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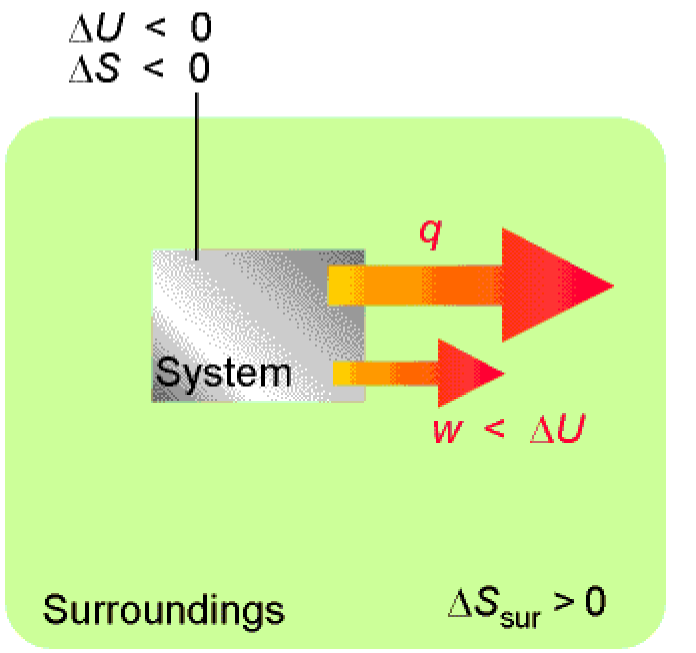
Standard Gibbs Energy Change



http://staff.um.edu.mt/jgri1/teaching/che2372/notes/04/g_a/F4_12.gif

Introduction

The criterion for spontaneous reaction can be demonstrated as follows:

$$\Delta S_{tot} = \Delta S_{sys} + \Delta S_{surr} > 0$$

$$\Delta S_{surr} = \frac{-\Delta H_{rxn}}{T}$$

$$\Delta S_{tot} = \Delta S_{sys} - \frac{\Delta H_{rxn}}{T} > 0$$

$$T * \Delta S_{sys} - \Delta H_{rxn} > 0$$

$$\Delta H_{rxn} - T * \Delta S_{sys} < 0$$

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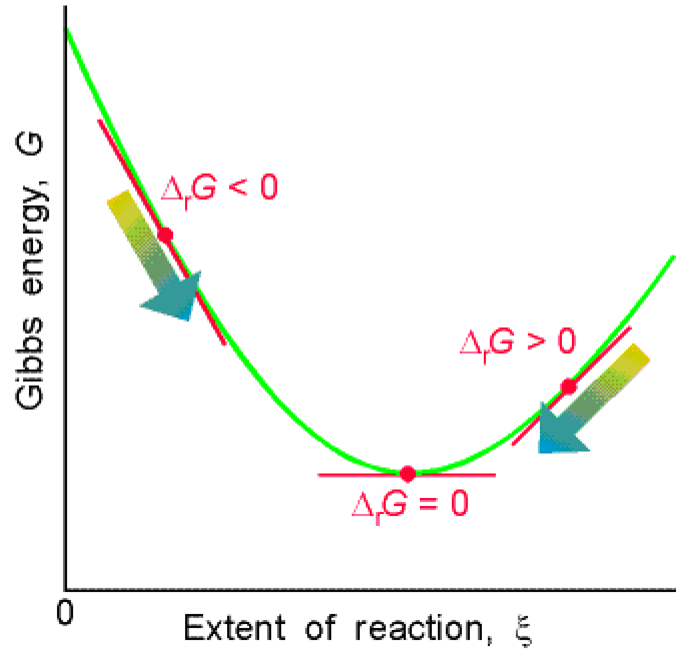

As the reaction advances (represented by motion from left to right along the horizontal axis) the slope of the Gibbs energy changes. Equilibrium corresponds to zero slope, at the foot of the valley.

http://staff.um.edu.mt/jgri1/teaching/che2372/notes/04/g_a/g_a.html

Definition

The standard Gibbs energy change (sometimes called “free energy”) combines change in entropy (in the system) and change in enthalpy (in the system) to have a simple expression for spontaneity at a given T temperature:

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

The Gibbs energy decreases in any spontaneous change.

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Standard Gibbs Energy Change

Standard Gibbs energy of formation		
Substance	State	kJ / mol
NH_3	g	-16.6
H_2O	l	-237.2
H_2O	g	-228.6
CO_2	g	-394.4
CO	g	-137
CH_4	g	-50.8

Standard Gibbs energy change

The standard Gibbs energy change (ΔG^0) is the Gibbs energy change per mole for conversion of reactants in their standard states into products in their standar states.

The criterion for spontaneous change is that

$$\Delta G^0 < 0$$

When applied to a formation reaction, the standard energy change of formation is related to the standard enthalpy and entropy changes of formation:

$$\Delta G_f^0 = \Delta H_f^0 - T * \Delta S_f^0$$

For elements, the standard Gibbs energy of formation is 0 by definition.

Standard Gibbs Energy Change

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Exercise

Calculate the standard Gibbs energy change of combustion of methane from the standard Gibbs energy changes of formation of the substances involved in the reaction.

Standard Gibbs energy changes of formation:

CO_2 ... -394 kJ/mol

H_2O (l) ... -237 kJ/mol

CH_4 (g) ... -50 kJ/mol

Solution

The reaction is:

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (l)$$

The change in Gibbs energy for this reaction will be:

$$\Delta G^0_{rxn} = \left[\Delta G^0_f\ (CO_2) + 2 * \Delta G^0_f\ (H_20)\right] - \left[\Delta G^0_f\ (CH_4) + 2 * \Delta G^0_f\ (O_2)\right]$$

$$\Delta G^0_{rxn} = \left[-394\ kJ/mol + 2 * (-237\ kJ/mol)\right] - \left[\ -50\ kJ/mol\right]$$

$$\Delta G^0_{rxn} = -818\ kJ/mol$$