Organic Chemistry-2

Semester-2, CBCS

Course: CEMA CC-2-3-TH

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Recommended texts:

**Halogenation of alkanes:**

**Introduction:**

Halogenation: Any reaction or process in which a halogen (fluorine, chlorine, bromine, or iodine) and no other elements are introduced into a molecule.

Under the influence of ultraviolet light or at high temperatures (250-400 degrees C), chlorine or bromine convert alkanes into chloroalkanes or bromoalkanes respectively; an equivalent of hydrogen chloride or hydrogen bromide is formed at the same time.

This is a substitution reaction where a hydrogen atom is replaced by a halogen atom:

\[
R - H + X - X \xrightarrow{\text{light or heat}} R - X + H - X
\]

# Chlorine is found to be more reactive than bromine in this case.
# The reaction is generally not very useful for fluorine and iodine.
# Fluorination is difficult to control as the reaction is strongly exothermic and fluorine is an extremely aggressive oxidising agent.
# Iodination, on the other hand, is unique among all these because of its endothermic nature. As such, iodine is least reactive of the four and iodoalkanes are never prepared in this way.

In general, chlorination is difficult to stop at the monohalogenation stage, if there is present more than one hydrogens that can be replaced.

For example, chlorination of methane affords a mixture of products. The reaction is thus synthetically less important, but industrially appealing as this mixture of chloroalkanes is used in various purposes:

\[
\text{CH}_4 + \text{Cl}_2 \xrightarrow{\text{hv, room temp.}} \text{CH}_3\text{Cl} + \text{CH}_2\text{Cl}_2 + \text{CHCl}_3 + \text{CCl}_4
\]

This polyhalogenation can somehow be managed if an excess of the starting alkane is used.

However, if there is present in the alkane more than one type of hydrogens that can be replaced, a mixture of different regiosomeric haloalkanes is obtained even for monohalogenation reactions:

\[
\text{Me} - \text{H}_a\text{Me} - \text{H}_b \xrightarrow{\text{Cl}_2, \text{hv, room temp.}} \text{Me} - \text{Cl} + \text{Me} - \text{Cl}_2 \text{or, t-butyl chloride} (33\%) + \text{Me} - \text{Cl}_2 \text{or, isobutyl chloride} (67\%)
\]

\[
\text{Me} - \text{H}_a\text{Me} - \text{H}_b \xrightarrow{\text{Br}_2, \text{hv, 127°C}} \text{Me} - \text{Br} + \text{Me} - \text{Br}_2 \text{or, t-butyl bromide (>99%)} + \text{Me} - \text{Br}_2 \text{or, isobutyl bromide (<1%)}
\]

Ha → tertiary H, attached to a tertiary centre (1 such)  
Hb → primary H, attached to a primary centre (9 such)
Halogenation of alkanes:

Introduction:

Two things are immediately apparent from the above scheme -

i) Chlorination reaction takes place under much milder condition than bromination does (look at the temperatures required for the two reactions).

ii) Bromination is much better in selectively preparing one regioisomer, the more substituted bromoalkane.

In other words, chlorine is definitely more reactive than bromine, so you can do the chlorination in room temperature. But, there is a trade-off. The higher reactivity of chlorine brings with it the burden of less selectivity. Bromination is more selective in nature in the sense that it affords mainly one of the two possible monohalogenated products with a very small amount of the other regioisomer, while chlorination is less so, providing a mixture where both compounds are present in significant amount, thereby demanding a rigorous separation after the halogenation process.

Chlorination proceeds more rapidly, i.e. chlorine is more reactive but less selective because it forms both regioisomers in appreciable amounts.

Bromination proceeds slowly, but forms one regioisomer much more selectively than others, i.e. bromine is more selective but less reactive.
Halogenation of alkanes:

**Mechanism of halogenation:**

We will take chlorination of methane as our working model here, considering only monochlorination.

\[
\text{H}_3\text{C} - \text{H} + \text{Cl} - \text{Cl} \xrightarrow{\text{heat or light}} \text{H}_3\text{C} - \text{Cl} + \text{H} - \text{Cl}
\]

There are certain observations for this halogenation reaction which strongly indicates that a radical mechanism operates when alkanes are halogenated under photochemical or thermal condition. These are outlined below:

i) When we mix chlorine and methane, and keep in the dark, without heating, no reaction takes place,

\[
\text{H}_3\text{C} - \text{H} + \text{Cl} - \text{Cl} \xrightarrow{\text{in dark, r.t.}} \text{X} \rightarrow \text{H}_3\text{C} - \text{Cl} + \text{H} - \text{Cl}
\]

ii) in the dark, reaction can be made to happen upon heating the mixture,

\[
\text{H}_3\text{C} - \text{H} + \text{Cl} - \text{Cl} \xrightarrow{\text{heat, in dark}} \text{H}_3\text{C} - \text{Cl} + \text{H} - \text{Cl}
\]

iii) this mixture, when exposed to light of the appropriate wavelength, reacts readily even without any heating,

\[
\text{H}_3\text{C} - \text{H} + \text{Cl} - \text{Cl} \xrightarrow{\text{r.t., in sunlight}} \text{H}_3\text{C} - \text{Cl} + \text{H} - \text{Cl}
\]

iv) the light which is found to initiate the chlorination is shown to independently induce a homolytic cleavage of a the weak (why weak?) Cl-Cl bond,

v) For each photon absorbed by the reaction mixture, millions of chloromethane molecules are produced, thus indicating a self-sustaining, cyclic mechanism typically found in radical reactions,

vi) known radical scavengers like molecular oxygen, if present, is shown to inhibit the reaction, strongly suggesting the presence of radical intermediates in the reaction medium that get quenched by these radical eaters. Molecular oxygen is triplet in its ground state and behaves like a diradical, eager to capture any radicals around. Each molecule of oxygen is capable of reacting with two radicals. As a result, radical reactions cannot proceed rapidly until all the available oxygen has been consumed.
Halogenation of alkanes:

Mechanism of halogenation:

Based upon the observations, and other experimental findings, the following mechanism of chlorination for methane is proposed. It is applicable to all photochemical or thermal halogenations of alkanes. Typical to any radical mechanisms, this one also has three phases:

i) Initiation step: New radicals are produced from non-radical precursors via photochemically or thermally induced homolysis of typically weak bonds:

\[
\text{Cl} - \text{Cl} \xrightarrow{\text{hv or heat}} \text{Cl}^\cdot + \text{Cl}^\cdot
\]

Alternative description:

\[
\text{Cl} - \text{Cl} \xrightarrow{\text{hv or heat}} \text{Cl}^\cdot + \text{Cl}^\cdot
\]

(showing only one fishhook arrow, the other is implied)

Note how two radicals emerge from a molecule of chlorine. Also note the use of the fishhook arrows (with half arrowheads) to delineate the movement of the single electron. A useful short cut when illustrating a radical reaction is to draw only the alternate arrows since each implies the existence of its partner in making or breaking of a bond, but conventionally, it is better that you use the full complement of these arrows to understand the symmetrical separation in the breaking of the Cl-Cl bond, in contrast to ionic reactions where there is definite sense of direction to the movement of the electrons.
Halogenation of alkanes:

Mechanism of halogenation:

ii) Propagation steps: In each of these steps, one radical is consumed while another one is produced, and more importantly, the one radical that is formed has to be used as a reactant in one of the subsequent propagation steps. This way, a cyclic mechanism starts to operate and there remains no further need for the initiation to occur. It, however, does occur occasionally and each time a new radical cycle spawns from the radicals that are generated from that initiation step.

The point is, propagation steps are self-sustaining and once set in motion, they can continue even without the initiation steps happening anymore. The propagation steps for chlorination of methane are:

- **propagation step-1 (hydrogen abstraction)**
  \[ \text{H}_3\text{C}-\text{H} + \cdot\text{Cl} \rightarrow \text{H}_3\text{C}^\cdot + \text{H}^-\text{Cl} \]

- **propagation step-2 (chlorine abstraction)**
  \[ \text{H}_3\text{C}^- + \text{Cl}^-\text{Cl} \rightarrow \text{H}_3\text{C}-\text{Cl} + \cdot\text{Cl}^-\text{Cl} \]

Notice that the propagation step-1 consumes a chlorine radical while step-2 regenerates a chlorine radical. In this way, one chlorine radical can ultimately cause thousands of molecules of methane to be converted into chloromethane (assuming enough Cl\(_2\) is present). Therefore, the reaction is called a *chain reaction*.

The cyclic nature of the mechanism is illustrated below:

Also note that propagation step-1 accounts for one of the products, HCl, and step-2 produces the other, methyl chloride.
If we sum up the two propagation steps, we have the overall reaction:

**propagation step-1**
\[ \text{H}_3\text{C}-\text{H} + \cdot\text{Cl} \rightarrow \text{H}_3\text{C}^\cdot + \text{H}^-\text{Cl} \]

**propagation step-2**
\[ \text{H}_3\text{C}^- + \text{Cl}^-\text{Cl} \rightarrow \text{H}_3\text{C}-\text{Cl} + \cdot\text{Cl}^-\text{Cl} \]

This now provides us with a new, refined definition for propagation steps. Specifically, the sum of the propagation steps gives the net chemical reaction. All other steps must be either initiation or something else (as we'll see in a while), not propagation.
Halogenation of alkanes:

Mechanism of halogenation:

The question is, how long does the propagation steps go on? Although a huge number of propagation steps follow after each initiation step, radical chain reactions cannot continue \textit{ad infinitum}. At some point of time, like all things, they must also end. The curtain call for a radical process are the termination steps.

Termination steps: radicals are, in general, highly reactive species. They run around, looking for partners to react with. And because of the high reactivity, they react with the first things they collide with. They do not have the patience to wait around for another radical to show up. In other words, a radical-radical collision is unlikely. But be that as it may, it is not impossible. And as the concentrations of the chlorine and methane gradually dwindle, the chances of radical seeing another radical increases. When they do meet, they \textit{quench} each other, that is to say, no new radicals are produced.

Several possibilities exist:

\[
\begin{align*}
\text{Initiation:} & & \text{Propagation:} & & \text{Termination:} \\
\text{Cl} + \text{Cl} & \xrightarrow{\text{hv or heat}} & \text{Cl}^\cdot + \text{Cl}^\cdot & & \text{Cl} + \text{Cl} & \xrightarrow{\text{hv or heat}} & \text{Cl}^\cdot + \text{Cl}^\cdot \\
\text{H}_3\text{C} + \text{Cl} & \rightarrow & \text{H}_3\text{C}^\cdot - \text{Cl} & & \text{H}_3\text{C} + \text{Cl} & \rightarrow & \text{H}_3\text{C}^\cdot + \text{HCl} \\
\text{H}_3\text{C} + \text{CH}_3 & \rightarrow & \text{H}_3\text{C} - \text{CH}_3 & & \text{H}_3\text{C} + \text{CH}_3 & \rightarrow & \text{H}_3\text{C} - \text{CH}_3 \\
\text{Cl}^\cdot + \text{Cl} & \rightarrow & \text{Cl} - \text{Cl} & & \text{Cl} + \text{Cl} & \rightarrow & \text{Cl} - \text{Cl} \\
\end{align*}
\]

As these reactions produce stable molecules and no new radicals, they signify the end of the chain reaction sequence.

As a general guideline, the propagation stage in a radical chain reaction such as chlorination of methane will occur approximately $10^8$ times after each initiation reaction before a termination step will occur. Therefore, it is easy to see why a few photons of light or a small input of energy is necessary to get a large extent of reaction.
Halogenation of alkanes:

Thermodynamics of halogenation of methane:

In the previous section, we explored the accepted mechanism for the chlorination of methane. We will now investigate how this reaction can be accomplished with other halogens as well. How easy, or difficult it is to achieve a radical fluorination, bromination, or iodination? To answer this question, we must look closely into the thermodynamic aspects of halogenation.

The quantity that determines whether or not a reaction is thermodynamically favourable is the $\Delta G^0$ (the change in Gibbs free energy between the reactants and the products). Recall that in order for a reaction to favour products over the reactants, the reaction must exhibit a negative $\Delta G^0$.

$$\Delta G^0 < 0 \quad \Rightarrow \quad K_{eq} > 1 \quad \Rightarrow \quad [P]_e > [R]_e$$

If $\Delta G^0$ is positive, the reactants will be favoured, and the reaction will not produce the desired products.

$$\Delta G^0 > 0 \quad \Rightarrow \quad K_{eq} < 1 \quad \Rightarrow \quad [P]_e < [R]_e$$

Let us use this information to determine whether halogenation can be accomplished with halogens other than chlorine. Recall that $\Delta G^0$ is comprised of two components - enthalpy and entropy:

$$\Delta G^0 = \Delta H^0 - T\Delta S^0$$

For halogenation of an alkane, the entropy component is assumed to be negligible because two molecules of reactants are converted into two molecules of products, i.e. the number of molecules on the reactant and product side are the same.

$$\text{H}_3\text{C}-\text{H} + \text{Cl}-\text{Cl} \xrightarrow{\text{heat or light}} \text{H}_3\text{C}-\text{Cl} + \text{H}-\text{Cl}$$

Since the entropy factor is neglected, we can assess the value of $\Delta G^0$ by analysing the enthalpy term alone.

Recall that we can calculate the enthalpy change of a reaction using the formula

$$\Delta H^0 = \sum \text{BDE}_\text{(bonds broken)} - \sum \text{BDE}_\text{(bonds formed)}$$

Therefore, to calculate the overall enthalpy change for a reaction:

$$\text{H}_3\text{C}-\text{H} + \text{X}-\text{X} \rightarrow \text{H}_3\text{C}-\text{X} + \text{H}-\text{X}$$

for $\text{X} = \text{F, Cl, Br, I}$

we are in need of the BDE values of the respective bonds.
Halogenation of alkanes:

Thermodynamics of halogenation of methane:

Calculating the enthalpy change for monohalogenation of methane:

Recall that we can calculate the enthalpy change of a reaction using the formula

$$\Delta H^0 = \Sigma \text{BDE}_\text{(bonds broken)} - \Sigma \text{BDE}_\text{(bonds formed)}$$

Therefore, to calculate the overall enthalpy change for a reaction:

$$\text{H}_3\text{C-H} + \text{X-X} \rightarrow \text{H}_3\text{C-X} + \text{H-X}$$

we are in need of the BDE values of the respective bonds.

These are tabulated below:

<table>
<thead>
<tr>
<th>Bond dissociation energies of relevant bonds (in kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
</tr>
<tr>
<td>--------</td>
</tr>
<tr>
<td>F</td>
</tr>
<tr>
<td>Cl</td>
</tr>
<tr>
<td>Br</td>
</tr>
<tr>
<td>I</td>
</tr>
</tbody>
</table>

Using these data, we can make the following calculation (do it):

$$\text{H}_3\text{C-H} + \text{X-X} \rightarrow \text{H}_3\text{C-X} + \text{H-X}$$

<table>
<thead>
<tr>
<th>X</th>
<th>$\Delta H^0$</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>- 431</td>
</tr>
<tr>
<td>Cl</td>
<td>- 104</td>
</tr>
<tr>
<td>Br</td>
<td>- 33</td>
</tr>
<tr>
<td>I</td>
<td>+ 55</td>
</tr>
</tbody>
</table>

These data support what we have discussed already.

All of the processes above have a negative value for $\Delta H^0$ and are exothermic except for iodination.

Iodination of methane has a positive $\Delta H^0$, which means that $\Delta G^0$ will also be positive for that reaction. As a result, iodination is not thermodynamically favorable, and the reaction simply does not occur, unless special measures are taken, like oxidizing the product HI with HNO$_3$ to I$_2$, whereby, the reactant alkane and I$_2$ reacts further to compensate for the loss of one of the products and accumulation of one of the reactants (Le Chatelier's principle).

The other halogenation reactions are all thermodynamically favourable, but fluorination is so exothermic that the reaction is too violent to be of practical use. Therefore, only chlorination and bromination are practical in the laboratory.