Programme: B.Sc. Microbiology Subject: Microbiology Paper Code: MID - 102 Paper Title: Microbial Physiology Unit: 01 Bioenergetics And Electron Transport Chain Module Name: Standard Free Energy Change and Equilibrium Constant, Coupled Reactions Module No: 02 Name of the Presenter: Ms. Johanna Gonsalves

Notes

## **Bioenergetics II**

A reacting system continues changing until equilibrium is reached. And at equilibrium, concentration of reactants and products, the rates of the forward and reverse reactions are equal and no further net change occurs in the system.

The concentrations of reactants and products at equilibrium define the **equilibrium constant**, **Keq.** 

For a reaction,

$$aA + bB \leftrightarrow cC + dD$$

where a, b, c, and d are the number of molecules of A, B, C, and D participating, the equilibrium constant is given by

$$\mathsf{Keq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

[A], [B], [C], and [D] are the molar concentrations of the reaction components at the point of equilibrium.

A driving force causes a reaction to move toward equilibrium, the magnitude of this force can be expressed as the free-energy change for the reaction, G.

The magnitude of the driving force, which moves the reaction towards equilibrium, can be expressed as the free-energy change for the reaction,  $\Delta G$ .

Under standard conditions (25°C), when reactants and products are initially present at **1** M concentrations at **1** atm pressure, the force driving the system towards equilibrium is defined as the standard free-energy change,  $\Delta G^{\circ}$ .

If we consider the reaction,  $aA + bB \leftrightarrow cC + dD$ 

 $\Delta G$  of the reaction, depends on the concentration of reactant and product, temperature and pressure.

At constant temperature and pressure the relation will be

$$\Delta \mathbf{G} = \Delta \mathbf{G}^{\circ} + \mathbf{RT} \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$$

 $\Delta G^{\circ}$  - Standard free energy change

R – Gas Constant

T – Absolute Temperature

[A], [B], [C], and [D] are the molar concentrations of the reaction components at the point of equilibrium.

For biochemical reactions, a standard state is defined as having a pH of 7.0, and termed as standard transformed constants and written with a prime (G' and Keq') to distinguish them from the untransformed constants used by chemists and physicists

The standard free-energy change can be calculated from the equilibrium constant Keq.

$$\Delta G = \Delta G^{\circ} + RT \ln \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$$

Since reaction is at equilibrium,  $\Delta G = 0$ ,

 $\frac{[\ ]_{eq} [\ ]_{eq}}{[A]_{eq} [B]_{eq}}$ CD  $\Delta G = \Delta G^{\circ} + RT In$ 

$$\Delta G^{\circ} = - RT \ln K'_{eq}$$

 $\Delta G^{\circ'}$  = standard transformed free energy change

## $\Delta G^{\circ}$ = - RT lnK'<sub>eq</sub> = -2.303 RT log K'<sub>eq</sub>

Each chemical reaction has a characteristic standard free-energy change, which may be positive, negative, or zero, depending on the equilibrium constant of the reaction.  $\Delta G^{\circ}$  indicates, in which direction and how far a given reaction must go to reach equilibrium under standard conditions. Thus G is a constant: it has a characteristic, unchanging value for a given reaction.

Keq	ΔG°	Reaction
>1.0	Negative	Proceeds forward
=1.0	Zero	At equilibrium
<1.0	Positive	Proceeeds Backward

## **Coupled Reactions**

The combined catabolic and anabolic processes constitute metabolism.For sequential reactions,

 $A \rightleftharpoons B$  and  $B \rightleftharpoons C$ , the  $\Delta G^{\circ}$  values are additive:

A→B ΔG1°′ B→C ΔG2°′

```
\Delta G^{\circ}Total= \Delta G1^{\circ} + \Delta G2^{\circ}
```

A thermodynamically unfavorable reaction (endergonic) can be driven in the forward direction by coupling it to an exergonic reaction.

Standard Free-Energy Changes are Additive Example:

## Synthesis of glucose 6-phosphate

Glucose + Pi  $\rightarrow$  glucose 6-phosphate + H2O  $\Delta G^{\circ}$  = 13.8 kJ/mol

(will not proceed spontaneously in this direction)

ATP + H2O  $\rightarrow$  ADP + Pi

ΔG°´ = - 30.5 kJ/mol

These reactions share the common intermediates Pi and H2O and may be expressed as:

1) Glucose + Pi  $\rightarrow$  glucose 6-phosphate + H2O

2) ATP + H2O  $\rightarrow$  ADP + Pi

Glucose + ATP  $\rightarrow$  ADP + glucose 6-phosphate  $\Delta G^{\circ}$  = 13.8 kJ/mol + (-30.5 kJ/mol) = -16.7 kJ/molThe overall reaction is exergonic.

Energy stored in ATP is used to drive the synthesis of glucose 6-phosphate, even though its formation from glucose and inorganic phosphate (Pi) is endergonic.