Alkyl halides (also known as haloalkanes) are compounds in which one or more hydrogen atoms in an alkane have been replaced by halogen atoms (fluorine, chlorine, bromine or iodine). We will only look at compounds containing one halogen atom. For example:

alkyl halides fall into different classes depending on how the halogen atom is positioned on the chain of carbon atoms. There are some chemical differences between the various types.

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**Primary alkyl halides**

In a primary (1°) halogenoalkane, the carbon which carries the halogen atom is only attached to one other alkyl group. Some examples of primary alkyl halides include:

\[
\begin{align*}
\text{CH}_3 - &\text{CH}_2 - \text{Br} & \text{CH}_3\text{CH}_2 - &\text{CH}_2 - \text{Cl} & \text{CH}_3\text{CH}_2 - &\text{CH}_2 - \text{I} \\
\end{align*}
\]

Notice that it doesn’t matter how complicated the attached alkyl group is. In each case there is only one linkage to an alkyl group from the CH₂ group holding the halogen. There is an exception to this: CH₃Br and the other methyl halides are often counted as primary alkyl halides even though there are no alkyl groups attached to the carbon with the halogen on it.

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**Secondary alkyl halides**

In a secondary (2°) halogenoalkane, the carbon with the halogen attached is joined directly to two other alkyl groups, which may be the same or different. Examples:

\[
\begin{align*}
\text{CH}_3 - &\text{CH} - \text{CH}_3 & \text{CH}_3 - &\text{CH} - \text{CH}_2\text{CH}_3 \\
\text{Br} & & \text{Cl} &
\end{align*}
\]

---

**Tertiary alkyl halides**

In a tertiary (3°) halogenoalkane, the carbon atom holding the halogen is attached directly to three alkyl groups, which may be any combination of same or different. Examples:

\[
\begin{align*}
\text{CH}_3 - &\text{C} - \text{CH}_3 & \text{CH}_3 - &\text{C} - \text{CH}_2\text{CH}_3 \\
\text{Br} & & \text{Cl} &
\end{align*}
\]
Physical properties of alkyl halides

Boiling Points

The chart shows the boiling points of some simple alkyl halides.

Notice that three of these have boiling points below room temperature (taken as being about 20°C). These will be gases at room temperature. All the others you are likely to come across are liquids. Remember:

• the only methyl halide which is a liquid is iodomethane;
• chloroethane is a gas.

The patterns in boiling point reflect the patterns in intermolecular attractions.

van der Waals dispersion forces

These attractions get stronger as the molecules get longer and have more electrons. That increases the sizes of the temporary dipoles that are set up. This is why the boiling points increase as the number of carbon atoms in the chains increases. Look at the chart for a particular type of halide (a chloride, for example). Dispersion forces get stronger as you go from 1 to 2 to 3 carbons in the chain. It takes more energy to overcome them, and so the boiling points rise.

The increase in boiling point as you go from a chloride to a bromide to an iodide (for a given number of carbon atoms) is also because of the increase in number of electrons leading to larger dispersion forces. There are lots more electrons in, for example, iodomethane than there are in chloromethane - count them!

van der Waals dipole-dipole attractions

The carbon-halogen bonds (apart from the carbon-iodine bond) are polar, because the electron pair is pulled closer to the halogen atom than the carbon. This is because (apart from iodine) the halogens are more electronegative than carbon. The electronegativity values are:

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronegativity</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>2.5</td>
</tr>
<tr>
<td>F</td>
<td>4.0</td>
</tr>
</tbody>
</table>
This means that in addition to the dispersion forces there will be forces due to the attractions between the permanent dipoles (except in the iodide case). The size of those dipole-dipole attractions will fall as the bonds get less polar (as you go from chloride to bromide to iodide, for example). Nevertheless, the boiling points rise! This shows that the effect of the permanent dipole-dipole attractions is much less important than that of the temporary dipoles which cause the dispersion forces. The large increase in number of electrons by the time you get to the iodide completely outweighs the loss of any permanent dipoles in the molecules.

Example 1: Boiling Points of Some Isomers

The examples show that the boiling points fall as the isomers go from a primary to a secondary to a tertiary halogenoalkane. This is a simple result of the fall in the effectiveness of the dispersion forces.

\[
\begin{align*}
\text{CH}_3\text{CH}_2\text{CH}_2\text{Br} & \quad 275 \text{ K} \\
\text{CH}_3\text{CH}_2\text{CH}_3 & \quad 364 \text{ K} \\
\text{CH}_3\text{C-CH}_2\text{Br} & \quad 346 \text{ K}
\end{align*}
\]

The temporary dipoles are greatest for the longest molecule. The attractions are also stronger if the molecules can lie closely together. The tertiary halogenoalkane is very short and fat, and won’t have much close contact with its neighbours.

Solubility

Solubility in water

The alkyl halides are at best only slightly soluble in water. For a halogenoalkane to dissolve in water you have to break attractions between the halogenoalkane molecules (van der Waals dispersion and dipole-dipole interactions) and break the hydrogen bonds between water molecules. Both of these cost energy.

Energy is released when new attractions are set up between the halogenoalkane and the water molecules. These will only be dispersion forces and dipole-dipole interactions. These aren’t as strong as the original hydrogen bonds in the water, and so not as much energy is released as was used to separate the water molecules. The energetics of the change are sufficiently "unprofitable" that very little dissolves.

Note

This energetic picture of solvation is not the whole story and entropic factors must also be considered to understand solvation properly.
Solubility in organic solvents

alkyl halides tend to dissolve in organic solvents because the new intermolecular attractions have much the same strength as the ones being broken in the separate halogenoalkane and solvent.

Chemical Reactivity

The pattern in strengths of the four carbon-halogen bonds are:

Notice that bond strength falls as you go from C-F to C-I, and notice how much stronger the carbon-fluorine bond is than the rest. To react with the alkyl halides, the carbon-halogen bond has got to be broken. Because that gets easier as you go from fluoride to chloride to bromide to iodide, the compounds get more reactive in that order. Iodoalkanes are the most reactive and fluoroalkanes are the least. In fact, fluoroalkanes are so unreactive that we shall pretty well ignore them completely from now on in this section!

The influence of bond polarity

Of the four halogens, fluorine is the most electronegative and iodine the least. That means that the electron pair in the carbon-fluorine bond will be dragged most towards the halogen end. Looking at the methyl halides as simple examples:

The electronegativities of carbon and iodine are equal and so there will be no separation of charge on the bond.

One of the important set of reactions of alkyl halides involves replacing the halogen by something else - substitution reactions. These reactions involve either:

- the carbon-halogen bond breaking to give positive and negative ions. The ion with the positively charged carbon atom then reacts with something either fully or slightly negatively charged.
- something either fully or negatively charged attracted to the slightly positive carbon atom and pushing off the halogen atom.

You might have thought that either of these would be more effective in the case of the carbon-fluorine bond with the quite large amounts
of positive and negative charge already present. But that's not so - quite the opposite is true! The thing that governs the reactivity is the strength of the bonds which have to be broken. If is difficult to break a carbon-fluorine bond, but easy to break a carbon-iodine one.

Contributors

• Jim Clark (Chemguide.co.uk)