Quadrant II – Transcript and Related Materials

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Module Name: Definition of ΔG and ΔGº

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Notes

Introduction:

Gibbs Free Energy Change (Δ G)

It is the maximum amount of non-expansion work that is extracted from a thermodynamically closed system and can be attained only in a completely reversible process.

The energy associated with a chemical reaction that is utilized to do work.

$$G = H - TS$$
$$\Delta G = \Delta H - T\,\Delta S$$

ΔG= Gibbs free energy change
ΔH= Enthalpy
ΔS= Entropy
T = Temperature in Kelvin

- The Gibbs free energy is a state function because it is defined in terms of thermodynamics properties that are state functions.
- For a spontaneous chemical reaction ΔG is negative and for that ΔH should be negative and TΔS should be positive.
- For a non-spontaneous chemical reaction ΔG is positive and for that ΔH should be positive and TΔS should be negative
- At equilibrium ΔG is zero, because both the opposing factors are exactly equal that is ΔH = TΔS, and the reaction is in the state of equilibrium, that is no net reaction would occur in any direction.

At equilibrium Gibbs free energy is minimum.

Standard Gibbs Free Energy (G^o)

It is the energy when all the reactants and products are at an initial concentration of 1.0 M and pressure of 1.0 atm.

$\Delta G^o = \Delta H^o - T\,\Delta S^o$
ΔG^o= Standard Gibbs free energy change
ΔH^o= Standard Enthalpy
ΔS^o= Standard Entropy
T = Temperature in Kelvin

- For a spontaneous chemical reaction ΔG^O is negative and for that ΔH^O should be negative and $T\Delta S^O$ should be positive.
- For a non-spontaneous chemical reaction ΔG^O is positive and for that ΔH^O should be positive and $T\Delta S^O$ should be negative.

Van't Hoff Isotherm

$\Delta G^o = -RT \ln K$
$\Delta G^o = - 2.303\, RT \log K$
ΔG^o = Standard Gibbs free energy
R= Gas Constant
T= Temperature in kelvin
K= Equilibrium constant

- This equation is used to calculate changes in free energy of a reaction in the standard state from the equilibrium constant and vice-versa.
- For a forward reaction to be spontaneous the standard Gibbs free energy is negative and log K must be positive.
- For a reverse reaction to be spontaneous the standard Gibbs free energy is positive and log K must be negative.
- At equilibrium the standard Gibbs free energy is zero and log K=0.

Free Energy for Non- Standard Condition

$\Delta G = \Delta G^o + RT \ln Q$
At equilibrium $\Delta G = 0$, so the above equation is written as
$\Delta G^o = - RT \ln Q$
ΔG = Gibbs Free Energy
ΔG^o = Standard Gibbs Free Energy
R = Gas Constant
T = Temperature in Kelvin
Q = Reaction quotient

Difference between ΔG and ΔG^{o}

1. ΔG is not a non- zero value while ΔG^{o} is a non- zero value: ΔG^{o} is a non- zero value and it can’t use ΔH^{o} or ΔS^{o} to find out ΔG because either ΔH^{o} or ΔS^{o} represents 100% completion.

2. At equilibrium ΔG is equal to zero while ΔG^{o} is negative: ΔG^{o} is negative at equilibrium and have lots of products at equilibrium and ‘Q’ is greater than 1 indicating that more product is accumulated. The term ‘RT’ in ‘Q’ gets increasingly positive and eventually adding that term to a negative ΔG^{o} makes $\Delta G = 0$, equilibrium will be established and no further change occurs.

3. Equilibrium is established when $\Delta G =0$ while lots of reactants at equilibrium when ΔG^{o} is positive: When ΔG^{o} is positive at equilibrium, lots of reactants at equilibrium (Q) needs to be smaller (less than 1) to approach K. As ‘Q’ gets smaller the term ‘RT Ln Q’ gets increasingly negative and eventually adding that term to a positive ΔG^{o} will make $\Delta G = 0$, equilibrium will be established and no further change occurs.

4. When $\Delta G =0$ at equilibrium, it will define which way the reaction proceeds while ΔG^{o} does not:

$\Delta G =0$ at equilibrium and ΔG^{o} is not equal to zero at equilibrium and the sign it generates by the combination of ΔG^{o} and ‘RT Ln Q’ will define which way the reaction proceeds.