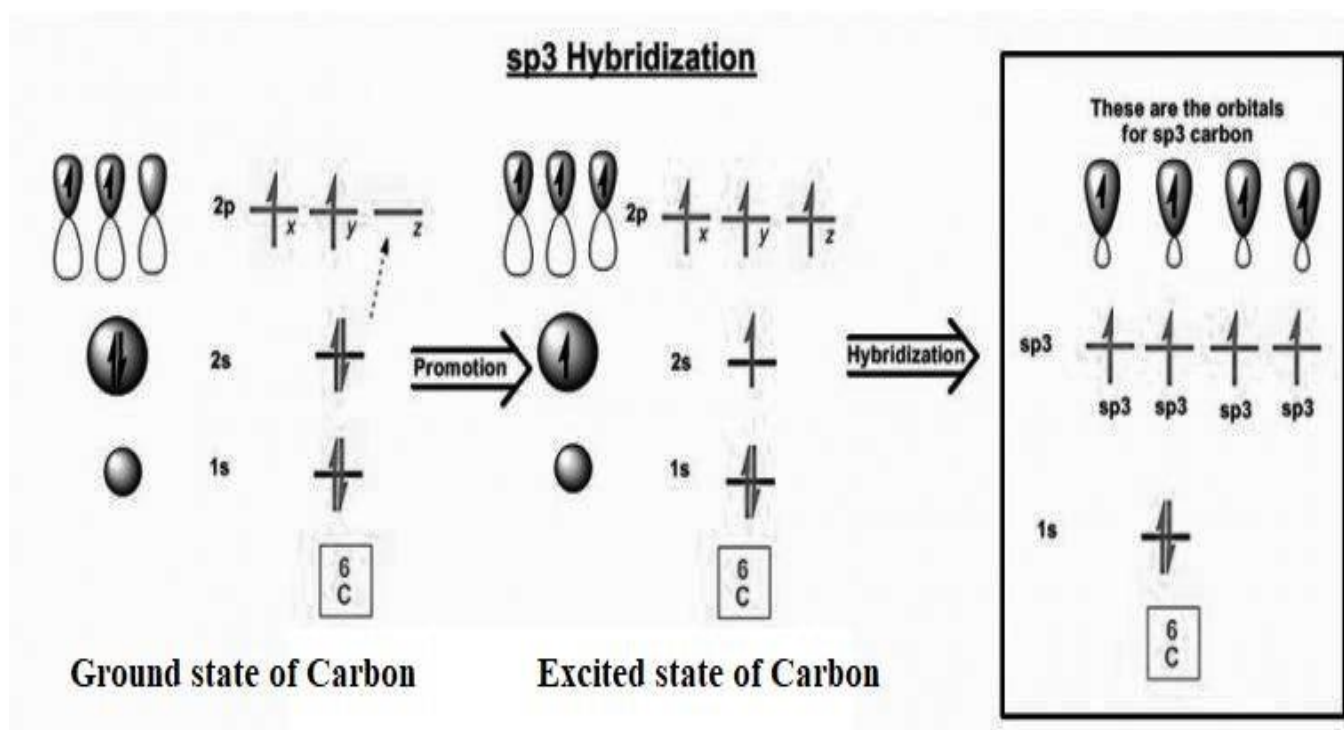


Orbital Shapes

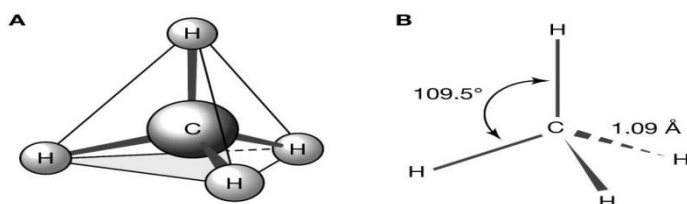


Sp³Hybridization

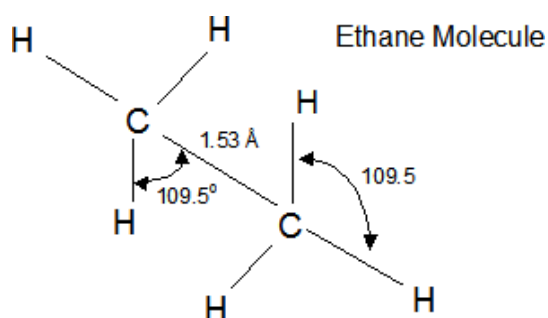
- The process of formation of 4 equivalent orbitals from hybridization of one s orbital and three p orbitals is known as sp³ hybridization. sp³ hybrid orbitals and properties of sigma bonds.
- **Characteristics:**
 - o sp³ has 25% s and 75% p character
 - o The 4 sp³ hybrids point towards the corners of a tetrahedron at 109.5° to each other
 - o Each sp³ hybrid is involved in a sigma bond.



- Bond Angle and Bond Length of Methane

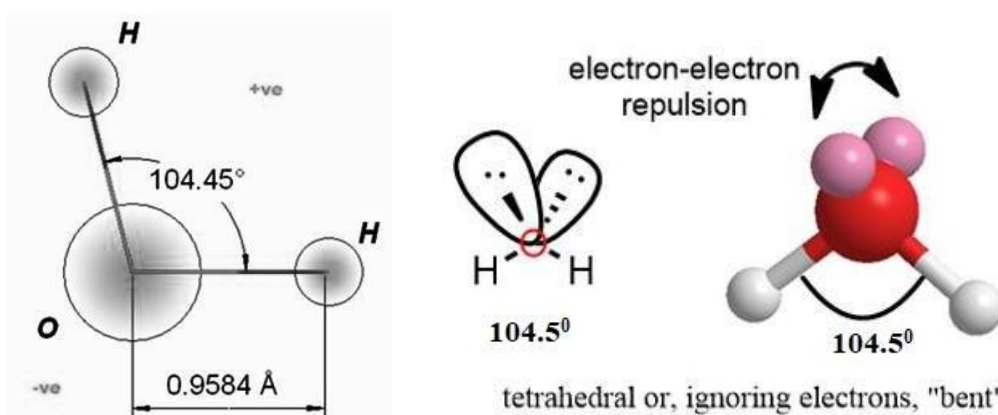


- **Bond Angle and Bond Length of Ethane**



- **Why does the bond angle go from 104.5° to 109.5° when water freezes to form ice?**

- o The reason water has an angle of 104.5° is because its two lone pairs are closer to the oxygen than the two bonding pairs. The electron pairs repel each other, which pushes the two bonding pairs closer together. This is what compresses the H-O-H bond angle in water.
- o In ice the molecules line up to form a network of hydrogen bonds. This means every lone pair has a hydrogen from a neighbouring molecule

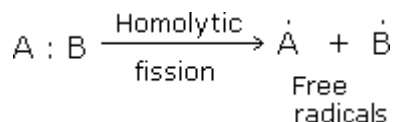


very close to it. This pulls the lone pairs away from the oxygen atom which increases the H-O-H bond angle back to 109.5°

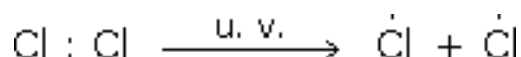
- o The oxygen in water is essentially tetrahedral. Since it is only singly-bonded, it must be sp^3 and have angles that approximate 109.5° . It does but since oxygen has two lone pairs of electrons that are diffuse (larger). These larger orbitals occupy greater volumes than covalent bonds to hydrogen, when only the geometry of bonds is considered, it is called bent. Organic chemists care about all attachments, particularly lone pairs of electrons.

Free Radical

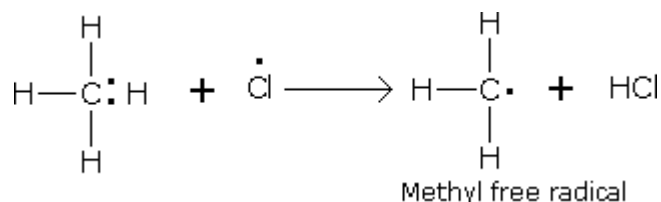
- A free radical may be defined as an atom or group of atoms having an unpaired electron. Free radicals are produced during the homolytic fission of a covalent bond.



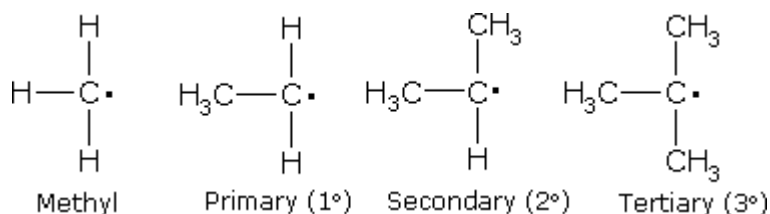
- Free radicals are very reactive as they have strong tendency to pair up their unpaired electron with another electron from wherever available. These pairs are very short lived and occur only as reaction intermediates during reactions.
- For example, dissociation of chlorine gas in the presence of ultra-violet light produces chlorine free radicals:



- The alkyl free radical may be obtained when free radical chlorine attacks methane.

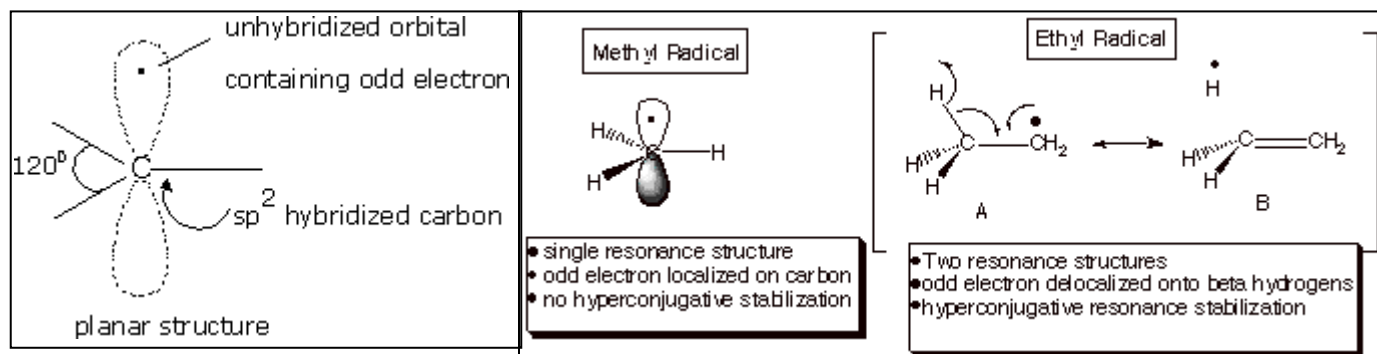


- Free radicals may be classified as *primary*, *secondary* or *tertiary* depending upon whether one, two or three carbon atoms are attached to the carbon atom carrying the odd electron:



- **Structure of alkyl free radical:**

The carbon atom in alkyl free radicals involves **sp² hybridization**.

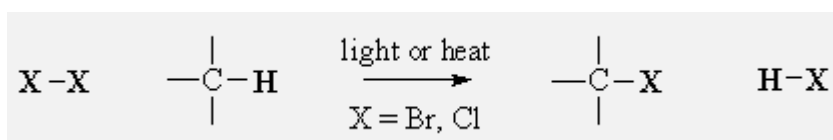


Therefore, it has a **planar structure**. Three hybrid orbitals are used in the formation of three **s-bonds** with three **H atoms** or **alkyl group**.

The unpaired electron is present in unhybridized **P orbital**.

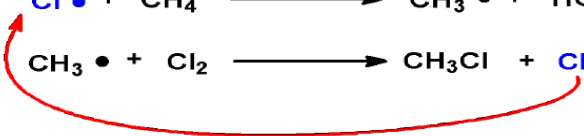
Halogenation of Alkanes (Free Radical Substitution Reaction)

- The reaction of a halogen with an alkane in the presence of ultraviolet (UV) light or heat leads to the formation of a haloalkane (alkyl halide). An example is the chlorination of methane.
- Radical Halogenation of Alkanes (Reaction type: Free Radical Substitution)



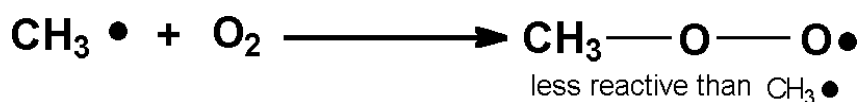
- **Summary:**

- o When treated with Br₂ or Cl₂, radical substitution of R-H generates the alkyl halide and HX.
- o Alkane R-H relative reactivity order: tertiary > secondary > primary > methyl.
- o Halogen reactivity F₂ > Cl₂ > Br₂ > I₂
- o Only chlorination and bromination are useful in the laboratory.
- o Bromination is selective for the R-H that gives the most stable radical.
- o Chlorination is less selective.

- Mechanism Of Halogenation	Radical chain mechanism for reaction of methane with Cl ₂
<p>Initiation Step</p> $X_2 \xrightarrow{uv} X\cdot + X\cdot$ <p>Propagation Steps</p> $RH + X\cdot \longrightarrow R\cdot + HX$ $R\cdot + X_2 \longrightarrow RX + X\cdot$ <p>Termination Steps</p> $X\cdot + X\cdot \longrightarrow X_2$ $X\cdot + R\cdot \longrightarrow RX$ $R\cdot + R\cdot \longrightarrow R-R$	<ul style="list-style-type: none"> Chain - initiating steps: steps in which the chain reaction starts $Cl-Cl \xrightarrow{\text{energy}} 2 Cl\cdot$ Chain propagating steps: steps which keep to the reaction going $Cl\cdot + CH_4 \longrightarrow CH_3\cdot + HCl$ $CH_3\cdot + Cl_2 \longrightarrow CH_3Cl + Cl\cdot$  Chain - terminating: steps which might occur which causes the chain reaction to stop $Cl\cdot + Cl\cdot \longrightarrow Cl_2$ $CH_3\cdot + CH_3\cdot \longrightarrow C_2H_6$ $CH_3\cdot + Cl\cdot \longrightarrow CH_3Cl$

- INHIBITORS

- o Inhibitor - a substance which slows down or stops a reaction even though the inhibitor is present in small amounts.
- o Inhibition period - time during which the inhibitor lasts.
- o Example: If oxygen is present during halogenation, the oxygen slows down the reaction.



- o This breaks the cycle (propagating steps) and slows down the reaction.
- o When the oxygen molecules are all reacted (inhibition period), the reaction then speeds up.

Why is the direct fluorination and iodination of alkanes via a free radical mechanism not possible?

- Fluorination

- o Fluorination (155 kJ/mol) seems to have relatively high activation energy. The initial reaction (chain initiation) - that is, the homolytic cleavage of a halogen molecule - must, however, occur only a few times. The subsequent reactions (chain propagation) between a

halogen radical and methane, and then between a methyl radical and a halogen molecule, yield another halogen radical. Therefore, one start reaction may initiate thousands of fluorination reactions. In addition, fluorination is very exothermic; the reaction enthalpy is -431 kJ/mol. As a result, the reaction itself provides enough energy for additional initiation reactions.

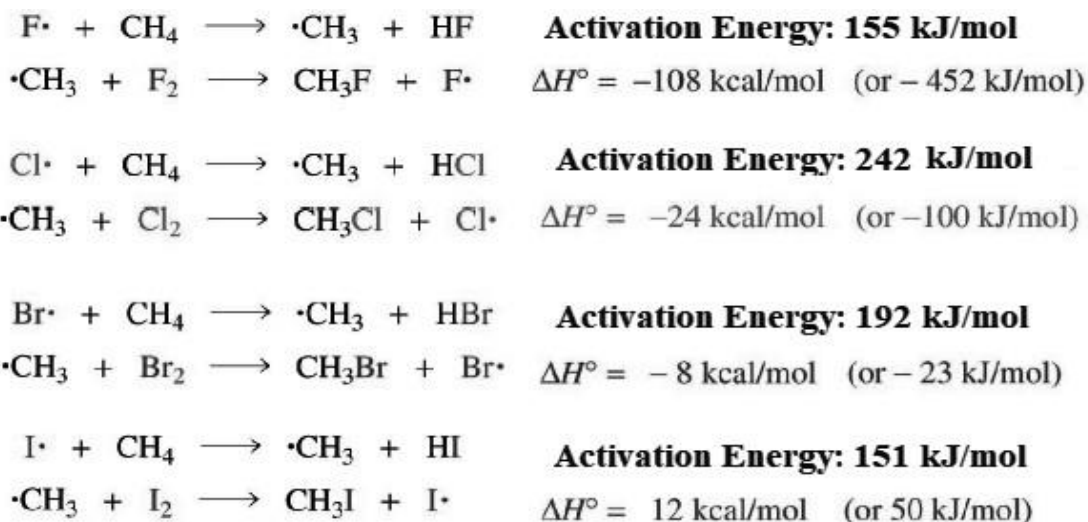
- o As a result, an explosion occurs.
 - o Nevertheless, methane fluorination may be carried out in a controlled reaction, so as to prohibit an explosion. Diluting the starting products with an inert gas or absorbing the reaction heat with copper granulate can help in this case.
- **Chlorination**
- o As a result of higher activation energy in chain initiation as well as a less exothermic character ($\Delta H^\circ = -115 \text{ kJ/mol}$) of the chain propagation, the reaction rate of methane chlorination is comparatively lower than that of the fluorination. Therefore, the reaction of methane chlorination is easier to control.

- Bromination

- o In the chain initiation of methane bromination, the activation energy is lower than that in chlorination. However, the chain propagation is far less exothermic, and the first reaction of the chain propagation is even much more endothermic (+75 kJ/mol) than in the case of chlorination (+8 kJ/mol). Therefore, the chain propagation proceeds extremely slowly, even at 300 °C, and bromine is by far less reactive than chlorine against methane.

- Iodination

- o In the chain initiation of methane iodination, the activation energy is even lower than it is in fluorination. Therefore, one could assume that methane iodination runs more rapidly than fluorination. However, this is not the case! The complete chain propagation (+54 kJ/mol), and, in particular, the first reaction (+142 kJ/mol), is very endothermic. As a result, the radical iodination of methane does not take place.



Factors That Affect Free Radical Stability

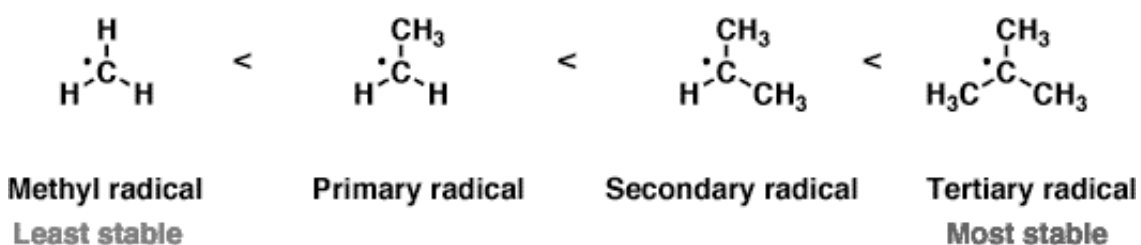
1. **Free radicals are electron-deficient species.** [Helpful to know: the factors which affect the stability of free radicals are the same which influence the stability of carbocations.] They can be stabilized through donation of electron

density by neighbours; for this reason, radical stability increases in the order methyl < primary < secondary < tertiary. [Radicals are also stabilized by adjacent atoms with lone pairs, such as oxygen and nitrogen].

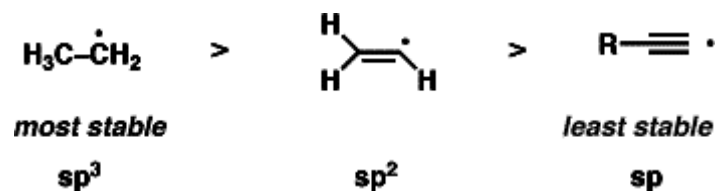
2. Free radicals decrease in stability as the % of *s-character* in the orbital increases [i.e. as the half-empty orbital becomes closer to the nucleus].

For that reason, free radical stability decreases as the atom goes from $sp^3 > sp^2 > sp$.

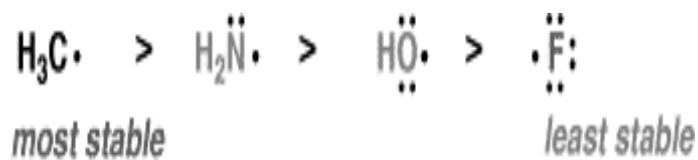
Radical stability increases in the order methyl < primary < secondary < tertiary



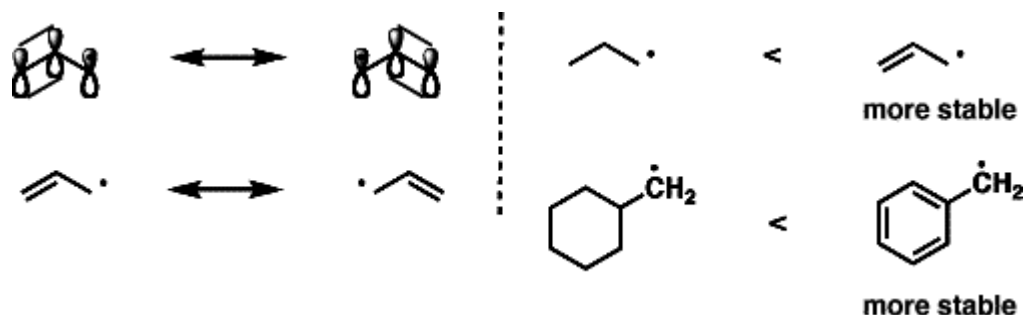
3. A second important factor which stabilizes free radicals is “delocalization” – that is, if the radical can be spread out over two or more carbons. A more familiar way of saying this is that free radicals are stabilized by resonance.



Across a row of the periodic table, free radicals decrease in stability as the electronegativity increases



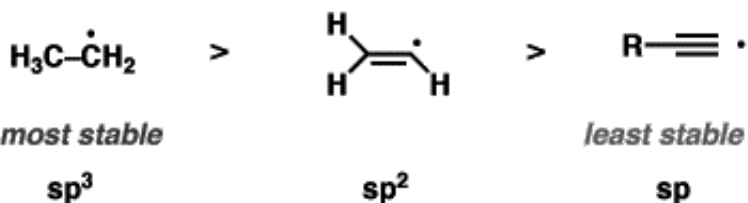
Free radicals increase in stability going down a column of the periodic table, $\text{F}\cdot < \text{Cl}\cdot < \text{Br}\cdot < \text{I}\cdot$ since the electron-deficient orbital is spread out over a greater volume



Factors Destabilize Free Radicals:

Factor 1: Hybridization

The stability of a free radical decreases as the orbital is held closer to the nucleus



Factor 2: Electronegativity

Across the periodic table, decrease in stability from left to right (i.e. with increasing electronegativity)

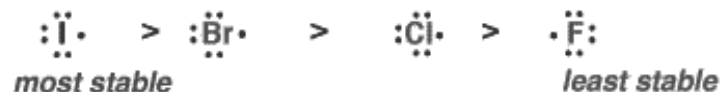


Why? A partially filled orbital becomes increasingly unstable the more tightly it is held to the nucleus

Nitrogen and oxygen centered radicals are quite unstable

Factor 3: Polarizability

Down the periodic table, increase in stability, since the partially filled orbital is on a larger atom - and the "deficiency" can be spread over a larger area

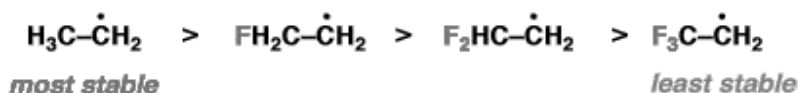


Why? A partially filled orbital becomes increasingly unstable the more tightly it is held to the nucleus

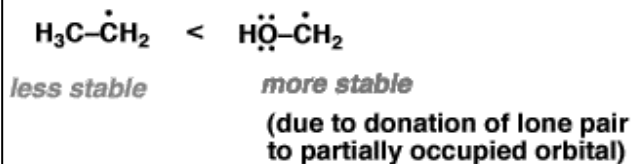
Nitrogen and oxygen centered radicals are quite unstable

Factor 4: Electron withdrawing groups

Removing electron density from a free radical makes it less stable



Watch out! For substituents with lone pairs of electrons, it is sometimes difficult to separate the electron-withdrawing effect of the substituent from resonance effects



Paraffin

- Paraffin's, more commonly referred to as alkanes, are the chemical family of saturated hydrocarbons.
- The general formula C_nH_{2n+2} , **C** being a carbon atom, **H** a hydrogen atom, and „**n**“ an integer.
- The paraffin's are major constituents of natural gas and petroleum.
- Paraffin's containing fewer than 5 carbon atoms per molecule are usually gaseous at room temperature, those having 5 to 15 carbon atoms are usually liquids, and the straight-chain paraffins having more than 15 carbon atoms per molecule are solids.

- Branched-chain paraffin's have a much higher octane number rating than straight-chain paraffin's and, therefore, are the more desirable constituents of gasoline.

Paraffin wax

- It is also known as American English paraffin, is a white wax obtained from petrol or coal. It is used to make candles and in beauty treatments.
- The term "**wax**" simply refers to saturated hydrocarbons that contain more than 16 carbon atoms in the paraffin series (**C₁₆-C₄₀**) and are in solid state at room temperature. Chemically, natural waxes are defined as long chain esters, monohydric (one hydroxyl group), or alcohols with long chain fatty acids. The majority of the waxes present in crude oil are considered synthetic paraffin waxes with non-oxidized saturated alkanes.

□ Uses of Paraffin

- Medicinal liquid paraffin, also known as *paraffinum liquidum*, is a very highly refined mineral oil used in cosmetics and for medical purposes.
- Liquid paraffin has many uses in the medical field. Because liquid paraffin passes through the body's intestinal tract without being absorbed, it can be used as a laxative to limit the amount of water removed from the stool and ease constipation.

- Liquid paraffin is considered to have a limited usefulness as an occasional laxative.
- Paraffin wax were developed, the most popular of which was giving hot wax baths to patients suffering from a variety of ailments, in particular rheumatism and joint pain. The wax would be used to soften the skin and the intense heat would soothe the muscles and ready them for massage treatment.

References:

- A textbook of organic chemistry, 22nd edition, Arun bahl, B S Bahl, S. Chand publishing. Pg no; 199-233, 234-271, 272-292, 293-318-365, 333-365.
- Organic chemistry; 7th edition; Robert Thornton Morrison, S.K. Bhattacharje, Pearson, Pg no: 120-291-326, 411-423 382-385, 507-510, 544-559. 570-597, 624-649