

**Energy in chemical reactions**

**Bioenergetics**

**Kinetics**

**Redox reactions**

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# Basic terms: system and energy

- A portion of universe separated from surroundings
  - **insulated** – no communication with surroundings
  - **closed** – exchange of energy
  - **open** – exchange of energy, matter, and information
- 
- **Energy** is the capacity of a system to do work
  - system = object, reaction mixture, cell, organism
  - Unit - joule (J), dimension  $J = \text{kg m}^2/\text{s}^2$

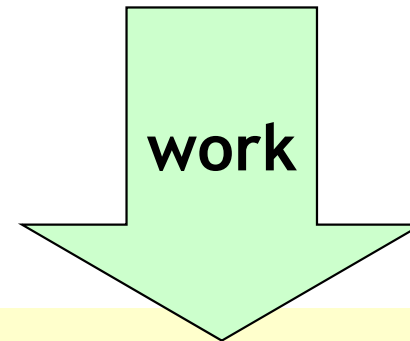
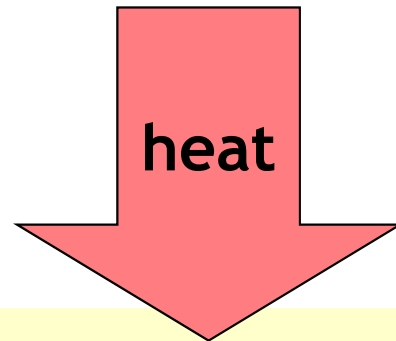
$$\Delta U = U_2 - U_1 = U_{\text{final}} - U_{\text{initial}}$$

$\Delta U > 0$  energy of system increases

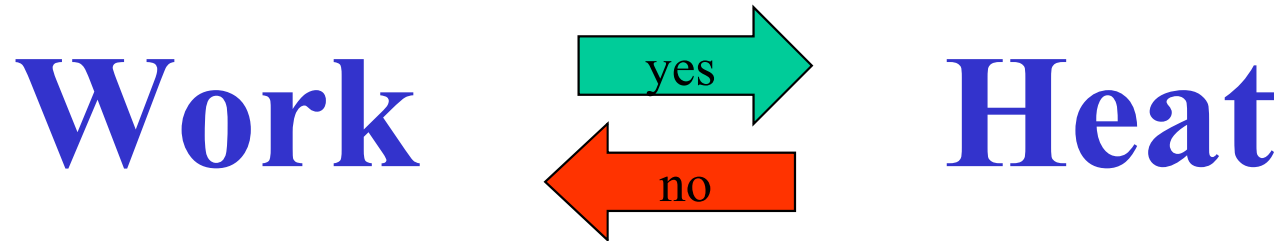
$\Delta U < 0$  energy decreases

## Internal energy ( $U$ )

- the sum of all forms of energy in the system
- the energy of all particles - atoms, molecules, ions
- the total energy  $U$  cannot be determined
- only the change ( $\Delta U$ ) during some event in the system



$$\Delta U = \Delta Q + \Delta W$$



- they both describe **energy in process** (the change of energy)
- **heat** (thermal energy) is associated with random motion of atoms and molecules, it is the transfer of energy between two bodies that are at different temperatures
- **work** – useful energy, can be transformed to other forms of energy, and also completely into heat
- the opposite is not possible  $\Rightarrow$  therefore heat is considered to be less utilizable form of energy (energy waste)

# Exchange of heat in the human body

**Exogeneous heat intake:** hot drinks, sun exposure ...

**Endogeneous heat production:** meal, physical exercises

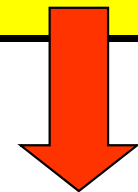
- human body cannot utilize heat, we need chemical energy of nutrients
- heat just contributes to keeping body temperature
- heat intake has limitations

## **Heat output:**

- radiation
- conduction + convection (facilitated by wind)
- sweat evaporation

# Heat output in various situations

Output way	Conditions (temperature, humidity)			
	20 °C / dry	30 °C / dry	36 °C / dry	36 °C / humid
Radiation	61 %	46 %	0 %	0 %
Conduction	26 %	27 %	0 %	0 %
Sweat evap.	13 %	27 %	100 %	0 %



Thermodynamically impossible heat transfer problem, body cannot remove heat, hyperthermia may be fatal.

## Hyperthermia (heat stroke)

- more heat is accumulated in body than released, body temperature increases,
- intensive exercise/work in hot humid weather, dehydration aggravates the condition
- symptoms vary from headache, tachycardia to coma
- treatment: careful and slow cooling the body

## Hypothermia

- excessive heat loss, decrease of body temperature below 35 °C
- stay in cold weather/water, wet clothing in wind etc.
- aggravating factors: older people, hypothyreosis, alcohol
- treatment: different rewarming methods

# Enthalpy corresponds to reaction heat

at constant external pressure

$$\Delta H = H_2 - H_1 = \Delta Q_p$$

It expresses the difference in the bond energy of reaction products and reactants.

$\Delta H < 0$  **exothermic reaction**, the enthalpy of the reaction products is lower (the bonds are more stable) than that of the reactants

$\Delta H > 0$  **endothermic reaction**, the enthalpy of the products is higher than that of reactants.



# Enthalpy changes in specific processes

- **Heat of formation**

elements  $\rightarrow$  1 mol of compound + heat

- **Heat of combustion**

1 mol of subst. + excess  $O_2 \rightarrow$  combustion products + heat

- **Heat of neutralization**

acid + hydroxide  $\rightarrow$  salt + water + heat

- **Heat of solution**

substance(s) + water  $\rightarrow$  solution  $\pm$  heat



- **Heat of dilution**

concentrated acid + water  $\rightarrow$  diluted acid + heat

Always pour acid into water, not the reverse way !!!

# Instant Cold/Hot Packs



- Based on **heat of solution**
- Consist of a pouch of water and a dry chemical
- Striking the pack leads to the pouch breaking, dissolving the chemical and enthalpy change: (up to **80 °C** or down to **0 °C**, 20 min)
- First-aid device at sports events etc.
- $\text{CaCl}_2(\text{s}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq})$        $\Delta H_{\text{soln}} < 0$  **hot pack**
- $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{NH}_4^{+}(\text{aq}) + \text{NO}_3^{-}(\text{aq})$        $\Delta H_{\text{soln}} > 0$  **cold pack**

# Entropy

**Entropy** is a thermodynamic property, a **measure of disorder**.

It is defined as the amount of energy (heat) in the system that cannot be transformed to work:  $S = Q_{\text{rev}} / T$ .

A change in entropy is  $\Delta S = S_2 - S_1$ .

in closed and open systems the entropy can either increase or decrease.

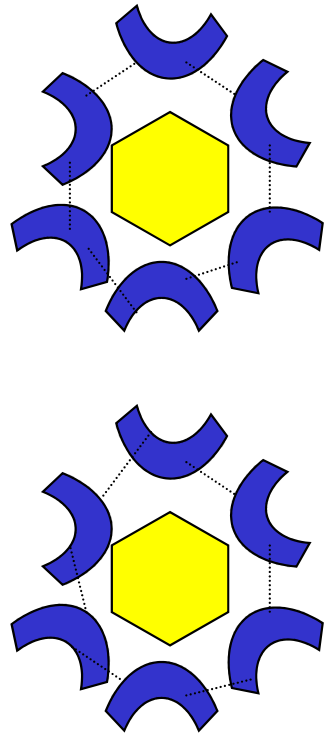
## Entropy increases

- Increasing temperature
- Melting a solid
- Evaporating a liquid
- Dissolving a solid in a liquid
- Mixing two substances
- Increasing the number of particles during a reaction (decomposition)

## Entropy decreases

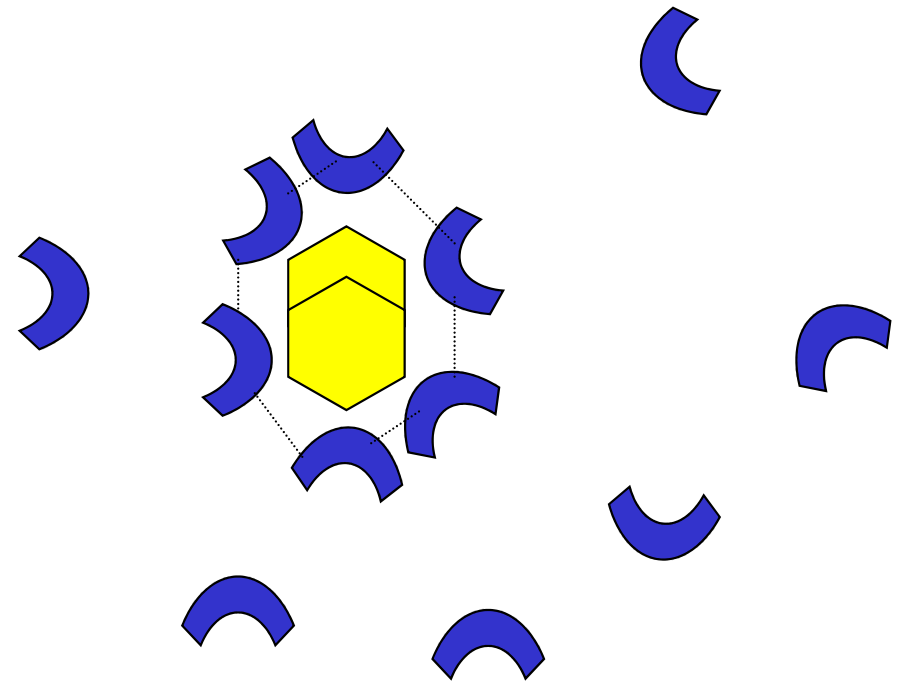
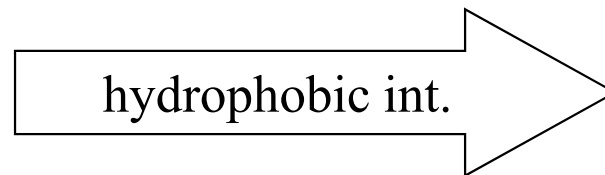
- Decreasing temperature
- Freezing a liquid
- Condensing a gas
- Precipitating a solid from sol.
- Separating two substances
- Decreasing the number of particles during a reaction (synthesis)

# Hydrophobic interactions in aqueous environment



two non-polar molecules  
in aqueous solution

Water molecules are organized  
(cluster) – they have **low entropy**



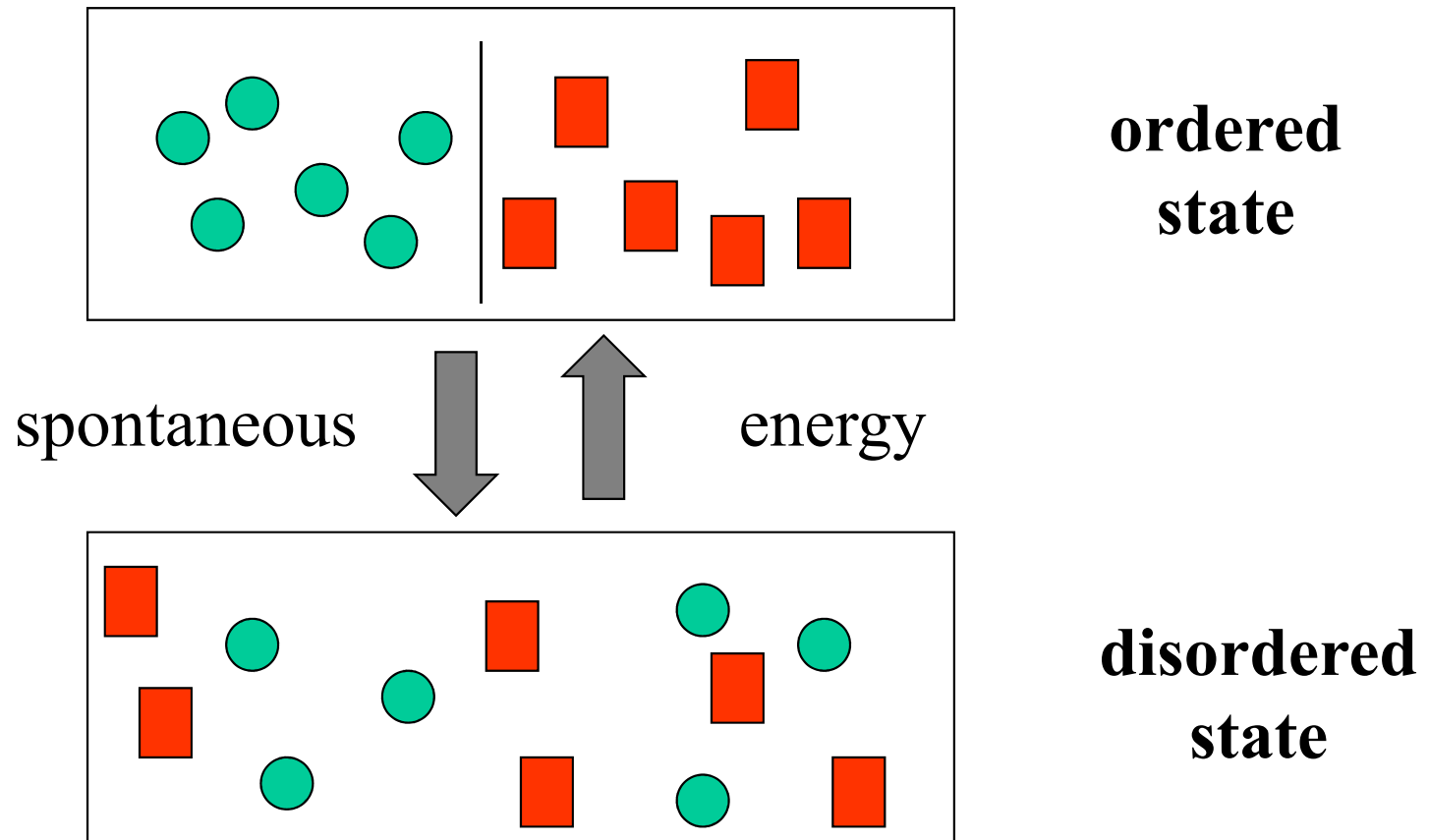
two non-polar molecules  
adhered to each other

Six water molecules are free  
– they have **high entropy**

Hydrophobic interactions increase the total entropy of a system

# Spontaneous processes are irreversible

They can happen without any continuing outside influence.



# Principal question of thermodynamics:

**Will a reaction proceed?**

**YES / NO**

**Neither enthalpy nor entropy can answer**

**In closed and open systems, the driving force of chemical reactions or physical changes is the free energy change  $\Delta G$**

The **Gibbs (free) energy  $G$**  of a system is the energy that is available to do useful work as the result of chemical or physical change at constant temperature and pressure.

The free energy change  $\Delta G (= G_2 - G_1)$  is defined as

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

useful work      enthalpy change  
(reaction heat)      heat lost or absorbed due  
to entropy change



# Overview of thermodynamic functions

## $\Delta H$ the heat of reaction

$\Delta H < 0$  exothermic reaction

$\Delta H > 0$  endothermic reaction

## $\Delta S$ the entropy change

$\Delta S > 0$  the final state is very probable

$\Delta S < 0$  a very low probability of the final state

## $\Delta G$ is the criterion of process feasibility

$\Delta G < 0$  exergonic reaction prone to proceed spontaneously

$\Delta G > 0$  endergonic reaction that cannot proceed spontaneously under given circumstances  
(the reverse reaction is spontaneous)

$\Delta G = 0$  the system is at the equilibrium state

# What is the driving force of spontaneous processes?

1. to get lowest energy
2. to get maximal disorder

$\Delta G$   
comprizes both aspects

## Examples of spontaneous processes:

- waterfall flows from up to down
- heat is transferred from warmer object to cooler one
- sugar crystals spontaneously dissolve in water
- gas expands into evacuated space
- elemental iron slowly gets rusty
- $\text{Na} + \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{H}_2 \uparrow + \text{NaOH}$

$\Delta S$	$\Delta H$	$\Delta G = \Delta H - T\Delta S$
<b>positive</b> (decompositions)	<b>negative</b> (exothermic)	always negative; the reaction is spontaneous, practically irreversible
	<b>positive</b> (endothermic)	as far as $T\Delta S > \Delta H$ (at higher temperatures), the reaction is spontaneous
<b>negative</b> (syntheses)	<b>negative</b> (exothermic)	as far as $T\Delta S < \Delta H$ (at lower temperatures), the reaction becomes more favourable
	<b>positive</b> (endothermic)	positive at all temperatures; the forward reaction cannot be spontaneous (the reverse reaction is always spontaneous)

spontaneous reactions are:

- the entropy increases and heat is released
- reaction is endothermic but accompanied with entropy increase
- reaction is highly exothermic so that it overpoises entropy decrease

# Standard state of substances

Defined by convention, indicated by symbol  $^{\circ}$

Substance	Standard state definition
Solid (s)	Pure solid at given temperature*
Liquid (l)	Pure liquid at given temperature*
Gas (g)	Pure gas, $p = 1 \text{ atm} = 101.3 \text{ kPa}$ , given temperature*
Solution	$c = 1 \text{ mol/l}$ , given $T$ , $p = 101.3 \text{ kPa}$ , $\text{pH} = 0.00$



Quite non-physiological conditions



\*often  $298 \text{ K} = 25 \text{ }^{\circ}\text{C}$

## Standard Gibbs free energy change $\Delta G^\circ$

for a reversible process represents the free energy change

between

the system that exists at the beginning of the process

in its **standard state** (the reaction quotient  $Q = 1$ ),

i.e. **all reactants**, both reactants and products are of unit activity,

**in aqueous solution their concentration  $c = 1 \text{ mol l}^{-1}$**

(if  $\text{H}^+$  is a reactant, then also  $[\text{H}^+] = 1$ ,  $\text{pH} = 0$ ),

at specified temperature (usually  $25^\circ\text{C}$  equal to  $298 \text{ K}$ ),

and atmospheric pressure  $101.3 \text{ kPa}$ ,

and the reaching the state with a minimal  $G$  value,

that is the **equilibrium state** of the system

in which the reactants and products has reached **the concentrations corresponding with the equilibrium constant  $K$** .

In biological systems, the standard state is defined by  $\text{pH} = 7.00$ ;  
then the free energy changes are marked as  $\Delta G^{\circ'}$

## The relationship between free energy and equilibrium

for a reaction  $aA + bB \rightleftharpoons cC + dD$

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

The  $\Delta G$  of a reaction depends on the **particular kind of reaction** (expressed by the  $\Delta G^\circ$  term) and the **initial concentrations** of reactants and products (expressed by the second term equivalent to  $Q$ ).

If the equilibrium concentrations are put in as initial ones, the system is in its equilibrium state,  $\Delta G = 0$ :

$$0 = \Delta G^\circ + RT \ln K \quad \text{and} \quad \Delta G^\circ = -RT \ln K$$

# Equilibrium

The general tendency of any spontaneous process is to reach an equilibrium state.

**The initial nonequilibrium concentrations** of the substances taking part in reaction  $a A + b B \rightleftharpoons c C + d D$  are used to calculate the **reaction quotient  $Q$** :

The expression takes the same form as the equilibrium constant but is used for a initial state, not for a reaction at equilibrium.

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

**The equilibrium state** is described by equilibrium constant  $K$ :

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

The value of  $Q$  indicates what changes will occur in reaching equilibrium:  
When  $Q < K$ , the reaction has a chance to proceed in the forward direction.  
When  $Q > K$ , the reaction has a chance to proceed in the reverse direction.

# Bioenergetics:

## Transformation of energy in the body

Living organisms are open systems that have to receive permanently **nutrients** – compounds of high enthalpy (energy) and low entropy (due to their complex structure).

Nutrients are transformed into **waste metabolites** of low enthalpy and high entropy (simplified structures -  $\text{CO}_2$ ,  $\text{H}_2\text{O}$ , urea).

The part of **free energy** gained by exergonic breakdown of nutrients drives endergonic reactions and processes (synthesis of complex molecules, performance of mechanical or osmotic work, etc.).

The remaining part of acquired energy is released as **heat** into the surroundings.



# Energy data of nutrients

Nutrient	Energy (kJ/g)	Thermogenesis
Lipids	38	4 %
Saccharides	17	6 %
Proteins	17	30 %

Thermogenesis is the heat production 3-5 h after meal.

It is expressed in % of the nutrient ingested. Thermogenesis is the consequence of digestion, absorption, and metabolism of nutrient.

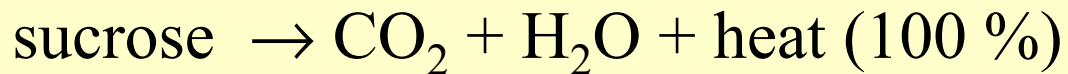
# Consider five sugar lumps

chemical energy content:

$$5 \times 2.8 \text{ (g)} \times 17 \text{ (kJ/g)} = 238 \text{ kJ}$$



## Combusted in calorimeter



## Consumed in a cup of coffee

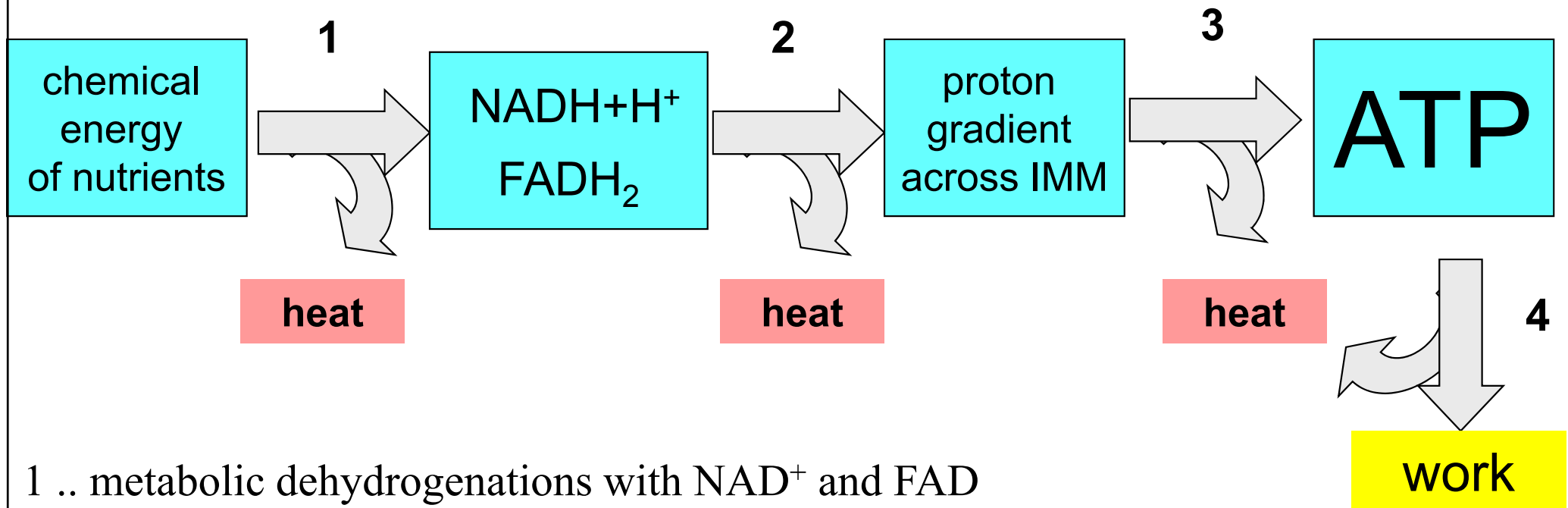


In both processes, oxygen excess is a necessary condition.

## Average fuel values of foods (kJ/100 g)

Water	0	Bread	900
Cucumbers	50	Roll	1100
Coca-Cola	180	Lentils	1200
Lemons	100	Rice	1300
Apples	200	Pastry	1400
Potatoes	280	Pork	1500
Bananas	400	Sugar	1700
Eggs	500	Chocolate	2200
Beef	600	Butter	3000
Ham	700	Oils	3800

# Energy transformations in the human body are accompanied with continuous production of heat



1 .. metabolic dehydrogenations with NAD<sup>+</sup> and FAD

2 .. respiratory chain (oxidation of reduced cofactors + reduction of O<sub>2</sub> to H<sub>2</sub>O)

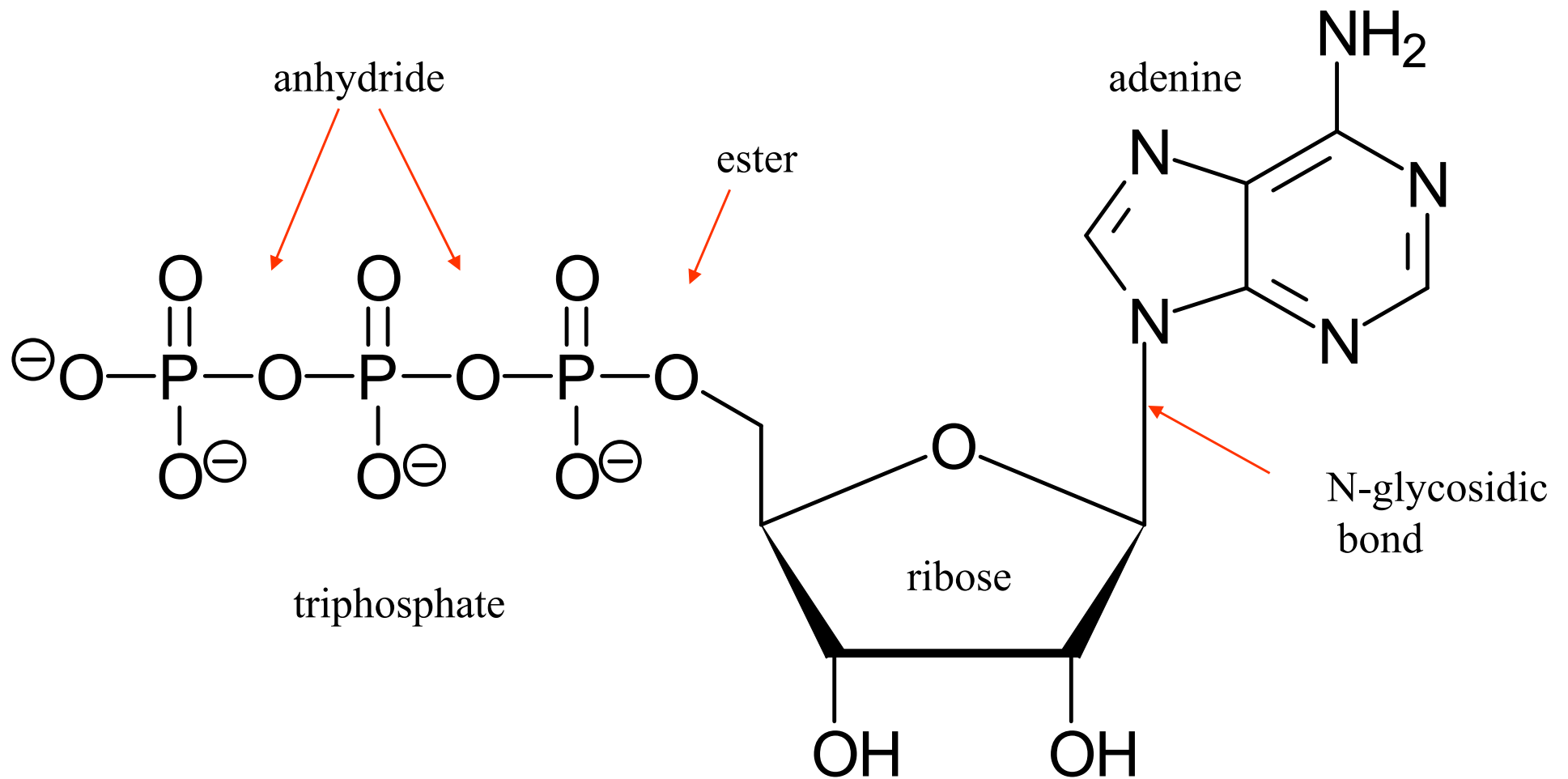
3 .. oxidative phosphorylation, IMM inner mitochondrial membrane

4 .. transformation of chemical energy of ATP into work + some heat

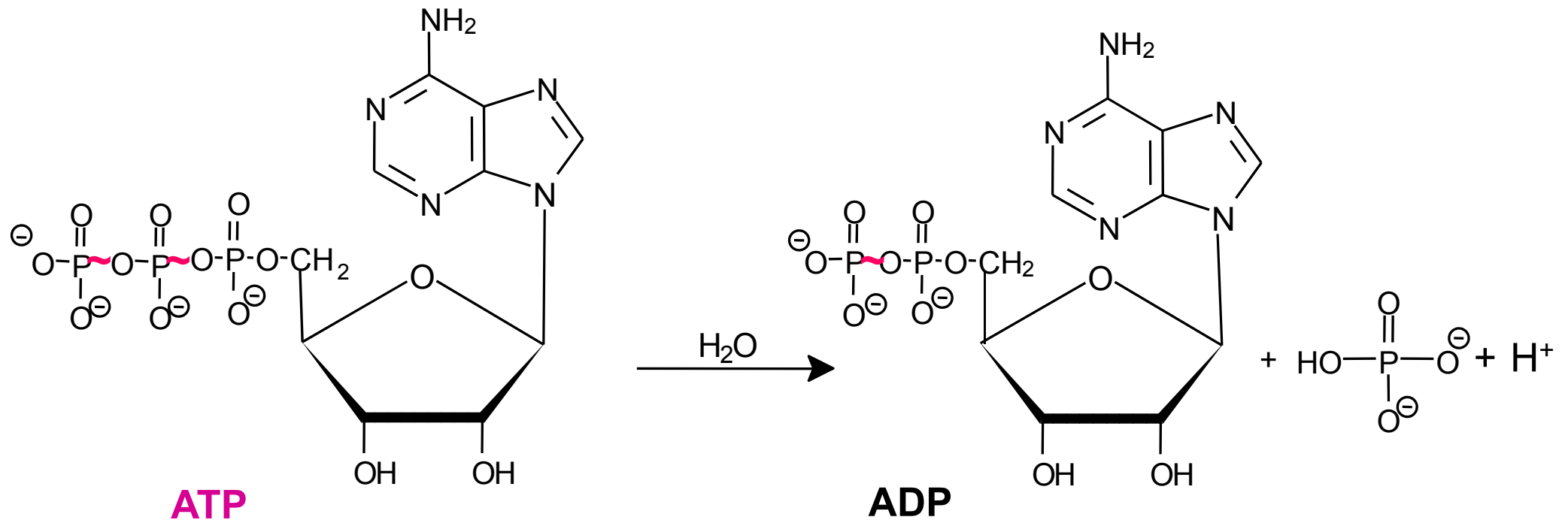
■ .. high energy systems

# ATP (adenosine triphosphate) is high energy compound

at physiological pH (7.4)



# Hydrolysis of ATP decreases the number of negative charges

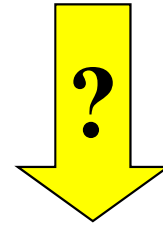


# Compare

**Hydrolysis of ATP is exergonic reaction**



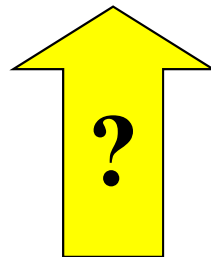
$$\Delta G^{\circ'} = -31 \text{ kJ/mol}$$



**ATP formation is endergonic reaction  
and requires energy investment**

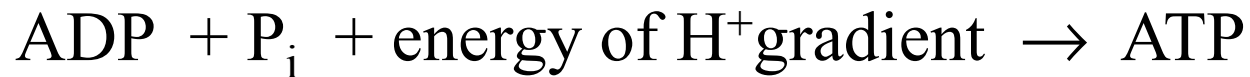


$$\Delta G^{\circ'} = 31 \text{ kJ/mol}$$



# Two ways of ATP formation in body

## 1. Oxidative phosphorylation in the presence of O<sub>2</sub> (~ 95 % ATP)

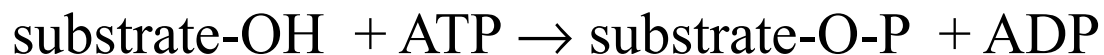


## 2. Substrate-level phosphorylation (~ 5 % ATP)



higher energy content than ATP

### Compare: Phosphorylation

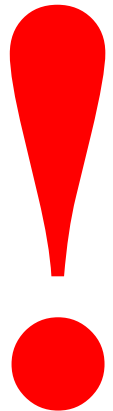


(e.g. phosphorylation of glucose, proteins, etc., catalyze kinases)



# Distinguish

Process	ATP is
Oxidative phosphorylation	produced
Substrate-level phosphorylation	produced
Phosphorylation of a substrate	consumed



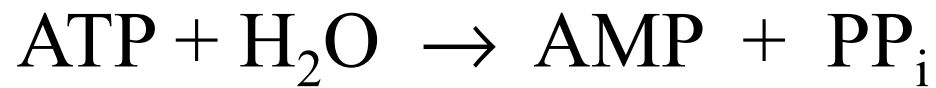
# ATP is used for many processes

- chemical work (endergonic biosyntheses)
- mechanical work
- electroosmotic work (ion transport across membranes)

## Average concentrations of ions (mmol/l)

Body fluid	Na <sup>+</sup>	K <sup>+</sup>	Ca <sup>2+</sup>	Cl <sup>-</sup>
ECF	140	4	2.5	110
ICF	10	160	0.0001	3

## Less usual hydrolysis of ATP means that two ATP are consumed



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summary:



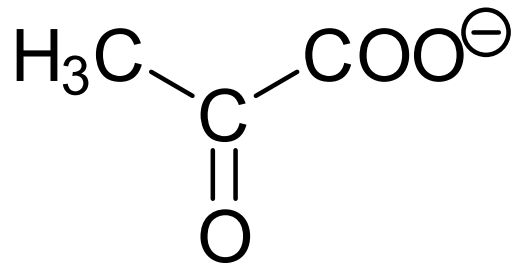
# High-energy phosphates

and  $\Delta G^{\circ'}$  of their hydrolysis (kJ/mol)

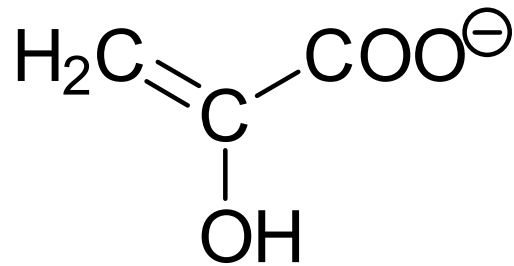
Compound	Phosphate derivative	$\Delta G^{\circ'}$
Phosphoenolpyruvate	enolester	-62
Carbamoylphosphate	mixed anhydride (acylphosphate)	-52
1,3-Bisphosphoglycerate	mixed anhydride (acylphosphate)	-50
Creatine phosphate	phosphoamide	-43
ATP	double anhydride	-31



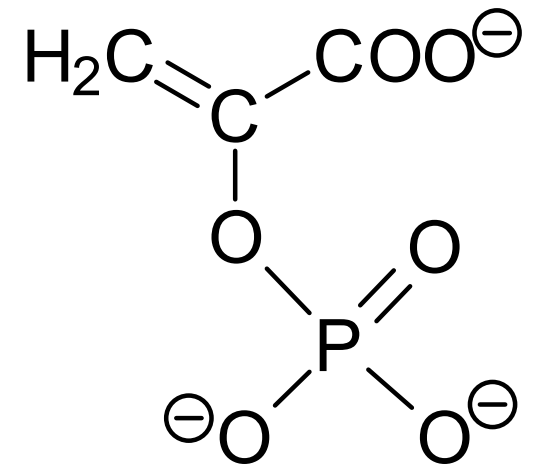
# Phosphoenolpyruvate



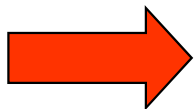
pyruvate



enolpyruvate

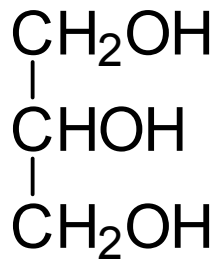


phosphoenolpyruvate

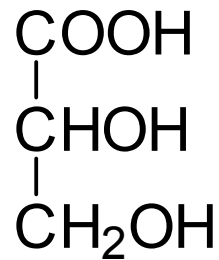


en + ol = hydroxyl attached to a double bond C=C

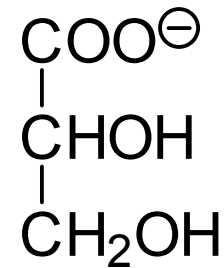
# 1,3-Bisphosphoglycerate



glycerol

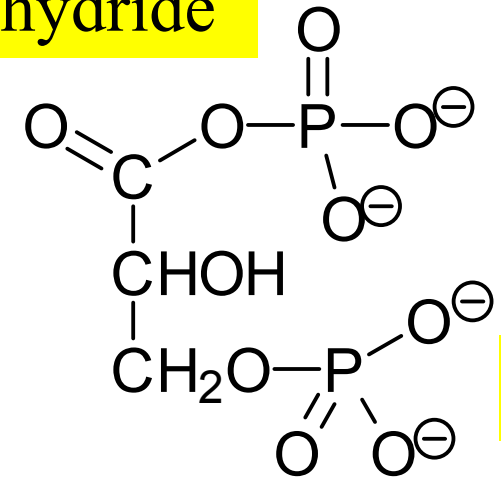


glyceric acid



glycerate

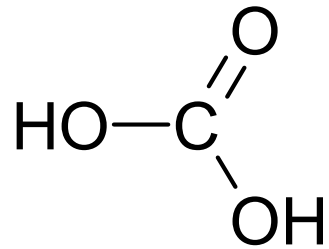
anhydride



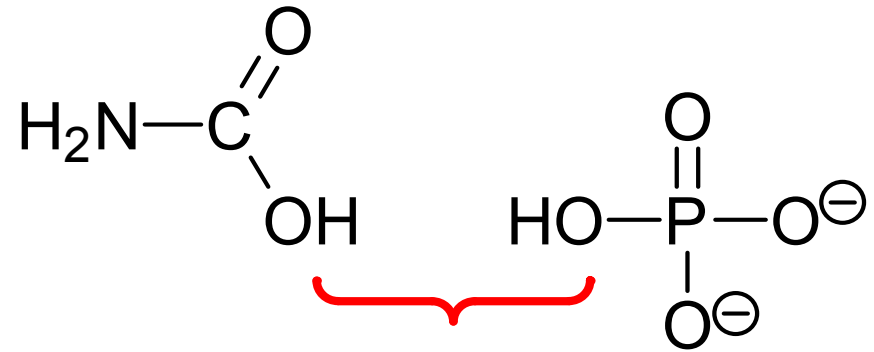
ester

1,3-bisphosphoglycerate

# Carbamoyl phosphate

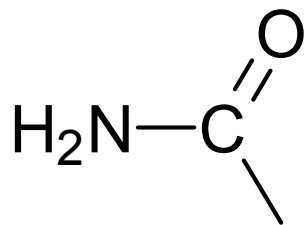


carbonic acid  $\text{H}_2\text{CO}_3$

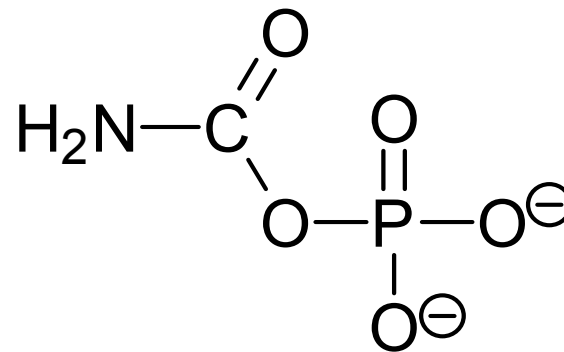


carbamic acid

phosphate

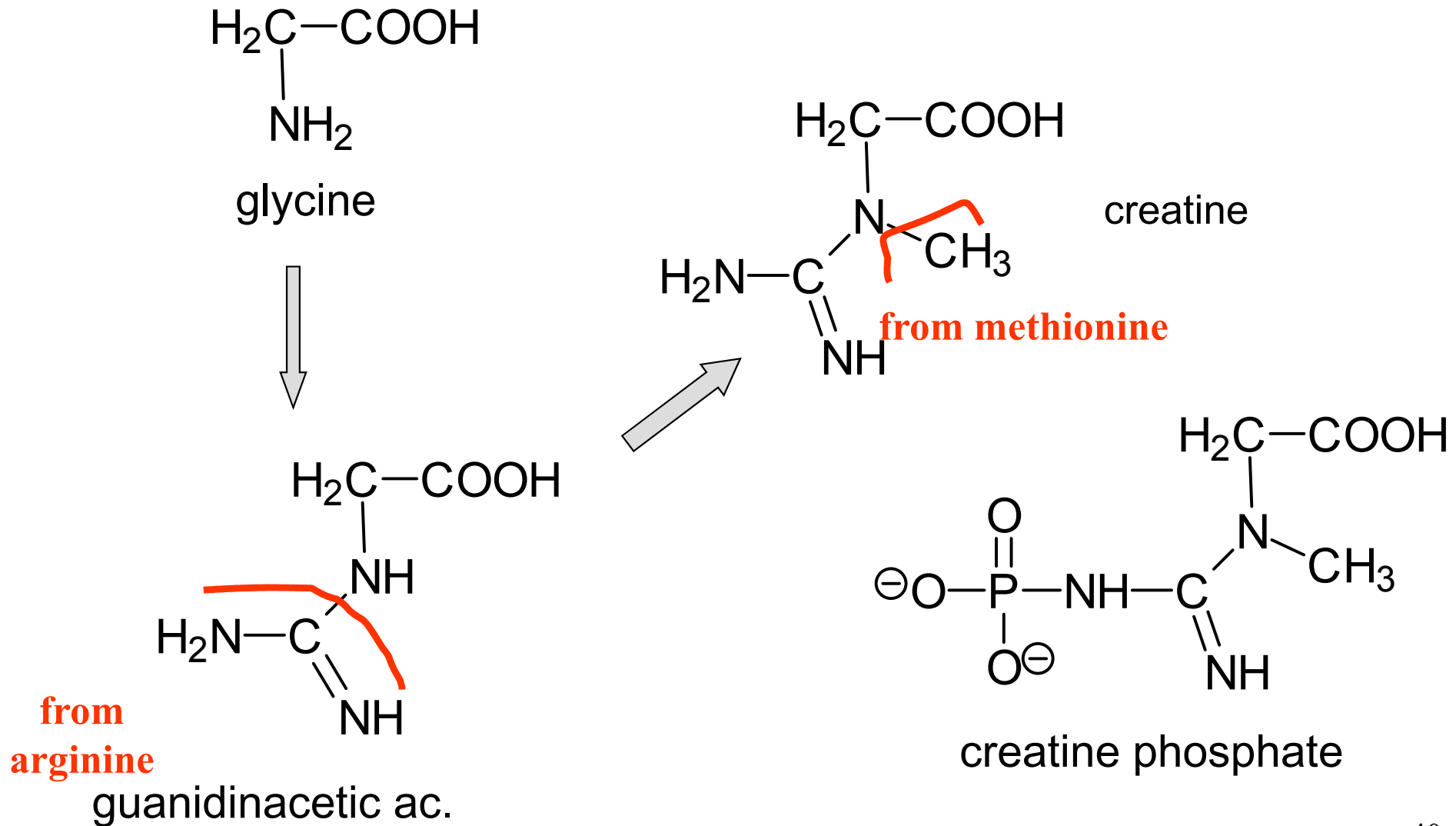


carbamoyl  
(acyl of carbamic ac.)



carbamoyl phosphate

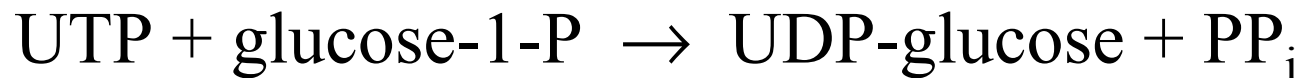
# Creatine phosphate





# Other high energy nucleoside triphosphates have specialized functions

**UTP** – activation of glucose and other sugars



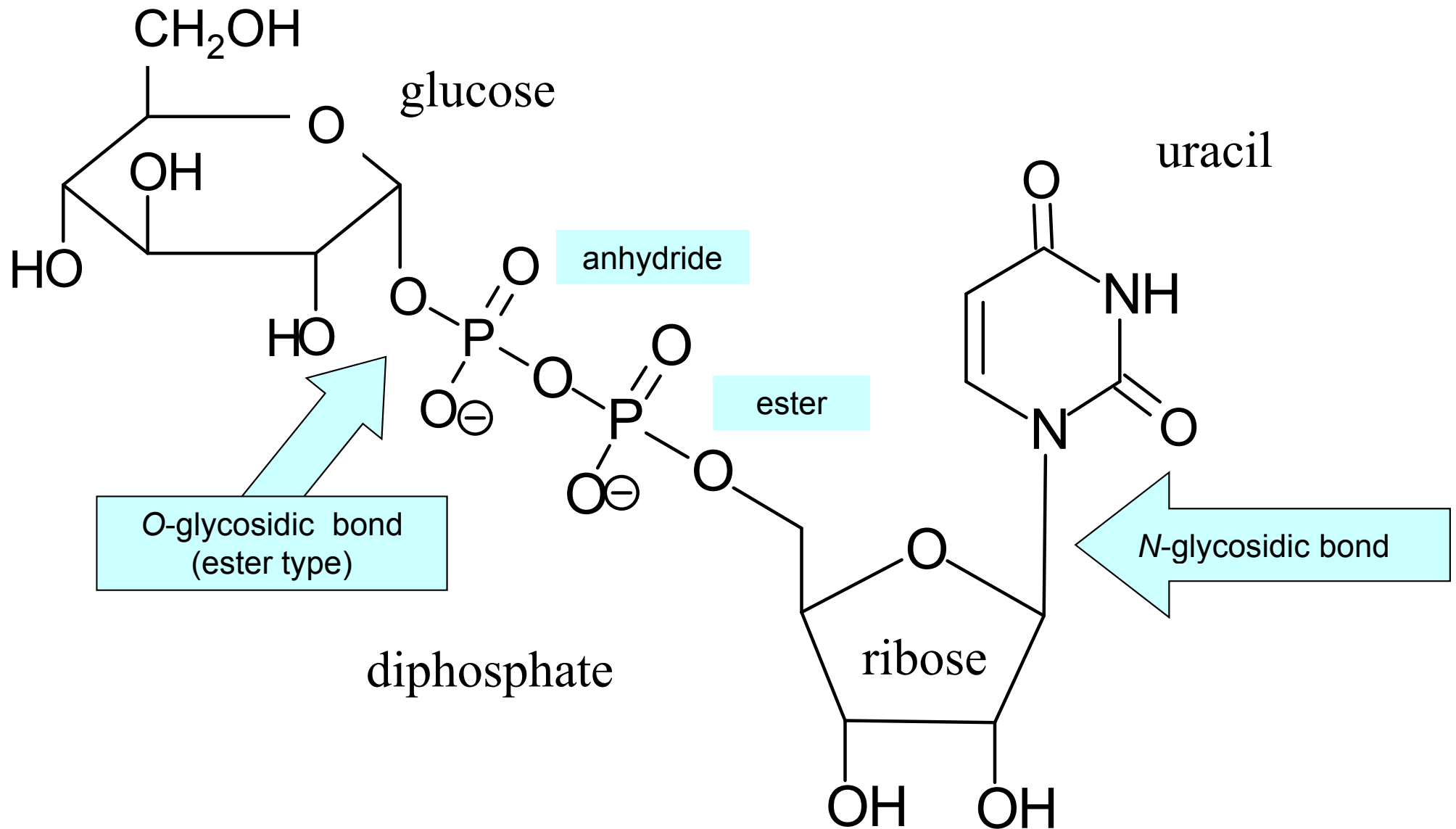
(synthesis of glycogen)

**GTP** – gluconeogenesis, proteosynthesis

**CTP** – activation of choline, ethanolamine

(synthesis of phospholipids)

# UDP-glucose is activated glucose

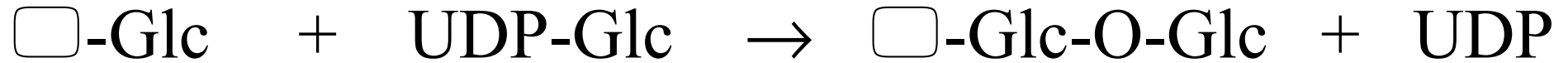


# Distinguish

Compound	Commentary
Glucose	metabolic fuel for most tissues, energy content = 17 kJ/g = 1700 kJ/100 g
Glucose-6-P	metabolically <i>active</i> form of glucose, enters various enzymatic reactions
UDP-glucose	<i>activated</i> form of glucose, energy-rich derivative for endergonic synthesis

# Compare

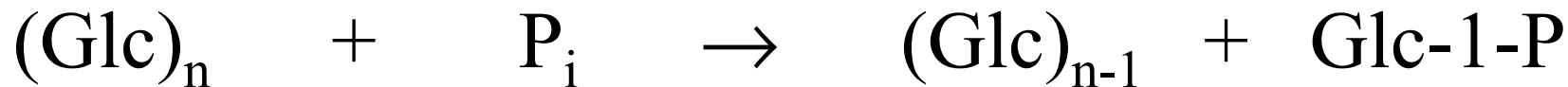
Synthesis of glycogen **requires** energy:



low-energy  
substrate  
(glucose unit)

high-energy  
reagent  
(activated glucose)

Cleavage of glycogen **does not require** energy:



high-energy  
substrate  
(glycogen  
macromolecule)

low-energy  
reagent  
(phosphate)

# Consider and distinguish:

low energy content = stable = non-reactive  
(CO<sub>2</sub>, H<sub>2</sub>O, NH<sub>3</sub>, P<sub>i</sub>)



high energy content = unstable = reactive  
thioesters (CoA-S-acyl),  
mixed anhydrides (acylphosphates),  
nucleoside triphosphates (ATP, GTP, UTP) etc.

# Low-energy compounds and their activated forms

Low-energy compound	High-energy activated form	Commentary
CO <sub>2</sub>	carboxybiotin	to make carboxybiotin, one ATP mol. is needed; the cofactor of carboxylation reactions, e.g.: pyruvate + carboxybiotin → oxalocetate + biotin
NH <sub>3</sub>	carbamoyl phosphate	two ATP molecules are needed for NH <sub>3</sub> activation; substrate for urea and pyrimidine synthesis
P <sub>i</sub>	ATP	there are two ways of ATP production in the body
R-COOH	R-CO-S-CoA (acyl-CoA)	unusual ATP decomposition occurs in activation: R-COOH + ATP + CoA-SH → R-CO-S-CoA + AMP + PP <sub>i</sub>

# Kinetics: Basic terms

- reaction:  $S \rightarrow P$  (S = substrate, P = product)
- Definition of reaction velocity (rate):

$$v = -\frac{\Delta[S]}{\Delta t} = \frac{\Delta[P]}{\Delta t} > 0 \quad \left[ \frac{\text{mol}}{\text{l.s}} \right]$$

this definition is for average rate,  
instantaneous rate:  $d[S]/dt$

# Initial velocity $v_0$

- The highest value of velocity
- It is not influenced by the decrease of substrate nor the reverse change of product
- **Determined from kinetic curves at the time  $t = 0$**



## Factors influencing reaction rate

- Concentration of substrate [S]
- temperature
- The presence of catalyst or inhibitor

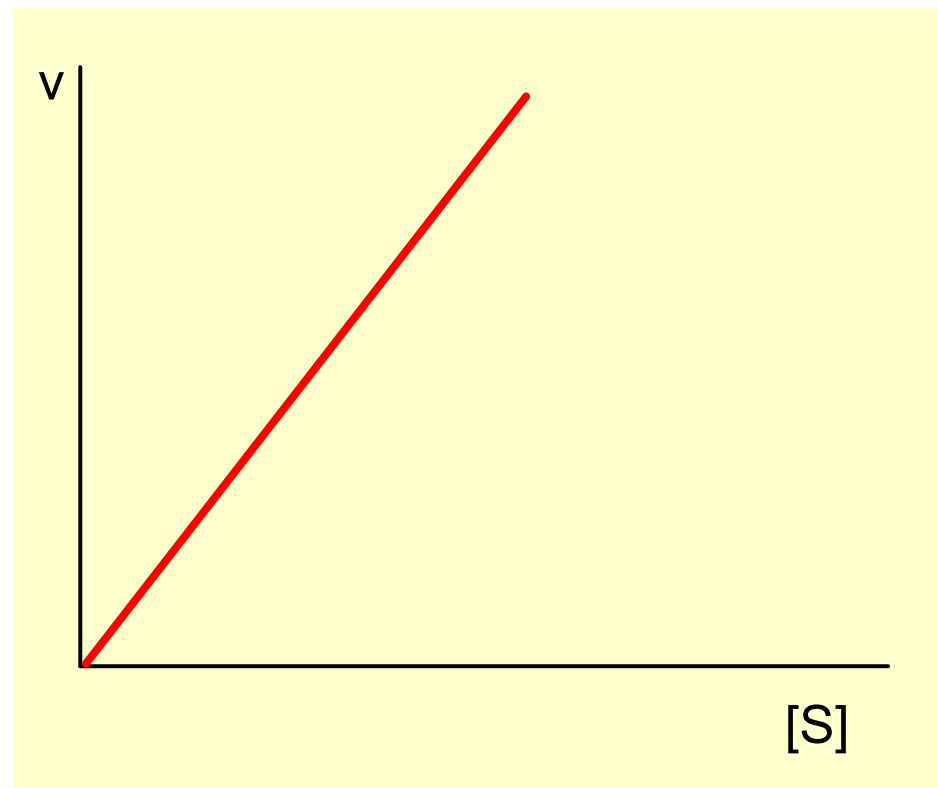
## In enzyme reactions

- Concentration of enzyme [E]
- pH

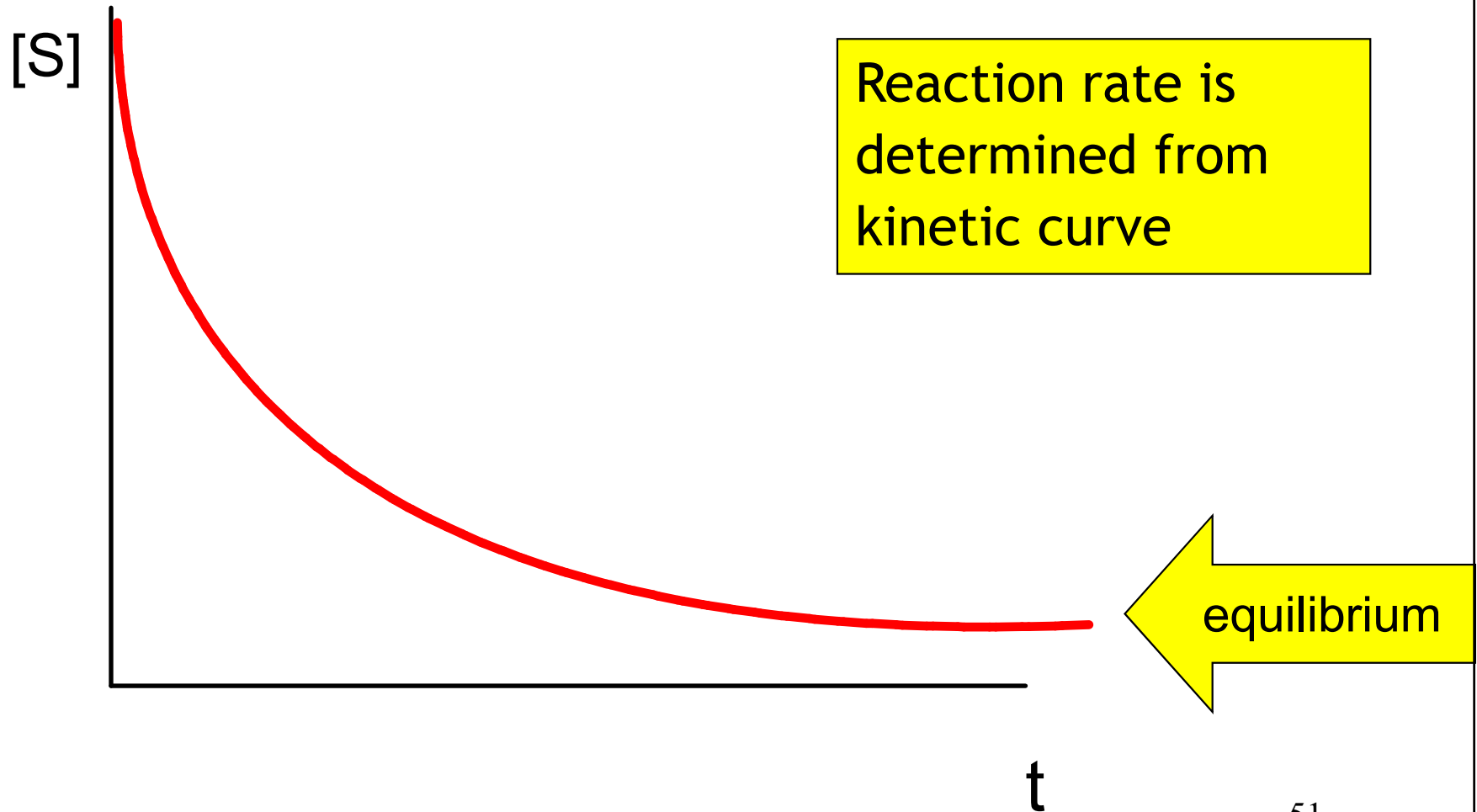
## Kinetic equation for reaction $S \rightarrow P$

$$v = k [S] = k [S]^1 \Rightarrow \text{reaction of 1. order}$$

$k$  = rate constant



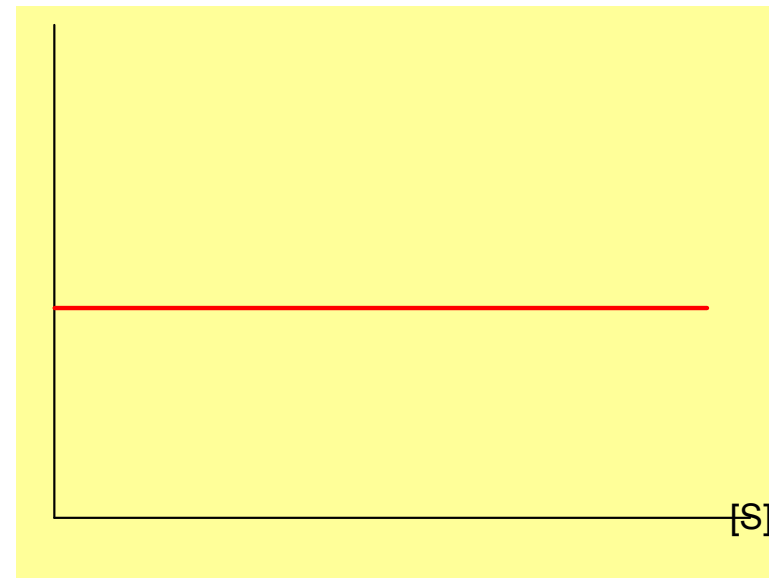
# Concentration of substrate drops down during reaction $\Rightarrow$ kinetic curve for substrate



# Reaction of 0. order

- Reaction rate does not depend on the substrate concentration
- $v = k [S]^0 = k \cdot 1 = k = \mathbf{constant}$
- At great excess of substrate S

in enzyme reactions only in laboratory conditions



# Redox reactions are based on electron transfer

Oxidation	Reduction
Loss of electrons	Gain of electrons
Dehydrogenation (loss of 2H) <sup>a,b</sup>	Hydrogenation (gain of 2H) <sup>c</sup>
Oxygenation (gain of O) <sup>d,e</sup>	Deoxygenation (loss of O)



<sup>a</sup> H = electron + proton = e<sup>-</sup> + H<sup>+</sup>

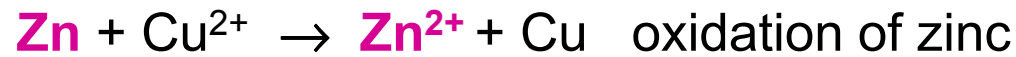
<sup>b</sup> also desaturation    <sup>c</sup> also saturation

<sup>d</sup> electronegative oxygen atom shifts bond electrons to itself =  
relative loss of electrons

<sup>e</sup> specific case is hydroxylation

# Different types of redox reactions – examples

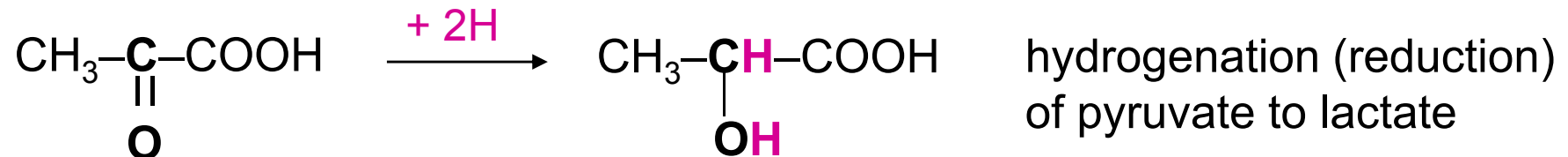
## – Loss and gain of electrons



## – Oxygenation and deoxygenation

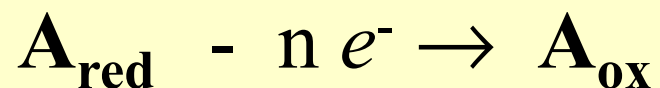


## – Dehydrogenation and hydrogenation

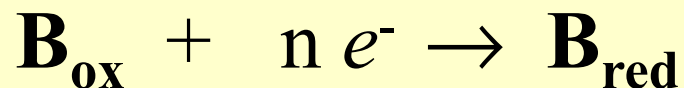


# Redox pairs $A_{\text{ox}} / A_{\text{red}}$ $B_{\text{ox}} / B_{\text{red}}$

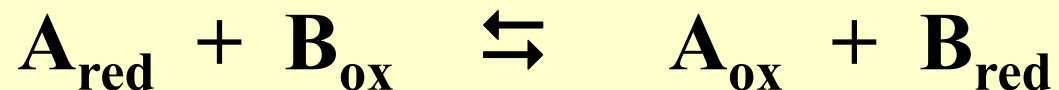
## Oxidation



## Reduction



Oxidation and reduction occur always simultaneously.



Redox pair = half-cell = partial reaction

Redox reaction = combination of two redox pairs

# Do not confuse redox pairs and conjugate pairs

A **redox pair** are two species, which differ each from other in the oxidation number of one or more atoms of the same element (mostly also in the number of electrons). One component of a redox pair is more oxidized and can give the second one by reduction (in a "half-reaction" of a particular redox reaction).

A **conjugate pair** is a couple that consists of an acid and a base (that differ just in one proton  $H^+$  ).

Conjugate pair:  $R-COOH / R-COO^-$   
carboxylic acid / carboxylate ion

Redox pair:  $R-COOH / R-CH=O$   
carboxylic acid / aldehyde



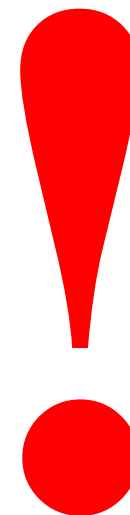
# Consider the difference

Substance donates  $H^+$  = acid

Substance donates  $H$  = reducing agent

Substance accepts  $H^+$  = base

Substance accepts  $H$  = oxidizing agent



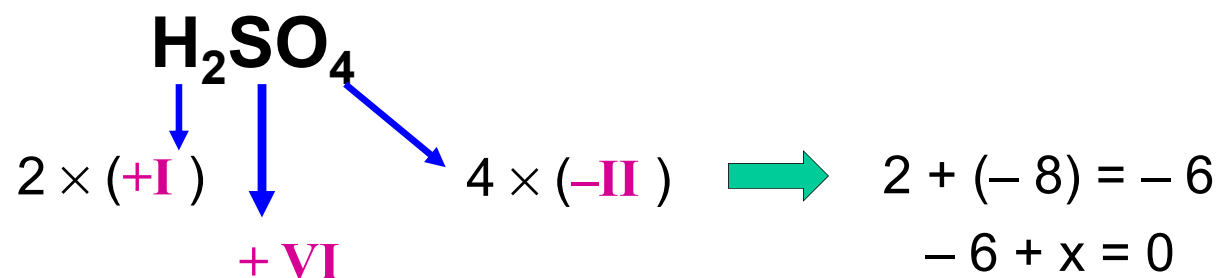
# Oxidation number

Is the charge an element has in a simple ion it forms or the it would hypothetically have, if the shared electron pairs in covalent bonds are assigned to the more electronegative element sharing the pair of electrons.

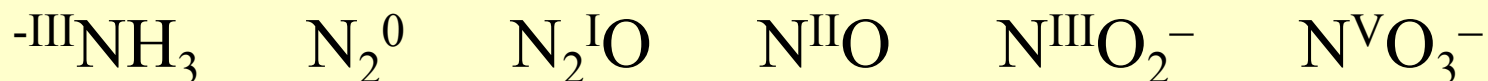
The algebraic sum of the oxidation numbers of elements in a molecular compound equals zero and, in a polyatomic ion, it must equal the charge on the ion.

## Examples:

Oxidation number of sulfur (x) in sulfuric acid:



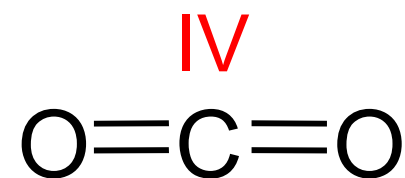
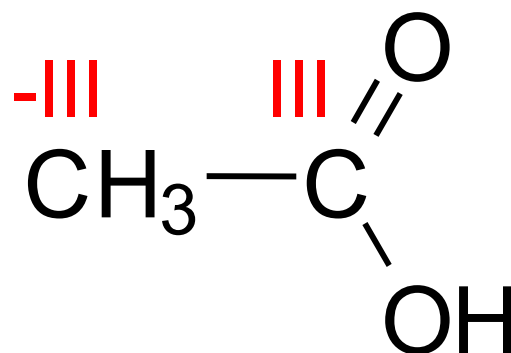
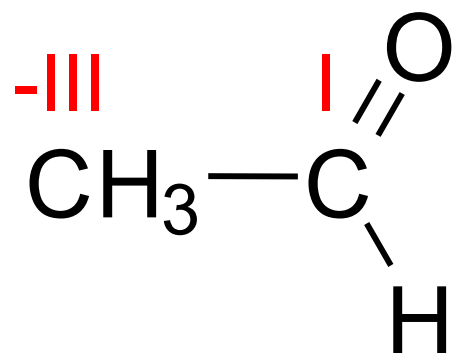
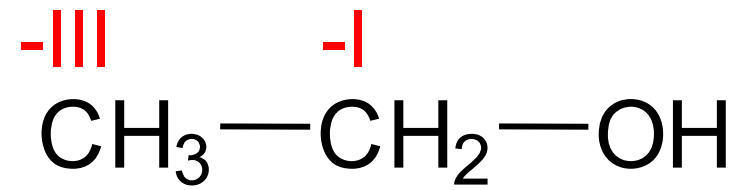
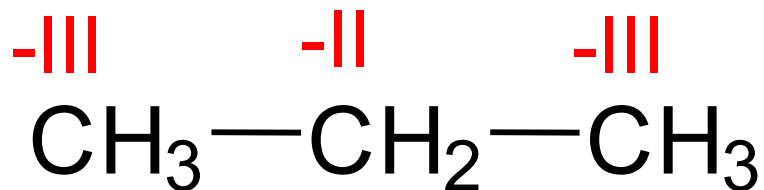
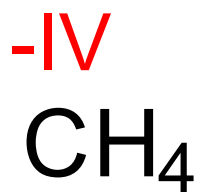
Different oxidation numbers of nitrogen:



# Rules to assign oxidation numbers

- The oxidation number of any **free element** is **zero**, even when the atoms are combined with themselves (e.g.  $O_2$ ,  $P_4$ ,  $S_8$ ).
- No regard is paid to covalent bonds between atoms of the same species.
- An element may have more than one oxidation number, if it forms a variety of compounds.
- The oxidation number of **hydrogen** in a compound or an ion is + I except in ionic hydrides (– I).
- The oxidation number of **oxygen** in a compound or in an ion is –II except in peroxides (it takes on a – I).
- **Metals** generally have only **positive** oxidation numbers in compounds.
- The oxidation number of alkali metals equals always + I, of alkaline earth metals always + II.
- **Nonmetals** have **negative** oxidation numbers when combined with metals, positive oxidation numbers when combined with more electronegative nonmetals.

# Oxidation numbers of carbon in organic compounds



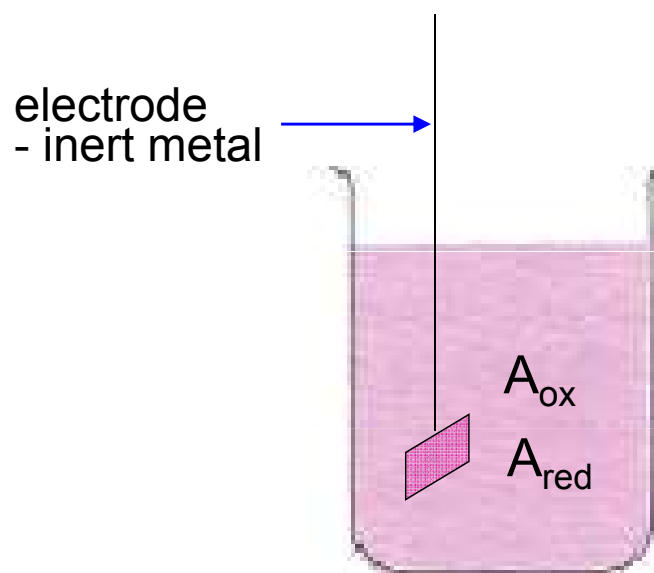
**The strength of oxidants and reductants** (their tendency to gain or lose electrons) is expressed for particular redox pairs by **standard electrode potentials  $E^\circ$** .

Standard electrode potential  $E^\circ$  is the potential of redox pair (both oxidized and reduced form at  $c = 1 \text{ mol/l}$ ) established relatively to the potential of 0.00 V of the standard hydrogen electrode ( $\text{H}^+/\text{H}_2$  pair under standard state conditions).

## Standard state of a redox pair:

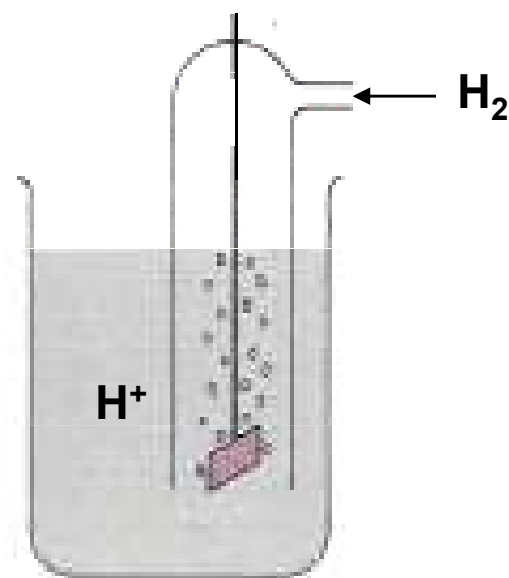
- both oxidized and reduced form of a redox pair at  $c = 1 \text{ mol/l}$
- specified **temperature**, usually **25 °C**
- atmospheric pressure 101.3 kPa is important only when there is a gaseous component of the redox pair

**Redox pair to be measured**  
in the standard state



$$[A_{\text{ox}}] = [A_{\text{red}}] = 1 \text{ mol/l}$$

**Standard hydrogen electrode**  
– reference electrode

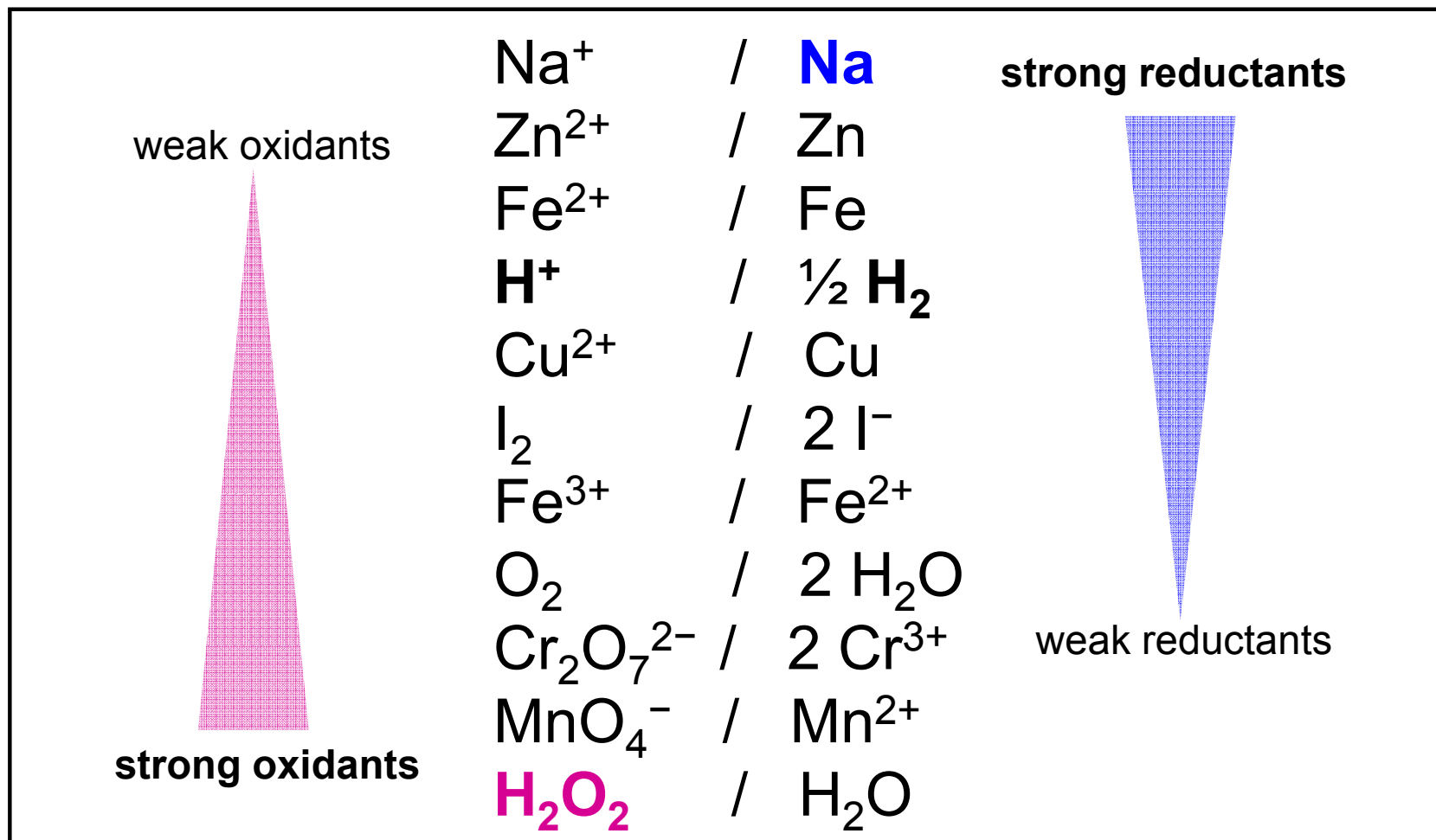


$$[H^+] = 1 \text{ mol / l}$$
$$pH_2 = 101.3 \text{ kPa}$$

$$E^0(H^+/H) = 0.000 \text{ V (25 °C)}$$

## Examples of standard electrode potentials (25 °C)

Half-reaction (Redox pair)	$E^\circ$ (V)
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Zn}^{2+} + 2 \text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{Fe}^{2+} + 2 \text{e}^- \rightarrow \text{Fe}$	-0.44
<b><math>2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2</math></b>	<b>0.00</b>
$\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu}$	0.34
$\text{I}_2 + 2 \text{e}^- \rightarrow 2 \text{I}^-$	0.54
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.76
$\text{O}_2 + 4 \text{H}^+ + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O}$	1.23
$\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 6 \text{e}^- \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$	1.33
$\text{MnO}_4^- + 8 \text{H}^+ + 5 \text{e}^- \rightarrow \text{Mn}^{2+} + 4 \text{H}_2\text{O}$	1.51
$\text{H}_2\text{O}_2 + 2 \text{H}^+ + 2 \text{e}^- \rightarrow 2 \text{H}_2\text{O}$	1.77



The guiding principle:

Under standard conditions, any oxidant will react with any reductant with a lower, more negative  $E^\circ$



If the difference  $\Delta E^\circ$  between both redox pairs is greater than 0.400 V, the **reaction is irreversible** (i.e. proceeds to completion) even under various non-standard concentrations of the reactants.

If the difference between both  $E^\circ$  is **less than 0.400 V**, then the reaction will reach equilibrium, the position of which depends on the initial concentrations of components of both redox pairs; the **direction of such a reaction may be reversed**.

## Electrode potentials $E$ under non-standard conditions

Nernst equation for a redox pair  $a \mathbf{A}_{\text{ox}} + n e^- \rightarrow b \mathbf{A}_{\text{red}}$

$E, E^\circ$  el. potentials in volts

$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$

$F = 96\,500 \text{ C mol}^{-1}$

$n$  = number of moles of electrons transferred

$[\mathbf{A}_{\text{ox}}]$  and  $[\mathbf{A}_{\text{red}}]$  relevant concentrations of reactants

$$E = E^\circ + \frac{RT}{nF} \ln \frac{[\mathbf{A}_{\text{ox}}]^a}{[\mathbf{A}_{\text{red}}]^b}$$

$$E = E^\circ + \frac{0.059}{n} \log \frac{[\mathbf{A}_{\text{ox}}]^a}{[\mathbf{A}_{\text{red}}]^b} \quad (\text{in volts; } t = 25 \text{ }^\circ\text{C})$$

# Redox reactions in biological systems



Most biological redox reactions are catalyzed by enzymes.

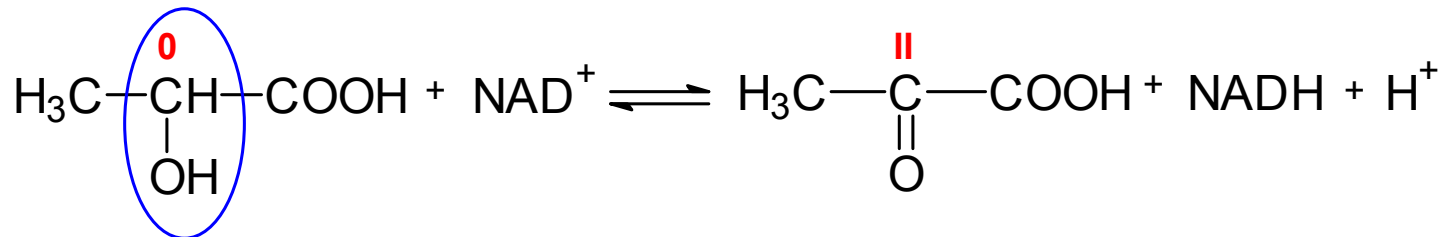
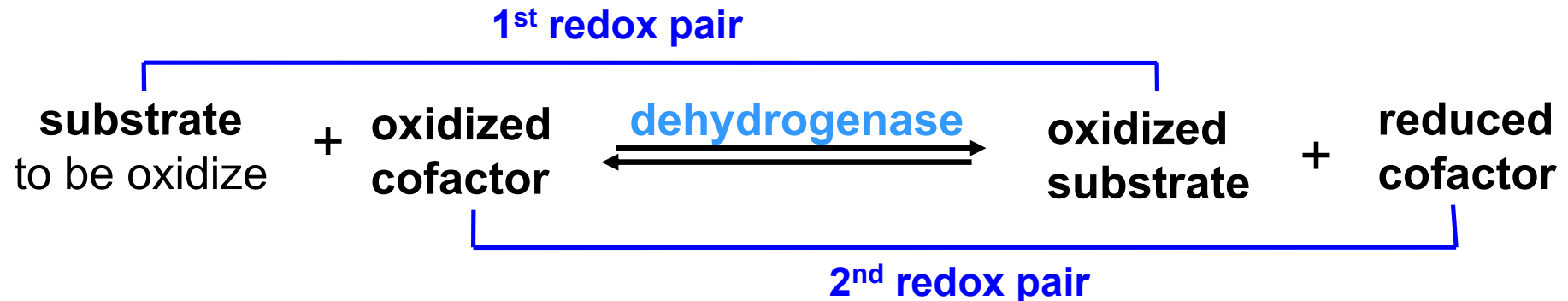
**Oxidative metabolism of nutrients** rich in hydrogen releases energy required to carry out any work of the body.

Some synthetic pathways, e.g. synthesis of fatty acids or cholesterol also include several redox reactions, but those are predominantly reductions (**reductive syntheses**).

Redox reactions serve to also other specialized purposes, e.g. hydroxylations of numerous compounds foreign to the cells (xenobiotics) and dehydrogenation of alcohols.

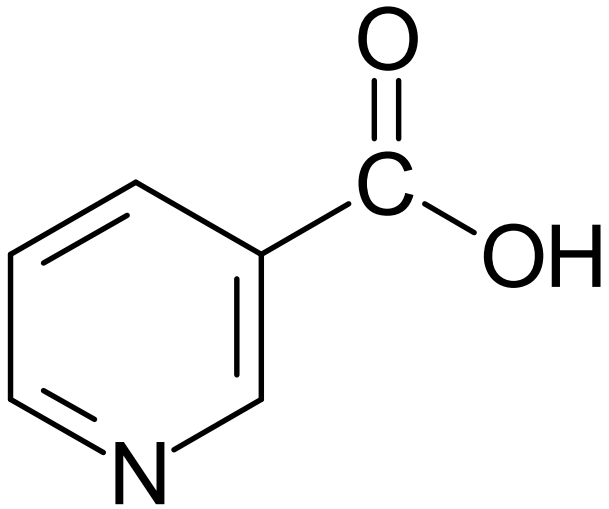
# The oxidation of nutrients proceeds through several dehydrogenation steps

Dehydrogenations are catalyzed by the enzymes dehydrogenases. The two atoms of hydrogen are taken off from substrates and are accepted by the oxidized forms of cofactors  $\text{NAD}^+$  or  $\text{FAD}$ .

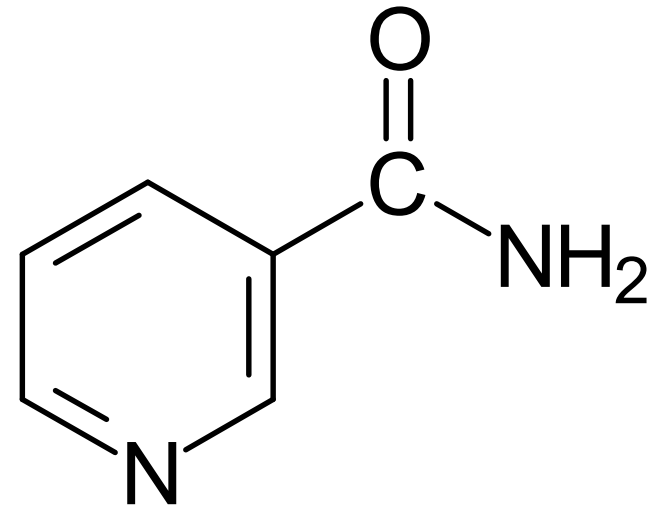


Reduced cofactors  $\text{NADH} + \text{H}^+$  or  $\text{FADH}_2$  are reoxidized in the respiratory chain within the inner mitochondrial membrane.

# Vitamin nicotinamide (niacin) is part of NAD<sup>+</sup>



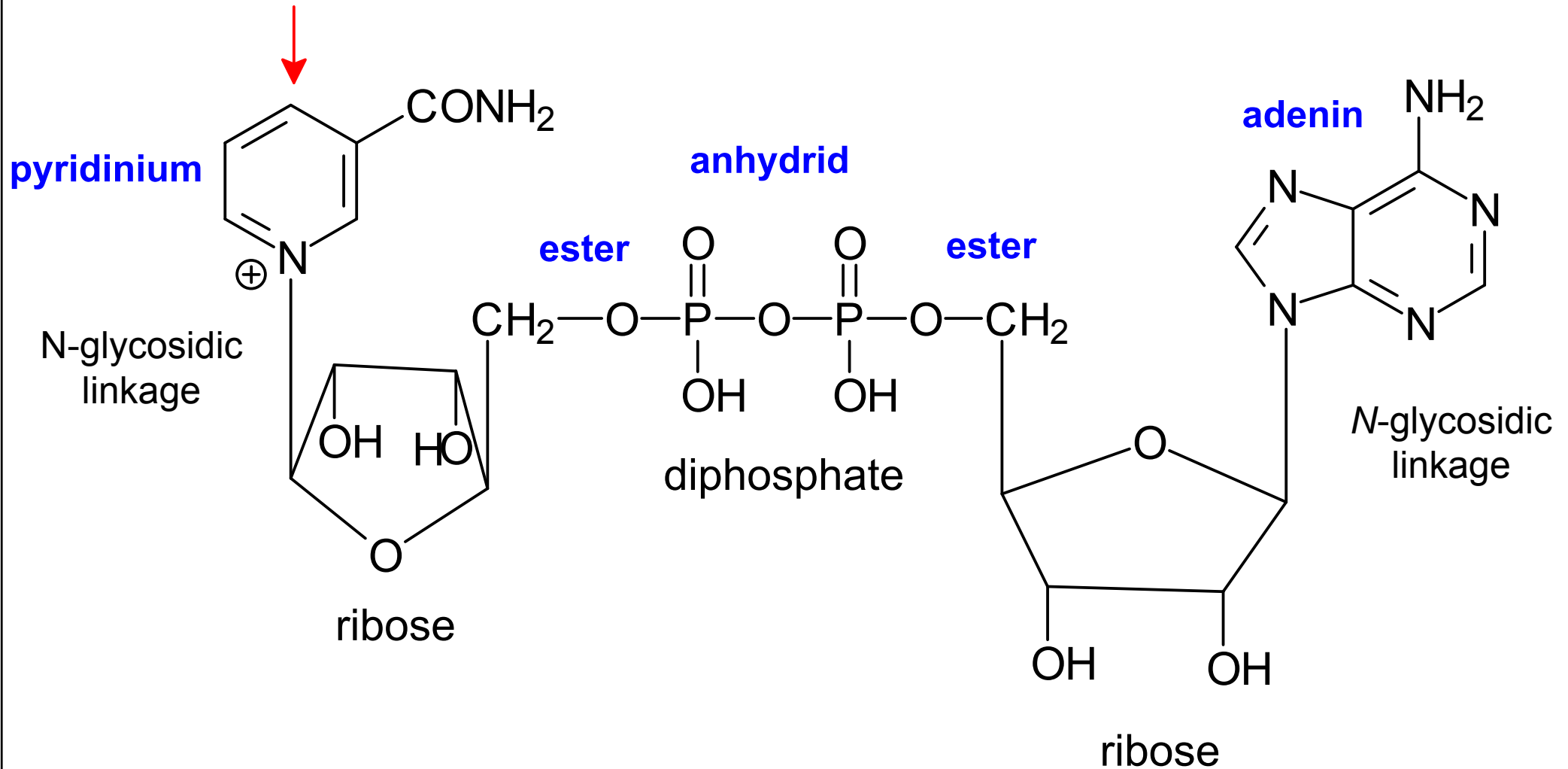
nicotinic acid  
(pyridine-3-carboxylic)



nicotinamide

# NAD<sup>+</sup> (nicotinamide adenine dinucleotide)

addition of hydride H<sup>-</sup> anion

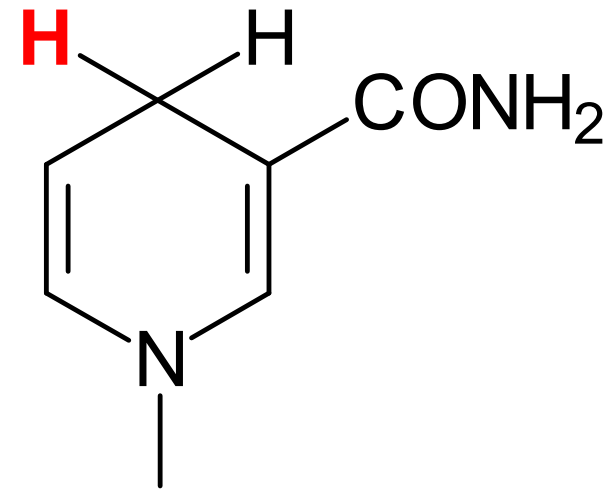
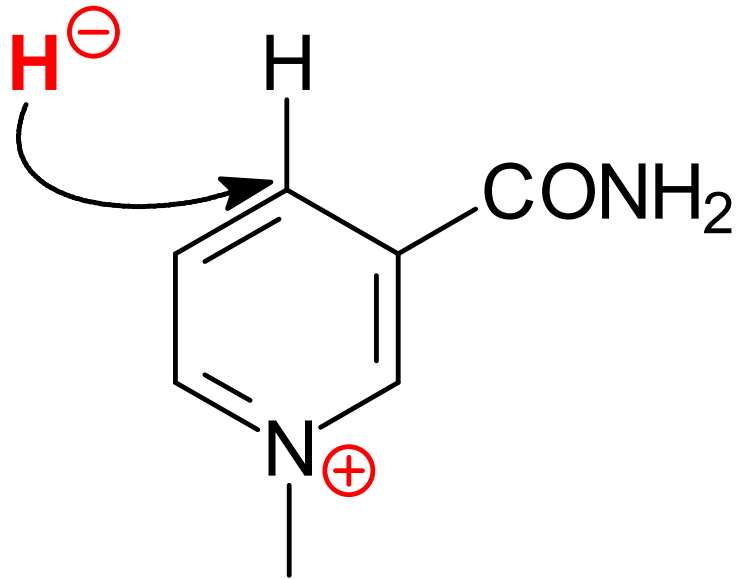


# NAD<sup>+</sup> is cofactor of dehydrogenases

- NAD<sup>+</sup> is oxidant – takes off **2H** from substrate
- one H adds as **hydride ion (H<sup>-</sup>)** into *para*-position of pyridinium cation of NAD<sup>+</sup>
- the second H is released as **proton (H<sup>+</sup>)** and binds to enzyme molecule



# Redox pair of cofactor



oxidized form NAD<sup>+</sup>

aromatic ring

positive charge on nitrogen

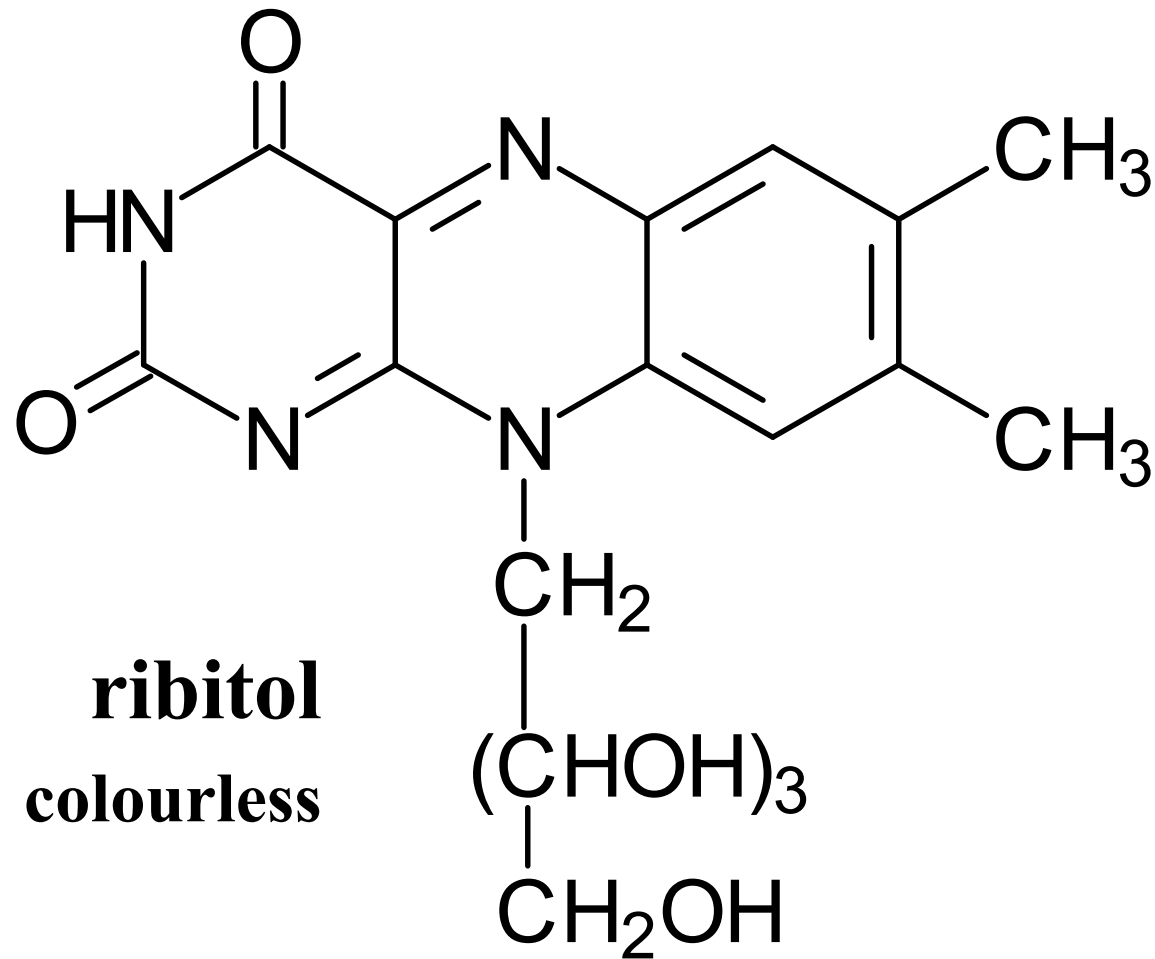
reduced form NADH

aromaticity disturbed

electroneutral species

**high-energy compound**

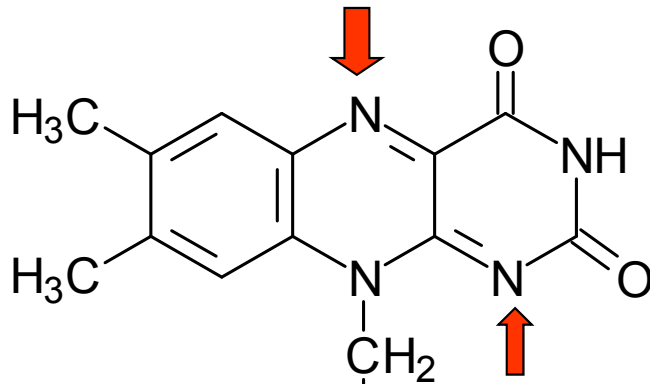
**Riboflavin (vitamin B<sub>2</sub>) comprises colourless ribitol and yellow dimethylisoalloxazine**





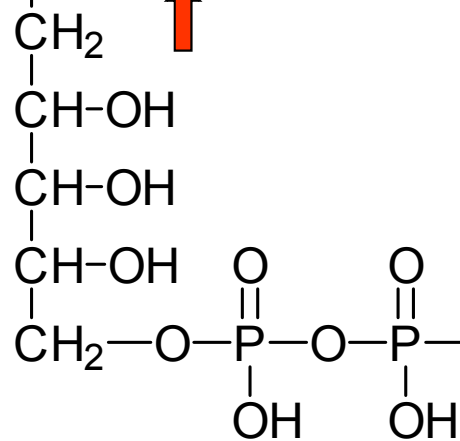
# FAD (flavin adenine dinucleotide)

dimethylisoalloxazin

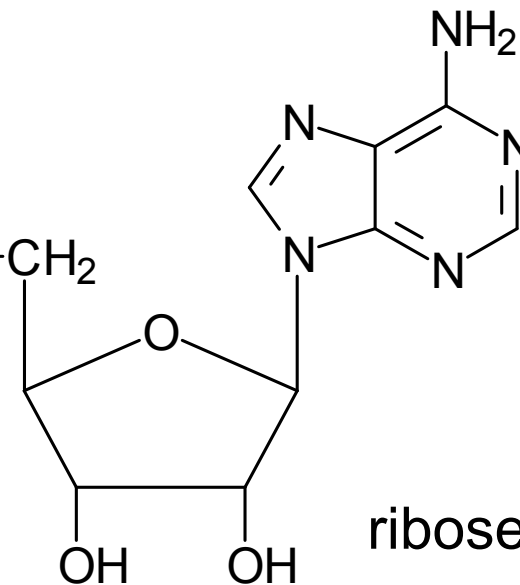


accepting two atoms

ribitol



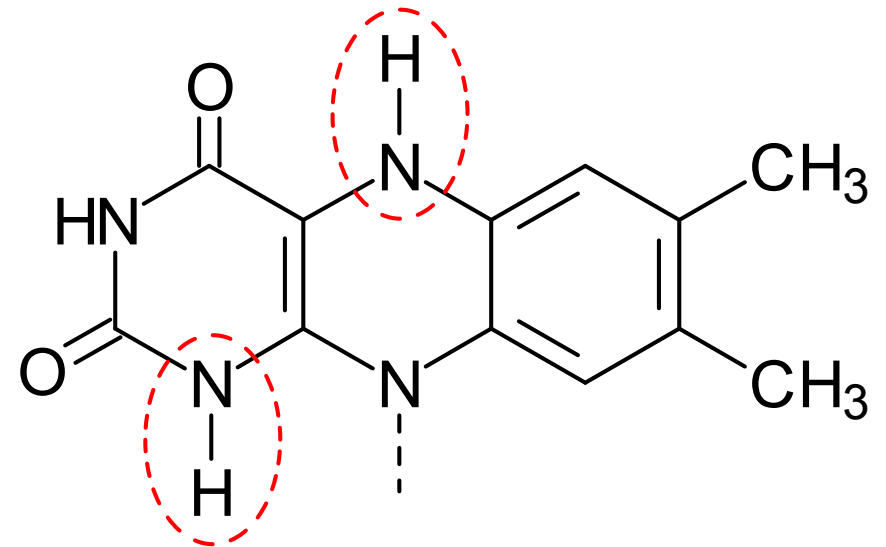
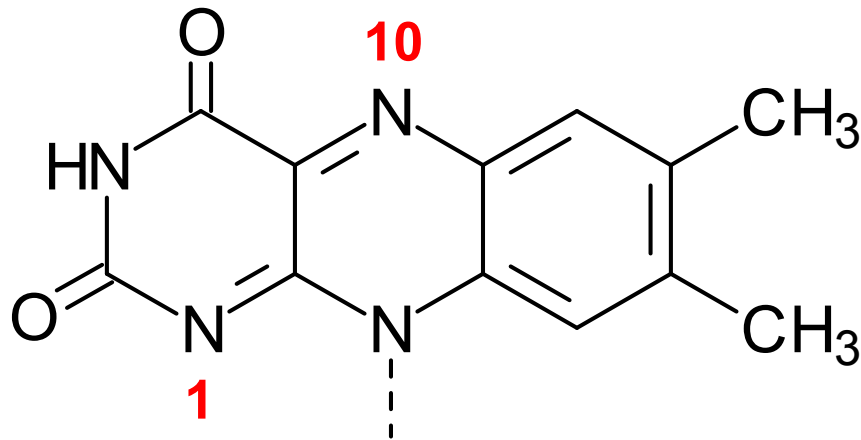
diphosphate



adenine

ribose

# Redox pair of cofactor



oxidized form FAD

aromatic system

electroneutral species

reduced form FADH<sub>2</sub>

aromaticity disturbed

electroneutral species

**high-energy compound**

# Electrode potentials in biological systems

are related to pH value 7.00 and temperature 30 °C;  
the symbols are  $E^{\circ'}$  and  $E'$  instead of  $E^{\circ}$  and  $E$ , respectively

Redox pairs in the respiratory chain	$E^{\circ'}$ (V)
$\text{NAD}^+ + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{NADH} + \text{H}^+$	- 0.320
$\text{FAD} + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{FADH}_2$	a)
$\text{FMN} + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{FMNH}_2$	a)
$2 \text{ cytochrome } b (\text{Fe}^{3+}) + 2 \text{e}^- \rightarrow 2 \text{ cytochrome } b (\text{Fe}^{2+})$	0.030
$\text{ubiquinone} + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{ubiquinol}$	0.100
$2 \text{ cytochrome } c (\text{Fe}^{3+}) + 2 \text{e}^- \rightarrow 2 \text{ cytochrome } c (\text{Fe}^{2+})$	0.235
$2 \text{ cytochrome } a_3 (\text{Fe}^{3+}) + 2 \text{e}^- \rightarrow 2 \text{ cytochrome } a_3 (\text{Fe}^{2+})$	0.385
$\frac{1}{2} \text{O}_2 + 2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2\text{O}$	0.816

a) Flavoproteins exhibit variable values of  $E^{\circ'}$  (0.003 – 0.091 V) which depend on the protein part of the enzyme.