Overview of Thermodynamics and Bioenergetics

Historical (Age of Enlightenment)

"in general, respiration is nothing but a ***slow combustion*** of carbon and hydrogen, which is **entirely similar to that which occurs in a lighted lamp or candle**, and that, from this point of view, **animals that respire are true combustible bodies** that burn and consume themselves….

--Antoine Lavoisier ~1780 (start of Chapter 13)

History links just fyi:

http://www.uh.edu/engines/epi728.htm (short history of Lavosier's execution)

http://www2.lifl.fr/~bossut/david_lavoisier.jpg (painting)

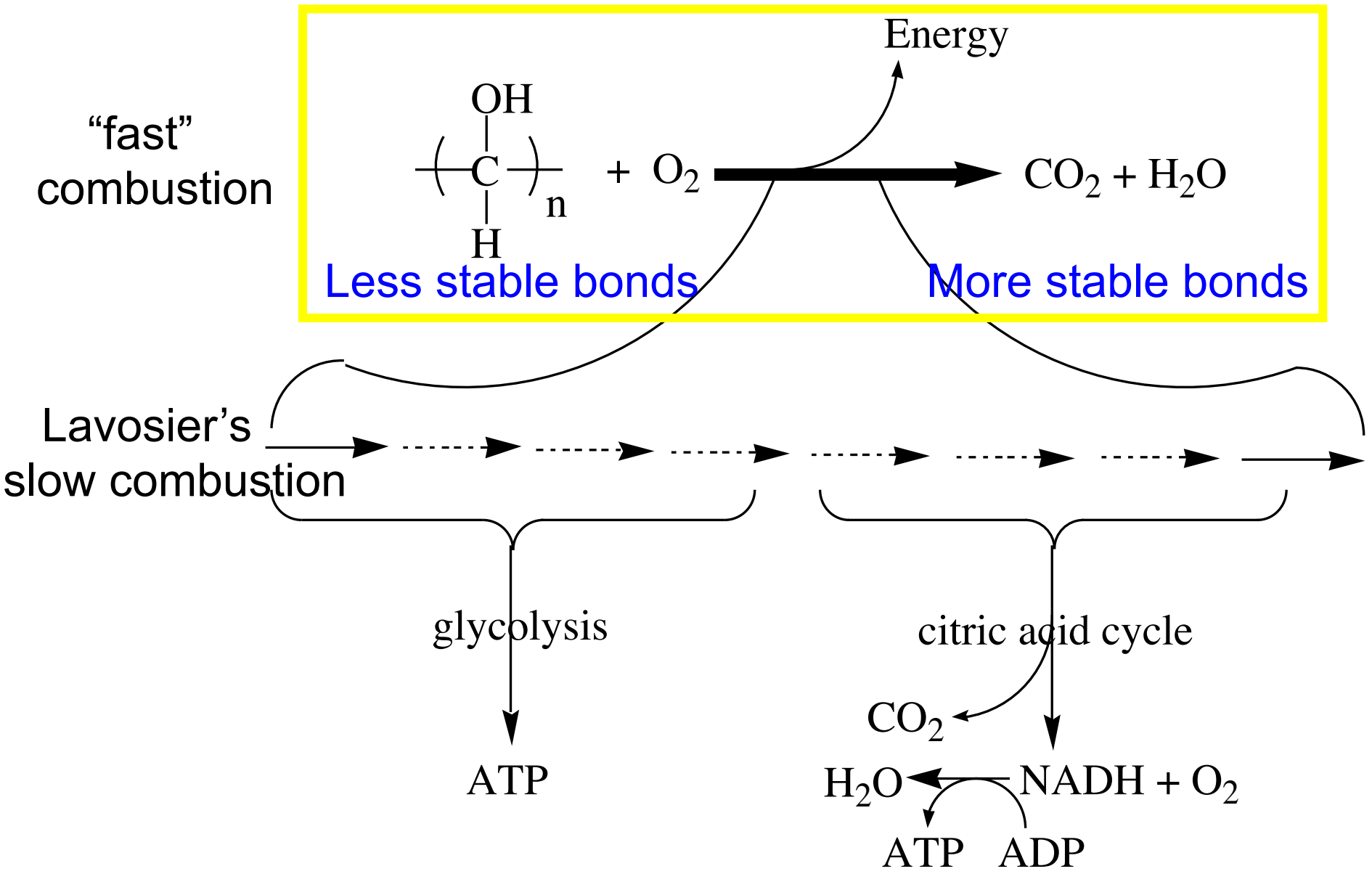
http://www.chemheritage.org/pubs/magazine/feature_lavoisier_p1.html (apparatus in painting used for key discoveries)

http://www.pbs.org/benfranklin/l3_inquiring_mesmer.html Lavosier & Franklin debunk Mesmer

Amusing story about Lavosier's collaborator on respiration experiments, Simon-Pierre de Laplace, Napoleon and astronomy.

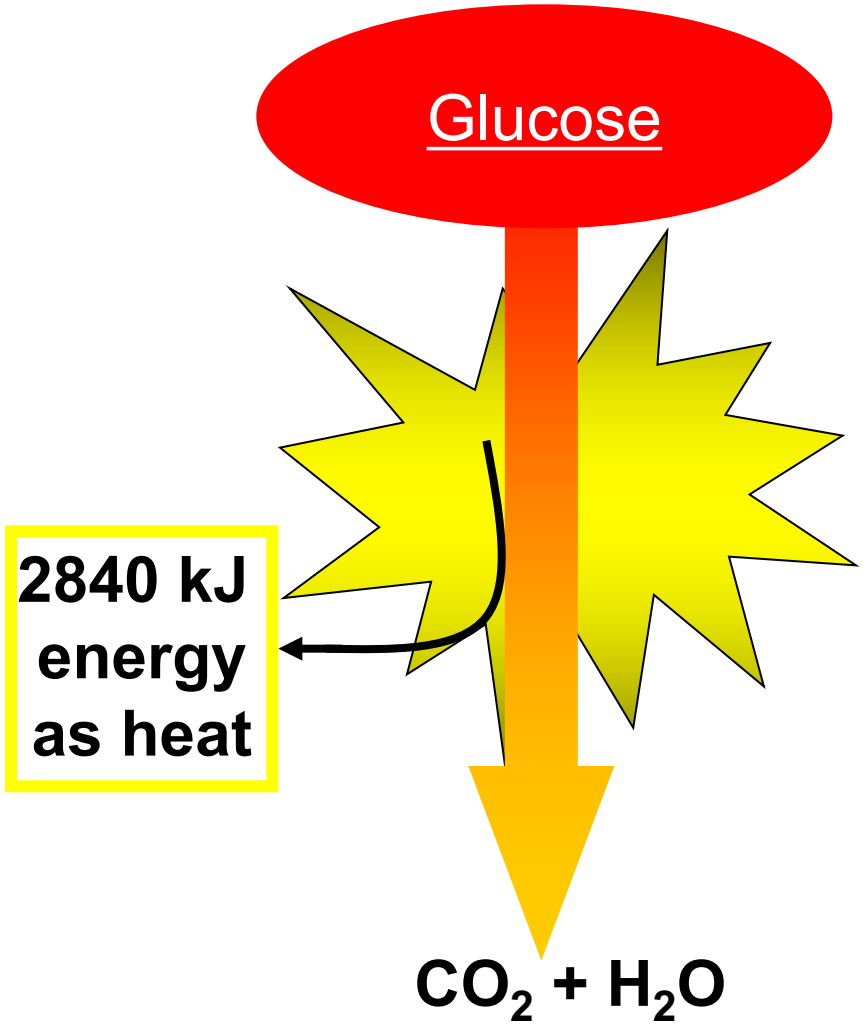
http://www.naturalhistorymag.com/universe/211420/the-perimeter-of-ignorance

Cells obtain most of their energy by oxidation reactions, but not by direct reaction with oxygen

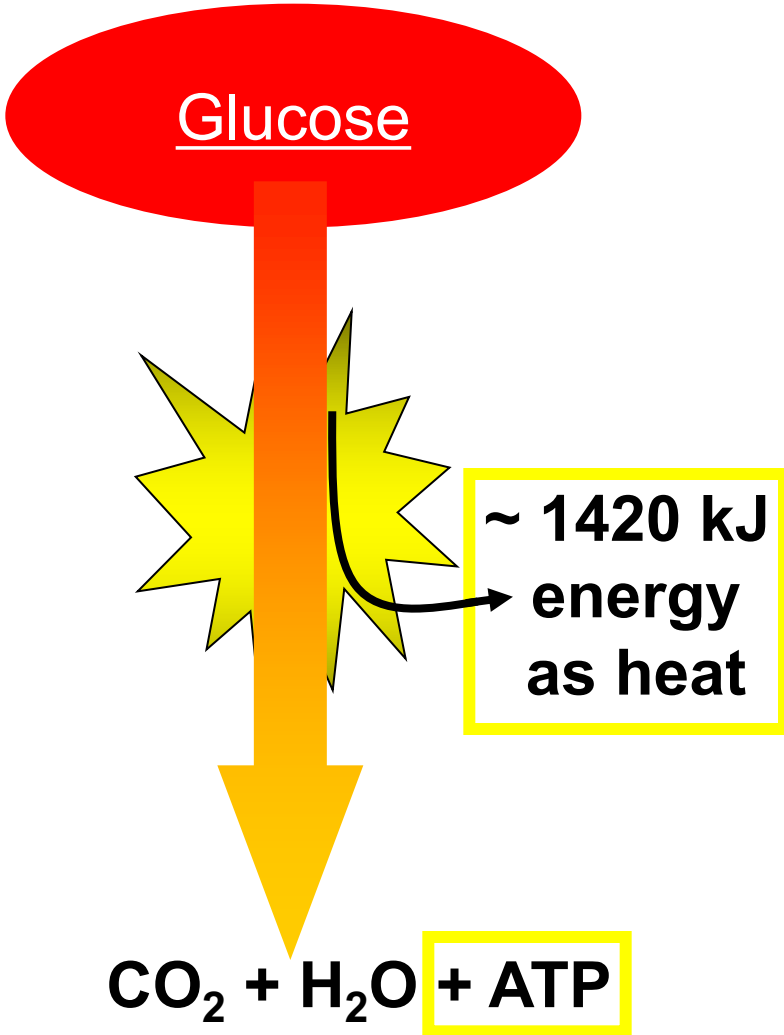


In cells, some energy of oxidations "saved" as ATP

Combustion of glucose in a bomb calorimeter



Conversion of glucose to CO_2 + ATP



- NET reactions "go" towards equilibrium

$$A + B \rightleftharpoons C + D$$

Reactants Products

$$K'_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Reactants & Products defined by left/right position; not NET direction of reaction.

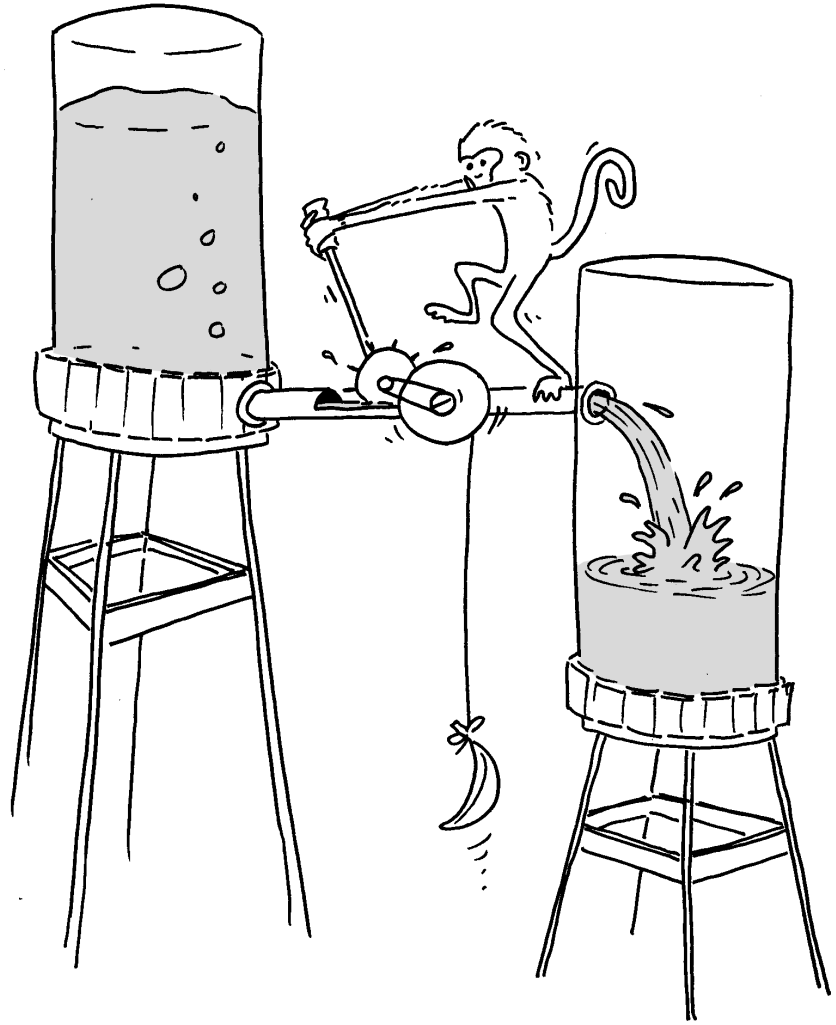
- Conditions can be on "either side" of equilibrium

For example, if equilibrium = 1.8 M A & B and 0.2 M C & D

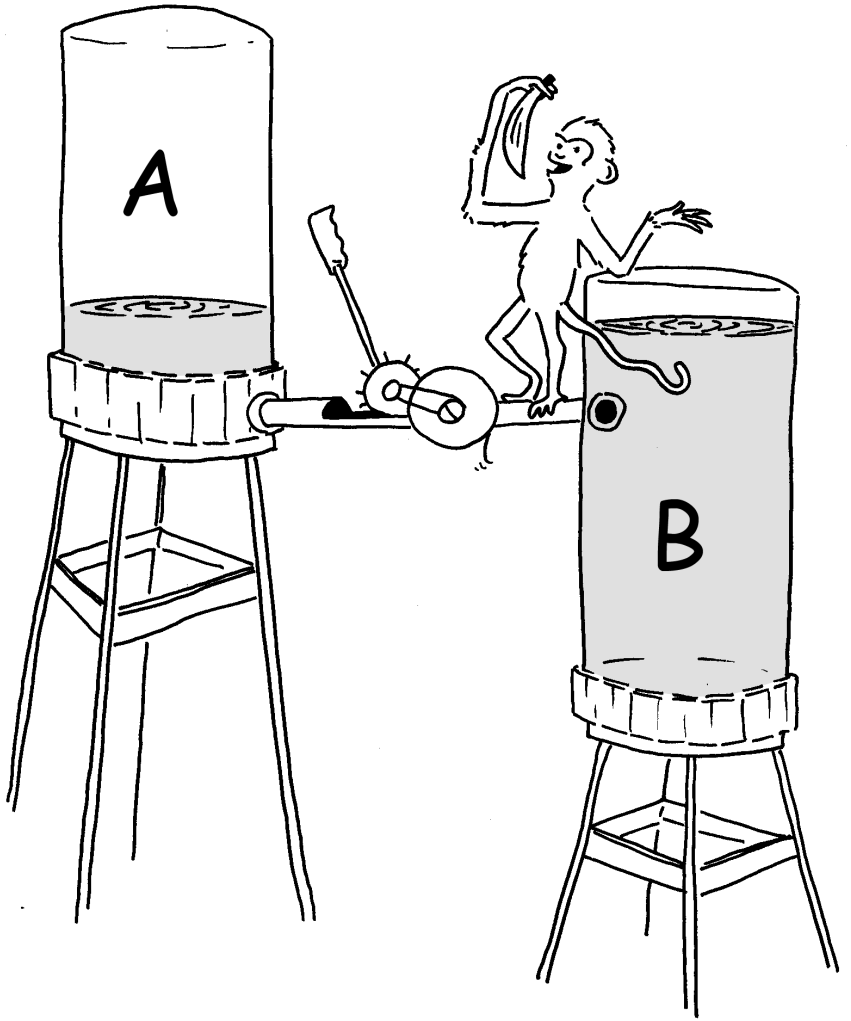
Starting with "standard" conditions:
1.0 M A, B, C, D; NET $A + B \leftarrow C + D$

Starting on the "other side" of equilibrium:
1.9 M A & B; 0.1 M C & D; NET $A + B \rightarrow C + D$

Concept: Energy available as equilibrium is approached.



Not at equilibrium



At equilibrium

Thermodynamics: how much energy released as a reaction proceeds towards equilibrium.

ΔG: Gibbs free energy

- Units of energy per mole (e.g., kJ/mol).
- Value of ΔG is a function of *how far from equilibrium*.

Negative ΔG (exergonic)
If rxn occurs, will proceed reactants → products.

Positive ΔG (endergonic)
If rxn occurs, will proceed products → reactants.

$\Delta G = 0$; reaction at equilibrium

Gibbs free energy has two components: **enthalpy** and **entropy** $\Delta G = \Delta H - T\Delta S$

1. Enthalpy Change (ΔH):

Chem: $\Delta H = \Delta E + P\Delta V$; *Biochem*: $\Delta H \approx \Delta E$, thus: $\Delta H \approx$ **difference in bond energies** between reactants and products.

- **exothermic** (releases heat): negative ΔH (more stable bonds formed & heat released – *e.g.*, candle burning).
- **endothermic** (heat input): positive ΔH (less stable bonds formed & heat absorbed).

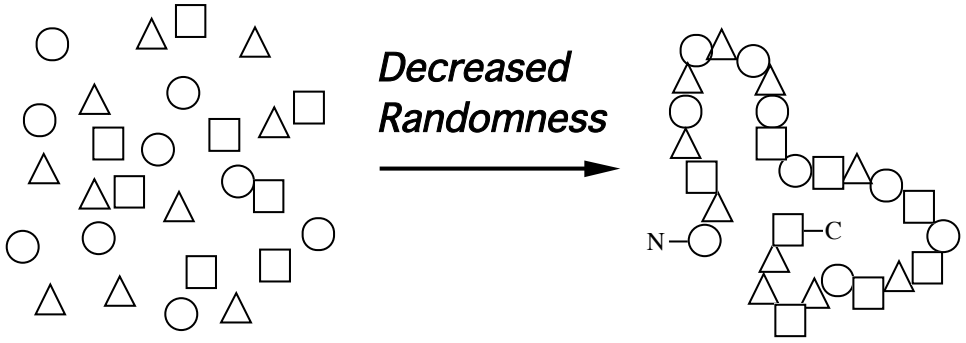
Exothermic (negative ΔH) contributes to a "favorable" (negative) ΔG. An endothermic reaction can also have a negative ΔG if there is a "over-riding" entropy term.

2. Entropy change (ΔS) = change in "randomness"

Increase in "randomness" is, by convention, a **positive ΔS** which contributes to a "favorable" (negative) ΔG.

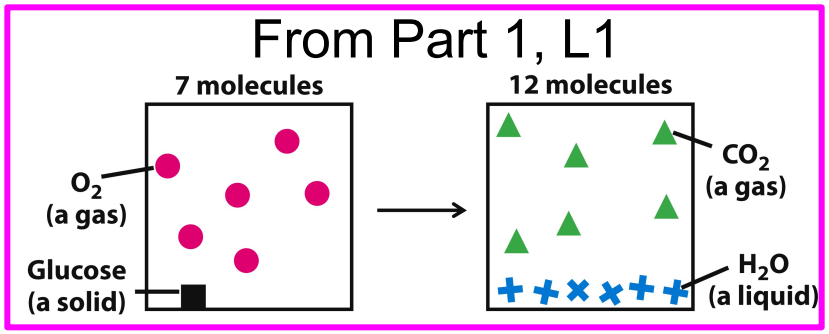
ΔS is multiplied by Temp (T) in eqn: $\Delta G = \Delta H - T\Delta S$.

- A system (e.g., a cell) can decrease local entropy only if a greater increase in entropy occurs in the surroundings. Example: protein synthesis requires other reactions in which overall entropy is increased (such as glucose oxidation & ATP breakdown).



Free Amino Acids

Folded polypeptide chain



ΔG and Keq relationship

Because ΔG is a measure of how far conditions are from equilibrium, Keq and ΔG are related; can calculate one from the other.

On exams will use simplified (and provided) formula:

$$\Delta G^o = -6 \log Keq$$

Notes:

- Standard conditions (1 molar; a "benchmark") indicated by superscript^o
- ΔG^o is energy change from *standard conditions to equilibrium*.
- ΔG for any other condition to equilibrium.
- Do not "worry" about this: Whether or not standard clearly indicated.

$\Delta G^o_{(\text{standard conditions})} = -RT \ln Keq$ (Equation 13-3 in text); assume constant temperature & natural ln to log10: $\Delta G^o = -5.69 \log Keq$ (in kJ/mol), "rounded" to $\Delta G^o = -6 \log Keq$

Relationship Between K_{eq} and ΔG^o

$$A + B \leftrightarrow C + D$$

$$K'_{eq} = \frac{[C]\,[D]}{[A]\,[B]}$$

$$\Delta G^o = -6 \log Keq$$

(negative)

If fwd rxn favored (C + D),
ΔG^o negative* and
Keq larger than 1*.
If reverse rxn favored(A + B),
ΔG^o positive# and
Keq less than 1#.

(like Table 13-2)

K_{eq}	ΔG° (kJ/mol)
#10^{-3}	#18
10^{-2}	12
10^{-1}	6
1	0.0
10^{1}	-6
10^{2}	-12
*10^{3}	*-18

Standard conditions: Why bother with a benchmark?

One of many reasons: Can calculate ΔG^{o} at standard conditions (Table 13-4) and determine actual ΔG at any concentration of reactants and products (next slide provides formula).

Please do not try to memorize any values from this table!

TABLE 13–4 Standard Free-Energy Changes of Some Chemical Reactions

Reaction type	$\Delta G'^{\circ}$ (kJ/mol)	$\Delta G'^{\circ}$ (kcal/mol)
Hydrolysis reactions		
Acid anhydrides		
Acetic anhydride + $H_2O \longrightarrow$ 2 acetate	−91.1	−21.8
ATP + $H_2O \longrightarrow$ ADP + P_i	−30.5	−7.3
ATP + $H_2O \longrightarrow$ AMP + PP_i	−45.6	−10.9
PP_i + $H_2O \longrightarrow$ 2P_i	−19.2	−4.6
UDP-glucose + $H_2O \longrightarrow$ UMP + glucose 1-phosphate	−43.0	−10.3
Esters		
Ethyl acetate + $H_2O \longrightarrow$ ethanol + acetate	−19.6	−4.7
Glucose 6-phosphate + $H_2O \longrightarrow$ glucose + P_i	−13.8	−3.3
Amides and peptides		
Glutamine + $H_2O \longrightarrow$ glutamate + NH_4^+	−14.2	−3.4
Glycylglycine + $H_2O \longrightarrow$ 2 glycine	−9.2	−2.2
Glycosides		
Maltose + $H_2O \longrightarrow$ 2 glucose	−15.5	−3.7
Lactose + $H_2O \longrightarrow$ glucose + galactose	−15.9	−3.8
Rearrangements		
Glucose 1-phosphate $\longrightarrow$ glucose 6-phosphate	−7.3	−1.7
Fructose 6-phosphate $\longrightarrow$ glucose 6-phosphate	−1.7	−0.4
Elimination of water		
Malate $\longrightarrow$ fumarate + H_2O	3.1	0.8
Oxidations with molecular oxygen		
Glucose + $6O_2 \longrightarrow 6CO_2 + 6H_2O$	−2,840	−686
Palmitate + $23O_2 \longrightarrow 16CO_2 + 16H_2O$	−9,770	−2,338

Table 13-4
Lehninger Principles of Biochemistry, Fifth Edition

Determining ΔG & Direction at Non-Standard Conditions

Actual ΔG depends on reactant and product concentrations:

- **$\Delta G = \Delta G^\circ + 6 \log$ [products]/[reactants]** (GIVEN)

actual ΔG — standard conditions from Table like 13-4 indicated by superscript^o — actual concentrations

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More formal treatment in book:

For a chemical reaction: $aA + bB \rightarrow cC + dD$

To calculate actual ΔG simply adjust for actual concentrations:

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

(Equation 13 - 4 in text)

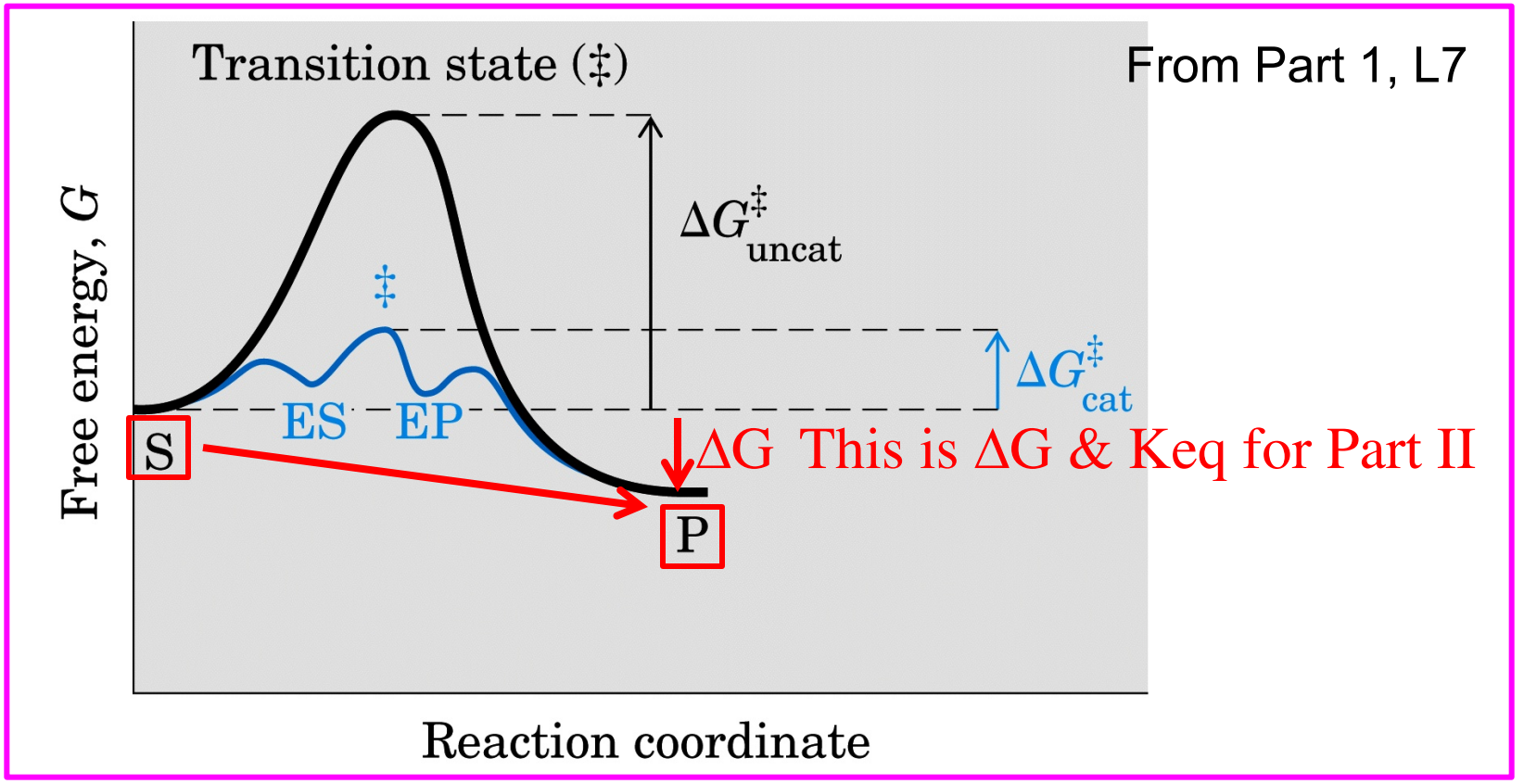
To review:

1) sign of ΔG reveals direction;
2) magnitude of ΔG indicates how far from equilibrium/how much energy will be released as reaction proceeds to equilibrium.

Thermodynamics (ΔG) does NOT predict how rapidly equilibrium is approached (rate).

As you know from Part 1, Enzymes change reaction rate (by lowering E_{act}), but do not change Keq.

Enzymes increase rate by lowering activation energy



The activation barrier for the enzyme catalyzed reaction is lower than the uncatalyzed reaction.

Enzymes may increase the rate of reaction by 10^5 - 10^{17} !

Enzymes change $\Delta G^{\ddagger}$ and thus reaction rate, but do NOT change ΔG or Keq.

Keq is related to **ΔG**; equilibrium state is a function of the inherent energy difference between S (reactants) & P (products).

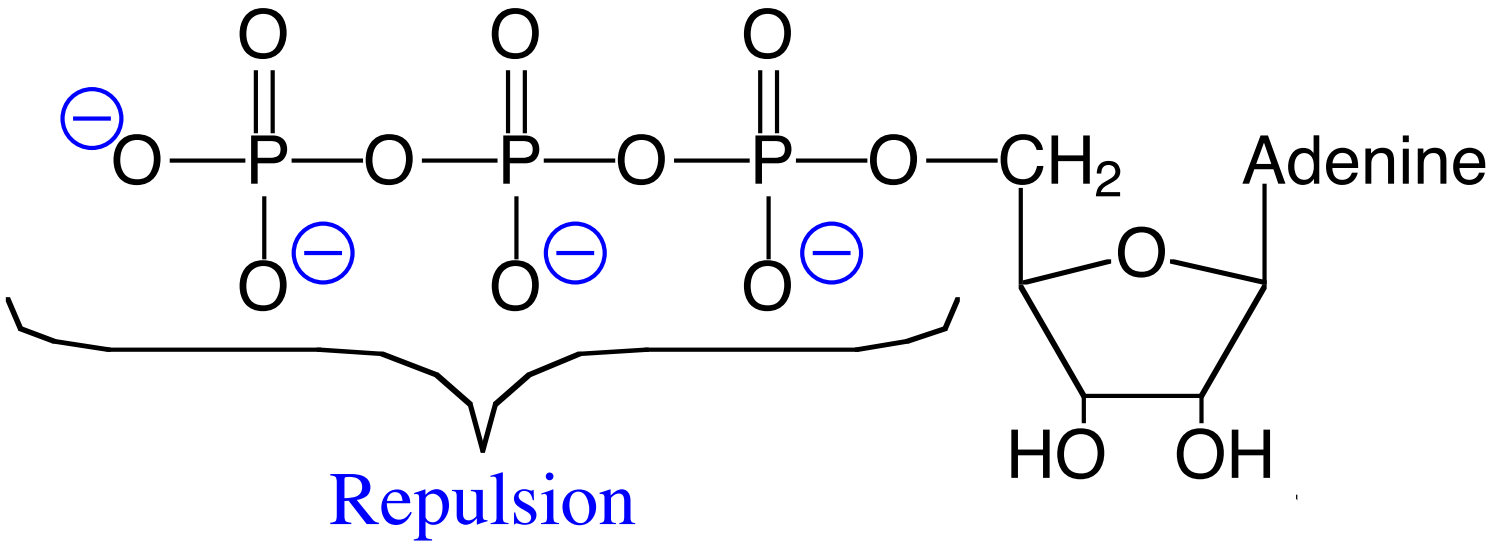
Formulas and Part II

- Examples of use of ΔG to calculate reaction direction in Problem Sets and Sample Exam.

- All formulas will be provided on the cover page of the second exam (see Sample Exam).

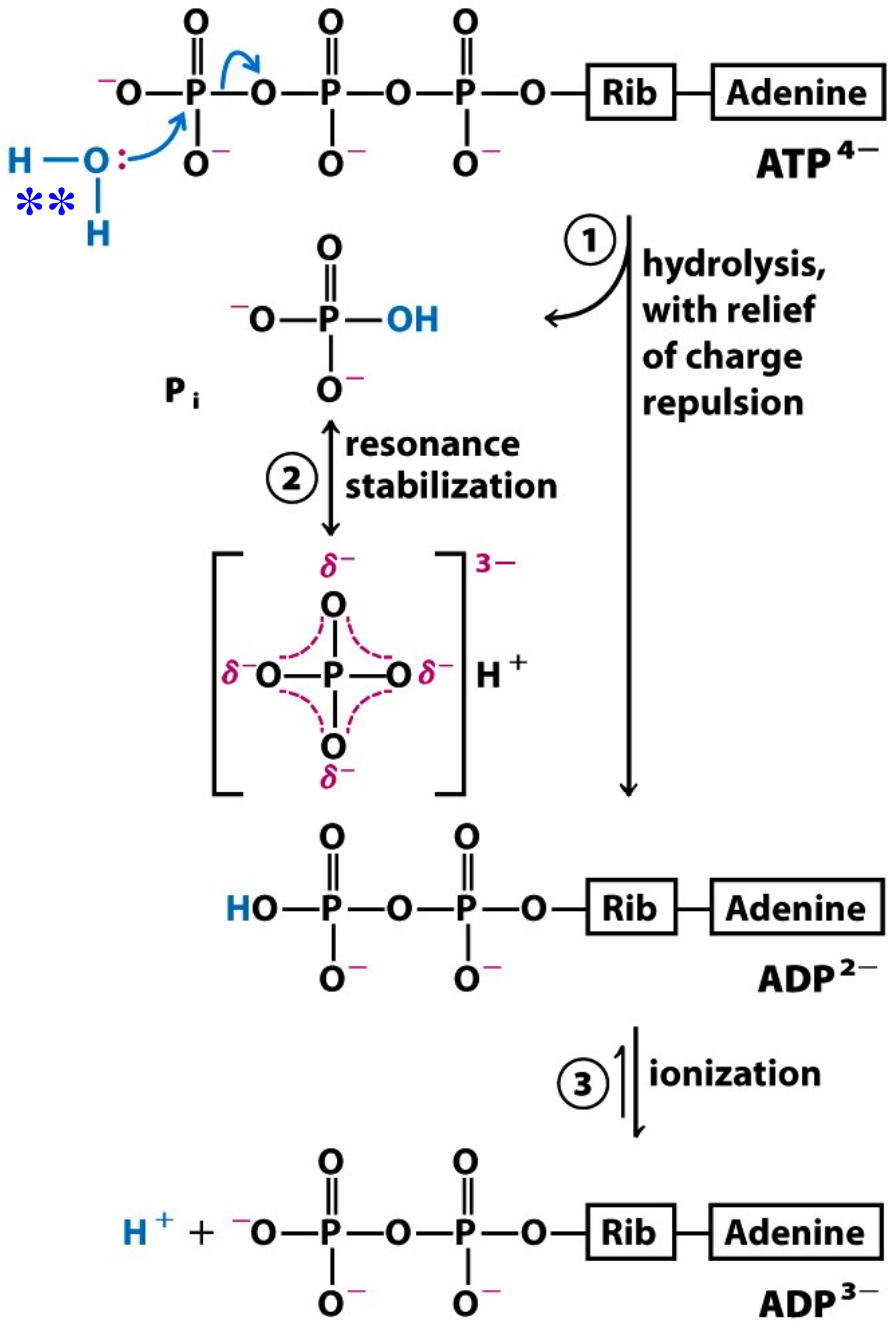
ATP - the Universal Energy Carrier

Similar to 13-11 & 13-12 in 5e; 13-12 in 6e; figure notes reference text position



Adenosine triphosphate
(ATP)

- Kinetically stable, little non-enzymatic breakdown



Basis of Large Free Energy Change (Negative ΔG) Associated With ATP hydrolysis** (–30.5kJ/mol)

Electrostatic repulsion between O^-

Resonance stabilization of products

Ionization stabilization and solvation

$$ATP^{4-} + H_2O \longrightarrow ADP^{3-} + P_i^{2-} + H^+$$
$$\Delta G'^\circ = -30.5\ kJ/mol$$

Figure 13-11
Lehninger Principles of Biochemistry, Fifth Edition

Figure 13-11 in all editions

In vivo conditions of ATP hydrolysis

ΔG^o of hydrolysis = -30.5 kJ/mol ($ATP + H_2O \rightarrow ADP + Pi$)

In a cell: [ATP] / [ADP] [Pi] = ~500

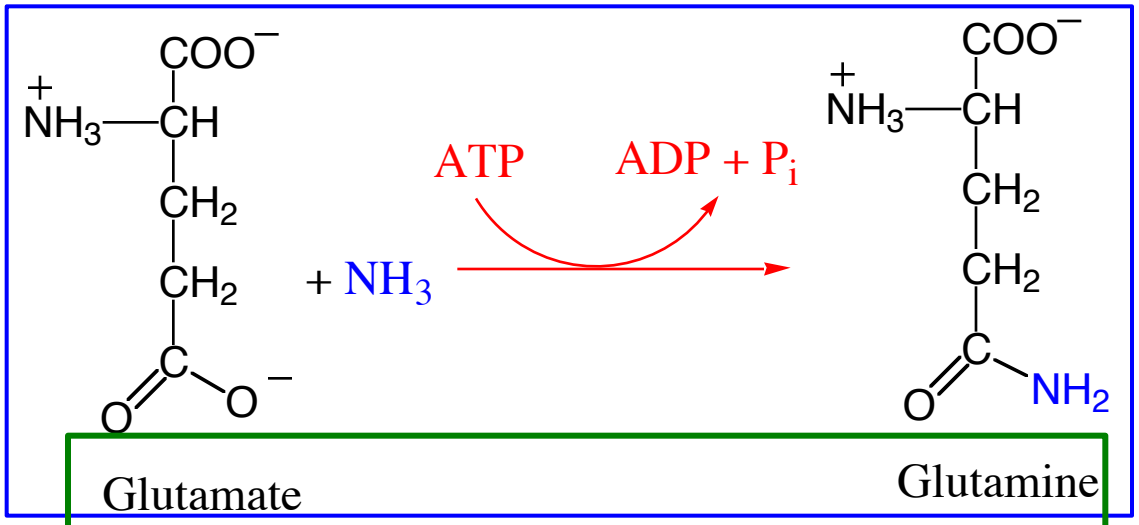
Actual ΔG of hydrolysis* in cell = ~-50 kJ/mol, thus a good energy source (corollary - requires ~+50 kJ / mol to produce ATP) Box 13-1 in 4e; Worked Example 13-2 in 5e/6e

Note: *direct* ATP hydrolysis not a common reaction; hydrolysis is used as a benchmark for comparison.

Muscle contraction among exceptions: Figs 5-30 & 5-31 in all editions

ATP Usually Provides Energy by Group Transfers, Not by Direct Hydrolysis

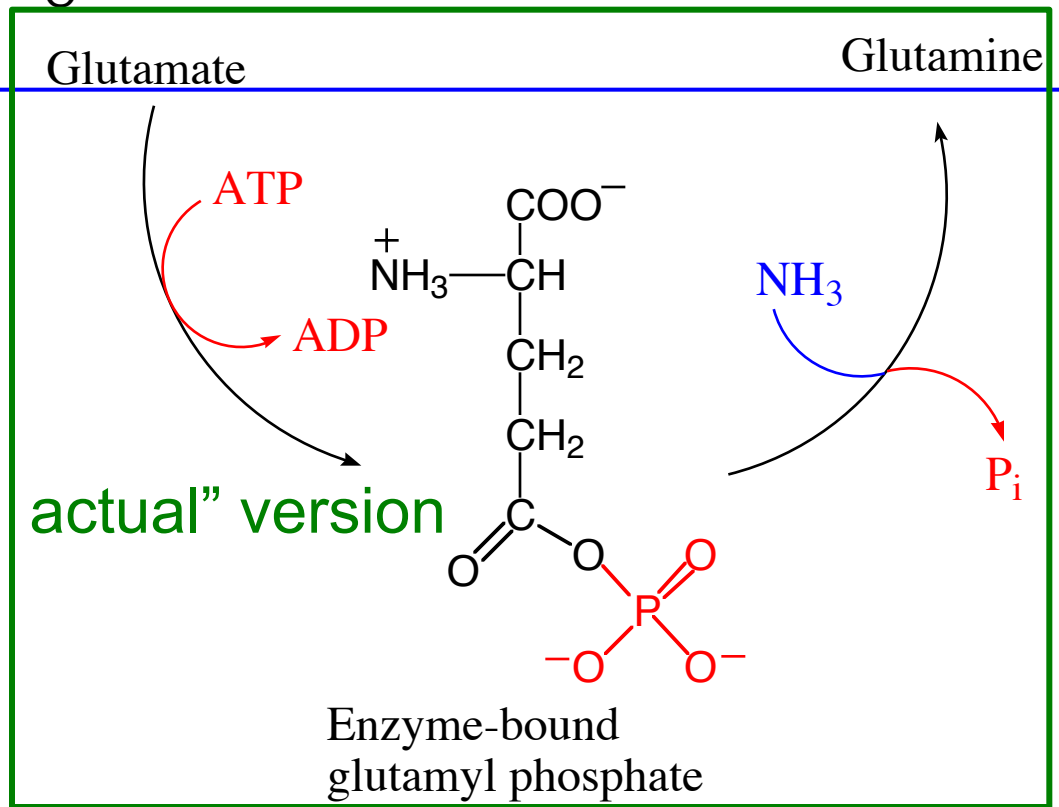
"shorthand" version



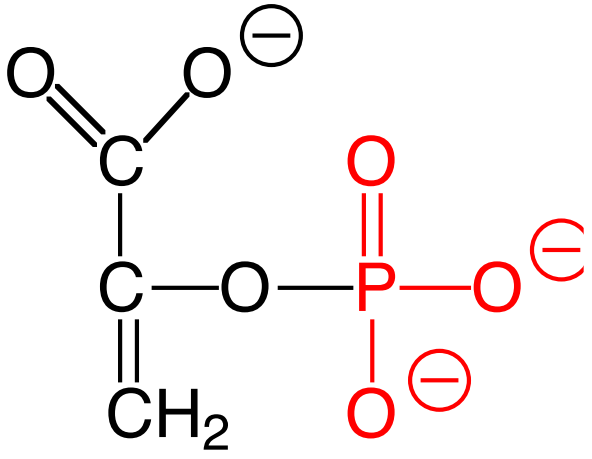
Group transfer "energizes" recipient.

"Coupling" to ATP breakdown renders amine formation favorable.

like Figure 13-8



Other P compounds with high energy of hydrolysis



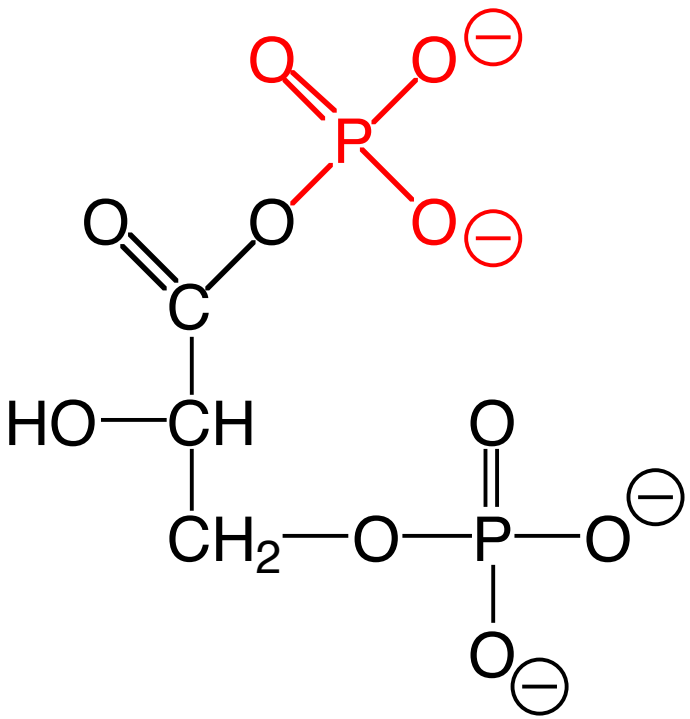
Phospho-enol pyruvate (PEP)

$\Delta G^{o} = -61.9$ kJ/mole: *higher energy than ATP*

(see Figure 13-13; p 504 5e—chemical discussion of why high energy)

PEP hydrolysis does not occur in glycolysis; hydrolysis is a benchmark for comparison.

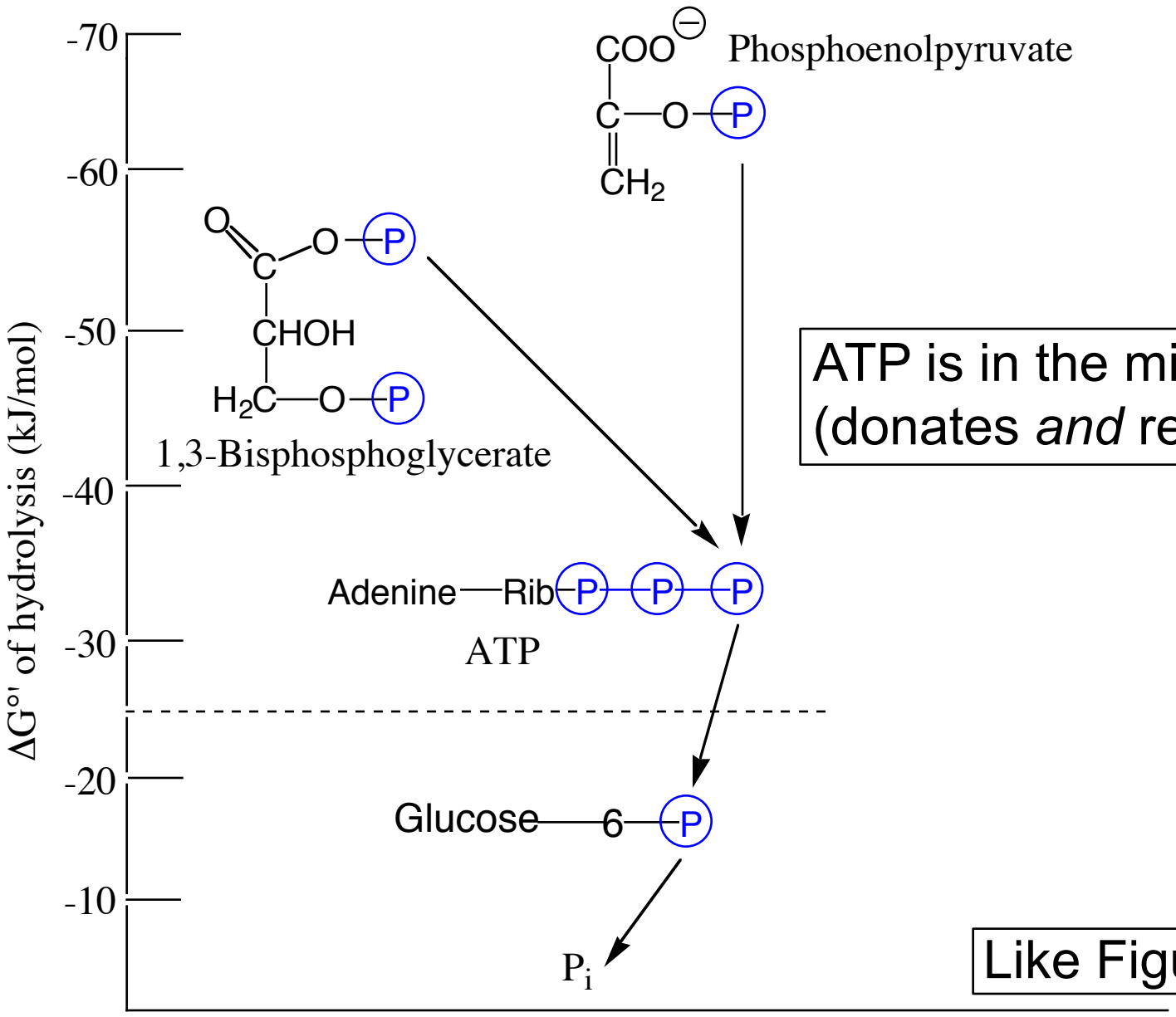
Other P compounds with high energy of hydrolysis



1,3-Bisphosphoglyceric Acid
(1,3-diPGA)

$\Delta G^{o} = -49.3$ kJ/mole (see Figure 13-14)

Much of metabolism involves the synthesis of high energy phosphate compounds



ATP is in the middle
(donates *and* receives **P**)

Like Figure 13-9