

Bioenergetics and Biologic oxidation

University of Anbar/College of Pharmacy

Second semester 2020-2021 / Biochemistry II / 3rd stage

References :

- 1- Harper's Illustrated Biochemistry
- 2- Lehninger Principles of Biochemistry

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Bioenergetics, or biochemical thermodynamics, is study of the energy changes associated with the biochemical reactions **or** is a field in biochemistry and cell biology that concerns energy flow through living systems.

Biologic systems are essentially isothermic and use chemical energy to power living processes

Some problems related with the energy:

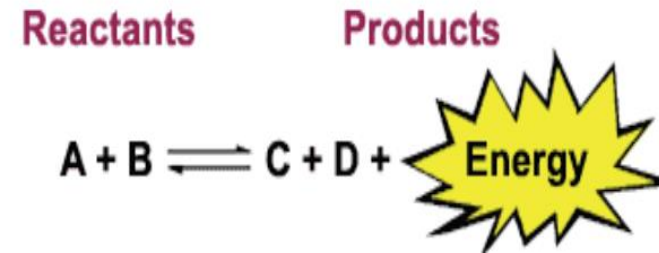
- A. Death from **starvation** occurs when available energy reserves are depleted
- B. Certain forms of **malnutrition** are associated with energy imbalance (**marasmus**).
- C. Thyroid hormones control the metabolic rate (rate of energy release) and disease results if they malfunction Excess storage of surplus energy causes **obesity**



THERMODYNAMICS OF METABOLISM

- Overall process of **catabolism** RELEASES energy
- Overall process of **anabolism** REQUIRES energy input

How do we define amount of energy?



The quantity of usable energy (chemical potential) in a reaction is called the Gibbs Free Energy (ΔG).

ΔG is the difference between the energy contained in the products of a reaction and the reactants:

$$\Delta G = (\text{energy of products}) - (\text{energy of reactants})$$



FREE ENERGY IS THE USEFUL ENERGY IN A SYSTEM

Gibbs change in free energy (ΔG) is that part of the total energy change in a system that is available for doing work—ie, the useful energy, also known as the chemical potential.

Gibbs free energy, G, expresses the amount of energy capable of doing work during a reaction at constant temperature and pressure.

Enthalpy, H, is the heat content of the reacting system. It reflects the number and kinds of chemical bonds in the reactants and products.

Entropy, S, is a quantitative expression for the randomness or disorder in a system .



Biologic Systems Conform to the General Laws of Thermodynamics

The first law of thermodynamics states that the total energy of a system, including its surroundings, remains constant.

- It implies that within the total system, energy is neither lost nor gained during any change.
- The energy may be transferred from one part of the system to another or may be transformed into another form of energy.

In living systems, chemical energy may be transformed into heat or into electrical, radiant, or mechanical energy.



The second law of thermodynamics states that the total entropy of a system must increase if a process is to occur spontaneously.

The relationship between the free energy change (ΔG) of a reacting system and the change in entropy (ΔS) is expressed by the following equation, which combines the two laws of thermodynamics

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

where ΔH is the change in enthalpy(heat) and T is the absolute temperature In biochemical reactions, because ΔH is approximately equal to ΔE , the total change in internal energy of the reaction, the above relationship may be expressed in the following way:

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



In biological system chemical reactions are classified as being either **exergonic** or **endergonic**. That just means that a reaction can either release energy useful for work (an exergonic reaction) or requires energy to proceed (an endergonic reaction)

The **spontaneous** reaction is an exergonic reaction and ΔG will be negative. Thus, a negative ΔG value tells you that the reaction is possible

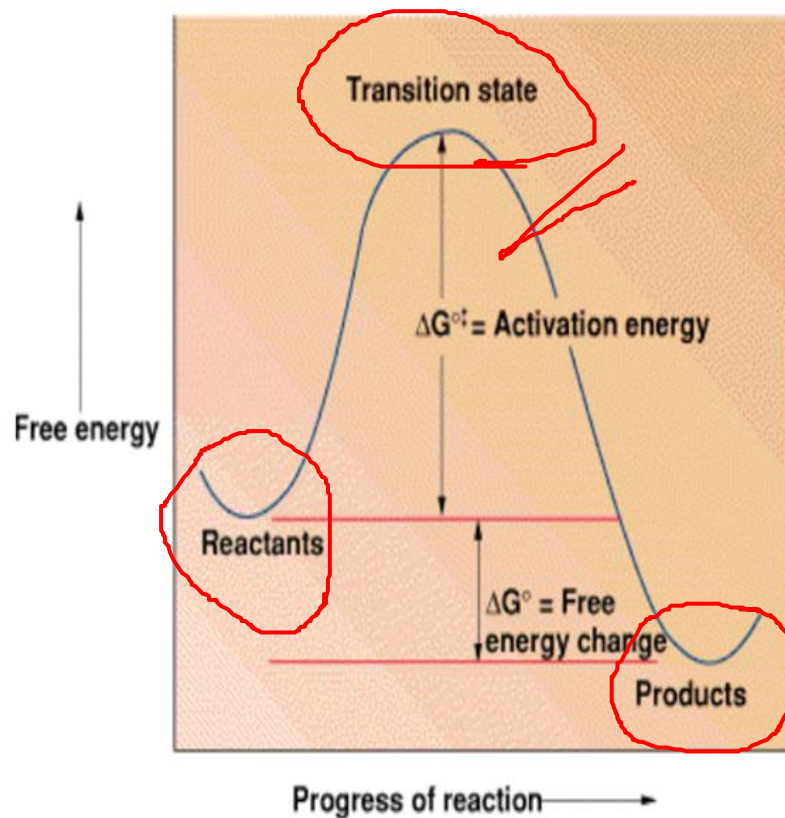


Table 14.5
Significance of ΔG° values

Value and Sign of ΔG°	Thermodynamic Consequences
$\Delta G^{\circ} = 0$	The reactants and the products are at the same energy level. The reaction under standard conditions is at equilibrium. No release of or requirement for energy.
$\Delta G^{\circ} < 0$ (negative values)	The reaction releases energy as it approaches equilibrium. The reactants are at a higher energy level than products. Useful energy is released and available to do work.
$\Delta G^{\circ} > 0$ (positive values)	The reactants are at a lower energy than products. The reaction requires an input of energy to proceed as written.

Table 14-5 Concepts in Biochemistry, 3/e
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If ΔG is **negative**, the reaction proceeds spontaneously with loss of free energy, that is, it is exergonic. If, in addition, ΔG is of great magnitude, the reaction goes virtually to completion and is essentially irreversible.

On the other hand, if ΔG is **positive**, the reaction proceeds only if free energy can be gained, that is, it is endergonic. If, in addition, the magnitude of ΔG is great, the system is stable, with little or no tendency for a reaction to occur.

If ΔG is **zero**, the system is at equilibrium and no net change takes place.



Exergonic Reaction

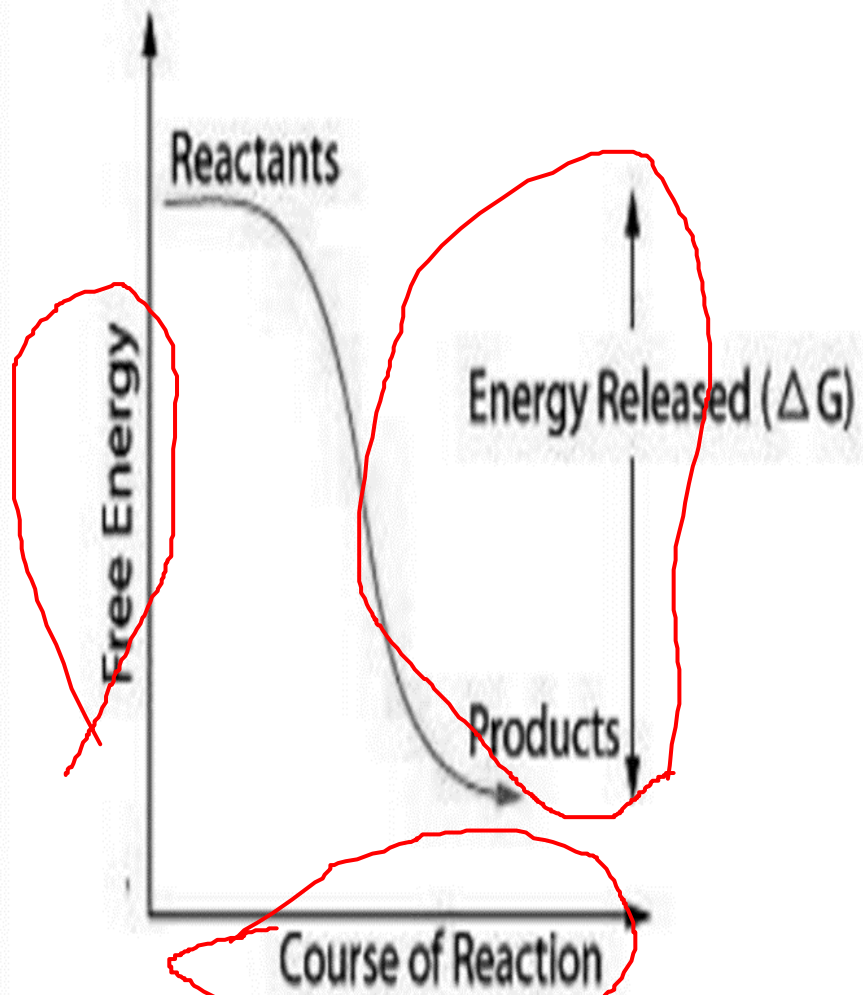
- Exergonic implies **the release of energy from a spontaneous chemical reaction** without any concomitant utilization of energy.
- The reactions are significant in terms of biology as these **reactions have an ability to perform work** and include most of the **catabolic reactions in cellular respiration**
- Most of these reactions **involve the breaking of bonds during the formation of reaction intermediates** as is evidently observed during respiratory pathways. The bonds that are created during the formation of metabolites are stronger than the cleaved bonds of the substrate
- $\Delta G = G_{\text{products}} - G_{\text{reactants}} < 0..$

Endergonic Reactions

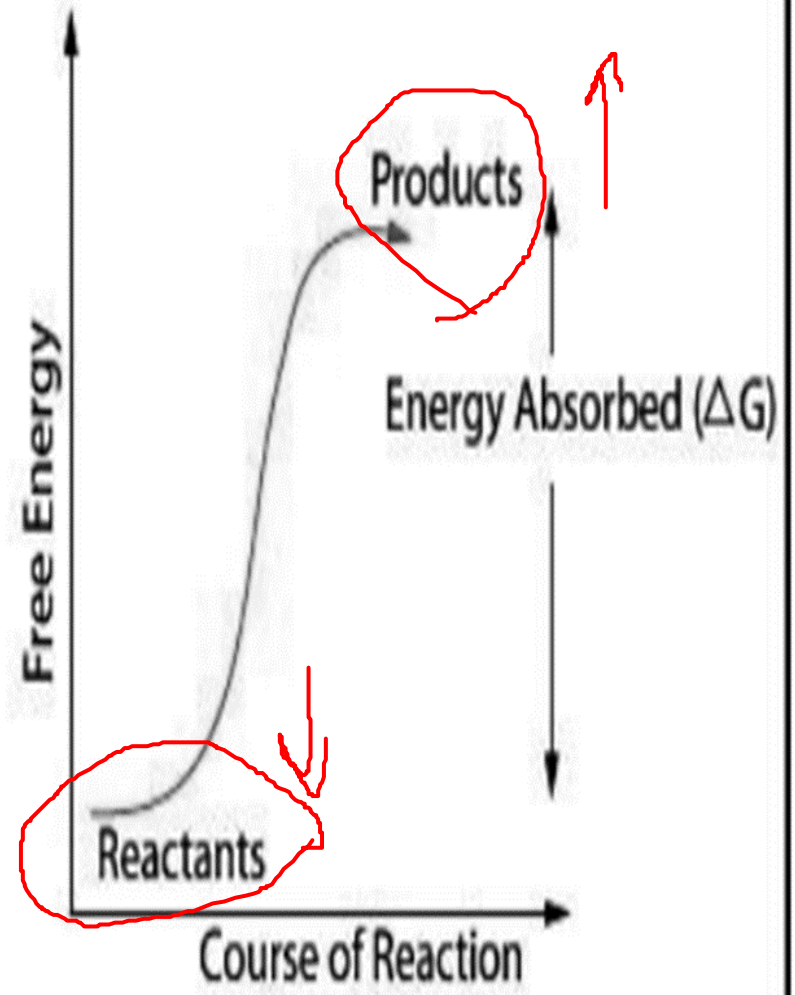
- Endergonic in turn is the opposite of exergonic in being **non-spontaneous** and requires an **input of free energy**.
- Most of the **anabolic reactions** like photosynthesis , DNA, and protein synthesis are endergonic in nature.
- $\Delta G = G_{\text{products}} - G_{\text{reactants}} > 0.$



Exergonic Reaction



Endergonic Reaction



Standard free energy change

The reactants are present in concentrations of 1.0 mol/L, ΔG° is the standard free energy change. For biochemical reactions, a standard state is defined as having a pH of 7.0. The standard free energy change at this standard state is denoted by $\Delta G^{\circ'}$. The standard free energy change can be calculated from the equilibrium constant K_{eq} .

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

Thermodynamic constants

Boltzmann constant, k = 1.381×10^{-23} J/K
Avogadro's number, N = 6.022×10^{23} mol⁻¹
Faraday constant, \mathcal{F} = 96,480 J/V · mol
Gas constant, R = 8.315 J/mol · K
(= 1.987 cal/mol · K)

Units of ΔG and ΔH are J/mol (or cal/mol)
Units of ΔS are J/mol · K (or cal/mol · K)
1 cal = 4.184 J

Units of absolute temperature, T , are Kelvin, K
25 °C = 298 K
At 25 °C, RT = 2.479 kJ/mol
(= 0.592 kcal/mol)



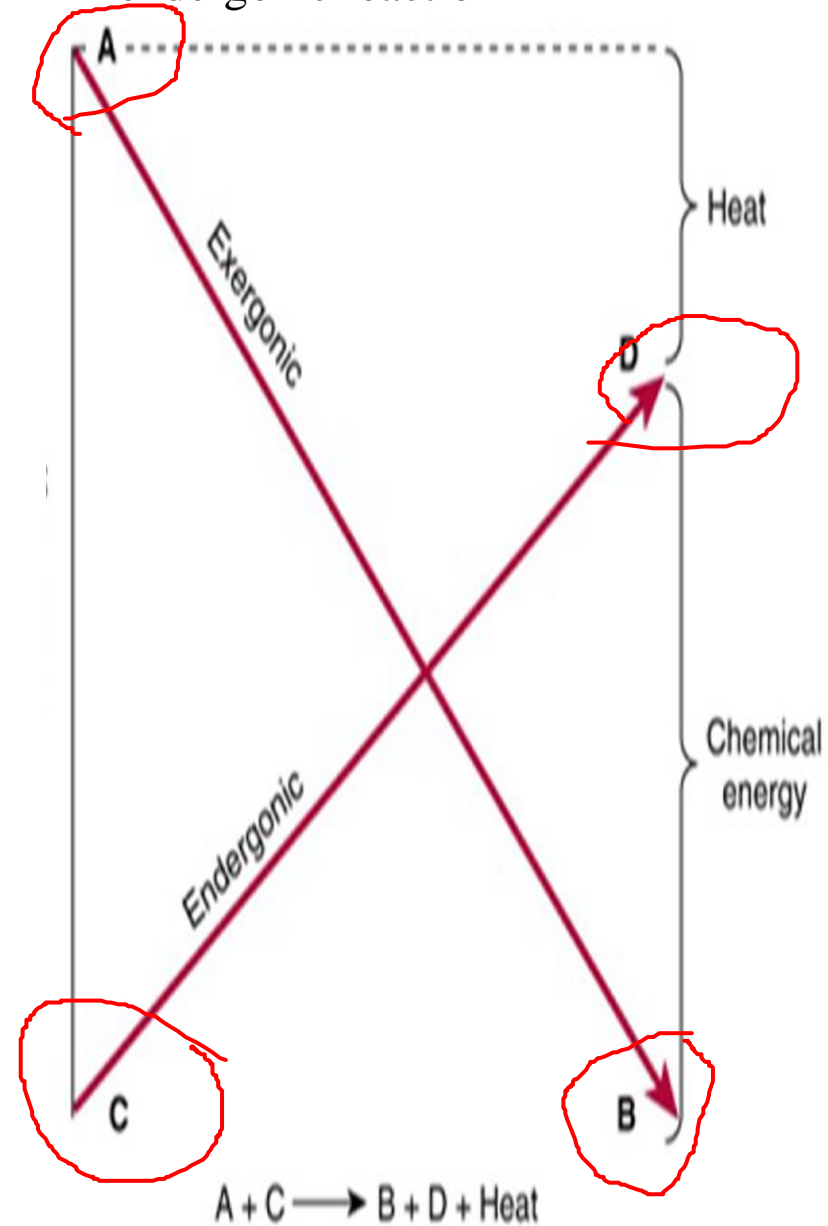
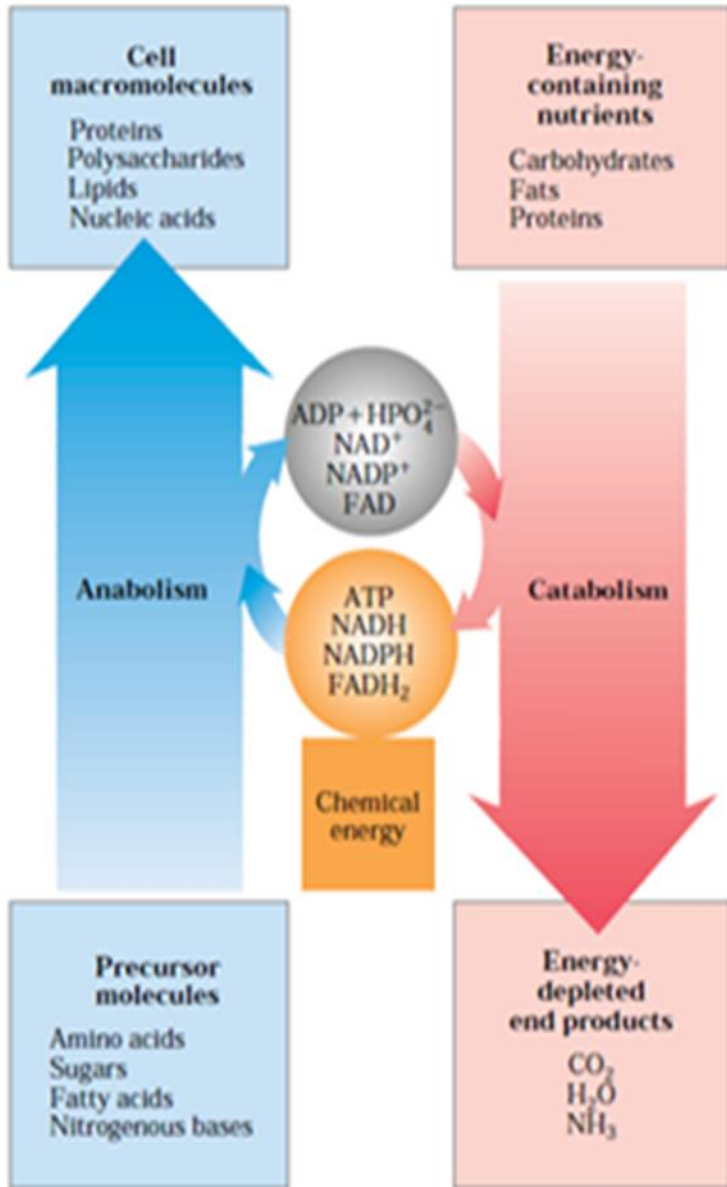
ENDERGONIC PROCESSES PROCEED BY COUPLING TO EXERGONIC PROCESSES

The vital processes, synthetic reactions, muscular contraction, nerve impulse conduction, and active transport—obtain energy by chemical linkage, or coupling, to oxidative reactions.

- The exergonic reactions are termed catabolism (generally, the breakdown or oxidation of fuel molecules)
- whereas the synthetic reactions that build up substances are termed anabolism. The combined catabolic and anabolic processes constitute metabolism.



Coupling of an exergonic to an endergonic reaction



HIGH-ENERGY PHOSPHATES PLAY A CENTRAL ROLE IN ENERGY CAPTURE AND TRANSFER

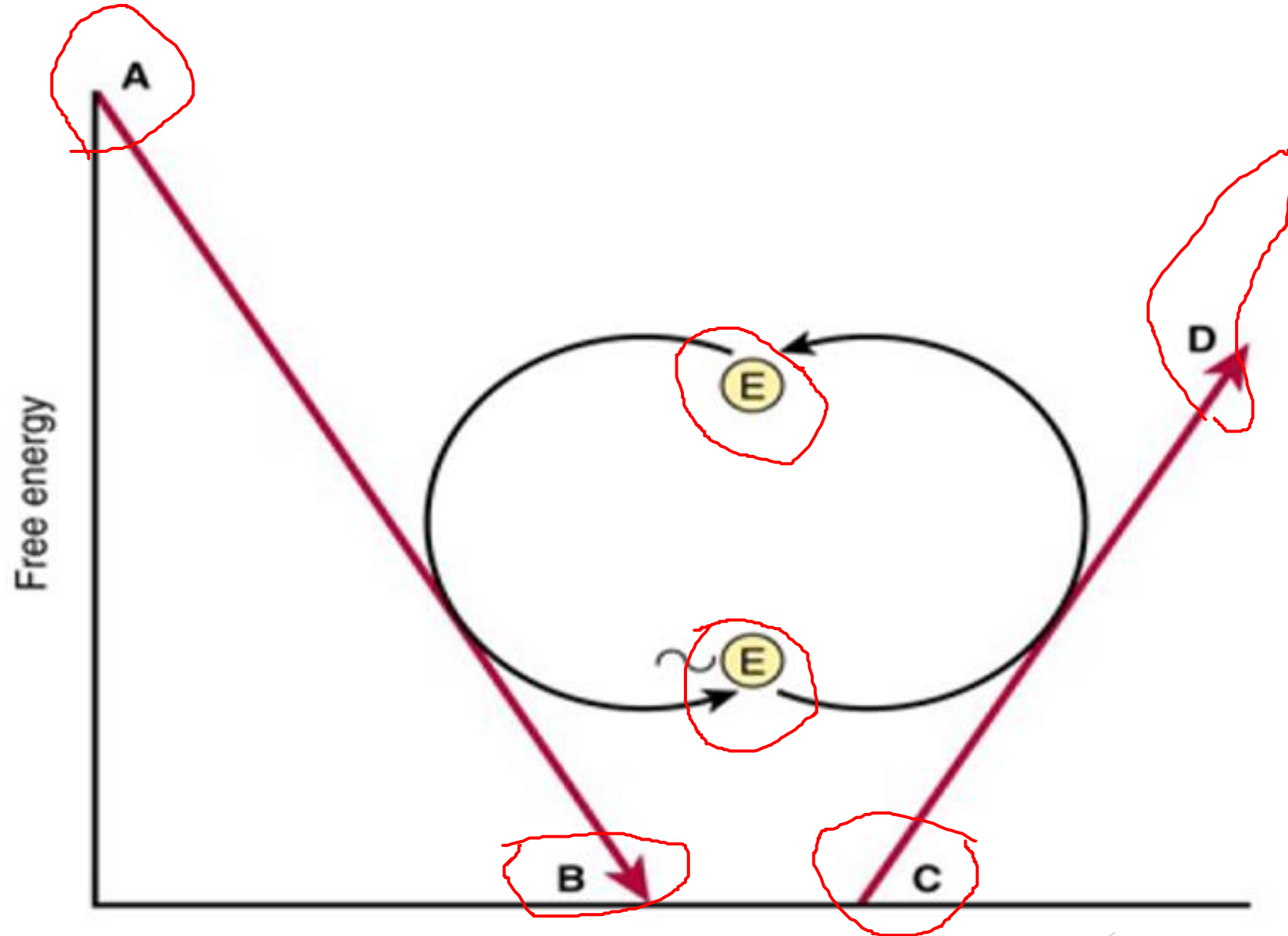
In order to maintain living processes, all organisms must obtain supplies of free energy from their environment.

Autotrophic organisms utilize simple exergonic processes; eg, the energy of sunlight (green plants), the reaction $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$ (some bacteria).

Heterotrophic organisms obtain free energy by coupling their metabolism to the breakdown of complex organic molecules in their environment. In all these organisms, **ATP plays** a central role in the transference of free energy from the exergonic to the endergonic processes (Figure below). high-energy intermediate compound ($\sim\text{E}$).



Transfer of free energy from an exergonic to an endergonic reaction via a high-energy intermediate compound



ATP is a nucleotide containing of the nucleoside adenosine (adenine linked to ribose), and three phosphate groups in its reactions in the cell, it functions as the Mg^{2+} complex

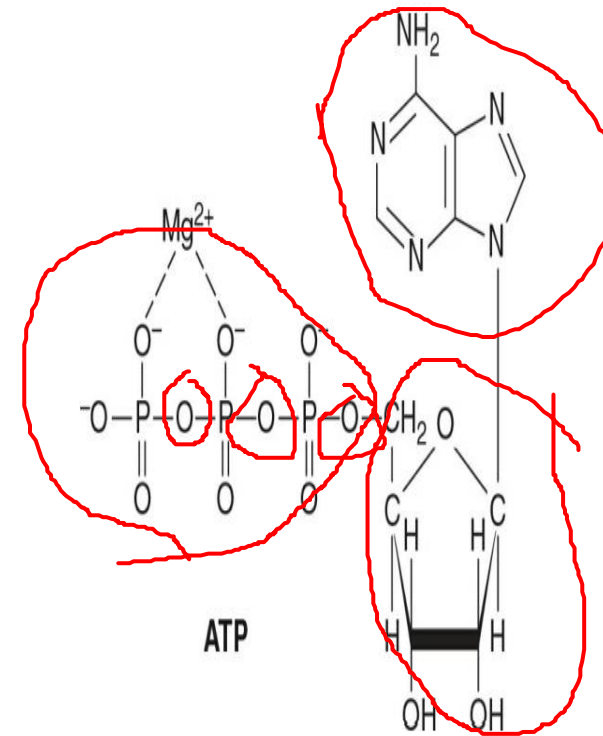
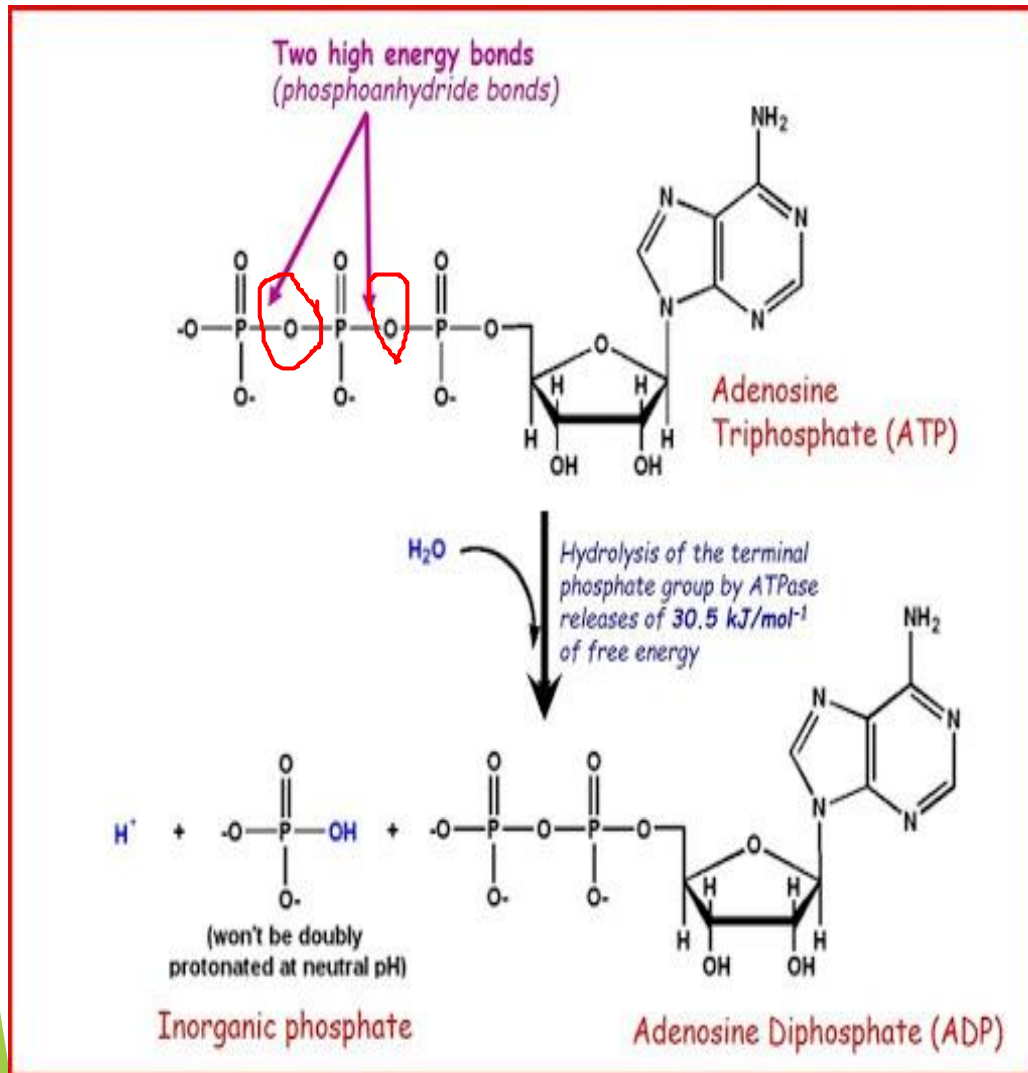


FIGURE 11-3 Adenosine triphosphate (ATP) is shown as the magnesium complex.



The standard free energy of hydrolysis of a number of biochemically important phosphates is shown in the Table

Compound	$\Delta G^{0'}$	
	kJ/mol	kcal/mol
Phosphoenolpyruvate	-61.9	-14.8
Carbamoyl phosphate	-51.4	-12.3
1,3-Bisphosphoglycerate (to 3-phosphoglycerate)	-49.3	-11.8
Creatine phosphate	-43.1	-10.3
ATP \rightarrow ADP + P _i	-30.5	-7.3
ADP \rightarrow AMP + P _i	-27.6	-6.6
Pyrophosphate	-27.6	-6.6
Glucose 1-phosphate	-20.9	-5.0
Fructose 6-phosphate	-15.9	-3.8
AMP	-14.2	-3.4
Glucose 6-phosphate	-13.8	-3.3
Glycerol 3-phosphate	-9.2	-2.2

¹P_i, inorganic orthophosphate.

²Values for ATP and most others taken from Krebs and Kornberg



Low-energy phosphates, exemplified by the ester phosphates found in the intermediates of glycolysis, have G' values smaller than that of ATP, while in **high-energy phosphates** the value is higher than that of ATP.

The symbol $\sim P$ indicates that **larger quantity of free energy**. For this reason, the term group **transfer potential**, rather than “high-energy Bond”. Thus, **ATP** contains **two** high-energy phosphate groups and **ADP** contains one, whereas the phosphate in **AMP** is of the low energy type since it is a normal ester link

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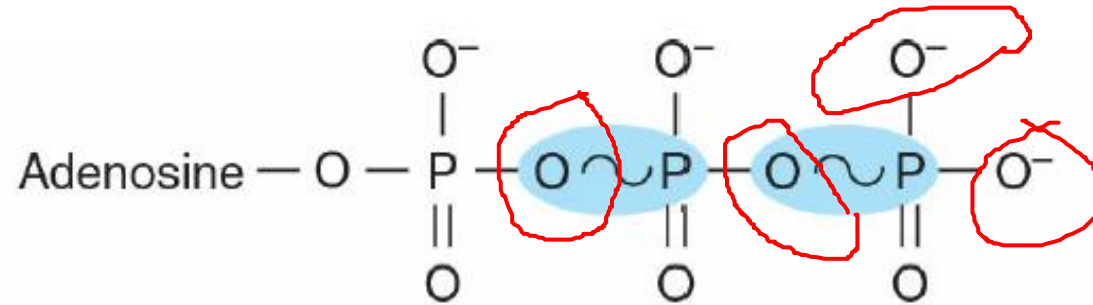


Reaction type	$\Delta G'^{\circ}$	
	(kJ/mol)	(kcal/mol)
Hydrolysis reactions		
Acid anhydrides		
Acetic anhydride + H ₂ O → 2 acetate	-91.1	-21.8
ATP + H ₂ O → ADP + P _i	-30.5	-7.3
ATP + H ₂ O → AMP + PP _i	-45.6	-10.9
PP _i + H ₂ O → 2P _i	-19.2	-4.6
UDP-glucose + H ₂ O → UMP + glucose 1-phosphate	-43.0	-10.3
Esters		
Ethyl acetate + H ₂ O → ethanol + acetate	-19.6	-4.7
Glucose 6-phosphate + H ₂ O → glucose + P _i	-13.8	-3.3
Amides and peptides		
Glutamine + H ₂ O → glutamate + NH ₄ ⁺	-14.2	-3.4
Glycylglycine + H ₂ O → 2 glycine	-9.2	-2.2
Glycosides		
Maltose + H ₂ O → 2 glucose	-15.5	-3.7
Lactose + H ₂ O → glucose + galactose	-15.9	-3.8
Rearrangements		
Glucose 1-phosphate → glucose 6-phosphate	-7.3	-1.7
Fructose 6-phosphate → glucose 6-phosphate	-1.7	-0.4
Elimination of water		
Malate → fumarate + H ₂ O	3.1	0.8
Oxidations with molecular oxygen		
Glucose + 6O ₂ → 6CO ₂ + 6H ₂ O	-2,840	-686
Palmitate + 23O ₂ → 16CO ₂ + 16H ₂ O	-9,770	-2,338

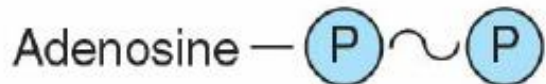


The Intermediate Value for the Free Energy of hydrolysis of ATP has Important Bioenergetic Significance

The high free-energy change on hydrolysis of ATP is due to relief of charge repulsion of adjacent negatively charged oxygen atoms and to stabilization of the reaction products, especially phosphate, as resonance hybrids Structure of ATP, ADP, and AMP showing the position and the number of high-energy phosphates (~ P).



Adenosine triphosphate (ATP)



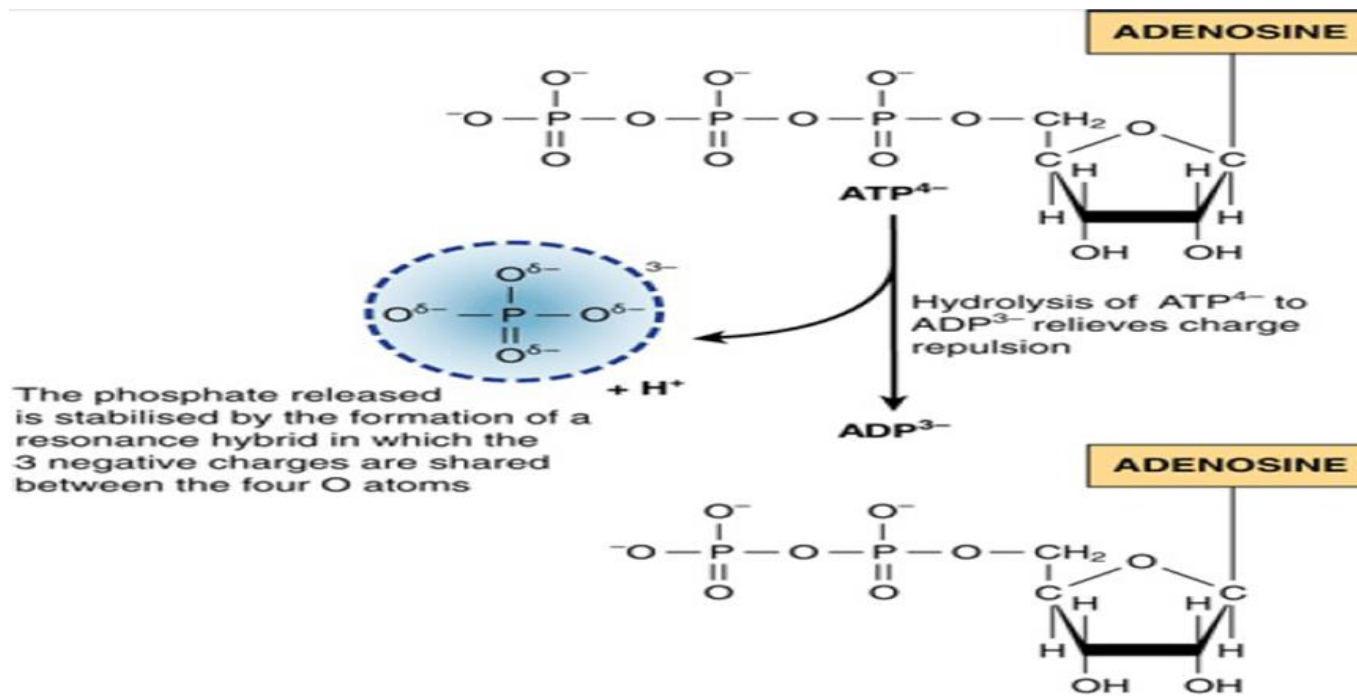
Adenosine diphosphate (ADP)

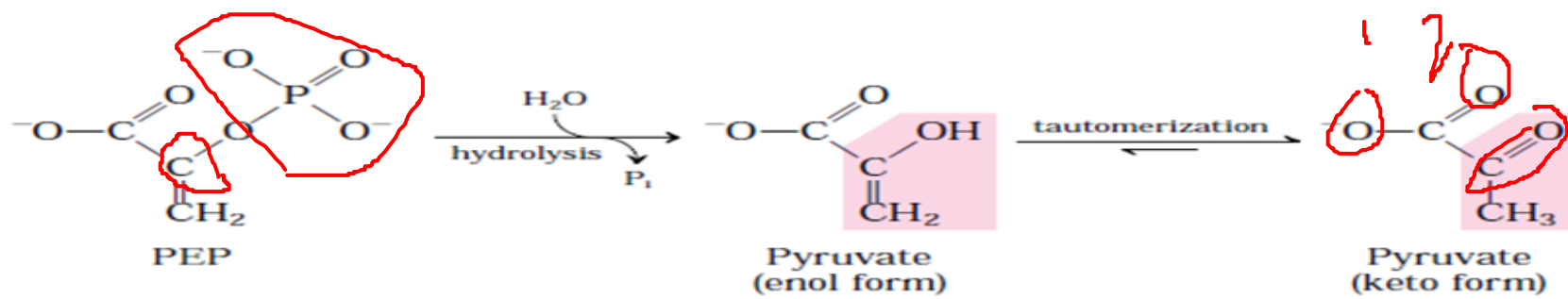


Adenosine monophosphate (AMP)



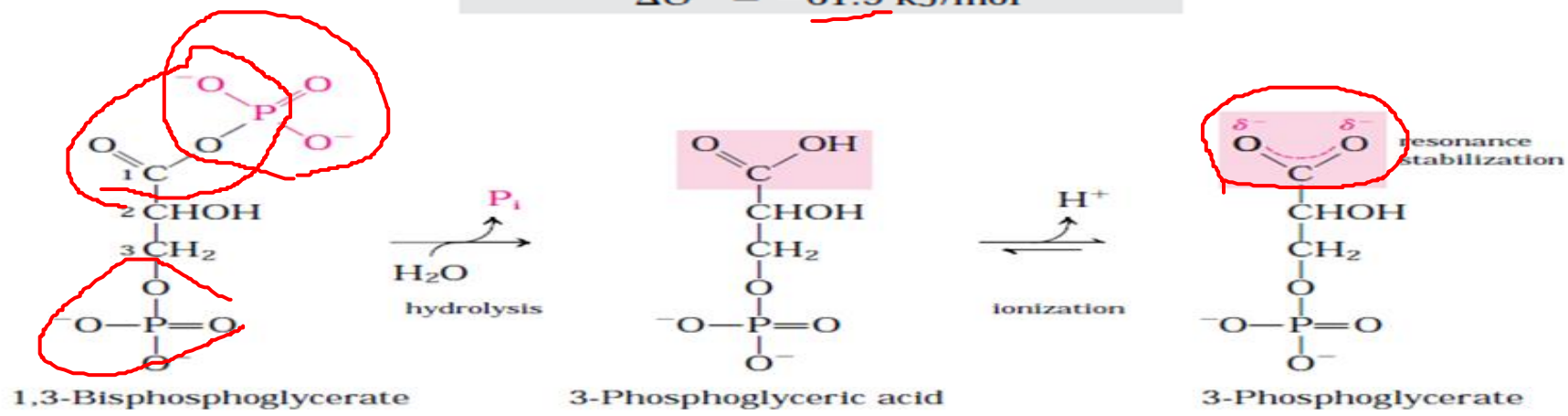
The value for the hydrolysis of the terminal phosphate of ATP divides the list into two groups. Low-energy phosphates, exemplified by the ester phosphates found in the intermediates of glycolysis, have G° values smaller than that of ATP, while in high-energy phosphates the value is higher than that of ATP. The components of this latter group, including ATP, are usually anhydrides (eg, the 1-phosphate of 1,3-bisphosphoglycerate), enolphosphates (eg, phosphoenolpyruvate), and phosphoguanidines (eg, creatine phosphate, arginine phosphate).





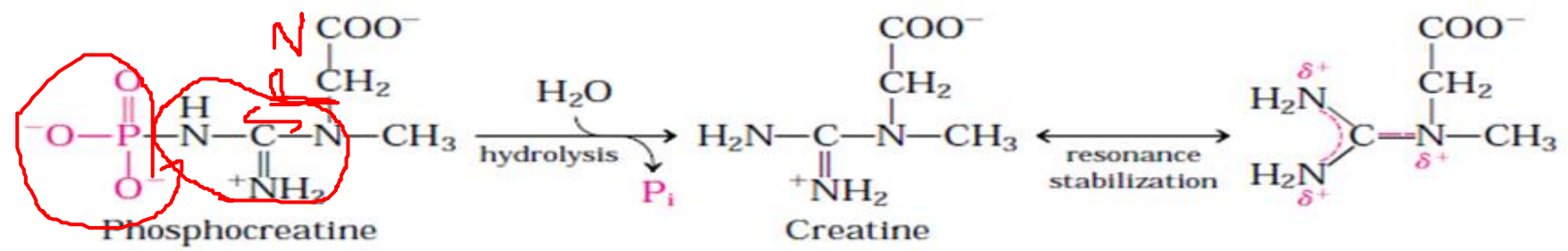
$$\text{PEP}^{3-} + \text{H}_2\text{O} \longrightarrow \text{pyruvate}^- + \text{P}_i^{2-}$$

$$\Delta G'^{\circ} = -61.9 \text{ kJ/mol}$$



$$1,3\text{-Bisphosphoglycerate}^{4-} + \text{H}_2\text{O} \longrightarrow 3\text{-phosphoglycerate}^{3-} + \text{P}_i^{2-} + \text{H}^+$$

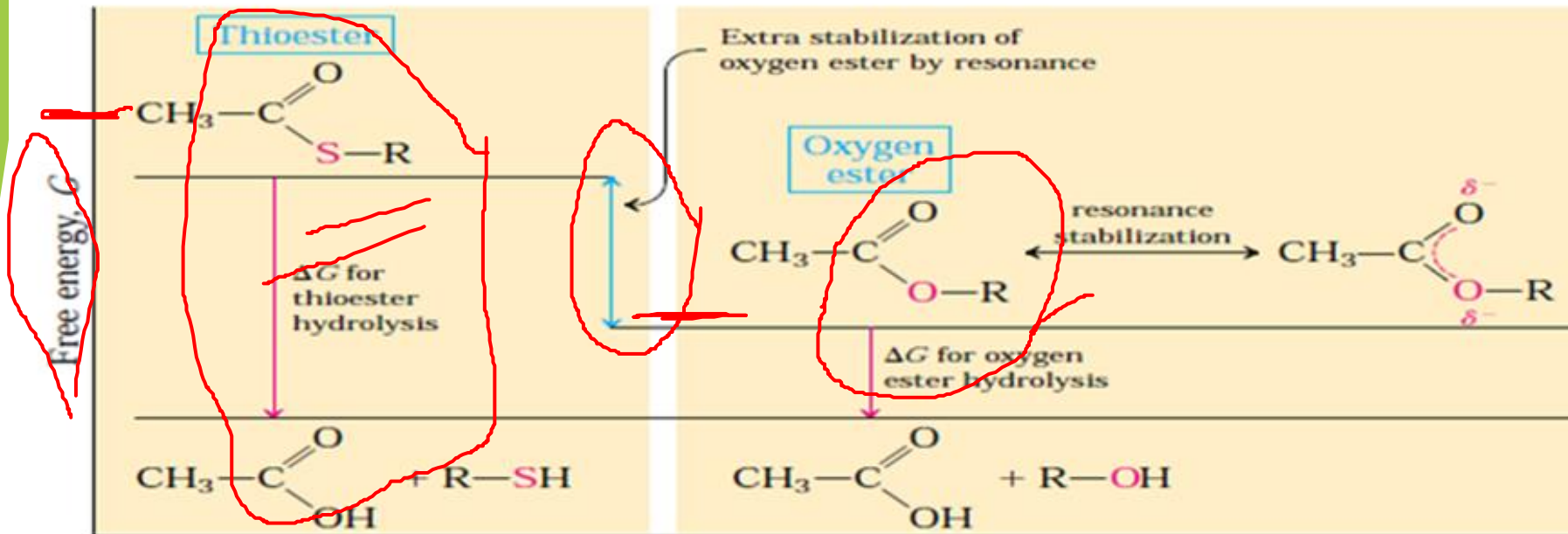
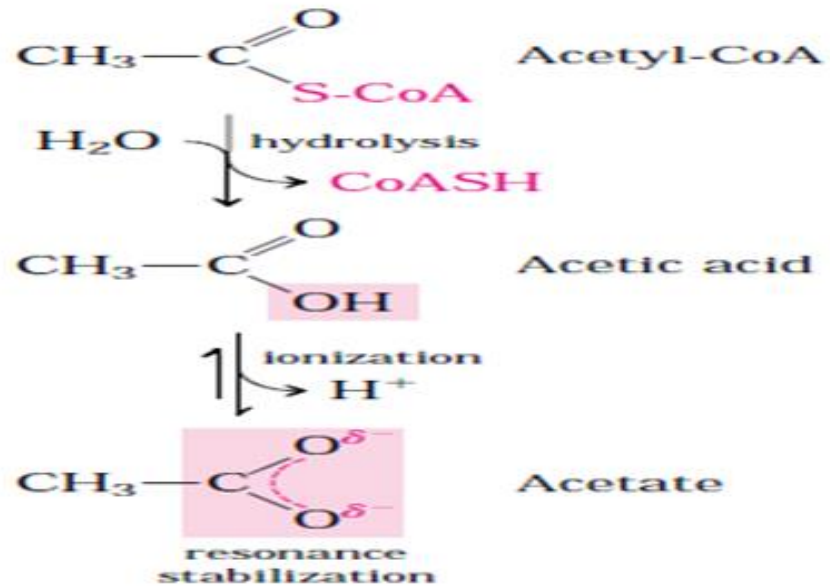
$$\Delta G'^{\circ} = -49.3 \text{ kJ/mol}$$



$$\text{Phosphocreatine}^{2-} + \text{H}_2\text{O} \longrightarrow \text{creatin} + \text{P}_i^{2-}$$

$$\Delta G'^{\circ} = -43.0 \text{ kJ/mol}$$





HIGH-ENERGY PHOSPHATES ACT AS THE ENERGY CURRENCY OF THE CELL

There are three major sources of ~ P taking part in energy conservation or energy capture:

1. Oxidative phosphorylation
2. Glycolysis.
3. The citric acid cycle.

Phosphagens act as storage forms of group transfer potential and include creatine phosphate, which occurs in vertebrate skeletal muscle, heart, spermatozoa, and brain, and arginine phosphate, which occurs in invertebrate muscle.



When ATP is rapidly being utilized as a source of energy for muscular contraction, phosphagens permit its concentrations to be maintained, but when the ATP/ADP ratio is high, their concentration can increase to act as an

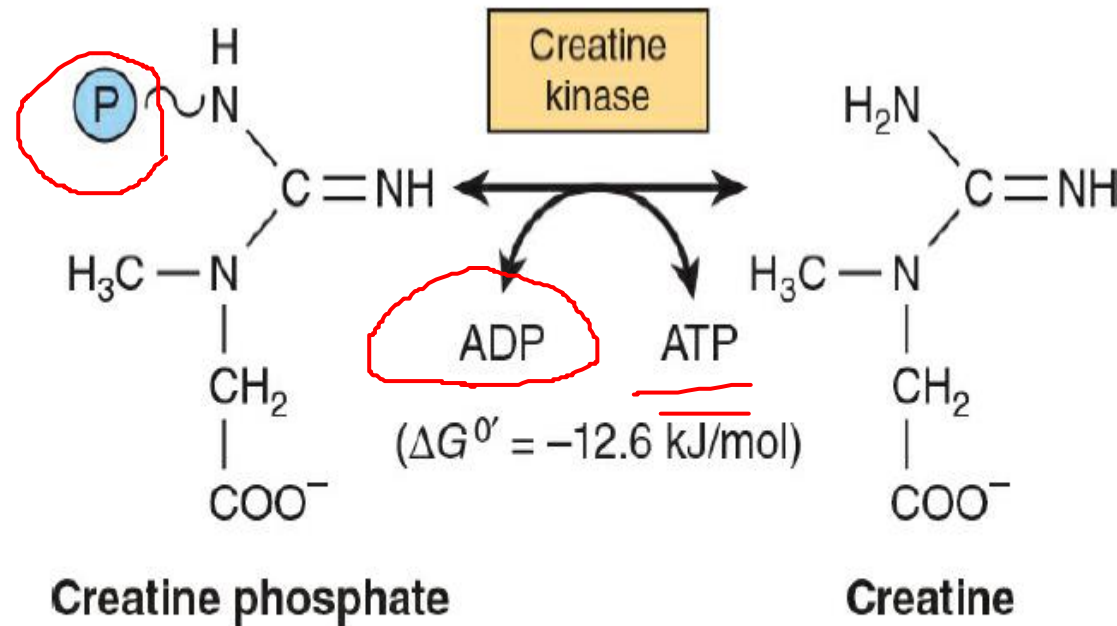
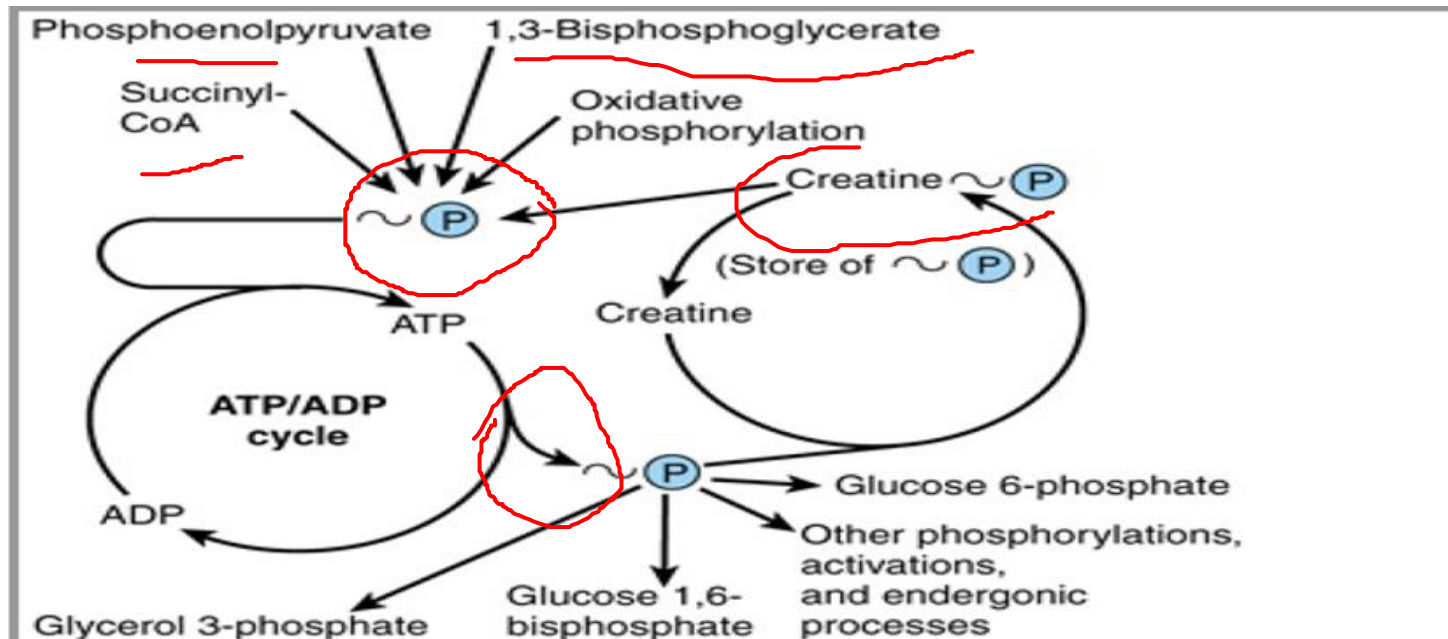


FIGURE 11-7 Transfer of high-energy phosphate between ATP and creatine.



HIGH-ENERGY PHOSPHATES ACT AS THE "ENERGY CURRENCY" OF THE CELL

ATP is able to act as a donor of high-energy phosphate to form those compounds below. The necessary enzymes, ADP can accept high-energy phosphate to form ATP . In effect, an **ATP/ADP cycle** connects those processes that generate $\sim P$ to those processes that utilize $\sim P$, continuously consuming and regenerating ATP. This occurs at a very rapid rate since the total ATP/ADP pool is extremely small and sufficient to maintain an active tissue for only a few seconds. In a cell, the ratio of ATP to ADP concentrations is known as the "energy charge" of the cell .



The cell can use this energy charge to relay information about cellular needs; if there is more ATP than ADP available, the cell can use ATP to do work, but if there is more ADP than ATP available, the cell must synthesize ATP via oxidative phosphorylation ATP allows the Coupling of thermodynamically Unfavorable reactions to Favorable Ones .



Adenylate Kinase (Myokinase) Interconverts adenine Nucleotides

Adenylate kinase is important for the maintenance of energy homeostasis in cells because it allows:

1. High-energy phosphate in ADP to be used in the synthesis of ATP.
2. The AMP formed as a consequence of activating reactions involving ATP to rephosphorylated to ADP.
3. AMP to increase in concentration when ATP becomes depleted so that it is able to act as a metabolic (allosteric) signal to increase the rate of catabolic reactions, which in turn lead to the generation of more ATP



THANK YOU

The background features abstract, overlapping geometric shapes in various shades of green, ranging from light lime to dark forest green. These shapes are primarily located on the right side of the frame, with some extending towards the center. The overall composition is clean and modern.