

LECTURE 1 - METABOLISM & BIOENERGETICS

Thursday, April 10, 2008
1:29 PM

MCB102 SPRING 2008 METABOLISM SYLLABUS -- Prof. Bryan Krantz

Class Time: MWF 11-12

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Lecture notes: Notes will be posted on bspace and at <http://mcb.berkeley.edu/labs/krantz/mcb102.html> in empty and filled-out form. Usually, the empty notes will be there the night or morning before lecture.

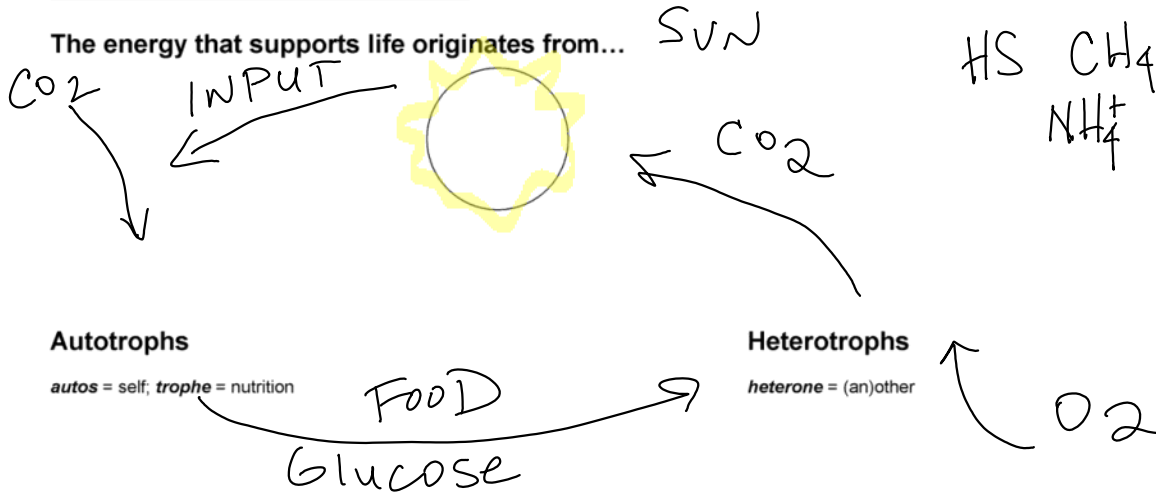
Bryan Krantz: University of California, Berkeley

MCB 102, Spring 2008, Metabolism Lecture 1

Reading: Ch. 13 of *Principles of Biochemistry*, "Bioenergetics and Biochemical Reaction Types."

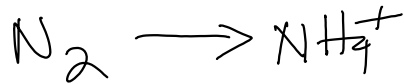
METABOLISM & BIOENERGETICS

The energy that supports life originates from...



Besides carbon, what other elements are converted to useable form(s) to support life?

Nitrogen Fixation



Metabolism - the set of chemical reactions that occur in the cell to maintain life.



Santorio Santorio (1561-1636)

"la respiration est donc une combustion"

Respiratory gas exchange is combustion, like a candle burning.



Antoine Lavoisier
1743-1794



Thermodynamics of life are consistent absolutely with chemistry.

"Ferments" inside yeast catalyze fermentation.

Enzymes come from cells – beginning of biochemistry



Louis Pasteur (1822-1895) / Eduard Buchner (1860-1917)

Enzymes are catalysts of Metabolic RXNS

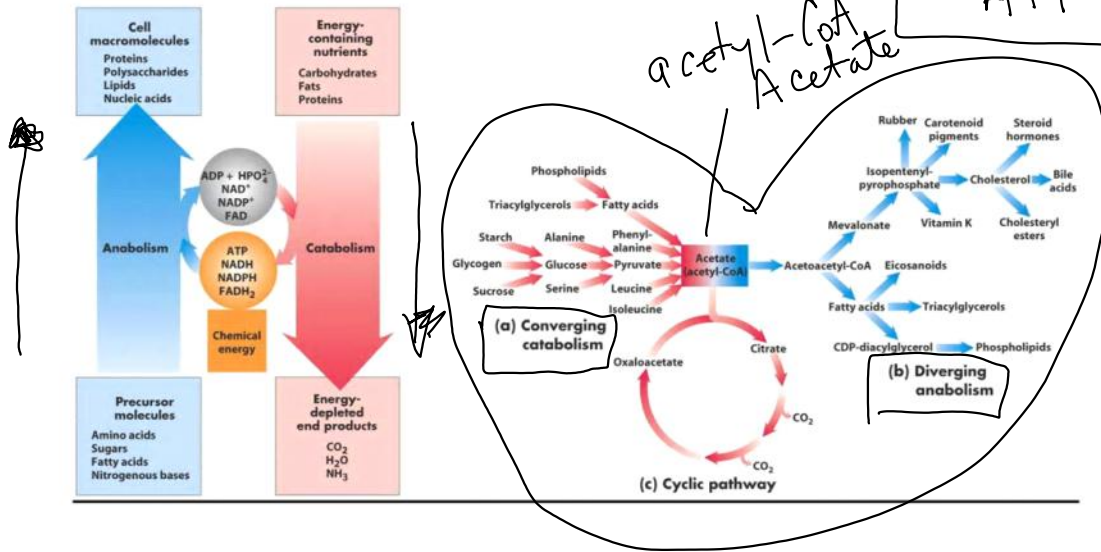
Metabolism - totality of the chemical reactions in the cell (biosphere) often catalyzed by protein enzymes.

(A) **Catabolism** - "down"

(B) **Anabolism** - "up"

Metabolites - small molecules

Energy currencies of the cell are ...



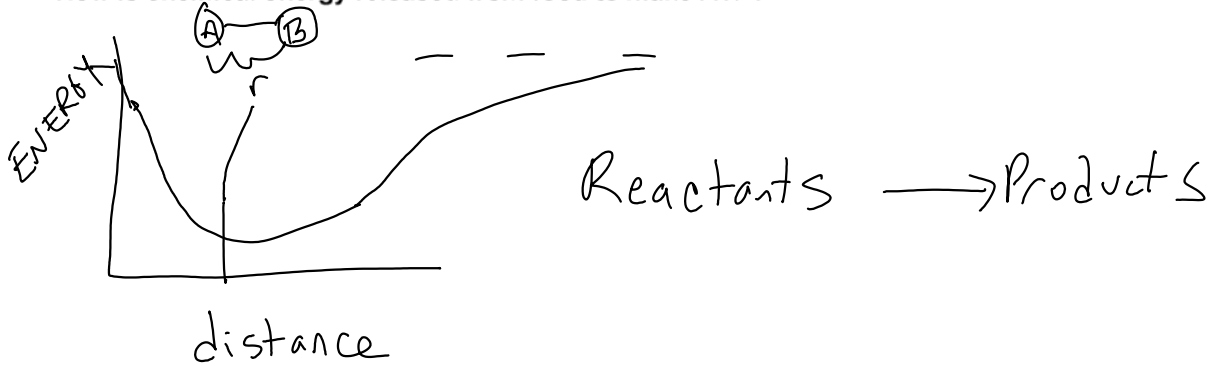
Bioenergetics & Thermodynamics – quantitative study of energy transduction in biology

Laws of Thermodynamics –

1st Matter & Energy are neither created nor destroyed.

2nd Universe tends toward disorder.

How is chemical energy released from food to make ATP?



Gibbs Free Energy Change, ΔG

Equilibrium $\Delta G = 0$

Positive $\Delta G > 0$
Non-spontaneous

Negative $\Delta G < 0$
Spontaneous

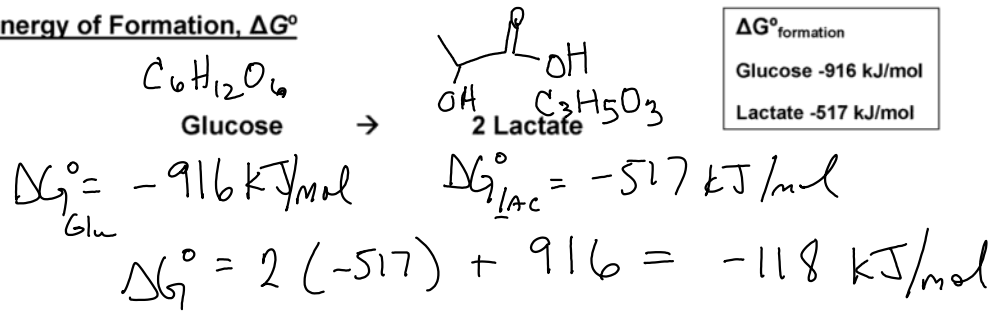
$\Delta G = \Delta H - T\Delta S$

UNITS
kcal/mol vs. kJ/mol
4.18 J = 1 cal
Kelvin

$\Delta G = RT \ln K_{eq}$ 8.31 J/mol/K

Why ΔG ? Why not K_{eq} ?

Standard Free Energy of Formation, ΔG°



$\Delta G^\circ_{\text{formation}}$
Glucose -916 kJ/mol
Lactate -517 kJ/mol

What is 'standard state'? And why do biochemists use, ΔG° ?

\rightarrow When all Reactant & Products are 1M. ΔG° is a way to make std. ΔG 's when @ neutral pH, 25°C, and 1mM $[Mg^{2+}]$. Basically, for std. Biochemical conditions.

Quantifying Free Energy Changes under Non-standard Conditions

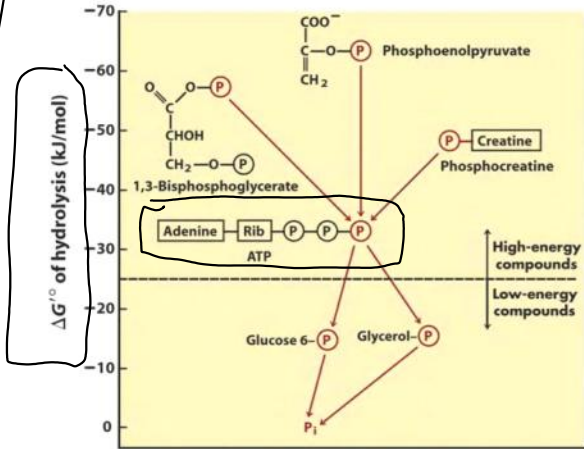
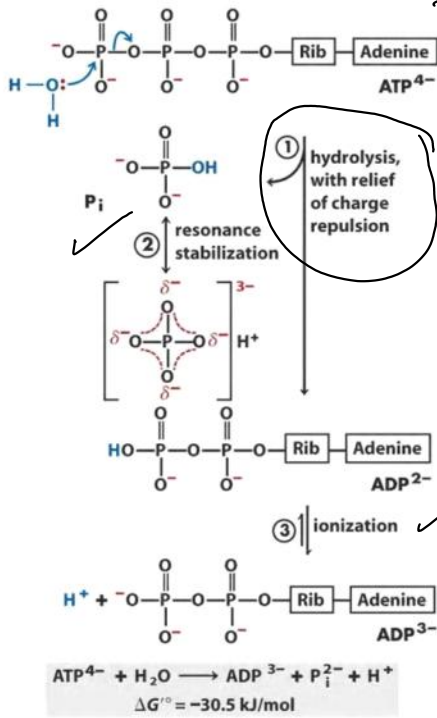
$$\Delta G = \Delta G^\circ - RT \ln Q$$



Q is the reaction quotient (mass-action ratio): $Q = ([C]^c [D]^d) / ([B]^b [A]^a)$

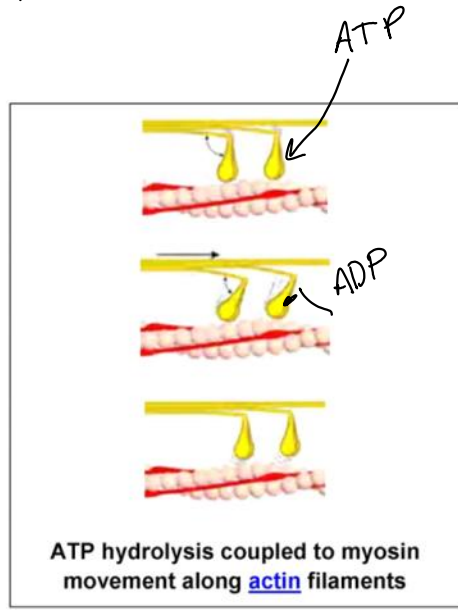
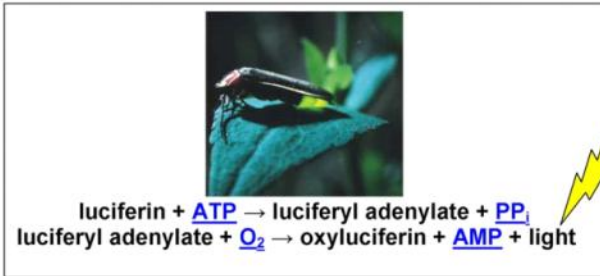
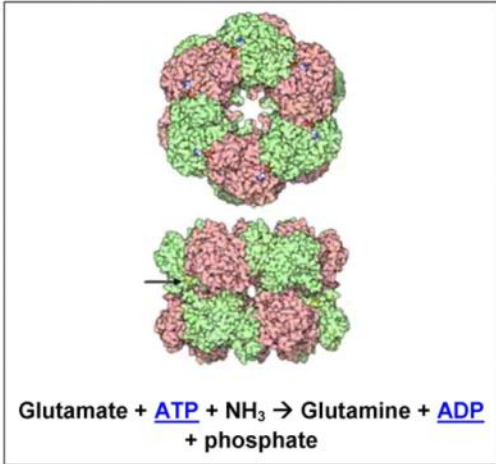
ATP

Why is ATP the energy currency in the cell?



How may ATP be used to do useful work?

anabolism $\left\{ \begin{array}{l} \text{phosphoryl transfer} \\ \text{adenylation} \end{array} \right.$
Motion Myosin
bioluminescence



Redox: Quantifying Metabolic Energy Transduction & Electron Flow

Reduction/oxidation reaction

Electron transport

Nernst Equation

$$E = E^{\circ'} - \frac{RT}{nF} \ln \frac{[e^- \text{ acceptor}]}{[e^- \text{ donor}]}$$

n = stoichiometry of electrons e^-

F = Faraday's Constant

E (the redox potential) is basically the voltage required to stop a chemical reaction

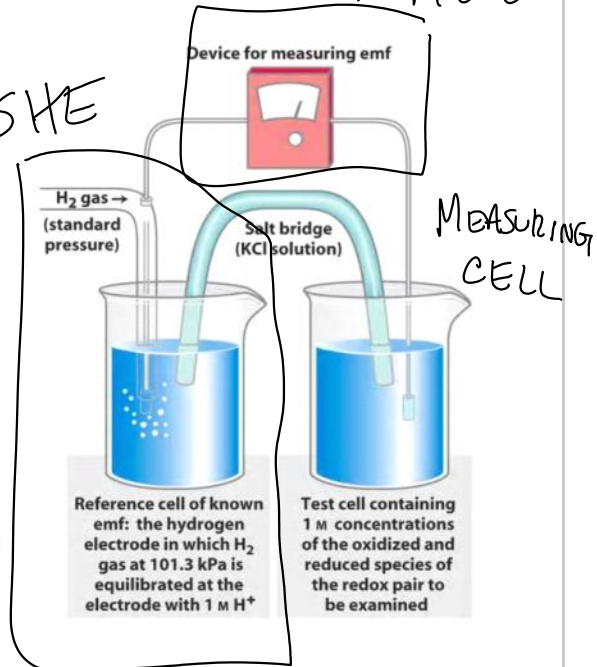
Relation of Standard Redox Potential Change to $\Delta G^{\circ'}$

$$\Delta G^{\circ'} = -nF \Delta E^{\circ'}$$

$E^{\circ'}$

SHE

Volt meter



Nicotinamide adenine dinucleotide (NAD⁺)

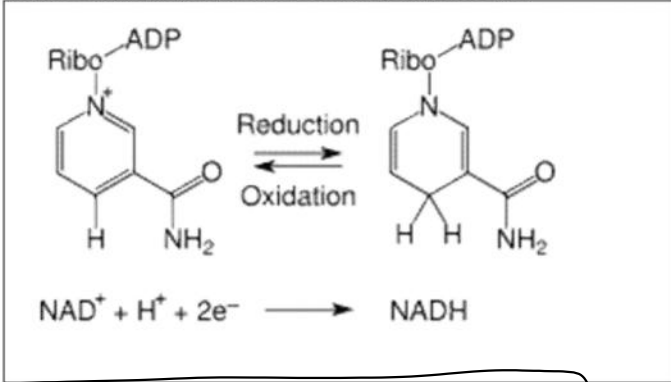


TABLE 13-7 Standard Reduction Potentials of Some Biologically Important Half-Reactions, at pH 7.0 and 25°C (298 K)

Half-reaction	E° (V)
$\frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2\text{O}$	0.816
$\text{Fe}^{3+} + \text{e}^- \longrightarrow \text{Fe}^{2+}$	0.771
$\text{NO}_3^- + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{NO}_2^- + \text{H}_2\text{O}$	0.421
Cytochrome <i>f</i> (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>f</i> (Fe^{2+})	0.365
$\text{Fe}(\text{CN})_6^{3-}$ (ferricyanide) + $\text{e}^- \longrightarrow \text{Fe}(\text{CN})_6^{4-}$	0.36
Cytochrome <i>a</i> ₃ (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>a</i> ₃ (Fe^{2+})	0.35
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2\text{O}_2$	0.295
Cytochrome <i>a</i> (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>a</i> (Fe^{2+})	0.29
Cytochrome <i>c</i> (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>c</i> (Fe^{2+})	0.254
Cytochrome <i>c</i> ₁ (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>c</i> ₁ (Fe^{2+})	0.22
Cytochrome <i>b</i> (Fe^{3+}) + $\text{e}^- \longrightarrow$ cytochrome <i>b</i> (Fe^{2+})	0.077
Ubiquinone + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ ubiquinol + H_2	0.045
Fumarate ²⁻ + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ succinate ²⁻	0.031
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ (at standard conditions, pH 0)	0.000

Source: Data mostly from Loach, P.A. (1976) In *Handbook of Biochemistry and Molecular Biology*, 3rd edn (Fasman, G.D., ed.), Physical and Chemical Data, Vol. 1, pp. 122-130, CRC Press, Boca Raton, FL.

* This is the value for free FAD; FAD bound to a specific flavoprotein (for example succinate dehydrogenase) has a different E° that depends on its protein environments.

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Half-reaction	E° (V)
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ (at standard conditions, pH 0)	0.000
Crotonyl-CoA + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ butyryl-CoA	-0.015
Oxaloacetate ²⁻ + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ malate ²⁻	-0.166
Pyruvate ⁻ + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ lactate ⁻	-0.185
Acetaldehyde + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ ethanol	-0.197
FAD + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ FADH ₂	-0.219*
Glutathione + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ 2 reduced glutathione	-0.23
$\text{S} + 2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2\text{S}$	-0.243
Lipoic acid + $2\text{H}^+ + 2\text{e}^- \longrightarrow$ dihydrolipoic acid	-0.29
$\text{NAD}^+ + \text{H}^+ + 2\text{e}^- \longrightarrow \text{NADH}$	-0.320
$\text{NADP}^+ + \text{H}^+ + 2\text{e}^- \longrightarrow \text{NADPH}$	-0.324
Acetoacetate + $2\text{H}^+ + 2\text{e}^- \longrightarrow \beta$ -hydroxybutyrate	-0.346
α -Ketoglutarate + $\text{CO}_2 + 2\text{H}^+ + 2\text{e}^- \longrightarrow$ isocitrate	-0.38
$2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$ (at pH 7)	-0.414
Ferredoxin (Fe^{3+}) + $\text{e}^- \longrightarrow$ ferredoxin (Fe^{2+})	-0.432

Source: Data mostly from Loach, P.A. (1976) In *Handbook of Biochemistry and Molecular Biology*, 3rd edn (Fasman, G.D., ed.), Physical and Chemical Data, Vol. 1, pp. 122-130, CRC Press, Boca Raton, FL.

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