AP15

Thermodynamics and Reactions

Chemical or physical changes are driven forward by

decreases in enthalpy, negative ΔH or both. increases in entropy, positive ΔS

Reactions are thermodynamically favored when **decrease Enthalpy** by converting potential energy into kinetic (release heat) energy and produce products with greater bond energies (products with stronger bonds.

ΔH = negative **ΔBond Energy = positive**

$H_{products} < H_{reactants}$ $BE_{reactants} < BE_{reactants}$

Reactions are thermodynamically favored when they **increase Entropy** by increasing the dispersion of energy or matter.

ΔS = positive

$S_{products} > S_{reactants}$

Noteworthy Information about ΔH

The absolute enthalpy of any substance, H, cannot be determined.

H = ?? kJ

Changes in energy can be observed and measured.

$\Delta H° = \Sigma n\Delta H°(\text{products}) - \Sigma n\Delta H°(\text{reactants})$

ΔH can be found since it reflects the **change in enthalpy** for a rxn.

By definition the ΔH_f of an element at standard conditions is 0

ΔH is normally measured in kJ/mol$_{rxn}$

Noteworthy Information about ΔS

ΔS + when:
- greater # of moles of gas formed
- gas > liquid > solid
- greater volume formed
- greater temperature
- greater number of moles formed

The absolute entropy, $S°$, of a substance can be determined and will always be greater than zero.

$S° > 0$ J/K

However, it is possible for the $\Delta S°$ of a reaction to be zero.

$\Delta S° = \Sigma nS°(\text{products}) - \Sigma nS°(\text{reactants})$

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ΔG Gibbs Free Energy – the ability to do work

$\Delta G < 0$ Thermodynamically favored $A \rightarrow B$
$\Delta G = 0$ Rxn at equilibrium $A \rightleftarrows B$
$\Delta G > 0$ Not thermodynamically favored $A \leftarrow B$

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

When ΔH° = neg ΔS°= pos	**ΔG° always neg $\rightarrow$ at all temps**	
When **ΔH° = pos ΔS°= neg**	**ΔG° always pos $\leftarrow$ at all temps w**	
ΔH° and ΔS° have same sign.	**ΔG° either pos or neg temperature dependent**	

Finding ΔG^o using formation table

$\Delta G^\circ = \Sigma n \Delta G_f^\circ\text{(prod)} - \Sigma n \Delta G_f^o\text{(react)}$

Get ΔG_f° from table

ΔG_f° of elements by definition = 0

K_{eq} and ΔG°

$\Delta G^\circ = -RT\ln K_{eq}$ $\quad K_{eq} = e^{-\frac{\Delta G^o}{RT}}$

ΔG° is almost always in kJ, so you will most likely want to change R to 0.008314 **kJ** /K mol

Under standard conditions, 25°C, RT = 2.5 kJ/mol (easy to remember and allows for fast solutions)

$K_{eq} = e^{-\frac{\Delta G^o}{2.5}}$ @ standard temp of 298K

Rule of 25 kilojoules

Starting at standard conditions° (reactants and products at 1 M and 1 atm)

ΔG° = -25 kJ means that K_{eq} is very large: reaction goes to completion $\rightarrow$

ΔG° = +25 kJ means that K_{eq} is very small: reaction will convert to just reactants $\leftarrow$