More on Gibbs Energy

$2NO_2(g) \rightarrow N_2O_4(g)$

- What do you predict for the sign of ΔS?
- ΔS° = –175.8 J/K.mol
- ΔH° = –57.2 kJ/mol
- What is ΔG°?
  A. – 4.8 kJ/mol
  B. 5.2×10^4 kJ/mol
  C. 4.8 kJ/mol
  D. -29 kJ/mol

$2NO_2(g) \rightarrow N_2O_4(g)$

- What do you predict for the sign of ΔS?
- ΔS° = –175.8 J/K.mol
- ΔH° = –57.2 kJ/mol
- What is ΔG°?
- ΔG°= ΔH° –TΔS°
- = – 57.2 kJ/mol – 298K(–175.8 J/K.mol x 1kJ/1000J)
- = – 4.8 kJ/mol

Careful of the units!

At what temperature will the reaction

- When $\Delta G° = 0$

When $\Delta G = 0$

- The system is at **equilibrium**
  - (for example a phase change)
- So $\Delta G = \Delta H - T\Delta S$ becomes
- $0 = \Delta H - T\Delta S$
- $\Delta H = T\Delta S$
- For water boiling $\Delta H° = 40.65$ kJ/mol
- What is $\Delta S°$?
- $\Delta S° = 109$ J/K.mol

$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$

This is a combustion reaction: what do you predict the sign of enthalpy change will ΔH be?

A. +
B. –
C. 0
D. Don't know

What about ΔS? ΔG?

$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$

ΔH is –, ΔS is +, and ΔG is – (at all temperatures)

This reaction is always favorable!

$6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$

ΔH is +, ΔS is –, and ΔG is + (at all temperatures)

This reaction is NEVER favorable!
But it happens – (how?)

"captured energy" + $6\ CO_2(g) + 6\ H_2O(g)$

As we will see next semester, the

Calculating $\Delta H°$, $\Delta S°$, and $\Delta G°$ from standard thermodynamic formation data.

- There are tables of these values that can be used to calculate overall changes for a reaction
- $\Delta H°_f$ (standard heat of formation)
  - Enthalpy change when a substance is formed from its elements in their standard states.
- For any reaction
- $\Delta H°_f = \Sigma \Delta H°_f$ products $- \Sigma \Delta H°_f$ reactants

$6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$

- $\Delta H°_f$ for:
- $CO_2(g) = -394$ kJ/mol
- $H_2O(g) = -242$ kJ/mol
- $C_6H_{12}O_6(s) = -1271$ kJ/mol
- $O_2(g) = 0$ kJ/mol (element in its standard state)

Third Law of Thermodynamics

- The entropy of a perfect crystal at 0K is 0
- This allows us to calculate actual entropies of materials (as opposed to changes in entropy)
- S° $H_2O(l)$ = 69.9 J/K.mol, $H_2O(g)$ = 189 J/K.mol
- And
- $\Delta S°_r = \Sigma S°$ products – $\Sigma S°$ reactants