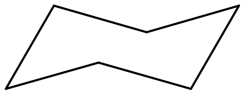
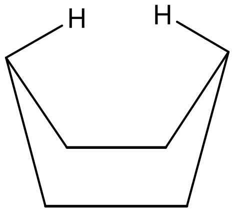
LECTURE 3

Shape: Conformations of Cyclohexanes

Cyclohexane has two limiting conformations, the **chair** and the **boat**:



Chair

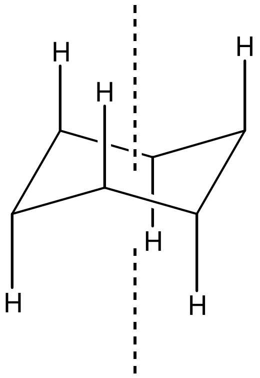


Boat

The boat conformation is much less stable than the chair for two reasons. Firstly, Newman projections along either of the horizontal C-C bonds in the boat form show eclipsing and, secondly, the two hydrogens shown, the so-called bowsprit hydrogens, come within such a short distance of one another that a repulsive force is set up. Thus, we will concentrate upon the chair which constitutes > 99.9% of cyclohexane molecules.

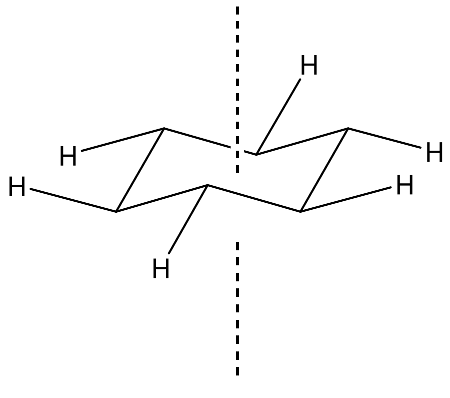
In the chair conformation there are two types of hydrogen. Six hydrogens are joined to their carbons by bonds which are parallel to an axis through the centre of the ring and are designated axial (ax) hydrogens:

Axial (ax)

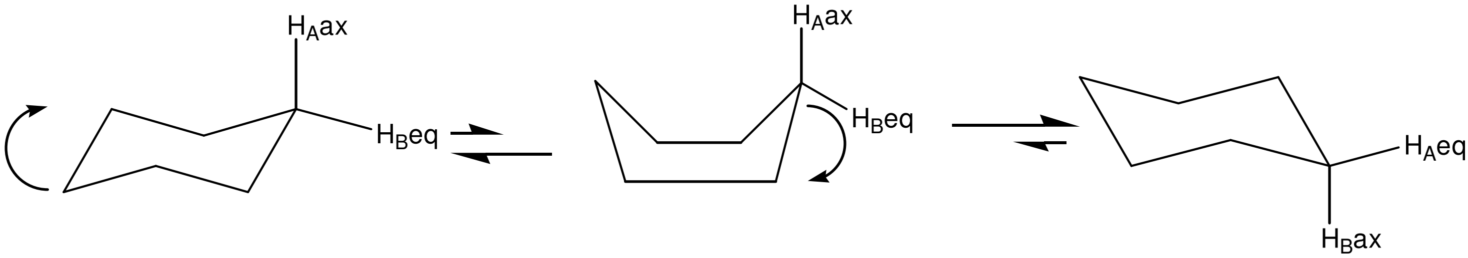


The remaining six hydrogens project around the equator of the axis and are designated equatorial (eq) hydrogens:

Equatorial (eq)



The chair form can flip through the boat to give another chair in which the axial and equatorial hydrogens have exchanged places. This occurs very rapidly at room temperature and explains why the ^1H NMR spectrum of cyclohexane only contains one peak, the average value for H(ax) and H(eq):



Reactivity of Alkanes, Alkenes and Alkynes

The great difference between alkanes and the other two classes of hydrocarbon is that the former only have σ-bonds which are relatively unreactive because the electrons are held in very stable orbitals between the carbon nuclei whereas the alkenes and alkynes have π-bonds in which the electrons reside in lower energy orbitals which are more readily available by virtue of their projection above and below the internuclear axis. Thus, in general terms alkanes are much less reactive than alkenes and alkynes and are often used as solvents. Alkanes will, however, react with very reactive species such as radicals and this is the first of our reactions to be discussed.

RADICAL SUBSTITUTIONS OF ALKANES

The general equation for this reaction is:

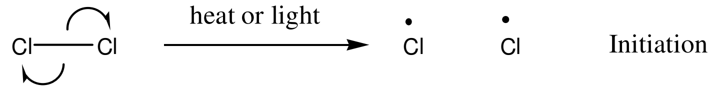
$$R_3C\text{-}H \quad + \quad X_2 \quad \rightarrow \quad R_3C\text{-}X \quad + \quad HX$$

(X = halogen)

(a) Mechanism

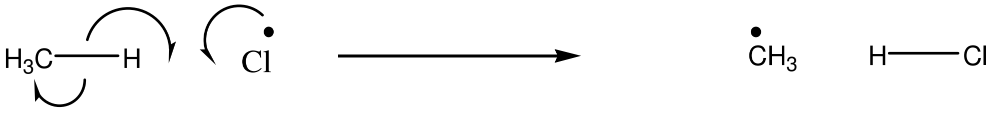
Alkanes only contain carbon and hydrogen and are fully saturated. Because the electronegativities of carbon and hydrogen are approximately equal there are no large permanent dipoles in alkanes i.e. they are non-polar molecules. Hence they do not interact with standard electrophiles or nucleophiles which tend to be attracted in the first instance to charges or partial charges in molecules. On the other hand, many free radicals have a high propensity to abstract hydrogen atoms in order to complete their coordination numbers and since alkanes are full of hydrogen atoms we may expect them to suffer ready attack by radical species. Some of the most common radicals are halogen atoms formed by homolytic cleavage of the halogen molecules and hence our mechanism starts there.

In the chlorination of methane, a mixture of methane and chlorine kept in the dark at room temperature does nothing; there is no reaction. Thus a **molecule** of chlorine is **not** the reactive species. If, however, that mixture is heated in the dark to >300 °C or if it is exposed to light then a violent reaction takes place. Obviously the heat or the light convert the chlorine molecule into the reactive species, namely, the chlorine radical (atom). Thus the first step in the chlorination of methane is:

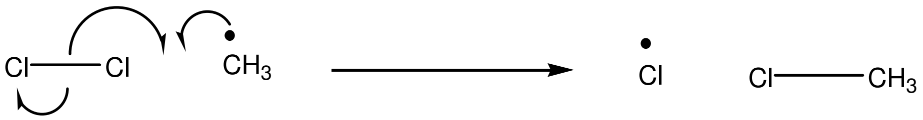


This step, called the Initiation Step, occurs by a process of bond homolysis i.e. each of the electrons in the sigma-bond go to separate atoms.

The next step is the abstraction of hydrogen from the methane to give HCl and the methyl radical. This is a radical substitution reaction occurring on hydrogen and designated S_H2 (substitution, homolytic, bimolecular):

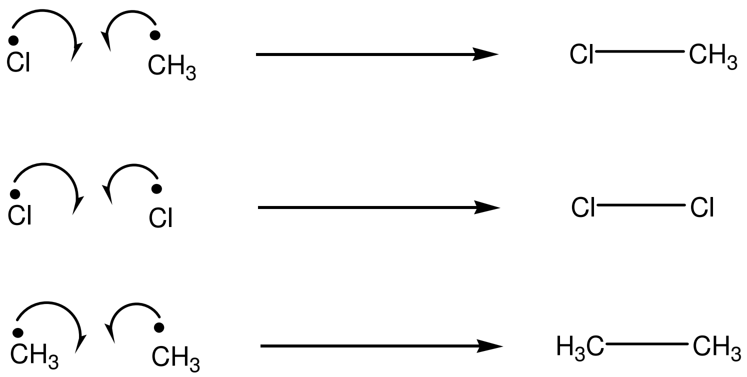


Then the methyl radical displaces a Cl from Cl_2 by another S_H2 reaction:



In this step the first molecule of organic product is formed as well as a regenerated chlorine atom; the latter can start the whole process again. Because of this recycling of the chlorine atom, these two S_H2 steps are called Propagating Steps and the whole mechanism is known as a Free Radical Chain Reaction because it is a chain of propagating steps leading to a build up of the alkyl halide product.

The chain stops when the radicals run out of starting materials with which to react and then they react with one another in a series of Termination Steps:



Thus the basic mechanism consists of a series of free radical substitutions and combinations. Note that, in theory, the Initiation Step need only happen once for the reaction to proceed; the heat or light can be turned off once reaction is underway, a characteristic feature of a radical chain reaction.

Now we must turn to some aspects of this basic mechanism and the first is to do with the comparison of halogen reactivity.

(b) Comparison of Halogen Reactivity

Fluorine reacts explosively with alkanes, even in the dark. Chlorine reacts violently on initiation, bromine reacts moderately and iodine does not react – how can we explain these differences in reactivity? This is done by looking at the thermodynamics of the two propagating steps, in particular the changes in free enthalpy when the various bonds in the molecules are either broken or made.

Step	Bond changes
$CH_4 + X^{\bullet} \rightarrow CH_3^{\bullet} + HX$	C-H lost; H-X gained
$CH_3^{\bullet} + X_2 \rightarrow CH_3X + X^{\bullet}$	X-X lost; C-X gained

The table summarises these values.

Table

Halogen	Bond Strength (kJmol^{-1})				Energy difference (kJmol^{-1})
	C-H	X-X	C-X	H-X	(HX + CX) – (CH + XX)
F_2	-439	-155	-460	-564	-430
Cl_2	-439	-242	-355	-430	-104
Br_2	-439	-192	-300	-364	-33
I_2	-439	-150	-238	-300	+51

We now have an explanation. The reaction with fluorine is exothermic to the tune of –430 kJ mol^{-1} whereas that with iodine is endothermic by +51 kJ mol^{-1}.

(c) Comparison of C-H Bond Reactivity

Not all C-H bonds have the same strength:

Bond	Strength (kJ mol^{-1})
$CH_3 – H$	-439
$MeCH_2 – H$	-409
$EtCH_2 – H$	-409
$iPrCH_2 – H$	-409
$Me_2CH – H$	-395
$Me_3C – H$	-389

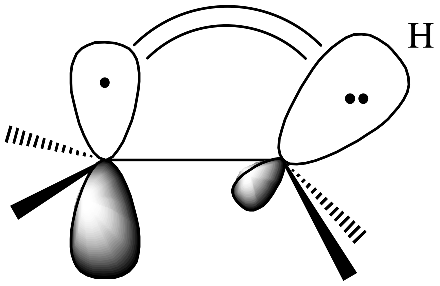
Note that all $CH_2 – H$ have the same strength irrespective of the structure of the rest of the molecule. The same is true of all $R_2CH – H$ and all $R_3C – H$.

How do we explain these differences? If we assume that the bond strength is related to the ease with which the bonds may be cleaved in a homolytic fashion then the following order of radical stability will provide the explanation:

Stability: $(CH_3)_3C^{\bullet} > (CH_3)_2CH^{\bullet} > CH_3CH_2^{\bullet} > CH_3^{\bullet}$

Thus, because the most substituted radical is the most stable it will be easier to form it from cleavage of the relevant C-H bond in the alkane i.e. that bond will be weaker. However, this explanation now leaves us with the question why should the radicals have that order of stability. For a rationalisation of that we need the theory of **hyperconjugation**.

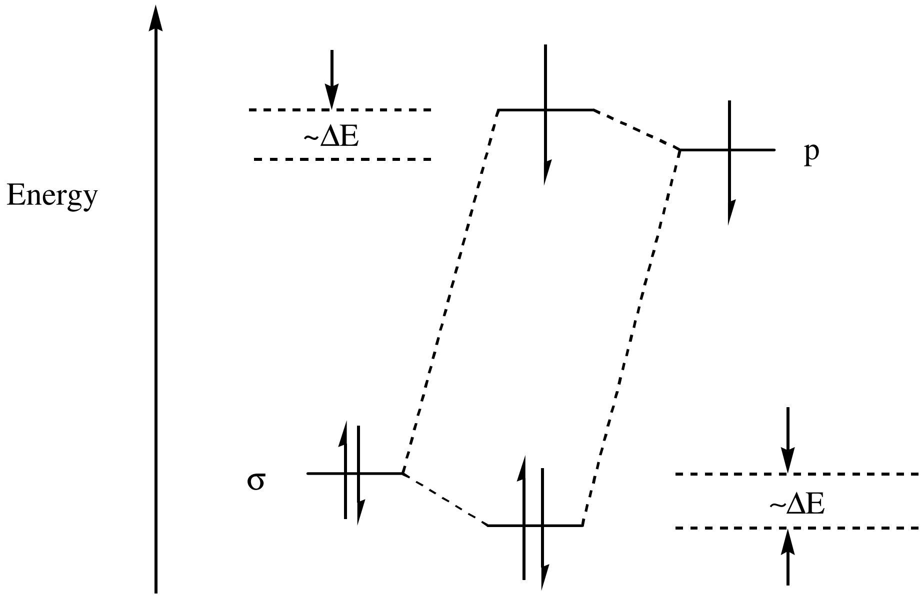
The orbital in which the single electron is located is a p-like, atomic orbital. We know that overlap of orbitals can stabilise species (formation of aromatic rings) but we normally consider the overlap of such a p orbital with another to form a π-system. However, in alkanes we have no such second p-like orbital to overlap with our radical orbital. The only available orbitals are σ, sp^3 C – H ones on the adjacent carbon:



Overlap involving σ-orbitals is called hyperconjugation, an unfortunate name since it suggests strong stabilisation through the use of the prefix "hyper".
In actuality the amount of conjugation (and hence stabilisation) by overlap of any pair of orbitals depends on two criteria:

- Physical overlap should be large i.e. the two orbitals should be close and ideally pointing towards one another
- The energies of the two orbitals should be similar (ideally the same)

By the first criterion the overlap is small in hyperconjugation because the large lobe of the sp^3 orbital is pointing away from the radical p-orbital. The second criterion also leads to the conclusion that the overlap is small as the following molecular orbital-energy diagram shows:



Sigma C-H orbitals are strong i.e. they have a low energy whereas p-orbitals are much higher in energy since they are essentially atomic orbitals with no stabilisation afforded by bonding. Hence the overlap by the second criterion is poor.

Nevertheless, there is a little stabilisation (2 x DE bonding electrons minus 1 x ΔE anti-bonding electrons = 1 x ΔE kJ mol^{-1}) and any such gain will be welcomed by the molecule. The amount of this stabilisation depends upon the number of adjacent σ-orbitals which can overlap, the greater the number the greater the accumulated stabilisation. Thus, returning to the radicals the degree of stabilisation should follow the expected order since the number of hyperconjugating C-H bonds on the adjacent carbon to the radical centre increases:

Radical:	$CH_3^\bullet$		$CH_3CH_2^\bullet$		$(CH_3)_2CH^\bullet$		$(CH_3)_3C^\bullet$
No of CH Bonds	0	<	3	<	6	<	9

(c) Comparison of C-H Bond Reactivity: Selectivity

Having established that more substituted radicals have greater stabilisation through hyperconjugation, we would now expect that a reaction such as halogenation which produces an intermediate radical would favour attack at that carbon which produces the more stable radical. Hence the reactivity of C-H bonds in free radical halogenation is:

R_3CH > R_2CH_2 > RCH_3 > CH_4
methine > methylene > methyl > methane

Two examples will illustrate this in practice.

$Cl_2 + CH_3CH_2CH_3 \rightarrow CH_3CH_2CH_2Cl + CH_3CH(Cl)CH_3$
1-chloropropane 2-chloropropane

In propane there are 6 methyl hydrogens and 2 methylene hydrogens and substitution of 1 each of these by chlorine leads to 1-chloropropane and 2-chloropropane respectively. If all the C – H bonds were of equal reactivity then we should expect to see the 1-chloropropane and 2-chloropropane to be formed in the statistical ratio of 6 : 2. The experimental value at 25 ^{o}C is 43 : 57 (1-chloro : 2-chloro). Clearly the methylene hydrogens are more reactive than the methyl ones, as expected, since we produce more of the 2-chloropropane than the 1-chloropropane. Indeed, the greater reactivity of the methylene hydrogens is overcoming the statistical factor which favours the methyl hydrogens. In order to remove the statistical factor from consideration and gain a better idea of how much more reactive methylene hydrogens are than methyl ones we need to calculate the relative reactivity **per hydrogen**, otherwise called the Selectivity. To do this we simply divide the experimental values by the number of hydrogens corresponding to each value:

Experimental ratio at 25 ^{o}C = 43 : 57
Selectivity = 43/6 : 57/2 = **1 : 3.76**

Thus a methylene C-H bond is nearly four times more reactive than a methyl C-H bond.
By looking at the chlorination of 2-methylpropane we can compare a methyl C-H with a methine C-H:

$Cl_2 + (CH_3)_2CHCH_3 \rightarrow (CH_3)_2CHCH_2Cl + (CH_3)_2C(Cl)CH_3$
Statistical Ratio = 9 : 1
Expt Ratio (25 ^{o}C) = 64 : 36
Selectivity = 64/9 : 36/1 = **1 : 5.06**

Thus, a methine C-H bond is approximately five times more reactive than a methyl C-H bond.
The next question is whether selectivities will be the same under all experimental conditions. The answer is no! If we repeat the chlorinations of propane and 2-methyl propane at 600 ^{o}C rather than 25 ^{o}C then we get the following selectivities:

CH_2 : CH_3 = 1 : 3
CH : CH_3 = 1 : 4
i.e. they approach the statistical values. This is because at higher temperatures the chlorine radical has much higher energy and is consequently much more reactive. Thus it tends to react with any hydrogen atom it meets rather than select the most reactive.

This is a general trend in chemistry, that selectivity is inversely related to reactivity; the more reactive a species the less selective it is and vice versa. Indeed, it can be seen well illustrated when comparing the reactivity/selectivity of the halogens:

Halogen	CH_3		CH_2		CH
F_2 (25 °C, gas)	1	:	1.2	:	1.4
Cl_2 (25 °C, gas)	1	:	3.76	:	5.06
Br_2 (98 °C, gas)	1	:	250	:	6300

Thus, the dangerously reactive fluorine is quite unselective whereas the moderately reactive bromine is very selective, even at higher temperatures. Despite its superb selectivity, the great disadvantage of bromine is that it is so slow. Hence industrial users tend to compromise upon selectivity and use the faster-reacting chlorine (which is also cheaper and more readily available than bromine). Fluorine is so dangerous that it is very rarely used.