# Section 5

## Reactions of Alkanes:  Radical Chain Halogenation

### Radicals and Homolysis

Radicals are molecules with unpaired electrons.  Radicals can be charged or neutral (e.g. $O_2^+$ radical and O radical).  Radicals can have just one or several unpaired electrons (e.g. $O_2$ has two unpaired electrons).  Here, we will be mostly concerned with the simplest systems:  Radicals that are neutral and that have one unpaired electron.  Radicals of this sort can have their unpaired electron either with *"spin up"* or with *"spin down"* (example methyl radical).  Because there are these two possibilities for the spin, we call these radicals *doublets.*  Closed shell systems have just one possibility for their spin; those systems are *singlets.*

Lewis notations.  The unpaired electron is shown as a dot (examples for the halogen radicals and so on).  The generic radical is denoted R· or X· depending on the EN of the radical carrying atom.

Formation by homolysis.  Homolysis describes the breaking of a bond to form **radicals**, that is, a molecule with an odd number of electrons.  The homolysis of a bond to form the free radicals requires the **bond dissociation energy**.  We will be mostly interested in C-H and C-X bonds.  Homolysis is indicated with two curved arrows in that each arrow points to the atom on which the electron will end up after the homolytic bond cleavage.
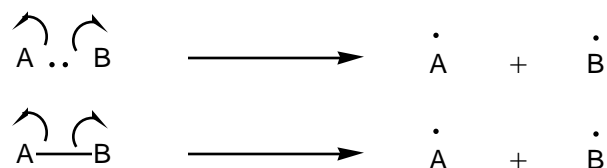


|  | **H** | **F** | **Cl** | **Br** | **I** |
|---|---|---|---|---|---|
| **H-** | 104 | 135 | 103 | 87 | 71 |
| **Me-** | 105 | 110 | 85 | 71 | 57 |
| **Et-** | 98 | 107 | 80 | 68 | 53 |
| **i-propyl** | 95 | 106 | 81 | 68 | 53 |
| **t-butyl** | 93 | 110 | 81 | 67 | 52 |

## Factors Affecting Radical Stability

Radicals basically are **electron deficient** species. Everything that helps to make up for the deficiency stabilizes the radical. Put the remaining electrons as close to the core as possible! Radical site C-atoms are sp$^2$ hybridized. In addition: One needs to add electron donors.

## (1) Allyl Radical - Stability due to Conjugation

Write the two main resonance forms of allyl radical. They are degenerate.

$$H_2C=CH-CH_2\bullet \qquad\qquad\qquad \bullet H_2C-CH=CH_2$$
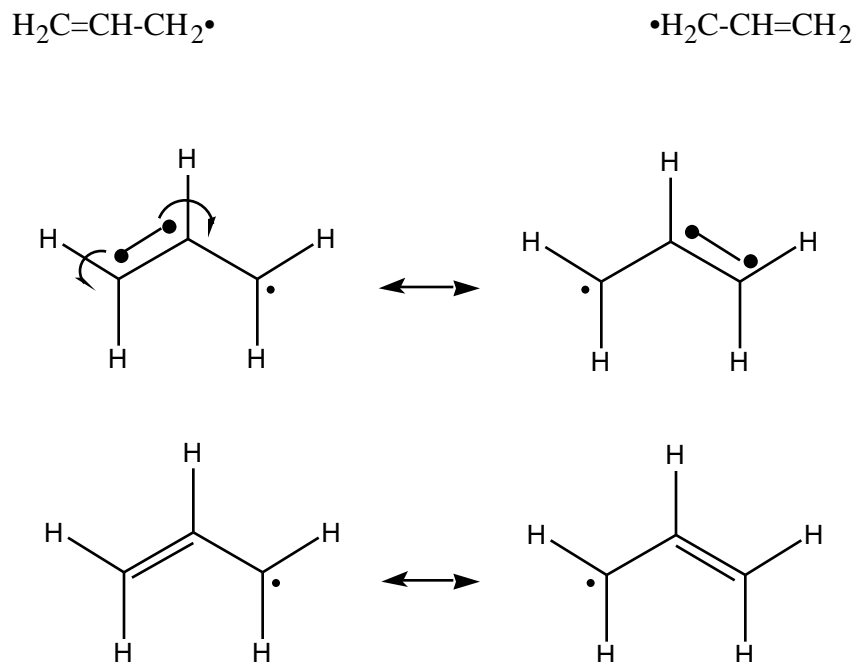


To generate a new resonane form of a radical: Disconnect the -bond as indicated by the curved arrows above. Then combine two unpaired electrins to form the new resoanance form. Usually we just think the the first line and we write the second line.

## (2) Benzyl Radical - Stability due to Conjugation
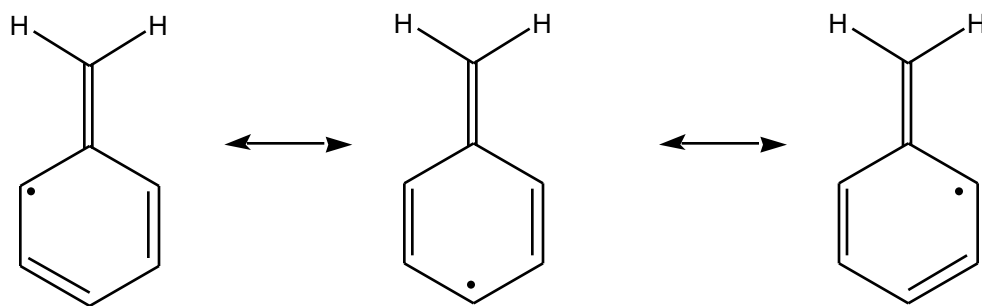
Write the main resonance forms of allyl radical. Point out which ones are degenerate.

two resonance forms with <u>exocyclic</u> radical site
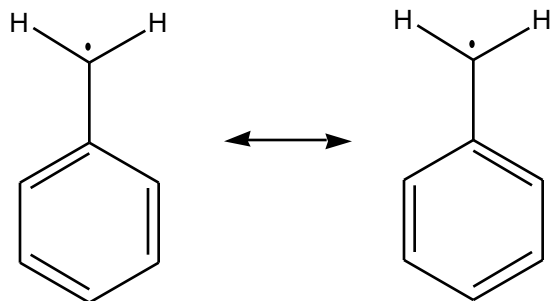
Resonance forms of radicals in which the radical sitye does not change are generated in the usual fashion by **shifting pairs of electrons** around.

three resonance forms with <u>endocyclic</u> radical position

Resonance forms that contain the radical site at different sites are generated in the same fashion as discussed for the allyl radical. Disconnect the -bond that is next to the radical site. The recombine unpaired electrons to get the new resonance form.

## (3) Radical Stability Increases with Alkyl Substitution - Hyperconjugation

| Methyl | primary | secondary | tertiary |
|--------|---------|-----------|----------|
| $H_3C\bullet$ | $H_3C\text{-}H_2C\bullet$ | $Me_2HC\bullet$ | $Me_3C\bullet$ |

"Hyperconjugation" refers to the notion that electron pairs from sigma bonds are being delocalized to make up for the electron deficiency of the radical site.

Draw ethyl radical and show the methyl group aligned such that the C-H bond is parallel to the p-AO of the radical. In this arrangement, there will be some overlap and some donation.

donation
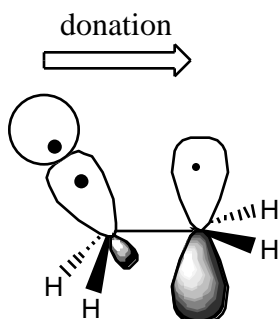
Symbolize with the "non-connected H-atom notation".
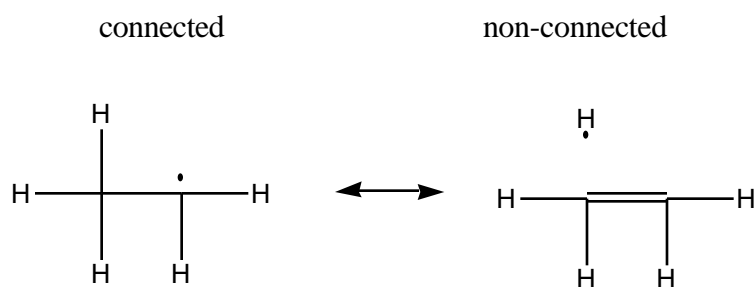
Formally, the hyerconjugation resonance forms are generated very much like the other resonance forms for conjugated radicals. Again we break a bond to the atom that is a next neighbor to the radical site. Then we recombine the electrons forming a new pi-bond and leaving the H-atom "unconnected".

## Potential Energy Surface Diagrams

Schematic plots of the change in energy as the reaction progresses.

The vertical axis is **energy**. The horizontal axis is **reaction progress**. This is often called the "reaction coordinate". Only in the simplest cases can a single internal coordinate be used to describe the reaction progress. For a simple bond dissociation, the bond length would be such a coordinate. In most reactions, several internal coordinates vary and some combination of all of these would have to be used to describe the progress.

Plot of a generic PES Diagram

A reacts to high energy intermediate B and in a fast reaction on to the exothermic product C.
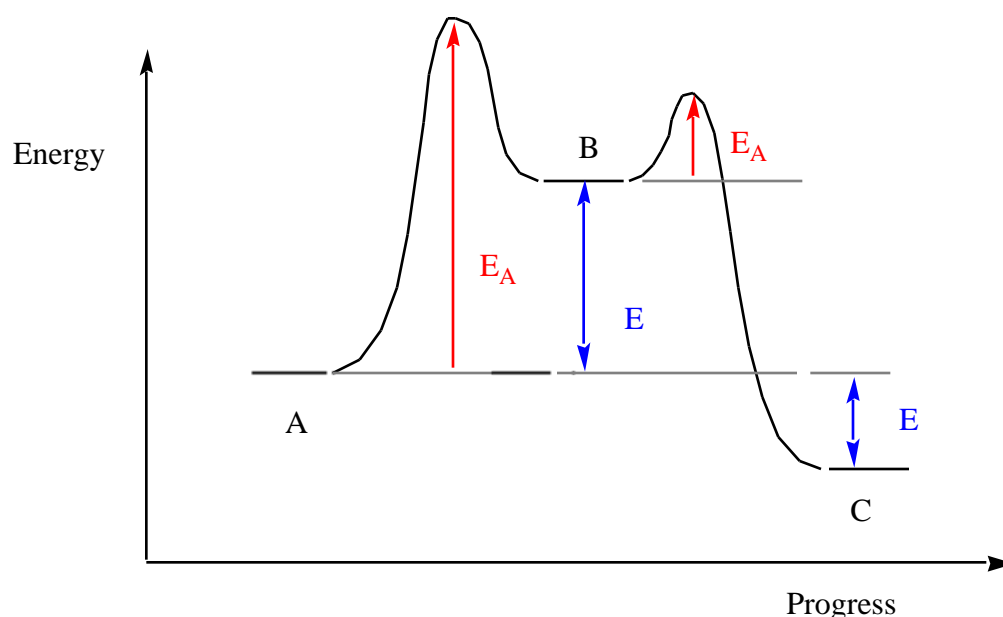


The **extrema** are of the most interest.

　　　Minimum (plural minima) correspond to **stable structures** and **intermediates**.
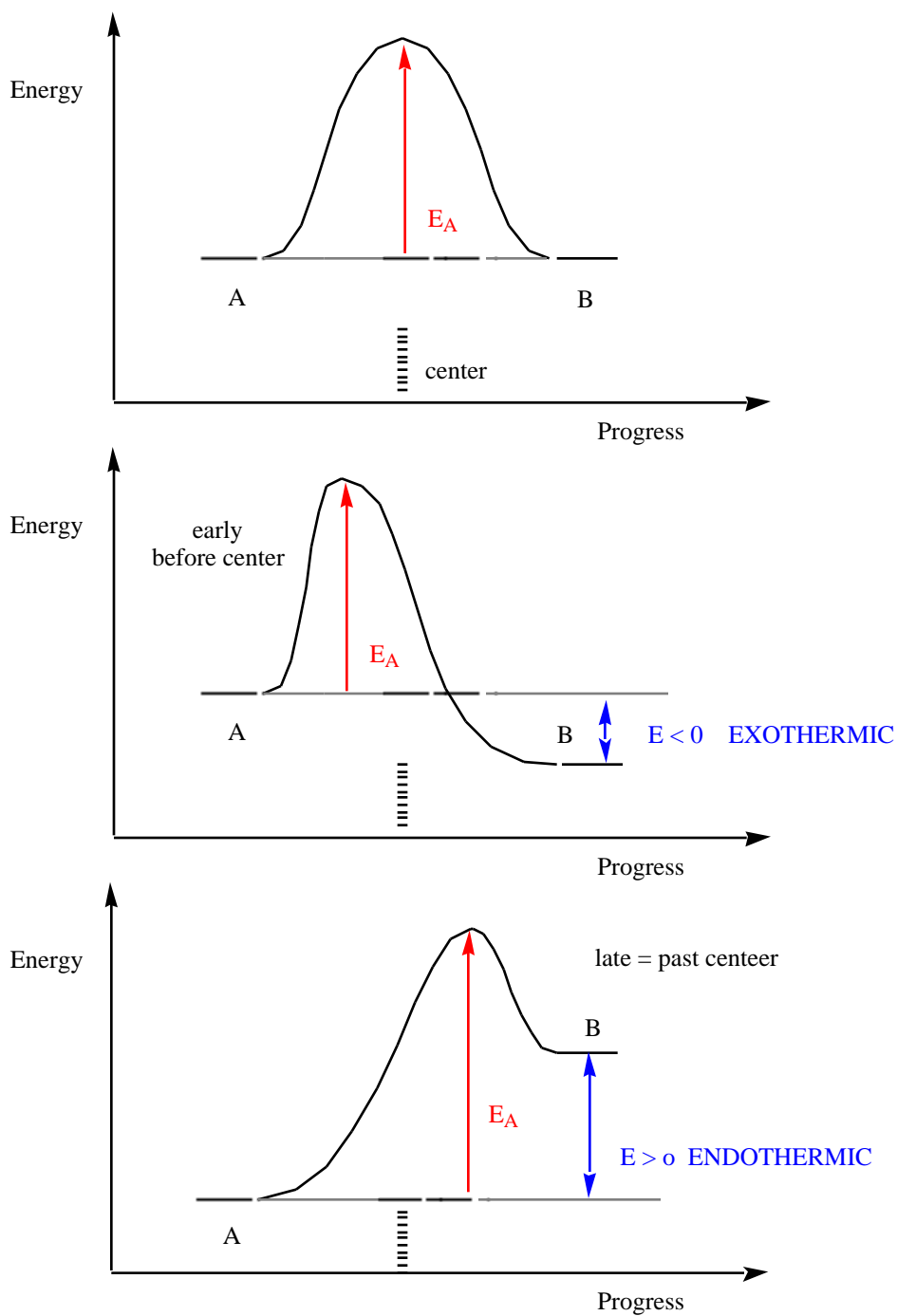　　　Maximum (plural maxima) correspond to **transition state** structures.

The difference between minima and maxima are **activation barriers $E_A$**. The control kinetics.
The differences between minima are **reaction energies** $\Delta E$. The control thermodynamics.

## Hammond's Postulate

The location of the transition states in potential energy surface diagrams can be approximated by the **Hammond Postulate**. The postulate states that the transition state is closer to the high-energy minimum. Early TS in exothermic reactions. Late TS in endothermic reaaction.

Energy

$E_A$

A                    B

center

Progress

Energy

early
before center

$E_A$

A                          B          $E < 0$   EXOTHERMIC

Progress

Energy

late = past centeer

B

$E_A$

$E > o$  ENDOTHERMIC

A

Progress

## Chlorination of Methane

**Overall Reaction**

Methane reacts with chlorine under the influence of UV light to form HCl and chloromethane.

$$CH_4 + Cl_2 \longrightarrow HCl + H_3CCl$$
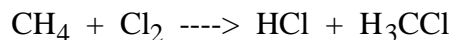
The reaction is exothermic.

| | |
|---|---|
| C-H bond in methane: 105 | C-Cl bond formed: 85 |
| Cl-Cl bond: 58 | H-Cl bond formed: 103 |
| NEED: 163 | OUT: 188 |

So, why does the reaction not proceed right away?  To asnwer this question, we need to look at the detailed **mechanism** of the reaction.

**Mechanism of reaction:**  A step-by-step description of the all of the changes in bonding and energy that occur along the **reaction path** or **reaction coordinate**.  The latter terms basically describe "progress of reaction".

The chlorination of alkanes involves a **chain reaction mechanism**.  Chain reactions are characterized by regenerating one of the molecules involved in the cation.   All in all, there are three types of steps:

(1) an *initiation* reaction to start the chain reaction,

(2) a *propagation* reaction sequence (2 or more reactions) to keep the reaction going, and
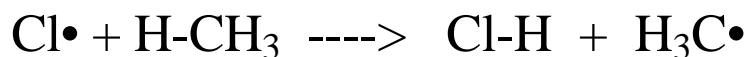
(3) a *termination* step (or several).

## Initiation Reaction - Make Reactive Radicals
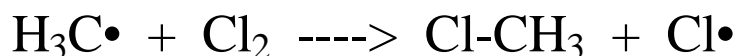
Homolysis of chlorine with heat (300 C) or UV light.

$Cl_2$ ----> 2 Cl•    requires 58 kcal/mol
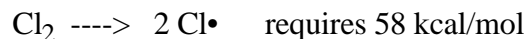
## Propagation Steps I and II - Productice Radical Reactions

Step 1: Cl• abstracts H from methane to form methyl radical and HCl

$$Cl• + H\text{-}CH_3 \ \text{----}> \ Cl\text{-}H \ + \ H_3C•$$

Step 2: $H_3C•$ reacts with $Cl_2$ and forms product and a new Cl•

$$H_3C• \ + \ Cl_2 \ \text{----}> \ Cl\text{-}CH_3 \ + \ Cl•$$

Steps 1 and 2 combined:  Cl• + $CH_4$ + $Cl_2$ ----> $H_3CCl$ + HCl + Cl•
The chlorine atom kind of acts as a catalyst!

The important thing to see here is that the propagation reaction yields the product, the chlorinated alkane and also a new chlorine radical.  That's why the initial generation of radicals needs to be done only for a few molecules.  Each initially formed Cl radical will end up producing lots and lots of product.  The number of cycles done per initial radical is called the **turnover** or **chain length**.

Kinetics of Step 1:  Endothermic by 2 kcal/mol!!
BDE(C-H) = 105 kcal/mol
is higher than
BDE(H-Cl) = 103 kcal/mol formed.
Draw PES picture below

Kinetics of Step 2:  Exothermic by 27 kcal/mol!!
BDE(Cl-Cl) = 58 kcal/mol
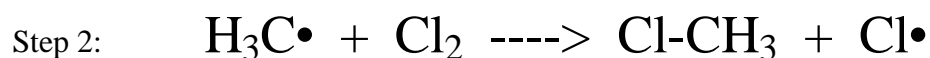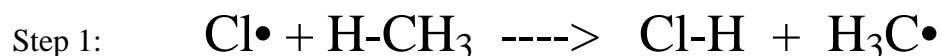is much lower than
BDE(C-Cl) = 85 kcal/mol formed.

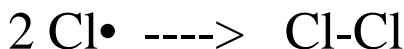Draw PES pictures for both of these steps!

## Chain Termination - Destroy Reactive Radicals

Any reaction that sabotages any of the propagation steps.

All good things must come to an end. Note that *in the propagation steps the radical reacts with a closed shell system* and because of that always a new radical is produced. This works quite well so long as there are many many more closed shell systems then radicals. If that is no longer true, then radicals will bump into radicals. When radicals collide then they will combine to form closed shell systems. There are two types of radicals involved; so we have three different possibilities for their pair collisions. Each possibility happens.
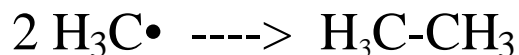
Step 1:     $Cl\bullet + H\text{-}CH_3 \ \text{----}> \ Cl\text{-}H \ + \ H_3C\bullet$

Step 2:     $H_3C\bullet \ + \ Cl_2 \ \text{----}> \ Cl\text{-}CH_3 \ + \ Cl\bullet$

Two chlorine atoms recombine

$$2 \ Cl\bullet \ \text{----}> \ Cl\text{-}Cl$$

Chlorine atom and methyl radical recombine. This forms product but not in the productive way!!

$$Cl\bullet + \ H_3C\bullet \ \text{----}> \ Cl\text{-}CH_3$$

Two methyl radicals combine.

$$2 \ H_3C\bullet \ \text{----}> \ H_3C\text{-}CH_3$$

### Formation of Mixtures

This is a big problem with all of these radical reactions.  That's why they are used mainly in industrial processes (with subsequent separation) but not so often in the lab (chemists are lazy just like anybody else and they rather get around separation problems *in a smart way*).

Note that in the propagation the chlorine reacts with an alkane.  Initially just the alkane is present and there is no problem.  As the reaction goes on for some time, we will have the alkane and also some already chlorinated alkane present.  The chlorine radical is so reactive that it can abstract a H from either one.  So, the initial product can become the substrate  --  the result is polychlorination.  And so on.  The reaction of methane thus yields $H_3CCl$, $H_2CCl_2$, $HCCl_3$, and $CCl_4$.

**Homework:**  Write down all the reaction steps that are involved in the chlorination of methane. (All that is required is to write the propagation steps with chlorinated alkanes.)  Also write down all of the termination steps that are possible with the polychlorinated species.  (Think first about the maximum number of termination reactions.  Then write.)

To prevent the formation of mixtures, we play with the concentrations to affect the statistics. For xeample:  High [alkane]  --->  high yield of monohalogenation.

## Other Halogenations of Methane

Fluorine, chlorine, and bromine all can halogenate alkanes.
Why can iodine not do this reaction?
Why does fluorine do it so explosively.

Does the reason lie with the initiation reaction?

$F_2$      37
$Cl_2$     58  hardest to break!
$Br_2$     46
$I_2$      36  easy to break!  same as $F_2$.

The reason lies with the propagation steps.

*Halogen radicals will <u>not always</u> abstract hydrogens from alkanes.*  The first step is exothermic only for F but not for the others.  The abstraction of a H-atom from methane is getting harder and harder in the order F > Cl > Br > I since the H-X bond strength weakens.

*Alkyl radicals <u>always</u> will cleave the $X_2$ molecule.*  All of the second steps are exothermic. The degree of exothermicity is reduced as the EN of X is decreased.  The second step always is exothermic but is not exothermic enough for the case of iodine.

Chlorination and bromination are the important reactions, and their energy profiles are shown schematically below.

|      | Step 1 | Step 2 | Overall |            |
|------|--------|--------|---------|------------|
| **F**  | -30  | -73    | -103    | Explosive!! |
| **Cl** | +2   | -27    | -25     | OK         |
| **Br** | +18  | -25    | -7      | mild!      |
| **I**  | +34  | -21    | +13     | NO GO!     |

For F, Cl, and Br, the second step provides enough energy for the first to happen and in those cases the overall reaction is thus exothermic.  But why does the reaction go in the first place? Well, a few of the X radicals will always have enough energy to do the reaction.  As they react they will liberate energy and that energy is used to help more X radicals to do the first step (T is increased) ....

## Chlorination of Higher Alkanes:  Relative Reactivity and Selectivity.

**Ethane reacts instead of Methane**

Pretty much the same with one exception:  Propagation Step 2 is now exothermic by 5 kcal/mol.
It requires only 98 kcal/mol to break the CH bond in ethane; 7 less than in methane.

**Propane**

Now there is a choice!!

Regiochemistry!

        Will the methyl group react?          6 H-atoms
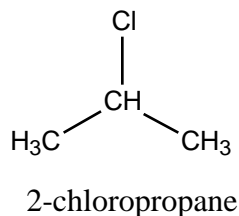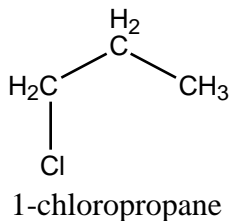
        Will the methylene group react?      2 H-atoms

        Statistics:  methyl:methylene = 3:1

        Actual outcome: methyl:methylene = 43:57

Regiochemistry!

        Will the methyl group react?         Less stable radical.    BDE = 98

        Will the methylene group react?     More stable radical.    BDE = 94.5    Winner



1-chloropropane                    2-chloropropane

<u>Relative Reactivity</u>

The definition of selectivity.

```
yield from methylene / number of methylene hydrogens    57/2

-------------------------------------------------- =  ----   =  4

yield from methyl  /  number of methyl hydrogens        43/6
```
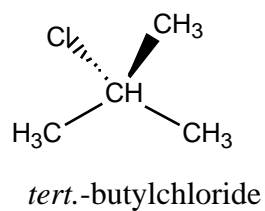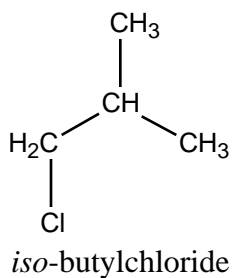
**Butane**

There are two isomers of butane.  The straight chain butane is pretty much like the case discusssed for propane except that threr are now two methylene groups.  The more interesting case is the branched butane.

Two products from branched butane:

        Isobutylchloride from reaction at methyl.  (IUPAC naame?)

        Butylchloride from reaction at tertiary site.  (IUPAC name?)

$CH_3$

$CH$

$H_2C$     $CH_3$

$Cl$

*iso*-butylchloride

$Cl$    $CH_3$

$CH$

$H_3C$    $CH_3$

*tert.*-butylchloride

Statistics:       Isobutyl vs butyl = 9 : 1

Actual:          Isobutyl vs butyl = 64 : 36   (at 25 C)    **Selectivity = 5.1**

Actual:          Isobutyl vs butyl = 80 : 20   (at 600 C)

Higher T reduces the selectivity and the ratio becomes more what statistics will tell you.

## 2-Methylbutane

Four products:

| | | |
|---|---|---|
| Chlorinate at one of the two methyl groups | 6H | 27% |
| Chlorinate at the unique methyl group | 3H | 14% |
| Chlorinate at methylene | 2H | 36% |
| Chlorinate at the tertiary site | 1H | 23% |

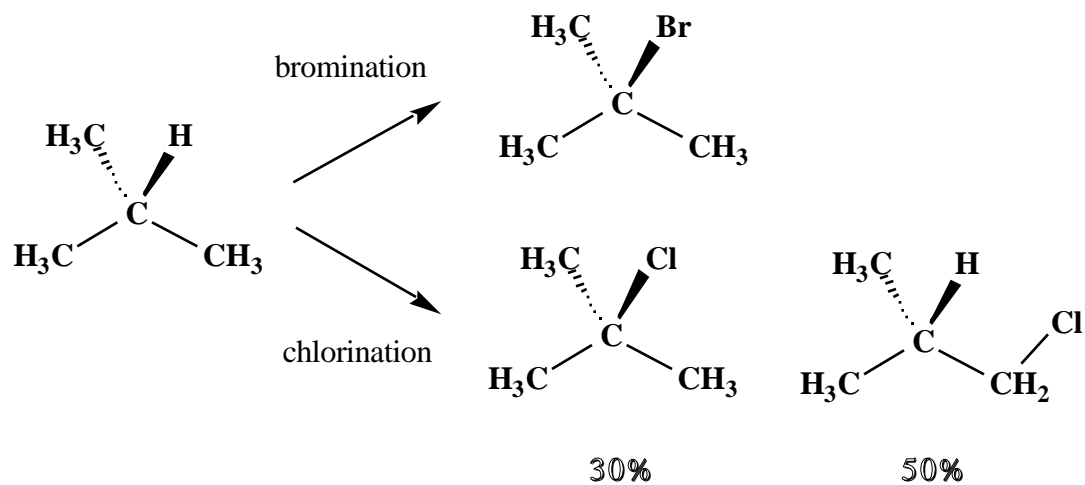Prim : Sec : Tert = (27+14)/9 : 36/2 : 23/1  =about=  1 : 4 : 5

The four products are shown and you can name them for your exercise.
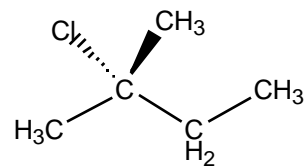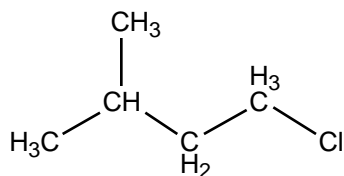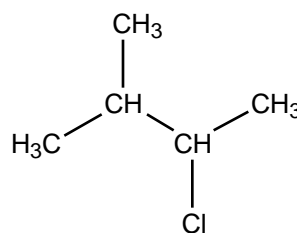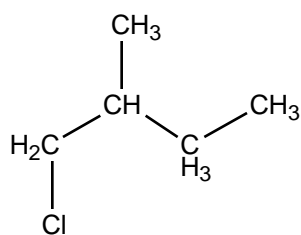
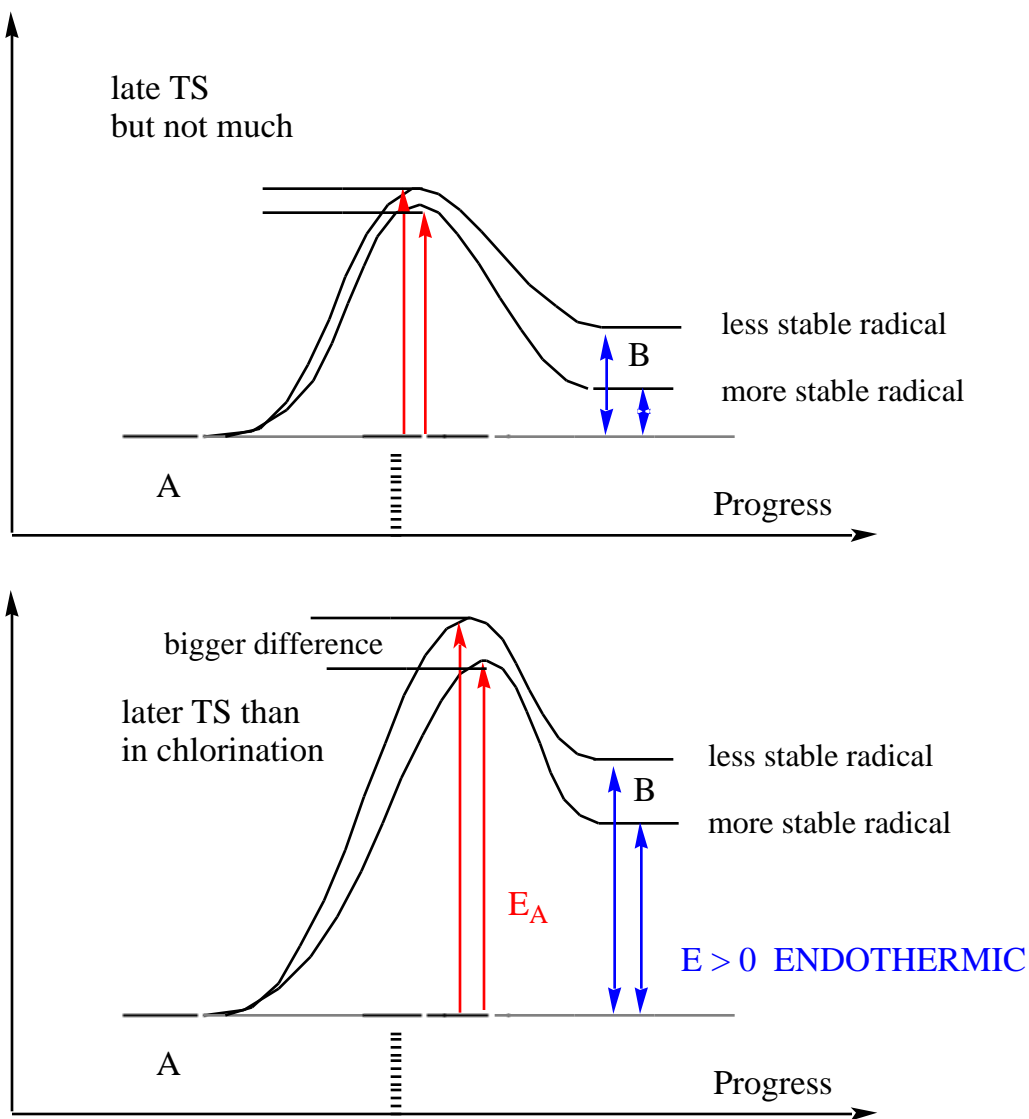## High Selectivity in Radical Halogenation with Bromine
### Tertiary Butane

Bromination of tert.-butane is highly selective producing tert.-butyl bromide > 99% and isobuty
bromide < 1%



The **rate-limiting step** is the H-abstraction from the alkane.  This step is **more endothermic for bromination** then for chlorination.  The potential energy surfaces for the first steps of the chlorination and bromination reactions are shown below to explain the selectivity.

Hammond: Endothermic reactions have late transition states. The late transition states for the two pathways more closely reflect the stability differences of the alternative radicals formed.



late TS
but not much

less stable radical

B

more stable radical

A

Progress



bigger difference

later TS than
in chlorination

less stable radical

B

more stable radical

$E_A$

$E > 0$  ENDOTHERMIC

A

Progress

The transition states leading to the more and less stable radicals both are later than in the case of the chlorination. The radical character of the forming radicals is more fully expressed in these late transition states. The difference in the two transition states equals a larger fraction of the energy difference of the isomeric radicals formed in this step. This is what causes the selectivity, the difference in the relative energies of the two transition state structures that lead to te isomeric products.

## Chlorination and Bromination of Propane

There are two types of Hs and they come in different numbers. Two are at secondary C, six are at primary Cs. Bromination is rather selective. Two products are formed in the chlorination (note that the product formed from the more stable radical dominates although there are fewer Hs of that type to be extracted). As before: The H-abstraction by bromine radical is more endothermic and the TS is late, that is, it more fully reflects the factors that stabilize the radical formed.
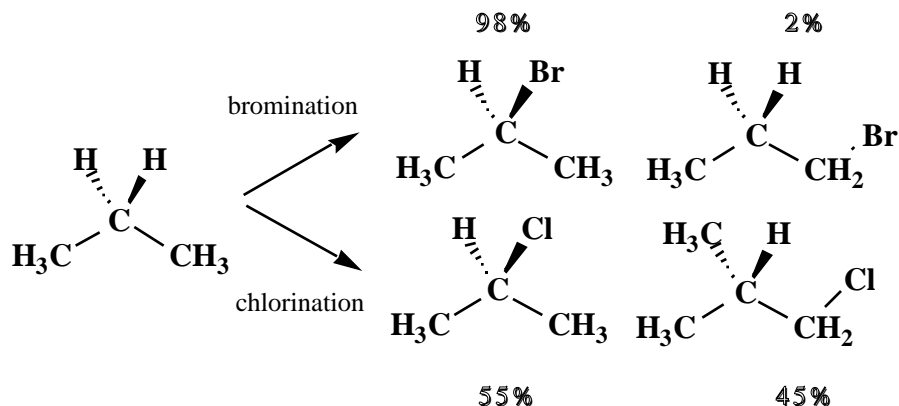
## Toluene and Ethylbenzene

There will not be any halogenation of the aromatic ring. Just the side chain will be halogenated and the rules are just the same as for the alkanes themselves. Toluene gives benzylhalides. Ethylbenzene gives selective bromination of the benzylic carbon. Chlorination of ethylbenzene also gives chlorination of the methyl group as a side product.
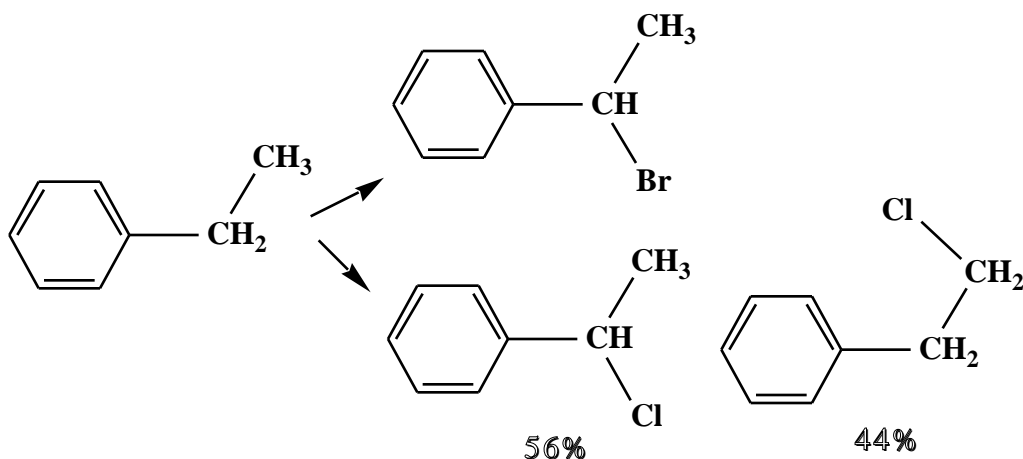
## Synthetic Uses of Radical Halogenations of Alkanes

**Example 1:** Methane

Used to be most important.

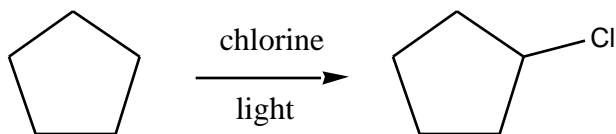To make chlorinated solvents.

> Tetrachloromethane, TETRA
>
> Chloroform, $HCCl_3$
>
> Methylenechloride, $H_2CCl_2$

The more chlorine is used the higher is the probability for polyhalogenation.

**Example 2:** Cyclopentane (using light, in 93% yield)

Only one product possible. No selectivity problems.



**Example 3:** Brominations in the research laboratory

Advantages: Bromine is a liquid, more selective reagent

Solvent: Chlorinated methanes (at least 2 Cl)

## Examples of Radical Reactions

### Allylic Halogenation

This is a matter of concentrations!  Can be done if low concentrations of the $X_2$ are used. A much better way is the reaction with <u>NBS, N-Bromosuccinimide</u>.  NBS produces bromine radicals in a small and constant amount that is just perfectly right to give only the allylic halogenation.  (NBS can also be used for bromination in benzylic positions but it is less needed for those do go as planned anyway.)
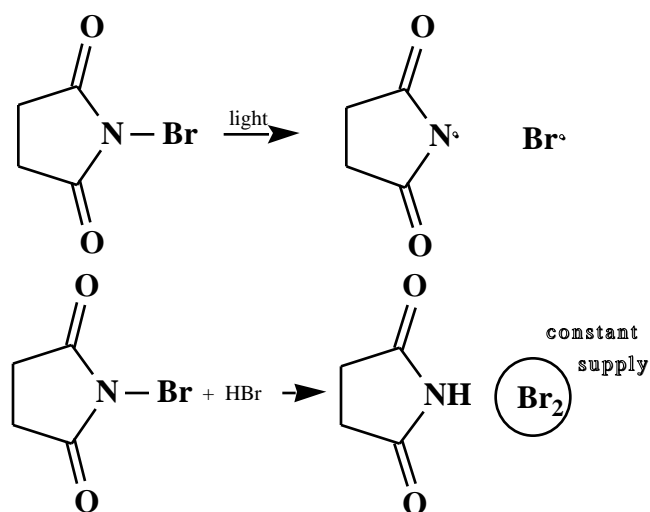


Note that the HBr is produced by the reaction of the Br radical with alkane.

## Pyrolysis

Thermal decomposition of alkanes at high temperatures without oxygen. For example: Butane. If you heat it high enough, then eventually a CC bond will break (initia-tion). Now you have two alkyl radicals. Two things can happen. Recombine the radicals; that is the boring option (a termination step). Use the radical to abstract an H from another butane. This leads to a smaller alkane and a butane radical. What can the butane radical do? Three things: Recombine with some radical. Break a CC bond and produce an olefine and a radical. Abstract a H from some other butane. In the end, you end up with small alkanes and small alkenes.

Homework: Write the equations down. Do the same with propane and pentane until you catch on.

## Oxygen in Autoxidations

The "rancid butter problem". Butter contains fats of various sorts. When exposed to light or heat and air butter goes bad. Why? Oxygen is a biradical and it can do all the types of reactions discussed. "Luckily" oxygen diradical is quite stable, so these reactions occur slowly.

**Initiation:** Not necessary because $O_2$ is a biradical as it is! $O_2$ takes an H and produces an *alkyl radical*.

**Propagation:** The alkyl radical adds an oxygen biradical to produce an hydroperoxide monoradical. The generated radical abstracts an H from the next fatty acid to produce a hydroperoxide and an *alkyl radical*. The propagation is carried on by the *alkyl radical*.

**Termination:** Any combination of radicals present.

The overall reaction is the insertion of $O_2$ into CH bonds to form hydroperoxides. These products are quite reactive themselves and go on to do many things such as producing acids (the bad smell is butanoic acid).

Other autooxidations include (i) the autooxidation of ethers (very dangerous!!) and (ii) the autooxidation of aldehydes to acids (peroxyacetic acid reacts with aldehyde to give to molecules of the acid.



from ether

from aldehyde
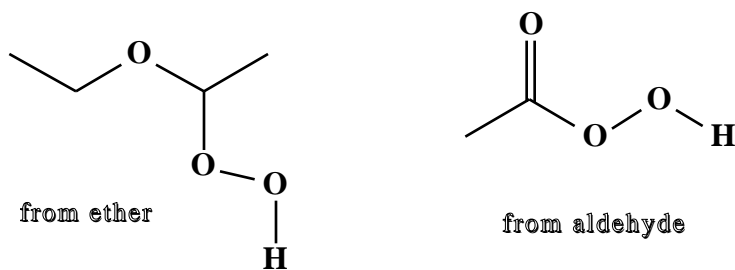
## Cumene Process

Very important process to produce phenol and acetaldehyde in large quantities.  The first step is the formation of a hydroperoxide by radical oxidation in air (requires heat).  Addition of acid protonates the terminal OH.  Water gets eliminated and the phenyl group moves to the remaining O.  The rearranged cation adds water, and loses a proton, and kicks out a phenoxide.
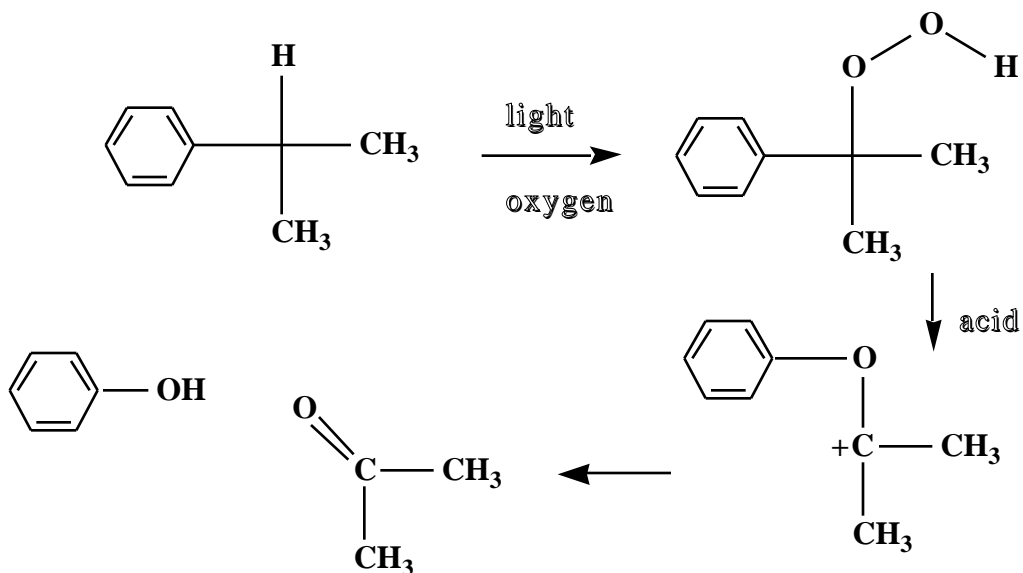
## Initiators and Inhibitors

**A free radical initiator is** anything that can produce radicals that can start a radical chain reaction. For the halogenations, *light* is the initiator. Light cleaves $X_2$ to produce two $X\cdot$ radicals. Chemicals with very small bond dissociation energies are often used. These just need to be heated to from radicals. Some of the more often used ones are

 (i) Benzoyl peroxide(ii) Peroxybenzoic acid        (iii) Azobisisobutyronitrile (AIBN).
Their structures and decomposition reactions are shown on the following page.

**An inhibitor is** anything that converts a reactive radical into a stable nonreactive radical. Most usually such compounds are phenols. The products are radicals that are very much stabilized by resonance. These compounds "trap" radicals. If they trap radicals that would otherwise lead to autooxidation, then they will prevent the oxidation. Thus, we call such inhibitors "antioxidants". Antioxidants are important as food preservatives. Let's look at how phenol acts as an inhibitor and then let's look at some specific phenols that are used as food preservatives.

Common Initiators of Free Radical Reactions
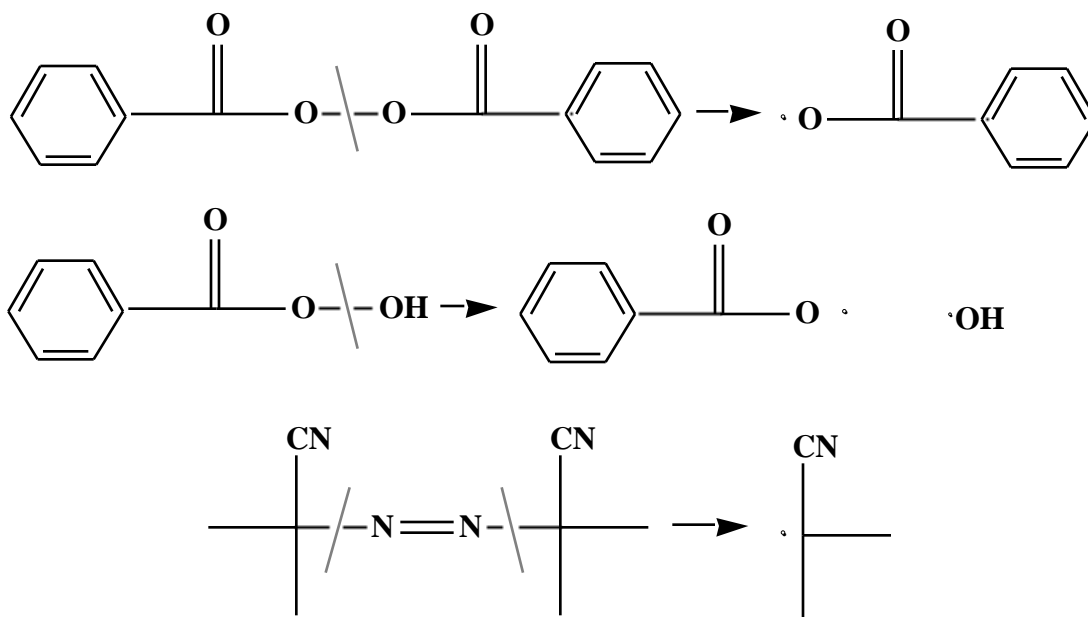
## The Phenol Trap

The phenol H is easily abstracted by a reactive radical. The reactive radical becomes unreactive in the process and the phenoxy radical is formed. The phenoxy radical is very stable. Many resonance forms. (Show them).

## BHT and Vitamin E

Synthetic and natural phenol based preservatives.



BHT
Butylated hydroxytoluene

Vitamin E;    tocopherol
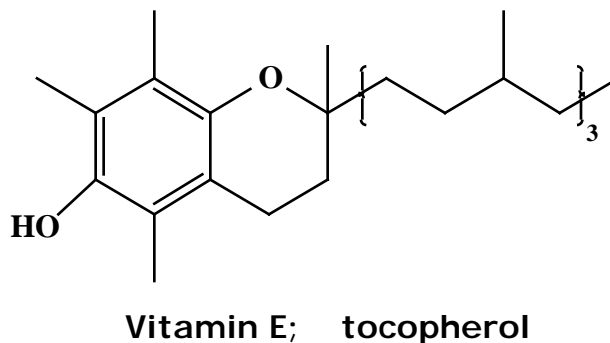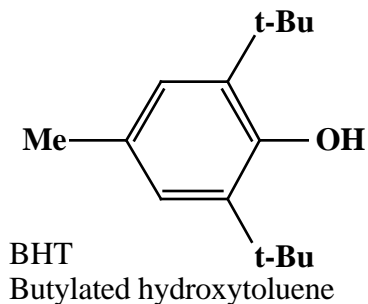
## Arylamines

Aromatic amines function just like aromatic alcohols. The reactive radical picks up the amino-H and leaves an unreactive resonance-stabilized radical. The N-phenyl-2-naphthylamine is used in rubber products. There are two aromatic rings to use for resonance. Write down the resonance forms (for fun).
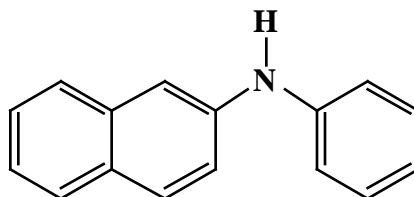
## Polymers

**Polymers** are large molecules with a periodically occurring a **repeating unit.** The **monomer** is the molecule that is polymerized, it characterizes the repeating unit. In some cases, the repeating unit may contain several monomers (co-polymerization). "Polymers" really are not one kind of molecule but they are mixtures of a variety of polymeric molecules with different chain lengths.

Polymers are classified depending on (a) their properties and on (b) the types of reactions that are used to make them. Among the types of reactions, we differentiate between **addition reaction** and **condensation reactions**. Each of these may occur with just one component or with several components (co-polymerization). **For the radical chain polymerizations, we usually deal with the simple case of an addition reaction of one monomer.**

A few examples are shown below. All of these reactions are radical chain reactions and, as such, they need to be initiated. Any of the discussed initiators together with heat (or light) will do. Say we use a peroxide as the initiator, then the reaction mechanism is as follows:

====================================================================

*Initiation:*     RO· adds to an olefine and produces a C radical.

*Propagation:*   The C radical adds to another olefine and produces a new C radical that is one
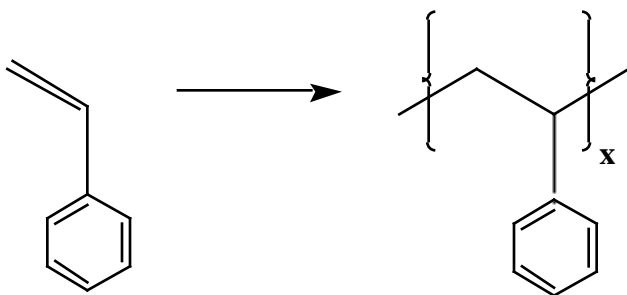                 repeating unit longer. And so on.

*Termination:*   Two C radicals will combine (not necessarily of the same length)

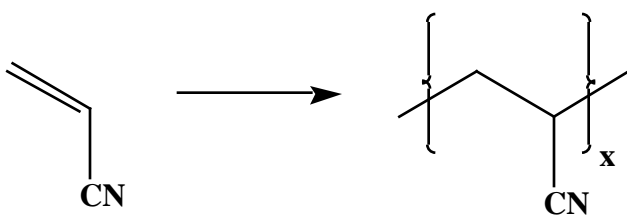====================================================================

Note that the end groups are different because of the radical starter. The propagation step is interesting with regard to stereochemistry. We will look at the stereochemistry of the polystyrene formation.
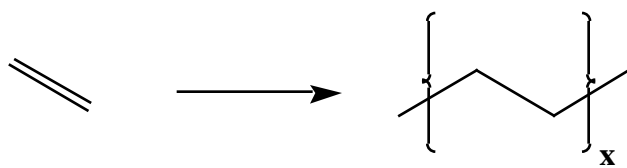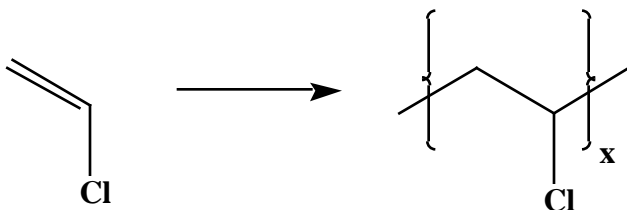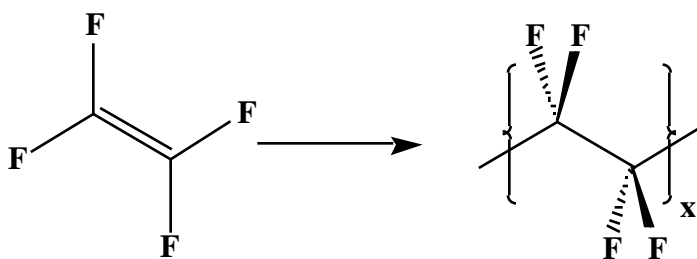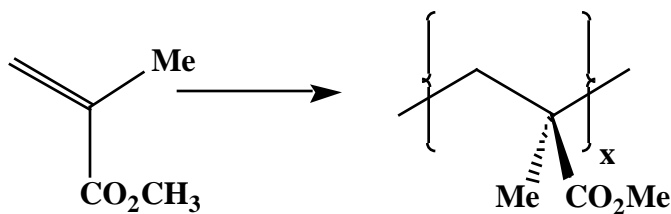
Polyethylene

Polystyrene

Orlon
Polyacryl
Polyacrylnitrile

PVC
Polyvinylchloride

Teflon
Polyperfluoroethylene

Plexiglas
polymethylmethacrylate

## Regiochemistry of the Polystyrene Formation  -  Head-to-Tail

The radical chain starter radical (RO• ) may do either of two things:  (a) add to the **head** or (b) add to the **tail**.  The head is the more substituted C of styrene.  Which is going to happen? Approach (a) leads to a 1º radical, approach (b) leads to a benzyl radical.  Approach (b) is preferred of course!  The same choice occurs for every reaction in the propagation and in each case we will form the more stable benzyl cation.  Thus, the phenyl groups will all end up in 1,3 positions.  In other words, all monomers are connected head to tail.

## Stereochemistry in Polypropylene  -  Tacticity

Aside from the regiochemistry, there is also a stereochemistry problem in the polymerizations of the substituted olefines.  The substituted C in the olefine becomes chiral in the polymerization.  Will it always be one configuration, or the other, or a mixture?  The answer is:  Polymer chemists today are so good at what they are doing, they can find conditions to achieve either one of these options!!  The resulting polymers are called **atactic** (random), **isotactic** (the same), and **syndiotactic** (alternating).
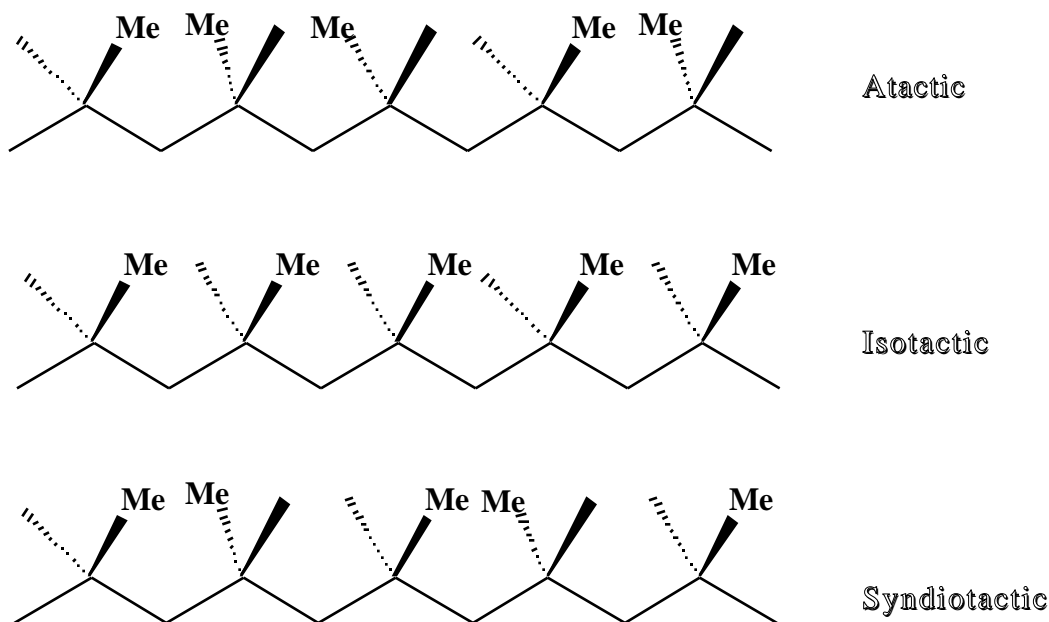
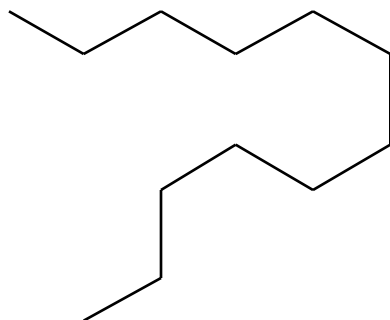**Secondary Structure - The Next Level of Stereochemistry**

This is a matter of **conformation**. For each of the CC bonds along the chain, there is a choice of *gauche* and *trans*. For the small molecules, we would say trans is clearly better and the reason is that we want to minimize the repulsions between the 1,2 substituents. That is still true in the polymers. However, in the polymers there might be other factors that would favor some gauche conformations to occur as well. What could these factors be? Two are important: (1) Entropy. The more disorder, the better. (2) Energy. Changing to a few gauche conformation surely will increase the energy somewhat but the structural change might result in a larger gain of energy. For example, if a polymer is all-trans, it will not get any vdW stabilization other than that which comes from other molecules. If, on the other hand, the molecule is not linear, then it will get *vdW stabilization from itself!*.(and from others).

Crystalline polymers prefer all-*trans*.
Non-crystalline polymers have both, *gauche* and *trans*.
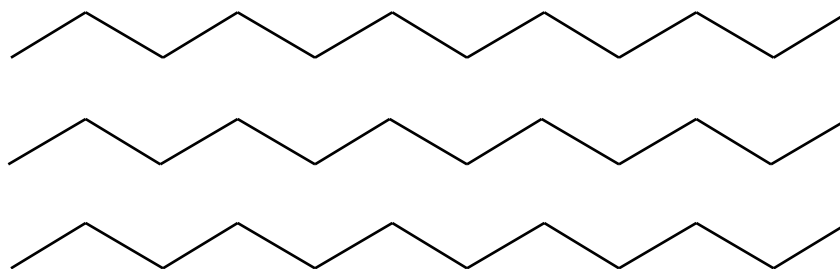
(micra) crystalline

noncrystalline