Heat of hydrogenation (symbol: \( \Delta H_{\text{hydro}} \), \( \Delta H^\circ \)) of an alkene is the standard enthalpy of catalytic hydrogenation of an alkene. Catalytic hydrogenation of an alkene is always exothermic. Therefore, heat of hydrogenation of alkenes is always negative.

eg:

\[
\text{CH}_2=\text{CHCH}_2\text{CH}_3 + \text{H}_2 \xrightarrow{\text{catalyst, transition metal}} \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3
\]

Standard enthalpy of this reaction is -30.3 kcalmol\(^{-1}\). Thus, heat of hydrogenation of 1-butene is -30.3 kcalmol\(^{-1}\).

Heat of hydrogenation of alkenes is a measure of the stability of carbon-carbon double bonds. All else being the same, the smaller the numerical value of heat of hydrogenation of an alkene, the more stable the double bond therein. Based on heats of hydrogenation of alkenes, the trend in the stability of carbon-carbon double bonds is tetrasubstituted > trisubstituted > disubstituted > monosubstituted > unsubstituted.

Heat of hydrogenation of alkenes is additive, provided that the double bonds are not conjugated.

eg:

<table>
<thead>
<tr>
<th>alkene</th>
<th>( \Delta H^\circ ), kcalmol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="1-butene" /></td>
<td>-30.3</td>
</tr>
<tr>
<td><img src="image" alt="1,3-butadiene" /></td>
<td>-60.8</td>
</tr>
</tbody>
</table>

\# monosubstituted double bonds in 1 = 1
\# monosubstituted double bonds in 2 = 2
\( \Delta H^\circ (2) \sim 2 \times \Delta H^\circ (1) \)

Heat of hydrogenation of an alkene in which the double bonds are conjugated is not additive because such an alkene is resonance stabilized, and, therefore, its Lewis structure, which the calculation would be based on, is a poor representation of the molecule.

eg: 1,3-butadiene (3)
Experimental heat of hydrogenation of $3 = -56.5 \text{ kcalmol}^{-1}$

# monosubstituted double bonds in the Lewis structure of $3 = 2$

# monosubstituted double bonds in $1 = 1$

Calculated heat of hydrogenation of $3 = 2 \times (-30.3 \text{ kcalmol}^{-1}) = -60.8 \text{ kcalmol}^{-1}$

$\Delta H^\circ(3, \text{calculated}) \neq \Delta H^\circ(3, \text{experimental})$

$\Delta H^\circ(3, \text{experimental}) - \Delta H^\circ(3, \text{calculated}) = \text{resonance energy of } 3$

Thus, resonance energy of $3 = -56.5 \text{ kcalmol}^{-1} - (-60.8 \text{ kcalmol}^{-1}) = 4.3 \text{ kcalmol}^{-1}$

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**Contributors**

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