

biochemistry



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SIXTH EDITION

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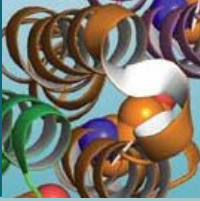
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Chapter 3

Thermodynamics of Biological Systems

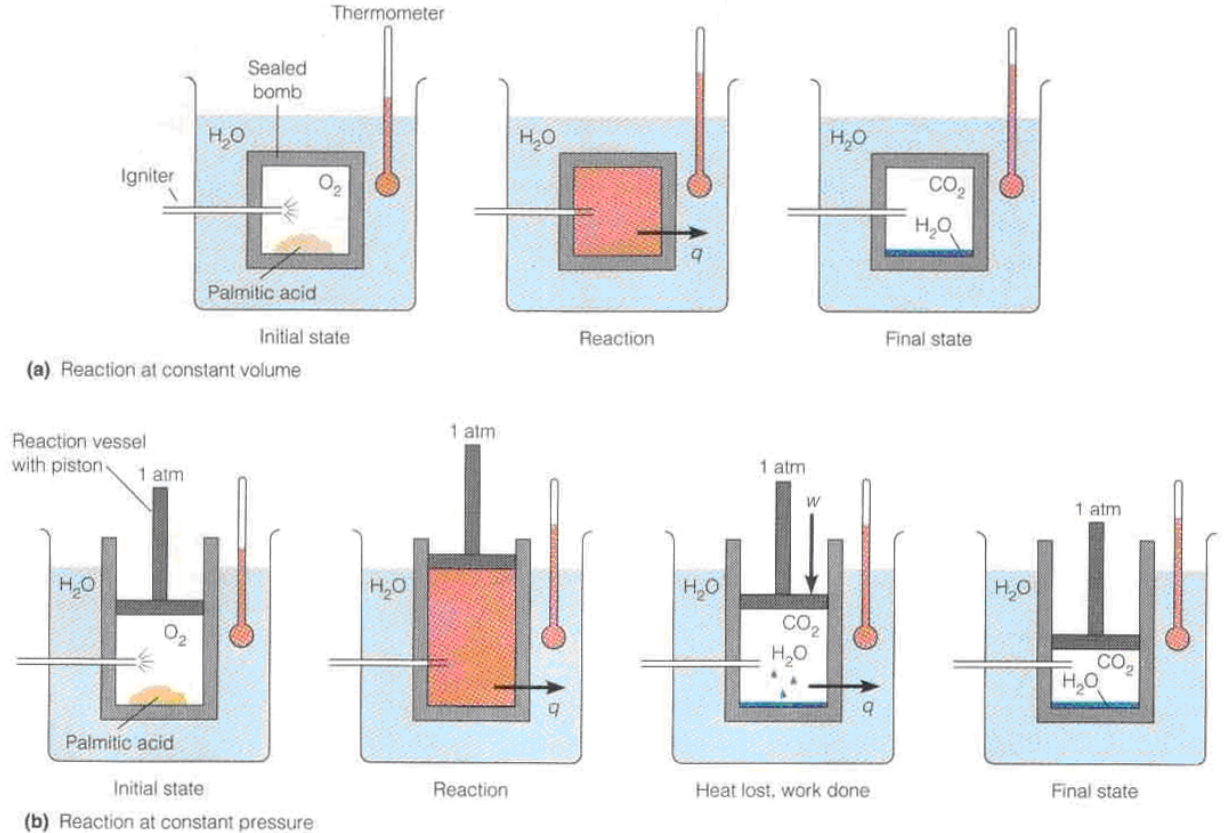
(Substitute Chapter 3 from Matthews et al. is
posted on eLearning)

Outline



- 1st & 2nd Laws of Thermodynamics
- Enthalpy & entropy
- Gibbs free energy equations for reactions
- Coupled reactions
- ATP, how it works as the energy carrier of the cell

Fig. 3.1 Exchange of heat & work in
(a) constant volume
and (b) constant
pressure reactions



$$\Delta E = q - w$$

At constant volume,

$$w = 0 \text{ thus } \Delta E_v = q_v = -9941 \text{ kJ/mol}$$

At constant pressure,

$$q_p = q_v + w = -9941 \text{ kJ/mol} - 17.3 \text{ kJ/mol}$$

$$q_p = \Delta E_v + P\Delta V = \Delta H \text{ (enthalpy)}$$

$$\text{since, } w = P\Delta V = \Delta nRT$$

Example Reactions Involving Weak Interactions

Reactants	ΔH	ΔG
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$H_2O +$

$LiCl$

-

-

$(NH_4)_2SO_4$

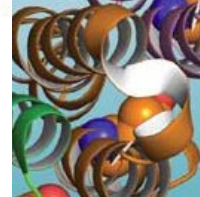
+

-

enthalpy change = $\Delta H = H_p - H_R$

entropy change = $\Delta S = S_p - S_R$

Gibbs Free Energy change = $\Delta G = G_p - G_R$



Entropy

$$S = k \ln W$$

where

$$k = 1.38 \times 10^{-23} \text{ J/K}$$

(Boltzmann Constant)

and

W = number of equal energy arrangements of a system

For entropy per mole, multiply by N (Avogadro's no.)

$$S = kN \ln W$$

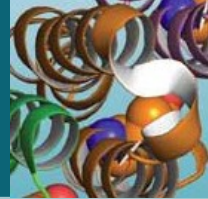
$$= R \ln W$$

$$= 2.303 R \log W$$

where

$$R = 8.314 \text{ J/K} \cdot \text{mol}$$

(Gas Constant)



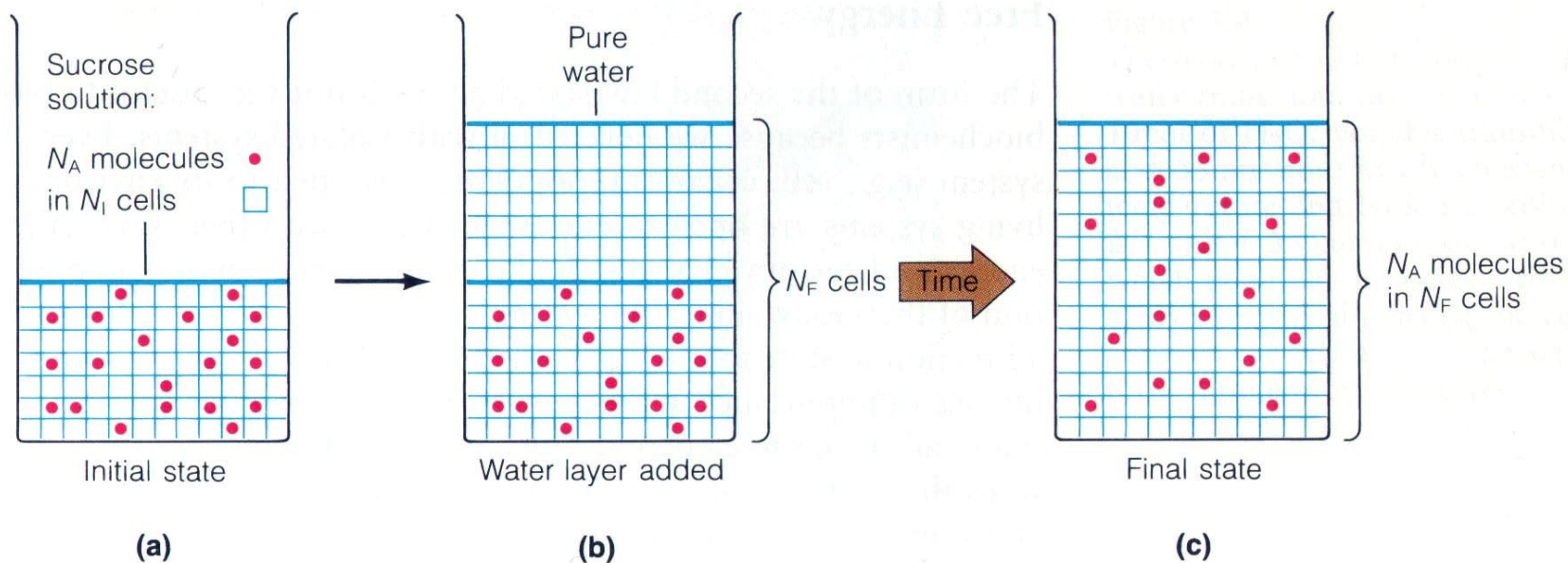
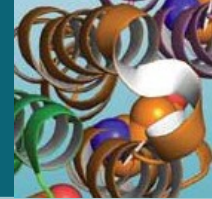


Figure 3.2
Diffusion as an entropy-driven process.

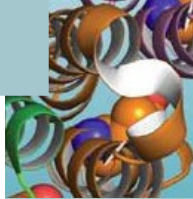
Entropy of Propane

type	amount (kcal/K.mol)
translational	36
rotational	23
vibrational	1
electronic	0

$$1 \text{ cal} = 4.184 \text{ J}$$



Example Reactions Involving Weak Interactions



Reactants ΔH $-(T\Delta S) = \Delta G$

$H_2O +$

$LiCl$

- - (-) -

$(NH_4)_2SO_4$

+ - (+) -

enthalpy change = $\Delta H = H_p - H_R$

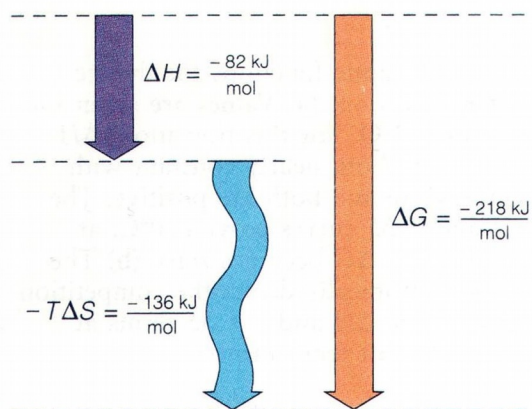
entropy change = $\Delta S = S_p - S_R$

Gibbs Free Energy change = $\Delta G = G_p - G_R$

Reactions are driven by enthalpy and entropy, alone or together



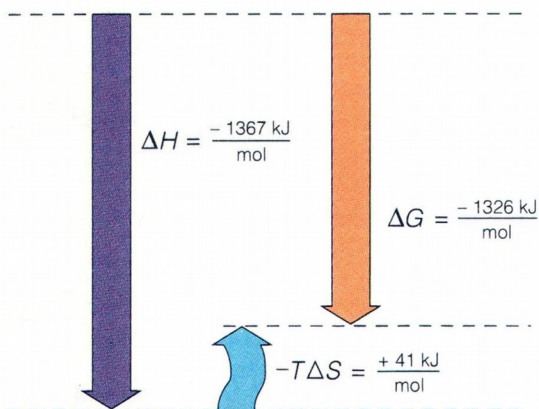
Figure 3.4



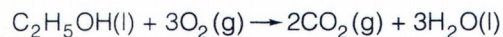
(a) Fermentation of glucose to ethanol



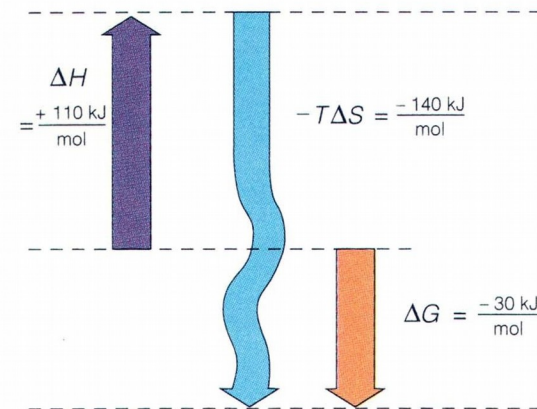
Both enthalpy and entropy changes favor the reaction.



(b) Combustion of ethanol



Enthalpy favors this reaction, but entropy opposes it. We could call this an "enthalpy-driven" reaction. If water *vapor* were the product, an entropy increase would favor the reaction as well.



(c) Decomposition of nitrogen pentoxide



This is a somewhat unusual chemical reaction in that it is "entropy-driven." The reaction actually absorbs heat but is favored by the large entropy increase resulting from the formation of gaseous products.

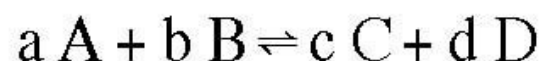
Gibbs Free Energy Equations

In terms of enthalpy and entropy:

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

In terms of reactant and product concentrations:

$$\overline{G}_A = G_A^\circ + RT \ln[A]$$



$$\Delta G = (\overline{G}_C + \overline{G}_D) - (\overline{G}_A + \overline{G}_B)$$

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

$$\text{At equilibrium} \quad \Delta G = 0 \quad \& \quad \ln \frac{[C]^c [D]^d}{[A]^a [B]^b} = \ln K_{eq}$$

$$\therefore \Delta G_{RX}^\circ = -RT \ln K_{eq}$$



Standard State

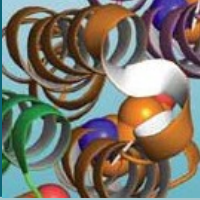
If all concentrations in a reaction are 1M,
or 1 atm for gases,

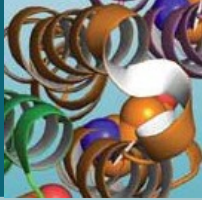
$$\Delta G = \Delta G_{RX}^{\circ} + RT \ln \frac{[1]^c [1]^d}{[1]^a [1]^b}$$
$$\therefore \Delta G = \Delta G_{RX}^{\circ}$$

In chemistry, these are taken as the
standard state conditions

But in biochemical systems, an aqueous
solution of pH 7 instead of pH 0 is defined
as the standard state, and the $\ln[H^+]$ value
of $\ln(10^{-7})$ is incorporated into ΔG_{RX}° ,
yielding the constant $\Delta G_{RX}^{\circ'}$

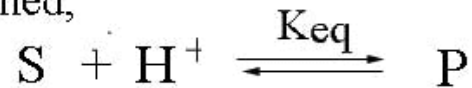
Also, if [Phosphate] either as Pi
or incorporated phosphate groups
is involved, the concentration is
the sum of all ionization states present,
i.e., $P^0 + P^{-1} + P^{-2} + \text{etc.}$





$$\Delta G^{\circ} \text{ vs. } \Delta G^{\circ'}$$

If H^{+} consumed,



$$K_{eq} = \frac{[P]_{eq}}{[S]_{eq} [H^{+}]_{eq}}$$

$$\Delta G = \Delta G^{\circ}_{RX} + RT \ln \frac{[P]}{[S][H^{+}]}$$

$$\Delta G = \Delta G^{\circ} - RT \ln [H^{+}] + RT \ln \frac{[P]}{[S]}$$

$$\Delta G' = \Delta G^{\circ'} + RT \ln \frac{[P]}{[S]}$$

At pH 7,

$$\Delta G^{\circ'} = \Delta G^{\circ} - RT \ln [10^{-7}]$$

In terms of K_{eq} ,

$$-RT \ln K_{eq}' = -RT \ln K_{eq} - RT \ln 10^{-7}$$

$$K_{eq}' = 10^{-7} K_{eq}$$

NB: Equations 3.18 and 3.21 on p.59 of Garrett & Grisham 3rd ed. are transposed.

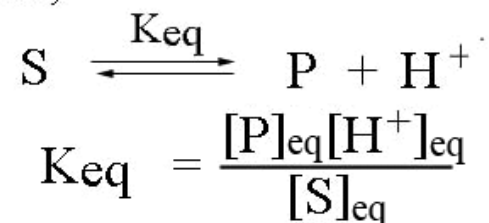
Equivalent equations 3.23 and 3.26 in 4th ed. are OK.

- These are equilibration concentrations
- These are variable concentrations



$$\Delta G^{\circ} \text{ vs. } \Delta G^{\circ'}$$

If H^+ produced,



$$\Delta G = \Delta G^{\circ}_{RX} + RT \ln \frac{[P][H^+]}{[S]}$$

$$\Delta G = \underbrace{\Delta G^{\circ}}_{\Delta G^{\circ'}} + RT \ln [H^+] + RT \ln \frac{[P]}{[S]}$$

$$\Delta G' = \Delta G^{\circ'} + RT \ln \frac{[P]}{[S]}$$

At pH 7,

$$\Delta G^{\circ'} = \Delta G^{\circ} + RT \ln [10^{-7}]$$

In terms of K_{eq} ,

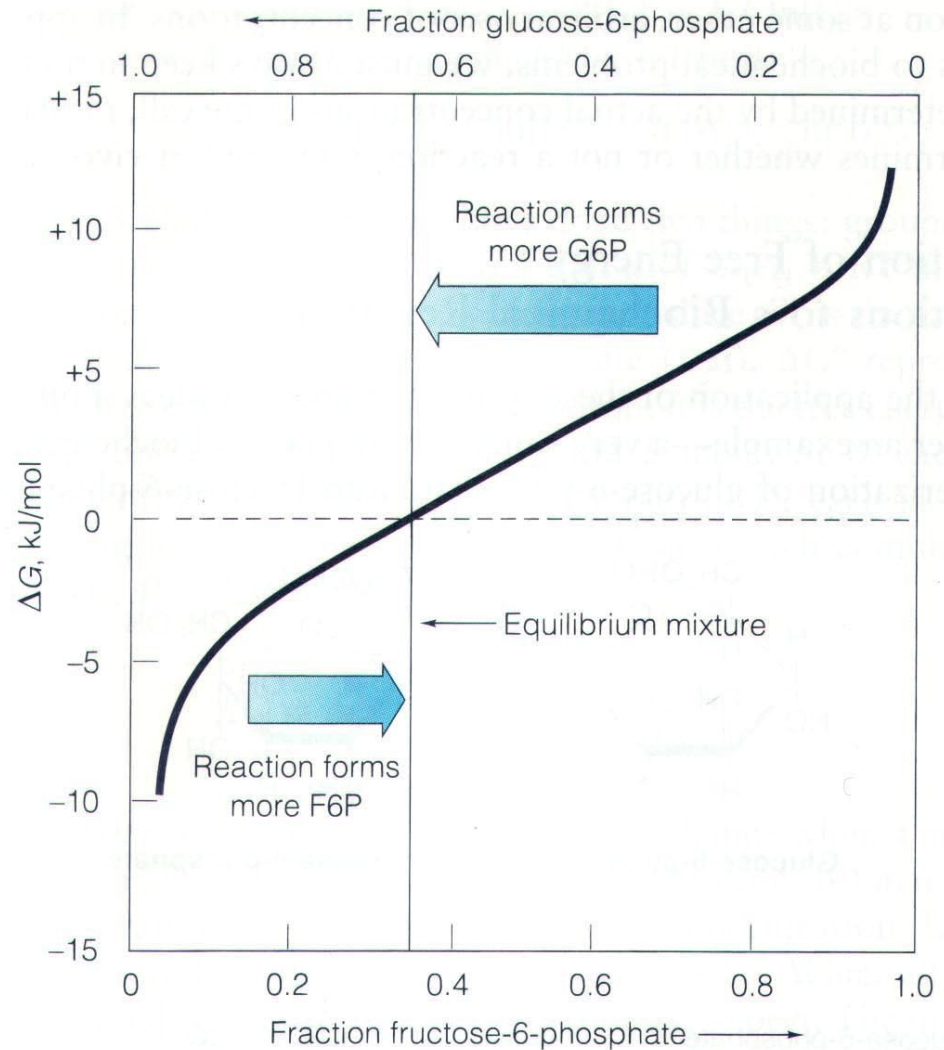
$$-RT \ln K_{eq}' = -RT \ln K_{eq} + RT \ln 10^{-7}$$

$$\ln K_{eq}' = \ln K_{eq} - \ln 10^{-7}$$

$$K_{eq}' = \frac{K_{eq}}{10^{-7}}$$

Fig. 3.6 Free energy change as a function of reaction mixture composition. The black curve shows ΔG for the conversion of 1 mole **glucose**-6-phosphate to **fructose**-6-phosphate as a function of the fraction of either component.

Figure 3.6



ΔG vs. [reactants]



$$K_{eq} = \frac{1}{2}$$

At equilibrium

$$\Delta G = \Delta G^\circ + RT \ln K_{eq} = 0$$

$$= \Delta G^\circ + 2.303 RT \log \frac{1}{2} = 0$$

If substrate G-G-P increased 10-fold

$$\Delta G = \Delta G^\circ + 2.303 RT \log \frac{1}{2 \cdot 10}$$

$$= \Delta G^\circ + 5.7 \text{ kJ/mol} \cdot \log \frac{1}{2} + 5.7 \text{ kJ/mol} \cdot \log \frac{1}{10}$$

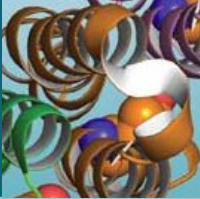
$$= -5.7 \text{ kJ/mol}$$

If product F-G-P increased 10-fold

$$\Delta G = \Delta G^\circ + 5.7 \text{ kJ/mol} \cdot \log \frac{10}{2}$$

$$= +5.7 \text{ kJ/mol} \cdot \log 10$$

$$= +5.7 \text{ kJ/mol}$$





Determination of ΔH° and ΔS°

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

$$-RT\ln K_{eq} = \Delta H^\circ - T\Delta S^\circ$$

$$\ln K_{eq} = -\frac{\Delta H^\circ}{R} \frac{1}{T} + \frac{\Delta S^\circ}{R}$$

Assuming that enthalpy and entropy are invariant with temperature, you can plot $\ln K_{eq}$ vs. $1/T$.

$$\text{Slope} = -\frac{\Delta H^\circ}{R}$$

$$\text{y-intercept} = \frac{\Delta S^\circ}{R}$$

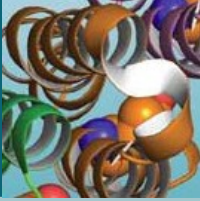
This is called a van't Hoff plot.



Coupled Reactions

- Bioreactions with unfavorable $\Delta G^{0'}$ are often driven in the forward direction by coupling the reaction to hydrolysis of a high-energy phosphate compound.
- The overall $\Delta G^{0'}$ for a coupled reaction is the sum of the two $\Delta G^{0'}$ values for the individual reactions.
- ATP is the most commonly used high-energy compound. It can be considered the energy currency of the cell.

ATP Contains Two Pyrophosphate Linkages



(from Garrett & Grisham, page 63)

Figure 3.4
ATP contains two pyrophosphate linkages. The hydrolysis of phosphoric acid anhydrides is highly favorable.

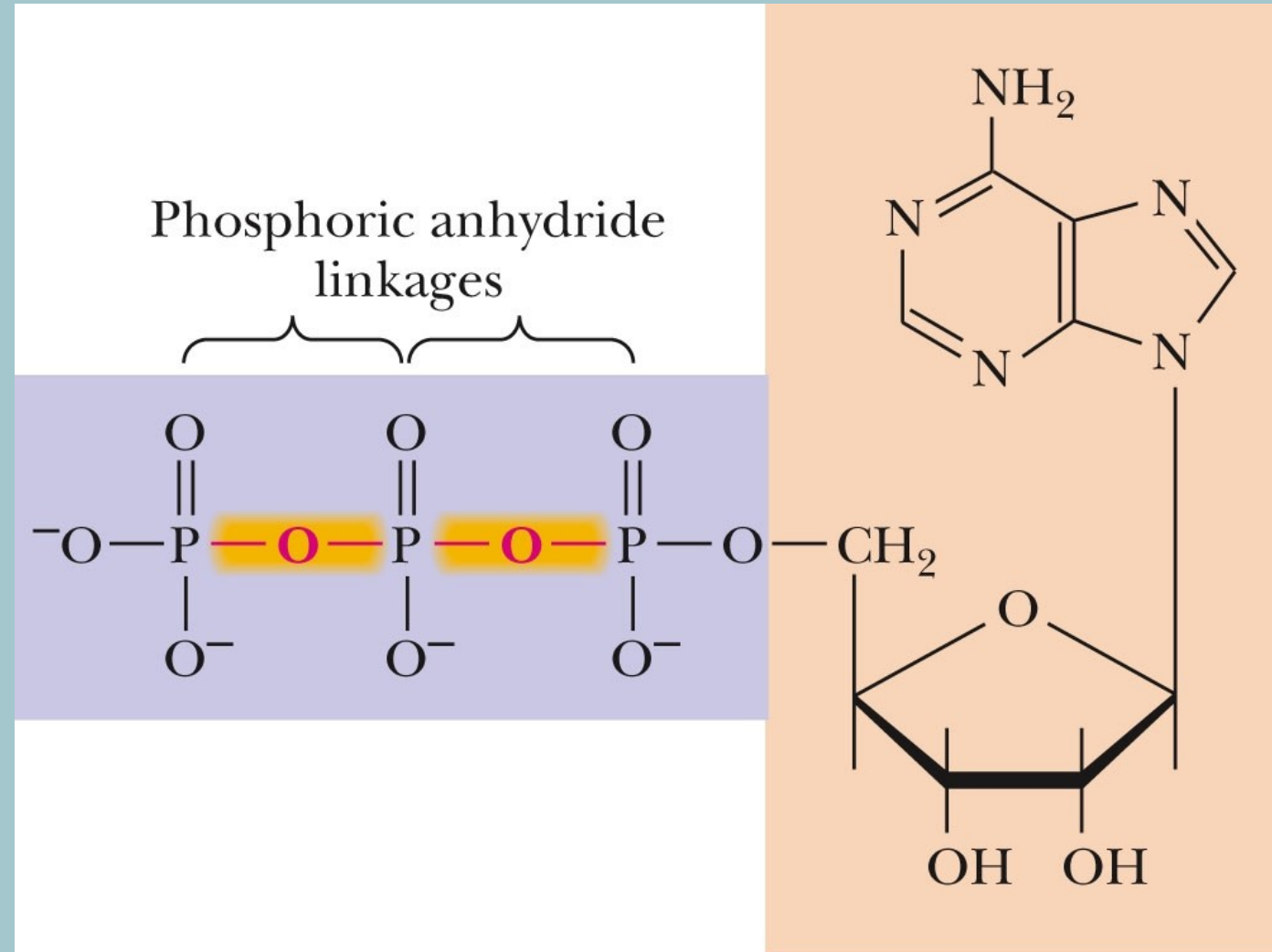
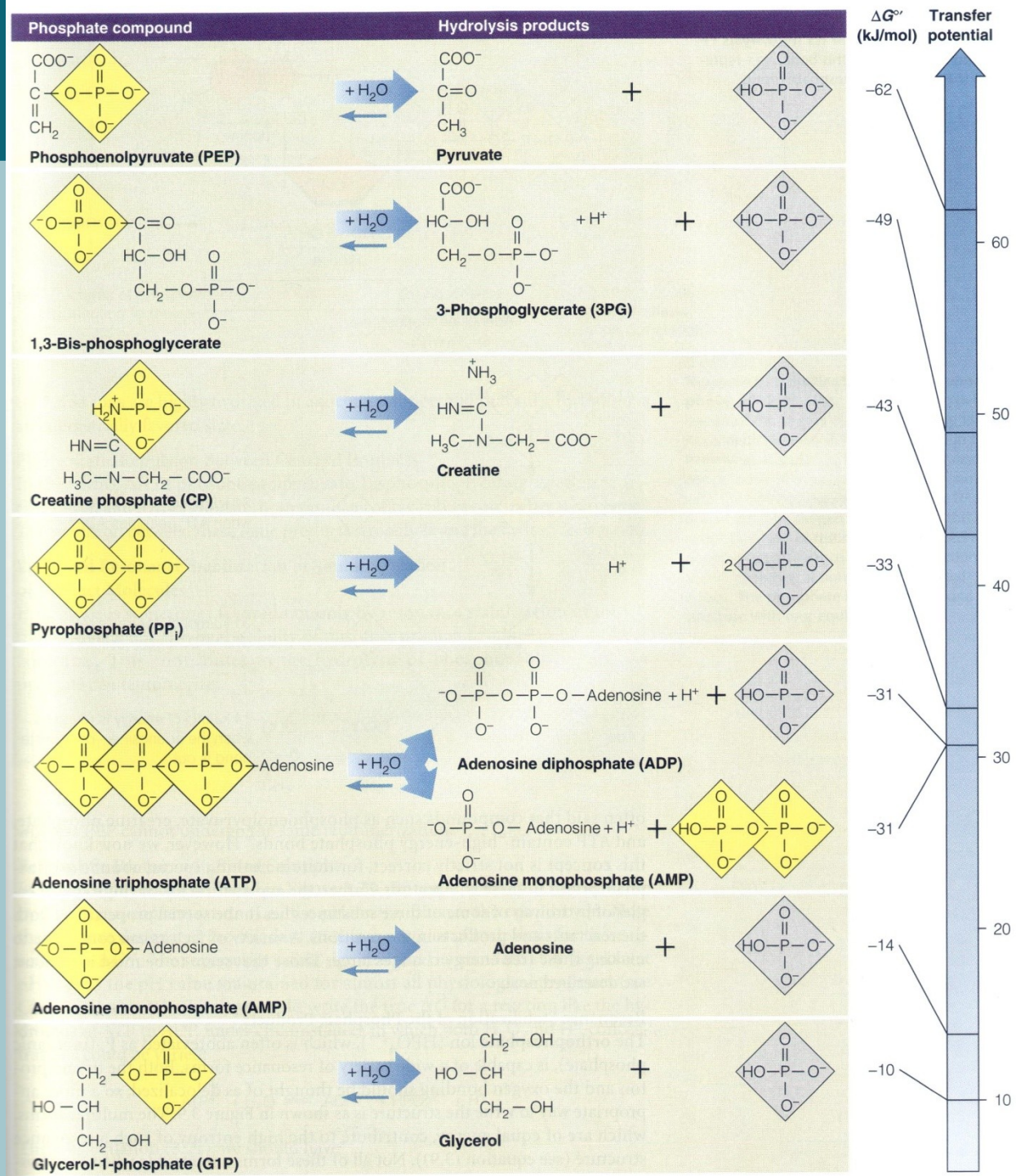
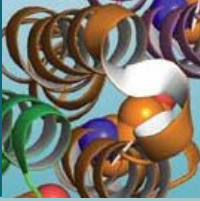


Fig. 3.7 Hydrolysis reactions for some biologically important 'high-energy' phosphate compounds. Labile phosphate groups highlighted in yellow; Pi products in blue. $\Delta G^{\circ'}$ values for the reactions (aka. transfer potentials for the labile phosphates) are shown to the right.



Concentration Affects the Free Energy of Hydrolysis of ATP



(from Garrett & Grisham, page 68)

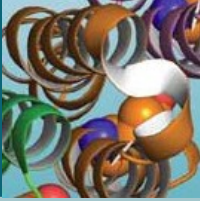
Eq. 3.28

$$\Delta G' = \Delta G^{\circ'} + RT \ln \frac{[\Sigma ADP][\Sigma Pi]}{[\Sigma ATP]}$$

In a typical bacterial cell at 25° C, [ATP], [ADP], & [Pi] are maintained at roughly 8 mM, 8 mM, & 1 mM.

Under these conditions, $\Delta G'$ for ATP hydrolysis is approximately - 47.6 kJ/mol.

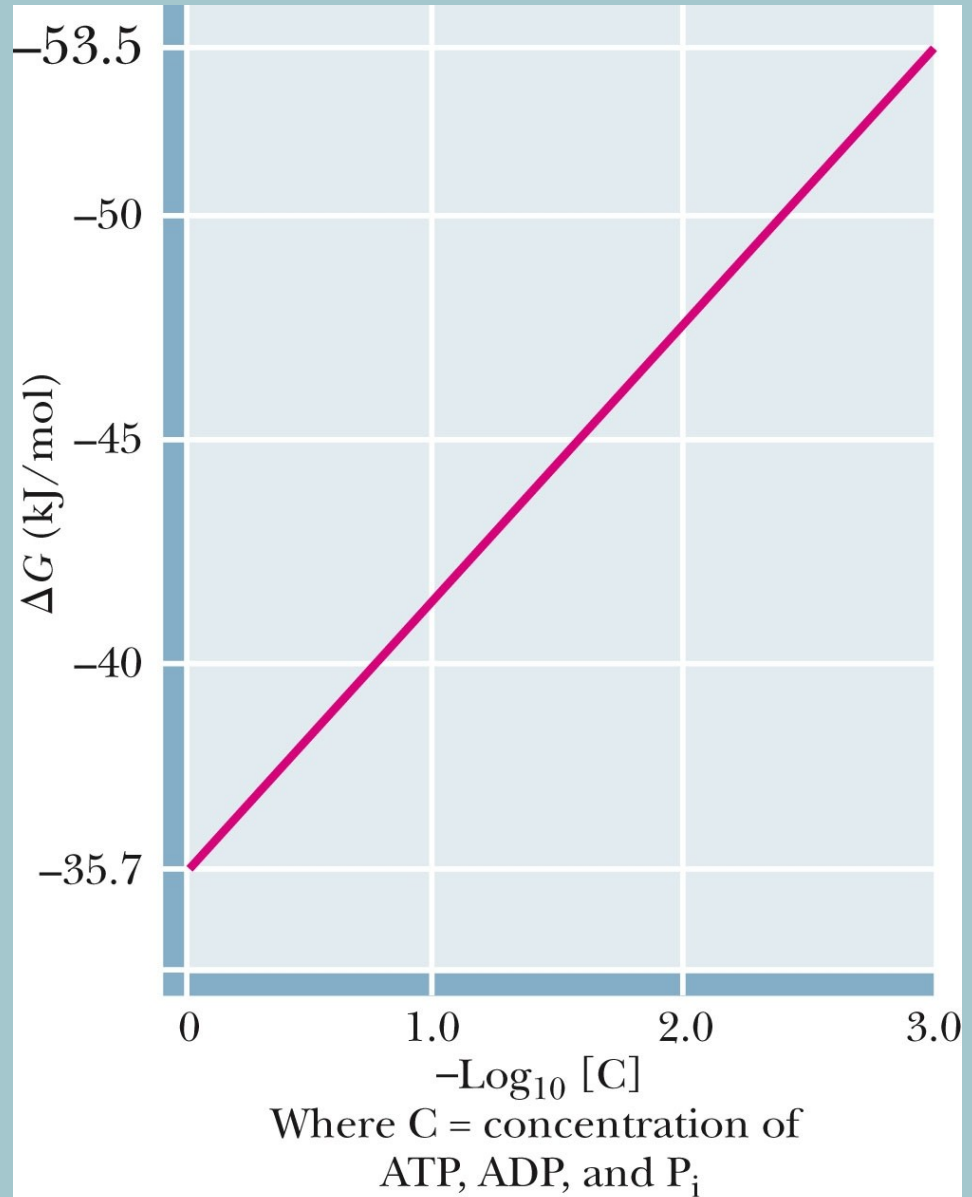
Concentration Affects the Free Energy of Hydrolysis of ATP



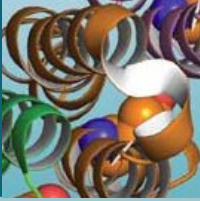
(from Garrett & Grisham, page 69)

Figure 3.14 The free energy of hydrolysis of ATP as a function of concentration at 38° C, pH 7.0.

The plot follows the relationship described in Eq. 3.28, with the concentrations [C] of ATP, ADP, and P_i assumed to be equal.



Phosphoric Acid Anhydrides



Why ATP does what it does!

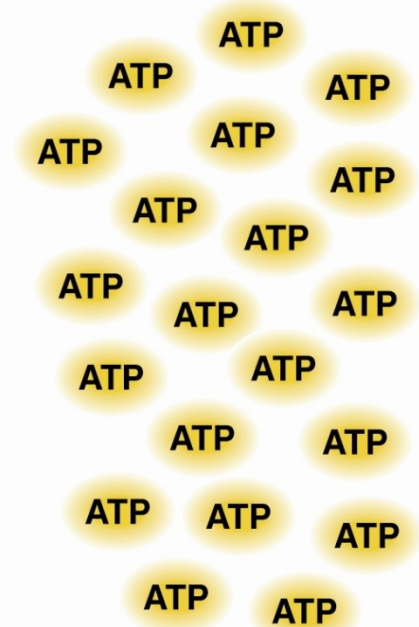
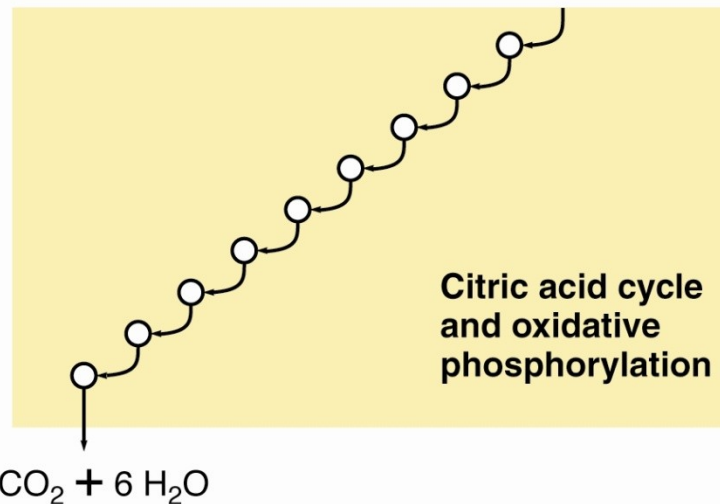
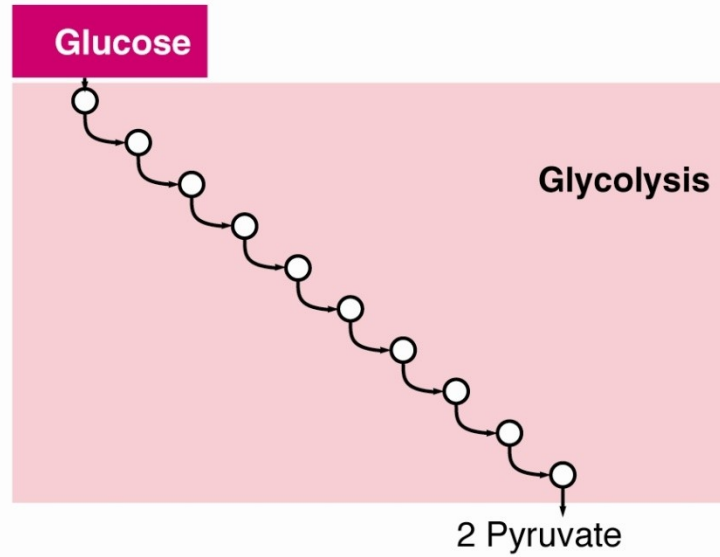
- ADP and ATP are examples of phosphoric acid anhydrides
- Large negative free energy change on hydrolysis is due to:
 - electrostatic repulsion
 - stabilization of products by ionization and resonance
 - entropy factors

1.4 – Properties of Biomolecules Reflect Their Fitness to the Living Condition



The combustion of glucose: $\text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2 \longrightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O} + 2870 \text{ kJ energy}$

(a) In an aerobic cell



30–38 ATP

Cells release the energy of glucose in a stepwise fashion, forming ATP