LECTURE 6

Bioenergetics

$A \xrightarrow{?} B$

First Law:

$\Delta H = \pm$ No ⟶ Non-favourable

$\Delta H = -$ Yes ⟶ Favourable

ΔH = -ve Yes
ΔS = +ve Yes
⟶ Spontaneous
$A \longrightarrow B$

Second Law:

$\Delta S = \oplus \uparrow$ Yes ⟶ favourable

$\Delta S = - \downarrow$ NO ⟶ Non-favourable

ΔH = -ve Yes
ΔS = -ve NO
⟶ ?

Gibb's Free energy

→ Entropy alone

→ Enthalphy alone

dnot decide spontaneouty

Predict direction of rxn.

The Gibbs free energy:

Energy available to do work

$G = H - TS$

H: Total Energy

TS: Energy cannot be used

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

? ? ?

A ⟶ B

ΔH, ΔS, ΔG

°C + 273 = K

Temp^{r} = Kelvin

- ΔG = -ve, ↓, Yes
- ΔG = +ve, ↑, NO
- ΔG = 0, equilibrium

Case A:

If ΔH = negative Yes

ΔS = positive Yes

ΔG = Negative Yes

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

-ve ← -ve

+ve

-ve

-ve

More Negative

A ⟶ B

- Always spontaneous at any temperature

eg Respiration

$$C_6H_{12}O_6 + O_2 \longrightarrow CO_2 + H_2O$$

Case B:

If ΔH = positive	NO
ΔS = negative	NO
ΔG = positive	NO

→ Non-spontaneous at any temperature.

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

$+ \quad \oplus \quad -$

$+$

$+$

$+\uparrow$

Example:

- Photosynthesis
- Protein biosynthesis

Case C: Example : Protein folding

SS DNA ⟶ DS DNA (Renaturation)

If ΔH = Negative
ΔS = Negative

[Yes / NO] ⟶ ΔG = -ve (Yes)

$T < \frac{\Delta H}{\Delta S}$

ΔG = +ve (NO) $T > \frac{\Delta H}{\Delta S}$

ΔH = -3000 J

ΔS = -10 J/K

$\Delta H/\Delta S = \frac{-3000}{-10} = 300$

299

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

$= (-3000) - (299 \times -10)$

$= (-3000) - (-2990)$

$= -3000 + 2990$

$\Delta G = -10$ J

301

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

$= (-3000) - (301 \times -10)$

$= (-3000) - (-3010)$

$= -3000 + 3010$

$\Delta G = +10$ J

Case D:

example : Denaturation of DNA or protein.

If ΔH = Positive → NO
ΔS = Positive → Yes
→ ? →

(Yes)
ΔG = −ve
$T > \frac{\Delta H}{\Delta S}$

ΔG = +ve
(NO)
$T < \frac{\Delta H}{\Delta S}$

NO ΔH = +10000 J
Yes ΔS = + 25 J/K

$\frac{10000}{25} = 400K$

minimum temp^r for rx^n to be spontaneous ?

- more than 400 K (400.1, 401 ...)

400 K − 273 = 127°C

ΔG does not provide information about rate of reaction

(fast or slow)

↓ Reaction direction

$A \xrightarrow{\Delta G = -100\ KJ/mole} B$ Forward

$X \xrightarrow{\Delta G = +200\ KJ/mole} Y$ Backward

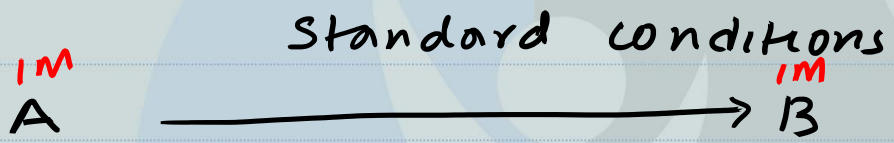
ΔG of the forward and backward reactions

$A \xrightarrow{\Delta G^{\circ} = -10\ KJ/mole} B$ Spontaneous

$B \xrightarrow{\Delta G^{\circ} = +10\ KJ/mole} A$ Non-Spontaneous

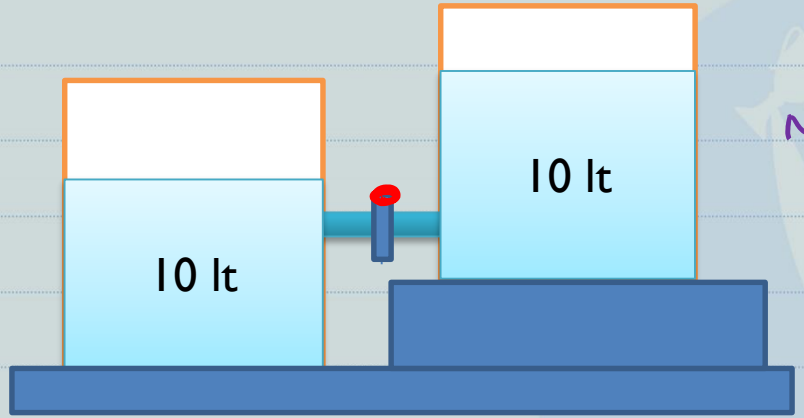
The Standard Gibbs free energy (ΔG^0)

- Energy available to do work under standard conditions



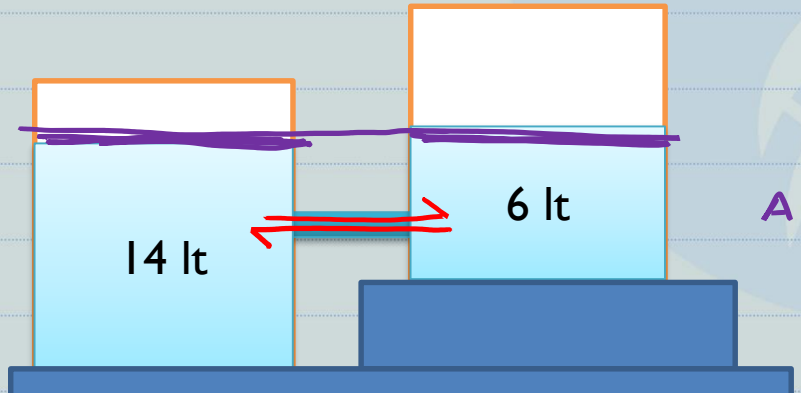
- Concentration of substrate & Product = 1 M
- Temperature = 25°C (298 K)
- Pressure = 1 Atm Pressure (760 mm Hg)
- pH = 0, (Biological system pH = 7)
  (ΔG^0) ($\Delta G^{0\prime}$)

Chemical equilibrium:



Not at eqbr

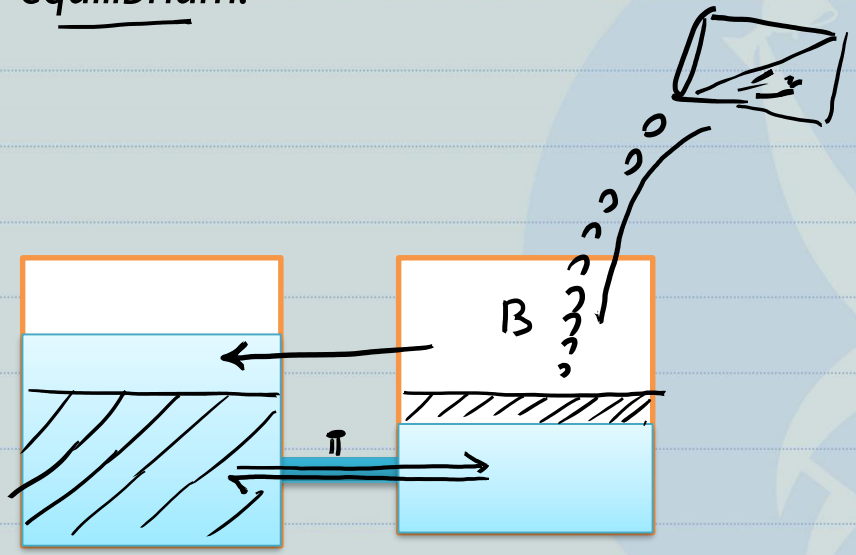
Equilibrium = Forward rate is equal to backward rate



At eqbr.

Le Châtelier's Principle:

Any deviation from equilibrium stimulates a process that tends to restore the system to equilibrium.



Keq = [P]/[S]

$$Keq = \frac{[P]}{[S]}$$

A ⟷ B

1 M 1 M (Standard)

	A	B
	1 M ⟵	1 M
Initial	1.8	0.2 M
Final	⟹	

finally 0.5M ⟶ 1.5M

$$Keq = \frac{0.2}{1.8} = \frac{1}{9} = 0.11$$

$$Keq = \frac{1.5}{0.5} = 3$$

9 | 10 | .1
9
10

Keq can be more than 1, less than 1 or 1

	A	B
Initially	1 M	1 M
Finally	1 M	1 M

$$Keq = \frac{1}{1} = 1$$

Keq > 1	Product ↑	Forward
Keq < 1	Substrate ↑	Reverse
Keq = 1	(P) = (S)	Eqbr.

Relationship between Keq and ΔG^0

$\Delta G^0 = -RT \ln Keq$

$\Delta G^0 = -2.303\ RT \log Keq$

Calone / mole

$R = 1.987$

$T = 298\ K$

$Keq > 1$ Forward $\Delta G^\circ = -ve$

$Keq < 1$ Backward $\Delta G^\circ = +ve$

$Keq = 1$ Eqbr $\Delta G^\circ = 0$

$\ln = 2.303 \log$

Keq>1 G=-ve

K_{eq} and ΔG^0

$2.303\ RT = 1.36$ K.cal/mole
$= 5.7$ KJ/mol

$\Delta G^0 = -1.36 \log Keq$ (for answer in Kcal/mole)

$\Delta G^0 = -5.7 \log Keq$ (for answer in KJ/mole)

$\Delta G^0 = -RT \ln Keq$

$\Delta G^0 = -2.303 RT \log Keq$

$\Delta G^0 = -1.36 \log Keq$

$= -1.36 \log [B]/[A]$

$= -1.36 \log 10$

$= -1.36 \times 1$

$= -1.36$ K.cal/mole

A ⟶ B ↑

Eqbr conc^n

A = 1 mM

B = 10 mM

ΔG^0 = ? (K.cal/mole)

Apply your mind:

$$K_{eq} = \frac{(P)}{(S)}$$

Equilibrium constant (K) of noncovalent interaction between two non-bonded atoms of two different groups was measured at 27°C. It was observed that K = 100 M^{-1}. The strength of this noncovalent interaction in terms of Gibbs free energy change is:

(1) ✗ 2.72 Kcal/mole
(2) ✓ - 2.72 Kcal/mole
(3) ✗ 0.27 Kcal/mole
(4) - 0.27 Kcal/mole

$$\Delta G^{\circ} = -1.36 \log K_{eq}$$
$$= -1.36 \log 100$$
$$= -1.36 \times 2$$
$$= -2.72$$

Apply your mind:

Equilibrium constant (K'eq) of reaction is a ratio of product to substrate concentration. The reaction between (K'eq) and free energy change in a reaction (ΔG') is follows

$$\Delta G' = -RT \ln K'eq$$

$= -1.36 \log K_{eq}$ (Kcal/mole)

$= -2.303\, RT \log K_{eq}$

Reaction A and reaction B have K'eq values of 10 and 100, respectively. Which of the following statement is correct with respect to ΔG'?

(1) ΔG' of A = ΔG' of B ×

(2) ΔG' of A > ΔG' of B ✓

(3) ΔG' of B > ΔG' of A

(4) ΔG' of A ≈ ΔG' of B ×

Rx A

$K_{eq} = 10$

$\Delta G^\circ = -1.36 \log 10$

$= -1.36 \times 1$

$\Delta G^\circ(A) = -1.36$

Rx B

$K_{eq} = 100$

$\Delta G^\circ = -1.36 \log 100$

$= -1.36 \times 2$

$\Delta G^\circ(B) = -2.72$

$\Delta G^\circ(A) > \Delta G^\circ(B)$

Apply your mind:

ΔG° for the base pairing of oligonucleotides (n=5) at 300 K is −17.1 kJ mol^{-1}. What would be the approximate value of the equilibrium constant K?

(1) 100 (2) 10

(3) 1000 (4) 1

$\Delta G^\circ = -17.1$ KJ/mole

$K_{eq} = ?$

$\Delta G^\circ = -5.7 \log K_{eq}$

$-17.1 = -5.7 \log K_{eq}$

$\frac{-17.1}{-5.7} = \log K_{eq}$

$3 = \log K_{eq}$

Anti log $3 = K_{eq}$

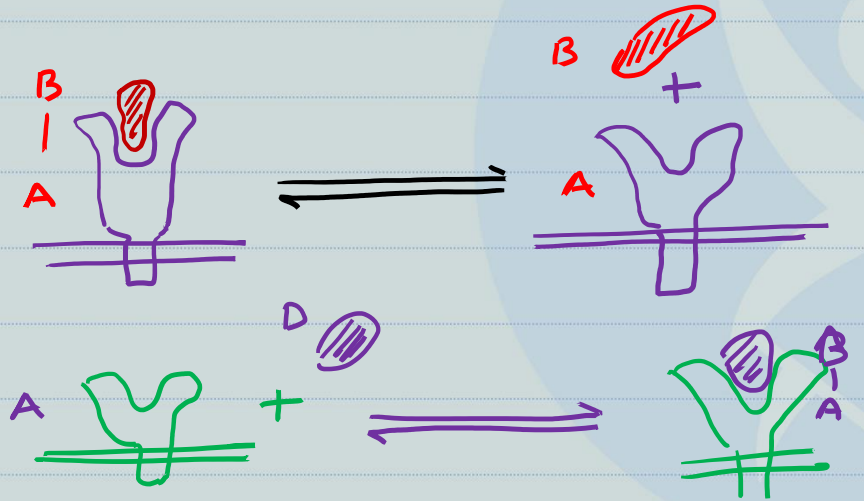
$1000 = K_{eq}$

Standard free energy change in molecular interactions:

- Dissociation constant (Kd)
- Binding or association constant (Ka)

K_{eq}

$\checkmark K_d = \frac{1}{K_a}$ $\checkmark K_a = \frac{1}{K_d}$



$K_d = \frac{(P)}{(S)} = \frac{(A).(B)}{(A-B)}$

$K_a = \frac{(P)}{(S)} = \frac{(A-B)}{(A).(B)}$

Apply your mind

K_d = less than 1

The Gibbs free energy of binding of a ligand with a protein is determined using calorimetric measurements at 25°C. The value of ΔG° thus determined is +2.72 kcal/mole. The binding constant for the ligand-protein association is:

(1) 0.01 (2) 0.10

~~(3) 1.00~~ ~~(4) 100~~

K_a

Antilog of $-2 = 1/100$

$= 0.01$

$$\Delta G^\circ = -1.36 \log K_a$$

$$+2.72 = -1.36 \log K_a$$

$$\frac{+2.72}{-1.36} = \log K_a$$

$$-2 = \log K_a$$

$$\text{Antilog } -2 = K_a$$

Van't Hoff Equation : Relationship bet^n keq, enthalpy & Entropy

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

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

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

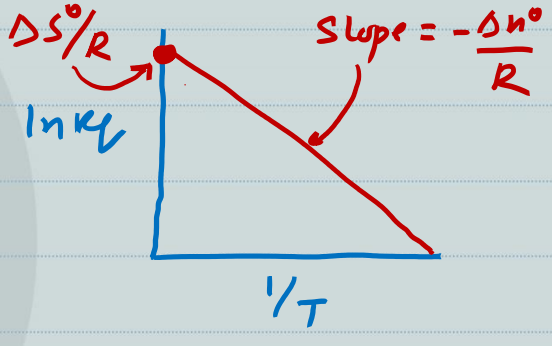
$$RT \ln K_{eq} = -\Delta H^\circ + T\Delta S^\circ$$

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

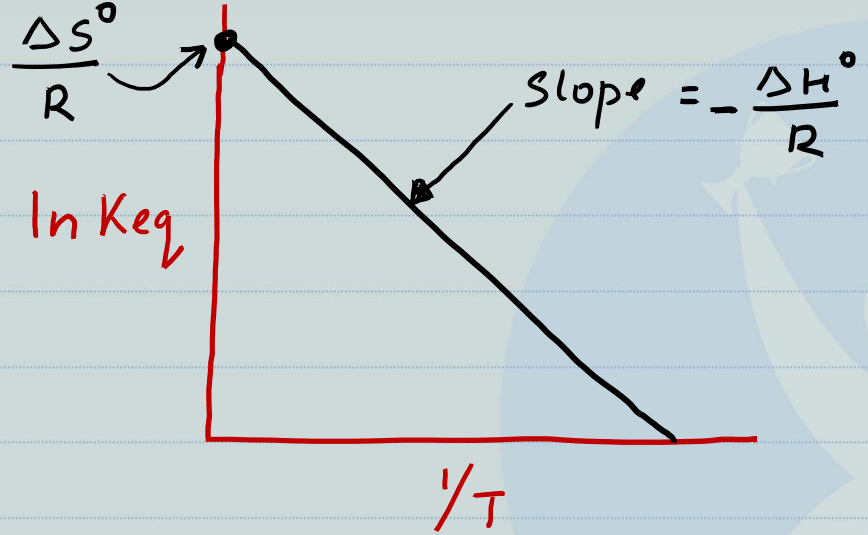
$$\ln K_{eq} = \frac{-\Delta H^\circ}{R} \cdot \frac{1}{T} + \frac{T\Delta S^\circ}{RT}$$

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

$Y = -m \cdot x + c$



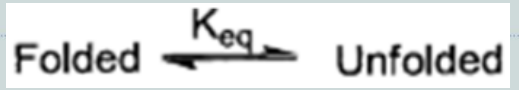
Van't Hoff Plot

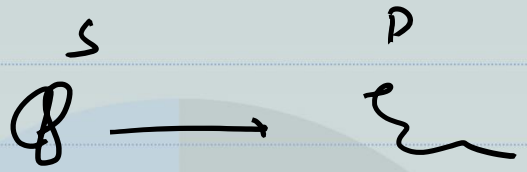


- Independent variable
  x- axis = 1/T
- Dependent variable
  Y- axis = ln Keq

Apply your mind:

Protein stability is represented as





Prior to development of sensitive calorimeters, thermodynamic parameters of processes were determined by following equation

$$\ln K_{eq} = \frac{-\Delta H^o}{R}\left(\frac{1}{T}\right) + \frac{\Delta S^0}{R}$$

ΔH^0 and ΔS^0 are standard changes m enthalpy and entropy, respectively. Which one of the following statements is correct for estimating ΔG, ΔH and ΔS?

(1) Determining the ratio of folded & unfolded protein at 37°C

(2) Plotting K_{eq} as a function of ΔH

(3) Plotting K_{eq} against ΔS

(4) Plotting ln K_{eq} against 1/temperature

Sequential and coupled reactions

Standard free-energy changes are additive for sequential and coupled reactions

Sequential :

$\Delta G^\circ = +1$ $\Delta G^\circ = -2$ $\Delta G^\circ = -1$

$A \longrightarrow B \longrightarrow C \longrightarrow D$

$A \longrightarrow D$ $\Delta G^\circ_{overall} = \Delta G^\circ_1 + \Delta G^\circ_2 + \Delta G^\circ_3$

$= 1 - 2 - 1$

$= -2$

Coupled :

$\Delta G^\circ = +2$

$A \xrightarrow{1} B$

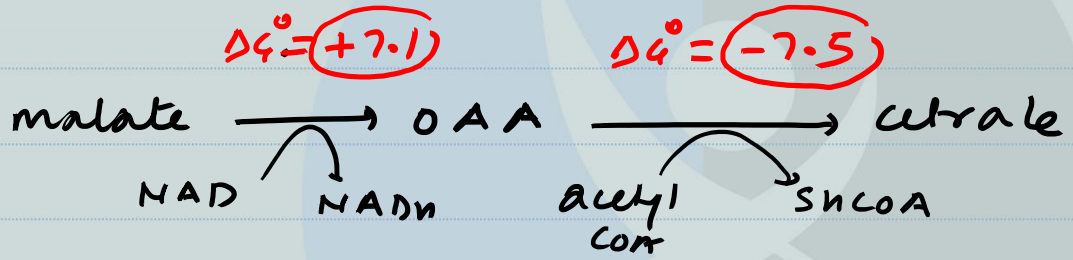
$C \xrightarrow{2} D$ $\Delta G^\circ = -5$

Rx^n		Rx^2		Overall ΔG°
2	−	5	=	−3

Sequential reactions

Malate + NAD+ → OAA + NADH + H+ ΔG^0 = + 7.1 kcal/mole

OAA + Acetyl CoA → Citrate + SHCoA ΔG^0 = - 7.5 kcal/mole



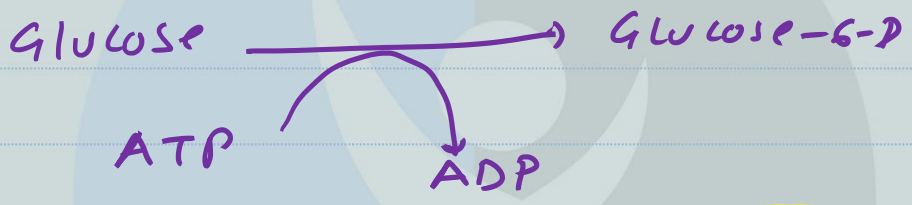
malate → citrate (favourable)

= (+7.1) + (−7.5)

= − 0.4 K.cal/mole

Coupled reactions

		ΔG°
Glucose + iP	⟶ Glucose-6-P	+ 5 KJ/mole
ATP	⟶ ADP + iP	− 30 KJ/mole



$\Delta G^\circ_{overall} = (Rxn) + (Rxn\ 2)$

$= (+5) + (-30)$

$= 5 - 30$

$= -25$

Apply your mind:

The reaction of glutamate and ammonia to glutamine and water has a value of + 14 kJ mol^{-1} for ΔG_o. This is coupled with the ATP reaction (ATP + H_2O → ADP + phosphate). The ΔG_o for this reaction is -30 kJ mol^{-1}.

The ΔG_o (KJ mol^{-1}) for the coupled reaction

Glutamate + NH_3 + ATP → Glutamine + ADP + Phosphate

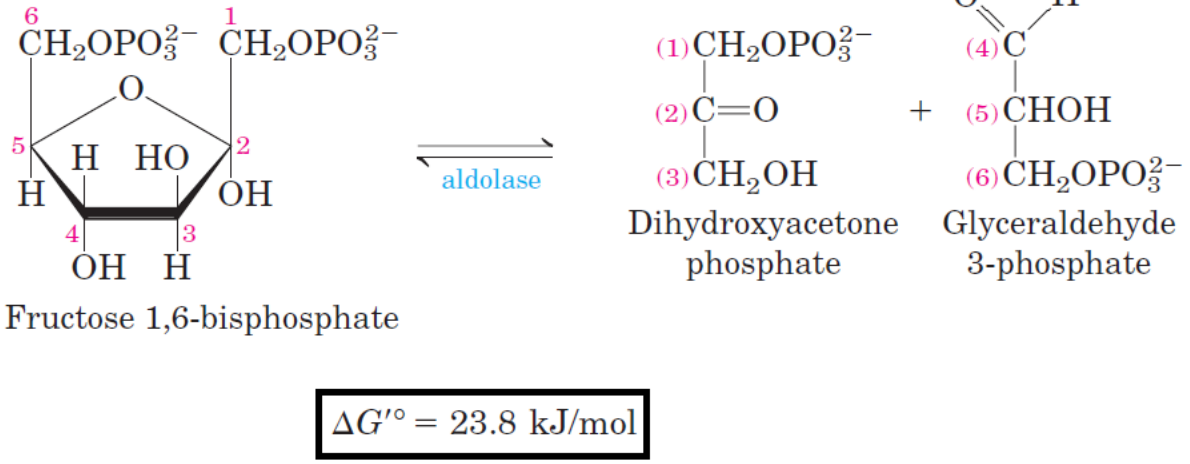
under equilibrium condition is (JUNE 2014)

(1) 16
(2) -44
(3) -16
(4) 44

+14
+ −30
−16

What will be direction of reaction under cellular conditions?

Forward/backward? under standard condition



Fructose 1,6-bisphosphate

$\Delta G'^{\circ} = 23.8$ kJ/mol

cannot be predicted until actual cellular conc^{r} of substrate & products are given.

ΔG° ⟶ standard condition
not about cellular condition

for cellular condition we need ΔG

Actual free-energy (ΔG) changes depend on reactant and product concentrations under physiological conditions

cellular conc^{n}

$$\Delta G = \Delta G^{\circ} + 2.303\, RT \log \frac{[P]}{[S]}$$

$$\Delta G^{\circ} \text{ (K cal/mole)} = \Delta G \text{ (K·cal/mole)} + 1.36 \log \frac{(P)}{(S)}$$

$$\Delta G^{\circ} \text{ (KJ/mole)} = \Delta G \text{ (KJ/mole)} + 5.7 \log \frac{(P)}{(S)}$$

Apply your mind:

$273 + 30 = 303\,K$

The standard free energy change (ΔG_o) per mole for the reaction A → B at 30°C in an open system is −1000 cal/mole. What is the approximate free energy change (ΔG) when the concentration of A and B are 100 micromolar and 100 millimolar, respectively? (JUNE 2016)

(1) 3160 cal/mole
(2) 316 cal/mole
(3) 31610 cal/mole
(4) -3160 cal/mole

$\Delta G^0 = -1000$ cal/mole

A ⟶ B

A: 100 μM, B: 100 mM

$$\Delta G = \Delta G^0 + 2.303\,RT \log \frac{(B)}{(A)}$$

$$= (-1000) + 1381 \log \frac{100 \times 10^{-3}\,M}{100 \times 10^{-6}\,M}$$

$= 2.303 \times R \times T$

$= 2.303 \times 1.987 \times 303$

$$= -1000 + 1381 \log 10^3$$

$$= -1000 + 1381 \times 3$$

$$= -1000 + 4143 = +3143$$

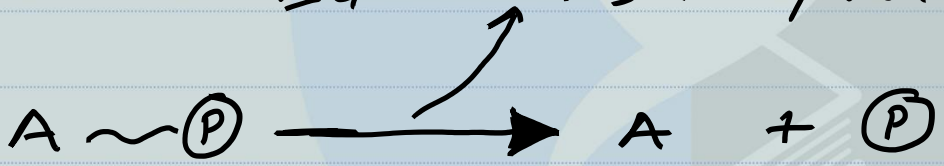
Energy Rich Molecules

- which has 1 or more energy rich bond

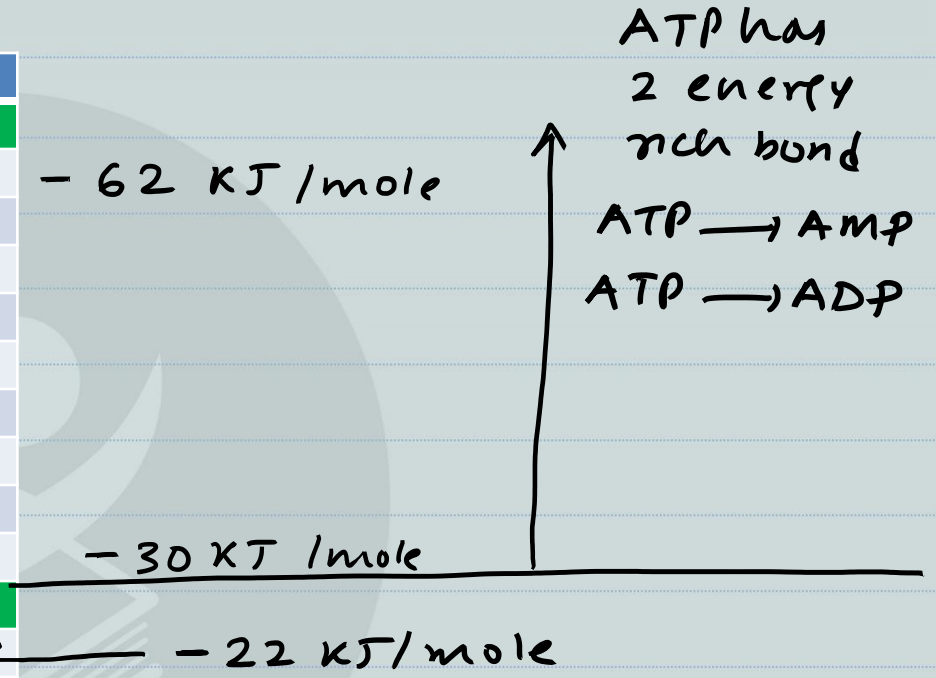
- Energy rich bond on hydrolysis

$\Delta G^{\circ} = -30$ KJ/mole or more negative

$\Delta G^{\circ} = -7.3$ K cal/mole or more negative



Compound	ΔG°′ (Kcal/mol)
High Energy Bonds	
1. Phosphoenolyruvate	– 14.8
2. Carbamoyl phosphate	– 12.3
3. 1,3-Bisphosphoglycerate	– 11.8
4. ATP (→ AMP + PPi)	– 10.9
5. Phosphate creatine	– 10.3
6. Acetyl phosphate	– 10.1
7. Arginine phosphate	– 8.0
8. Acetyl CoA	– 7.5
9. ATP (→ ADP + Pi)	– 7.3
Low Energy bonds	
1. Glucose 1-phosphate	– 5.0
2. Pyrophosphate (→ 2Pi)	– 4.0
3. Fructose 6-phosphate	– 3.8
4. AMP (→ Adenosine + Pi)	– 3.4
5. Glucose 6-phosphate	– 3.3
6. Glycerol 3-phosphate	– 2.2



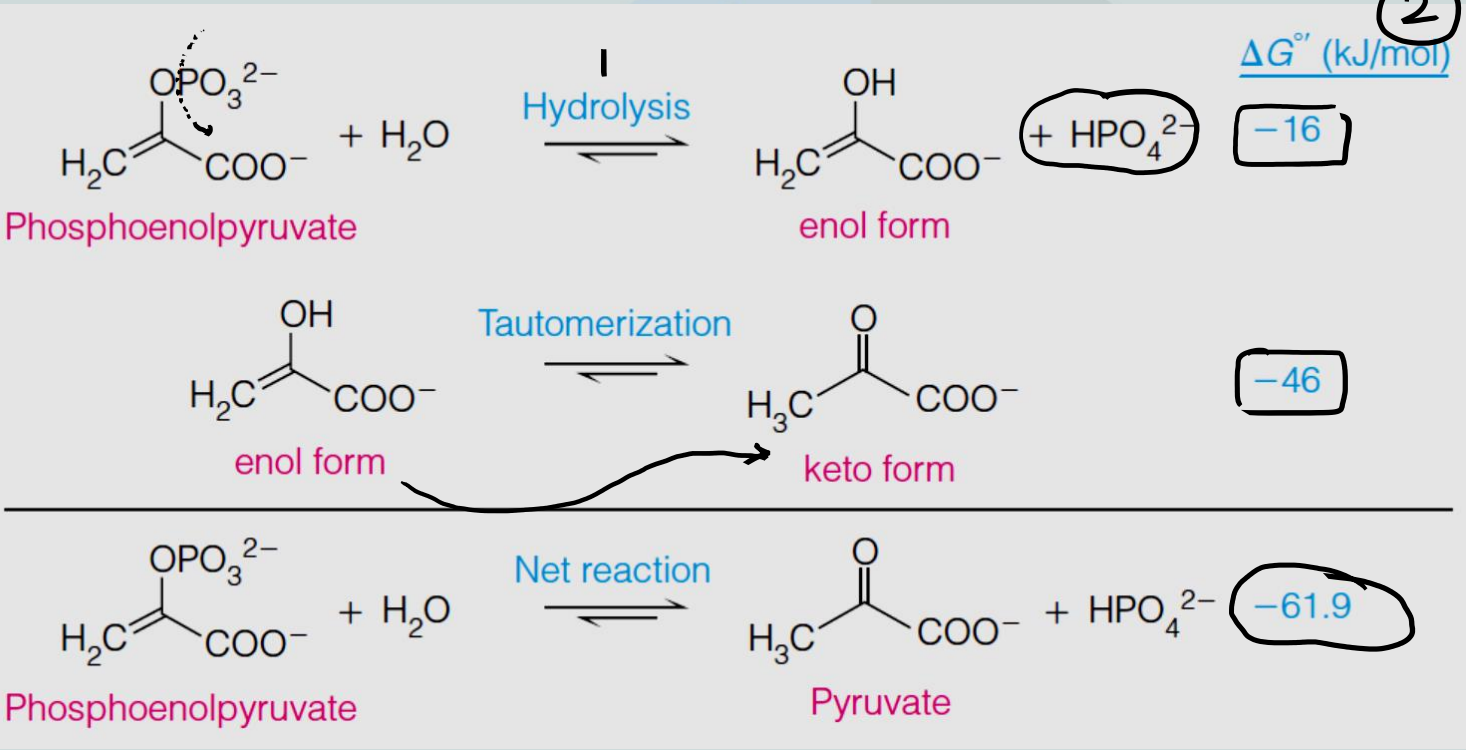
−16

1 Hydrolysis of Phosphate

2 Tautomeric Shift

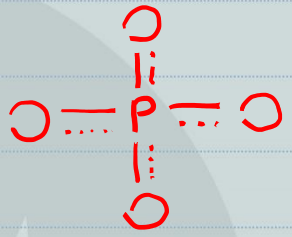
−46

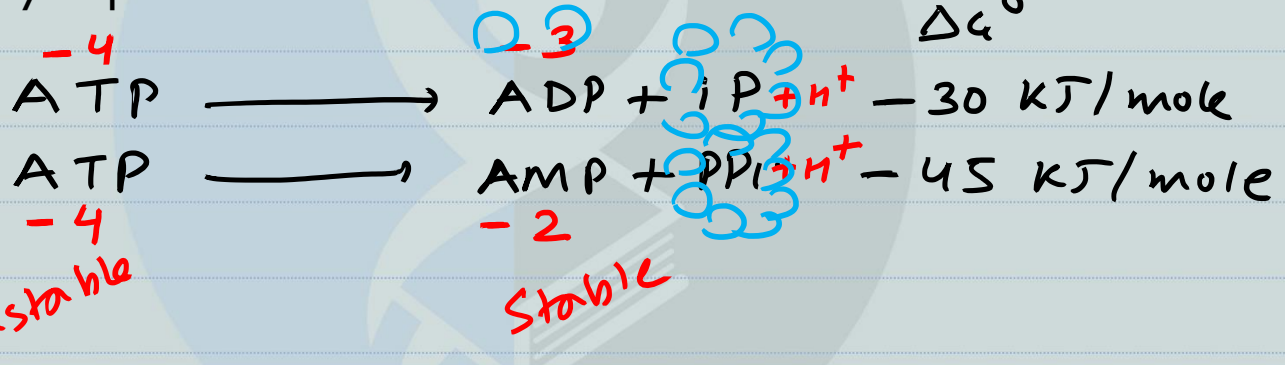
Chemical basis for large free energy change with PEP hydrolysis



Chemical basis for large free energy change with ATP hydrolysis

- Charge repulsion in substrate
- Resonance stabilization of production
- Ionization favors forward reaction
- Hydration stability of products





HAPPY LEARNING

THANKS

IfAS

India's No **1** for NEET, IIT-JAM, GATE and NET Exams

www.ifasonline.com

9172266888