IV. Chemical Equilibrium

A) Law of Chemical Equilibrium

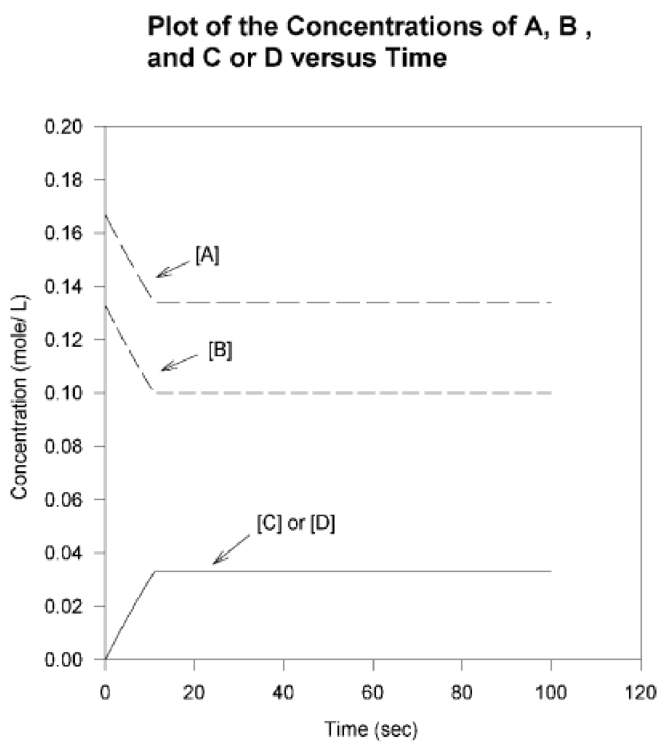
Concentration expressed as molarity

$$Molarity\ of\ A = [A] = \frac{moles\ of\ A}{volume\ in\ liters} \qquad units: \frac{mole}{L} = M$$

Consider the hypothetical reaction

$$A\ (g) + B\ (g) \rightarrow C\ (g) + D\ (g)$$

Compound A is allowed to react with B and the concentrations of the reactants and products are plotted as a function of time.



Note: After eleven seconds the concentrations do not change with time.

Chemical equilibrium is the condition that exists when the concentrations of the reactants and products no longer change with time. Chemical equilibrium is a dynamic state. It occurs when the rate of the forward reaction (A + B $\rightarrow$ C + D) equals the rate of the reverse reacion (A + B $\leftarrow$ C + D).

$$A\ (g) + B\ (g) \rightleftharpoons C\ (g) + D\ (g)$$

indicates equilibrium

At equilibrium there is a specific relationship between the concentrations of the reactants and products.

Consider three experiments for the reaction

$$A(g) + B(g) \rightleftharpoons C(g) + D(g)$$

in which the initial concentrations of A and B vary from experiment to experiment.

Table of Equilibrium Concentrations

Exp. #	**$[A]_{eq}$**	**$[B]_{eq}$**	**$[C]_{eq}$**	**$[D]_{eq}$**
1	3.00	2.00	1.00	1.00
2	9.60	10.0	4.00	4.00
3	0.500	3.00	0.500	0.500

The following relationship fits the data in each experiment.

$$\frac{[C]_{eq}\ [D]_{eq}}{[A]_{eq}\ [B]_{eq}} = 0.167 = constant = K_c$$

For the general reaction

$$m\,A(g) + n\,B(g) + \cdots \rightleftharpoons p\,C(g) + q\,D(g) + \cdots$$

K_c is defined

$$\frac{[C]^p_{eq}\ [D]^q_{eq}\ \cdots}{[A]^m_{eq}\ [B]^n_{eq}\ \cdots} = K_c$$

The magnitude of the constant K_c depends on the nature of the reaction and the temperature.

The units of K_c depend on the specific reaction.

Example: $A + 3\,B \rightleftharpoons 2\,C$

$$K_c = \frac{[C]^2_{eq}}{[A]_{eq}\ [B]^3_{eq}} \qquad units: \frac{\left[\frac{mole}{L}\right]^2}{\left[\frac{mole}{L}\right]^4} = \left[\frac{L}{mole}\right]^2$$

Example: $A + B \rightleftharpoons C + D$

$$K_c = \frac{[C]_{eq}\ [D]_{eq}}{[A]_{eq}\ [B]_{eq}} \qquad units: \frac{\left[\frac{mole}{L}\right]^2}{\left[\frac{mole}{L}\right]^2} = unitless$$

Terms:

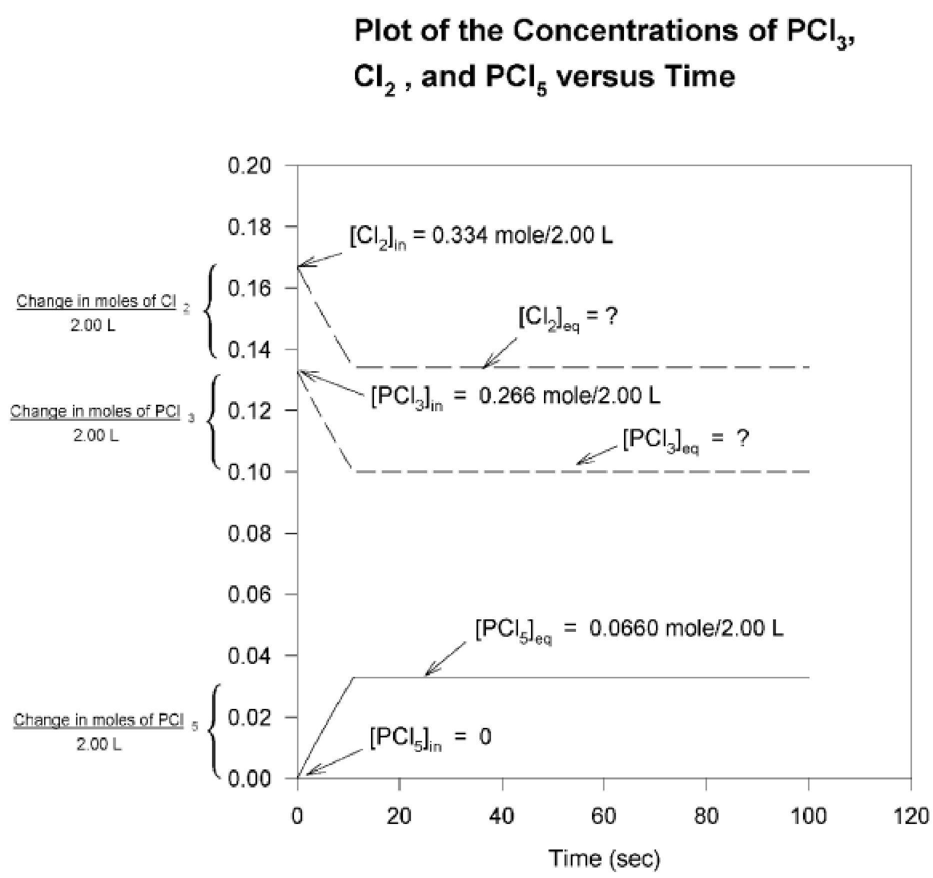
When $K < 1$, then the equilibrium is said to lie to the *left*.
When $K > 1$, then the equilibrium is said to lie to the *right*.

B) Equilibrium Calculations

Example #1: Initially, 0.2660 mole of PCl_3 and 0.3340 mole of Cl_2 are placed in a 2.00 L container at 300°C. If the moles of PCl_5 are found experimentally to be 0.0660 at equilibrium, calculate K_c for the reaction

$$PCl_3 + Cl_2 \rightleftharpoons PCl_5$$

at 300°C.



where $[A]_{in}$ is the initial (at time = 0) molarity of A and $[A]_{eq}$ is the equilibrium molarity of A

Important observations from the above plot

• *Initial moles of a substance + change in moles of substance = equilibrium moles of substance*

Example:

initial moles of PCl_5 + change in moles of PCl_5 = equilibrium moles of PCl_5

0.0 moles of PCl_5 + 0.0660 moles of PCl_5 = 0.0660 moles of PCl_5

• *The changes in number of moles of reactants and products are related by* ***mole ratios****.*

These observations can be utilized in an Equilibrium Table to solve equilibrium problems.

Equilibrium Table

$$PCl_3 + Cl_2 \rightleftharpoons PCl_5$$

	PCl_3	Cl_2	PCl_5	
initial:	0.2660	0.3340	0.0	mole
change:	-0.0660	-0.0660	0.0660	mole
equil:	0.2000	0.2680	0.0660	mole

Change in moles of PCl_5 = (equilibrim moles PCl_5) - (initial moles of PCl_5)

Negative sign indicates moles of Cl_2 are consumed

Change in moles of Cl_2 = (change in moles PCl_5)(1 mole Cl_2 /1 mole PCl_5)

Equilibrium moles of Cl_2 = (initial moles Cl_2) + (change in moles Cl_2)

Substituting the equilibrium moles into the equilibrium expression

$$K_c = \frac{[PCl_5]_{eq}}{[PCl_3]_{eq}\ [Cl_2]_{eq}} = \frac{\left[\frac{0.0660\,mole}{2.00\,L}\right]}{\left[\frac{0.200\,mole}{2.00\,L}\right]\left[\frac{0.268\,mole}{2.00\,L}\right]} = 2.46\ \frac{L}{mole}$$

Example #2: Initially, 3.00 moles of PCl_3 and 3.00 moles of Cl_2 are placed in a 5.00 L container at 300°C. Calculate the equilibrium concentration of PCl_5, $[PCl_5]_{eq}$. K_c is 2.46 L/mole.

Unknown: $[PCl_5]_{eq}$, Let x = moles of PCl_5 at equilibrium.
Knowns: initial moles of PCl_3 = 3.00 mole, initial moles of Cl_2 = 3.00 mole, volume of container =

5.00L, K_c = 2.46 L/mole

Concepts: chemical equilibrium, molarity

Substitute known values into equilibrium table.

	PCl_3	+	Cl_2	$\rightleftharpoons$	PCl_5	
initial	3.00		3.00		0	mole
change						mole
equil					x	mole

Use the concepts presented above to fill in the table.

	PCl_3	+	Cl_2	$\rightleftharpoons$	PCl_5	
initial	3.00		3.00		0	mole
change	- x		- x		x	mole
equil	3.00 - x		3.00 - x		x	mole

Substituting equilibrium moles into the equilibrium expression

$$K_c = \frac{[PCl_5]_{eq}}{[PCl_3]_{eq}\ [Cl_2]_{eq}} = \frac{\left[\frac{x}{5.00L}\right]}{\left[\frac{3.00-x}{5.00L}\right]\left[\frac{3.00-x}{5.00L}\right]} = 2.46\ \frac{L}{mole}$$

$$2.46x^2 - 19.7\overline{6}x + 22.1\overline{4} = 0$$

For the quadratic equation

$$ax^2 + bx + c = 0$$

the solutions are

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Substituting

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(-19.7\overline{6}) \pm \sqrt{(-19.7\overline{6})^2 - 4(2.46)(22.1\overline{4})}}{2(2.46)}$$

$$x = 1.35,\ 6.67$$

Although both roots (1.35 and 6.67) are mathematically correct, only one root satisfies the physical constraints of the problem and that root is 1.35. If root 6.67 is used, then the moles of PCl_3 and Cl_2 at *equilibrium* would be - 3.67 (3.00 - 6.67) and the *equilibrium* concentrations of PCl_3 and Cl_2 would be - 0.734 M. *A negative or zero equilibrium concentration for reactants or products is physically impossible.*

$$x = 1.35\ mole = moles\ of\ PCl_5\ at\ equilibrium$$

$$[PCl_5]_{eq} = \frac{x}{5.00\,L} = \frac{1.35\,mole}{5.00\,L}$$

$$[PCl_3]_{eq} = \frac{(3.00 - x)}{5.00\,L} = \frac{(3.00 - 1.35)}{5.00\,L} = \frac{1.65\,mole}{5.00\,L}$$

$$[Cl_2]_{eq} = \frac{1.65\,mole}{5.00\,L}$$

The degree of dissociation, α, of B

$$B \rightleftharpoons 2\ A$$

is defined as

$$\alpha = \frac{change\ in\ moles\ of\ B}{initial\ moles\ of\ B}$$

and the percent dissociation is (α)(100).

Example #3: If 3.00 moles of B are placed in a 1.00 L flask at 25°C, calculate the degree of dissociation, α, of B. K_c for B at 25°C is 0.100 mole/L.

Unknown: change in moles of B. Let x = change in moles of B
Knowns: initial moles of B = 3.00 mole, volume of flask = 1.00 L, K_c = 0.100 mole/L
Concepts: chemical equilibrium, degree of dissociation

Substitute known values into the equilibrium table.

	B	⇌	2 A	
initial	3.00		0	mole
change	- x			mole
equil				mole

The reaction must go to the right in order to produce A. Moles of B will be consumed and thus, the change in moles of B will be -x.

Use the concepts presented above to fill in the table.

	B	⇌	2 A	
initial	3.00		0	mole
change	- x		2 x	mole
equil	3.00 - x		2 x	mole

Substitute equilibrium moles into the equilibrium expression.

$$K_c = \frac{[A]^2_{eq}}{[B]_{eq}} = \frac{\left[\frac{2x}{1.00\,L}\right]^2}{\left[\frac{3.00-x}{1.00\,L}\right]} = 0.100\ \frac{mole}{L}$$

$$4.00x^2 + 0.100x - 0.300 = 0$$

$$x = 0.261,\ -0.287$$

x = 0.261 mole = change in moles of B

$$\alpha = \frac{change\ in\ moles\ of\ B}{initial\ moles\ of\ B} = \frac{x}{3.00\,mole} = \frac{0.261\,mole}{3.00\,mole} = 8.70\cdot 10^{-2}$$

C) Factors Affecting Equilibrium Concentrations at Constant Temperature

Le Châtelier's principle - when a system is in equilibrium , a change in any one of the factors upon which the equilibrium depends will cause the equilibrium to shift in such a way as to diminish the effect of the change.

To determine if the reaction

$$a\,A\ (g) + b\,B\ (g) \rightleftharpoons c\,C\ (g) + d\,D\ (g)$$

will go to the left or to the right to reestablish the equilibrium, Q is defined as

$$Q = \frac{[C]^c_{in}\ [D]^d_{in}}{[A]^a_{in}\ [B]^b_{in}}$$

If $Q < K_c$, then the reaction will go to the *right* to reestablish the equilibrium. Moles of reactants A and B will be consumed and moles of products C and D will be produced.

If $Q > K_c$, then the reaction will go to the *left* to reestablish the equilibrium. Moles of products C and D will be consumed and moles of reactants A and B will be produced.

1) Addition or Removal of Moles of Reactants or Products

Example #4: If 1.00 mole of PCl_5 is added to the equilibrium in Example #2, what will be the NEW equilibrium concentration of PCl_5?

Unknown: new equilibrium molarity of PCl_5 after the addition of 1.00 mole of PCl_5
Knowns: *equilibrium* moles of PCl_3, Cl_2, and PCl_5 (1.65, 1.65, and 1.35 mole) *before* the addition of 1.00 mole of PCl_5 (see Example #2); K_c = 2.46 L/mole; volume of the container = 5.00 L

Concepts: chemical equilibrium, Q, molarity

Step #1: Evaluate Q and determine whether the reaction will go to the left or right to reestablish the equilibrium.

	PCl_3	+	Cl_2	$\rightleftharpoons$	PCl_5
initial	1.65		1.65		1.35 + 1.00 mole

The moles in the table are the equilibrium values from **Example #2** and the 1.00 mole of PCl_5 added.

$$Q = \frac{[PCl_5]_{in}}{[PCl_3]_{in}\ [Cl_2]_{in}} = \frac{\left[\frac{(2.35\,mole)}{5.00\,L}\right]}{\left[\frac{(1.65\,mole)}{5.00\,L}\right]\left[\frac{(1.65\,mole)}{5.00\,L}\right]} = 4.31\ \frac{L}{mole}$$

$$Q = 4.31\,\frac{L}{mole} > 2.46\,\frac{L}{mole} = K_c$$

Since $Q > K_c$, the reaction will go to the left to reestablish the equilibrium. Moles of PCl_5 will be consumed and moles of PCl_3 and Cl_2 will be produced.

Step #2: Solve the equilibrium problem.

Let x = change in moles of PCl_5. Since the reaction goes to the left to reestablish the equilibrium, moles of PCl_5 will be consumed and the entry for PCl_5 in the row labeled "change" of the equilibrium table will be -*x*.

	PCl_3	+	Cl_2	$\rightleftharpoons$	PCl_5	
initial	1.65		1.65		2.35	mole
change	x		x		- x	mole
equil	1.65 + x		1.65 + x		2.35 - x	mole

$$K_c = \frac{[PCl_5]_{eq}}{[PCl_3]_{eq}\ [Cl_2]_{eq}} = \frac{\left[\frac{2.35 - x}{5.00\,L}\right]}{\left[\frac{1.65 + x}{5.00\,L}\right]\left[\frac{1.65 + x}{5.00\,L}\right]} = 2.46\ \frac{L}{mole}$$

$$0.492x^2 - 2.62\overline{4}x + 1.01\overline{0} = 0$$

x = 0.361 mole = change in moles of PCl_5

$$[PCl_5]_{eq} = \frac{(2.35\,mole - x)}{5.00\,L} = \frac{(2.35\,mole - 0.36\,mole)}{5.00\,L} = \frac{1.99\,mole}{5.00\,L}$$

2) Change in Volume of the Container

Example #5: If the volume of the container in Example #2 is decreased from 5.00 L to 2. 00 L, will the moles of PCl_5 at the NEW equilibrium be greater or less than the moles of PCl_5 in Example #2?

Unknown: new equilibrium moles of PCl_5 after the volume of the container in Example #2 is reduced to 2.00 L
Knowns: equilibrium moles of PCl_3, Cl_2, and PCl_5 in Example #2, new volume = 2.00 L, K_c = 2.46 L/mole
Concept: Q, molarity

Step #1

	PCl_3	+	Cl_2	⇌	PCl_5	
initial	1.65		1.65		1.35	mole

$$Q = \frac{[PCl_5]_{in}}{[PCl_3]_{in}\ [Cl_2]_{in}} = \frac{\left[\frac{(1.35\,mole)}{2.00\,L}\right]}{\left[\frac{(1.65\,mole)}{2.00\,L}\right]\left[\frac{(1.65\,mole)}{2.00\,L}\right]} = 0.992\ \frac{L}{mole}$$

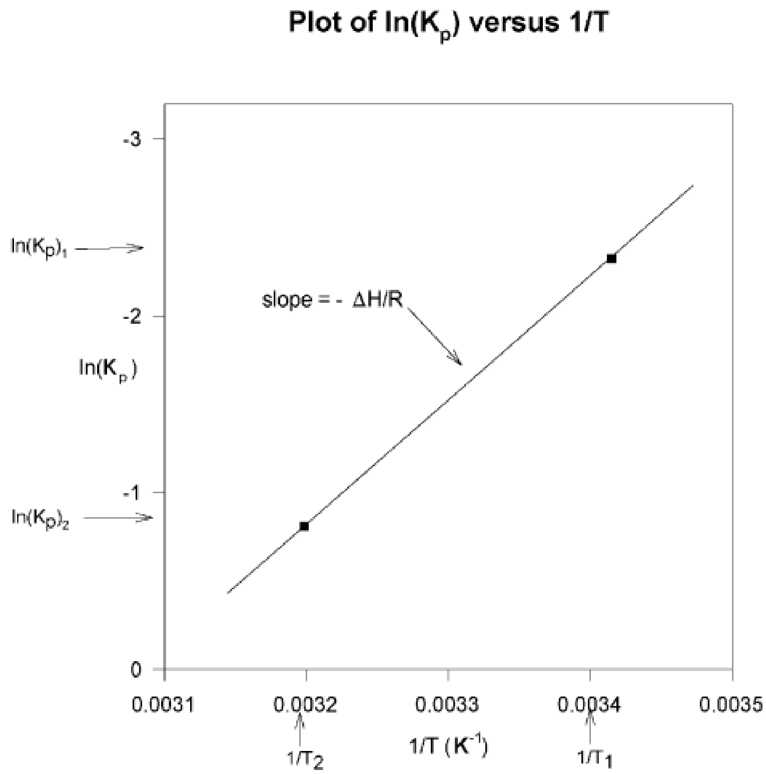
$$Q = 0.992\,\frac{L}{mole} < 2.46\,\frac{L}{mole} = K_c$$

Since Q < K_c, the reaction will go to the right to reestablish the equilibrium. Moles of PCl_3 and Cl_2 will be consumed and moles of PCl_5 will be produced. The moles of PCl_5 at the NEW equilibrium will be greater than the moles of PCl_5 in Example #2.

D) Dependency of K on Temperature

Plot of ln(K_p) versus 1/T for the reaction

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



Equation for a straight line is

$$y = (slope)x + (intercept)$$

$$\ln(K_p) = \left[\frac{-\Delta H}{R}\right]\left[\frac{1}{T}\right] + C$$

where C is a constant.

Consider K_2 at T_2 and K_1 at T_1 for a specific reaction, then

$$\ln(K_p)_2 = \left[\frac{-\Delta H}{R}\right]\left[\frac{1}{T_2}\right] + C$$

$$-\left[\ln(K_p)_1 = \left[\frac{-\Delta H}{R}\right]\left[\frac{1}{T_1}\right] + C\right]$$

$$\ln(K_p)_2 - \ln(K_p)_1 = \left[\frac{-\Delta H}{R}\right]\left[\frac{1}{T_2} - \frac{1}{T_1}\right]$$

Rearranging

$$\mathbf{\ln(K_p)_2 = \left[\frac{\Delta H}{R}\right]\left[\frac{T_2 - T_1}{T_2 T_1}\right] + \ln(K_p)_1}$$

When $\Delta H > 0$, heat is absorbed and the reaction is said to be *endothermic*.
When $\Delta H < 0$, heat is evolved and the reaction is said to be *exothermic*.

Example: If K_p = 0.09823 at 20°C and 0.4554 at 40°C, calculate ΔH for the reaction.

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

$$\ln(K_p)_2 = \left[\frac{\Delta H}{R}\right]\left[\frac{T_2 - T_1}{T_2 T_1}\right] + \ln(K_p)_1$$

Rearranging

$$\Delta H = R\ln\left[\frac{(K_p)_2}{(K_p)_1}\right]\left[\frac{T_2 T_1}{T_2 - T_1}\right] = \left(8.314\frac{J}{K}\right)\ln\left[\frac{(0.4554)}{(0.09823)}\right]\left[\frac{(313\,K)(293\,K)}{20\,K}\right] = 5.85\cdot 10^4\,J$$

E) Entropy and Free Energy

The thermodynamic function *entropy,* S, is a measure of the disorder or randomness of a system. *Entropy is a state function*.

The second law of thermodynamics states that an increase in entropy of the universe is associated with every spontaneous process. The change in entropy of the universe is the sum of the changes in the entropy of the system and surroundings.

$$\Delta S_{univ} \quad = \quad \Delta S_{sys} \; + \; \Delta S_{surr} \qquad (1)$$

For a process to proceed spontaneously ΔS_{univ} must be a positive number. If ΔS_{univ} is a negative number, then the process will proceed spontaneously in the reverse direction and when $\Delta S_{univ} = 0$, then the system is at equilibrium.

The change in the entropy of the surroundings, ΔS_{surr}, is dependent on the heat that is transferred between the system and the surroundings. For a constant pressure and constant temperature process

$$\Delta S_{surr} \quad = \quad \frac{q_{surr}}{T} \quad = \quad \frac{-(q_p)_{sys}}{T} \quad = \quad \frac{-\Delta H}{T} \qquad (2)$$

As an example consider the reaction

$$N_2\,(g) \;+\; 3\,H_2\,(g) \;\rightarrow\; 2\,NH_3\,(g) \qquad \mathbf{\Delta H^\circ = -\,92.4kJ,\quad \Delta S^\circ = -\,198.5\ J/K}$$

The change in the entropy of the surroundings with the formation of two moles of NH_3 is

$$\Delta S_{surr} \quad = \quad \frac{-\Delta H}{T} \quad = \quad \frac{-(-92.4\,kJ)\left[\frac{1000\,J}{1\,kJ}\right]}{298\,\boldsymbol{K}} \quad = \quad 310\,\frac{J}{\boldsymbol{K}}$$

and the change in the entropy of the universe is

$$\Delta S_{univ} \quad = \quad \Delta S_{sys} \;+\; \Delta S_{surr} \quad = \quad -198.5\,\frac{J}{\boldsymbol{K}} \;+\; 310\,\frac{J}{\boldsymbol{K}} \quad = \quad 111\,\frac{J}{\boldsymbol{K}}$$

The entropy of the system, $\Delta S^\circ = -\,198.5$ J/**K**, decreases primarily because there are fewer product molecules than reactant molecules. The reduction in number of molecules results in a lowering of the positional disorder. Although ΔS° for the reaction is a negative number, $\Delta S_{univ} > 0$ due to the large increase in the entropy of the surroundings and the reaction proceeds spontaneously to the right.

For a constant pressure and constant temperature process, eq 2 can be substituted into eq 1.

$$\Delta S_{univ} \quad = \quad \Delta S_{sys} \;+\; \frac{-\Delta H}{T} \quad = \quad \Delta S \;+\; \frac{-\Delta H}{T} \qquad (3)$$

Multiply eq 3 by - T and rearrange.

$$-T\Delta S_{univ} = \Delta H - T\Delta S \qquad (4)$$

Gibbs free energy, G, is defined as

$$G \equiv H - TS$$

For a constant pressure and constant temperature process

$$\boldsymbol{\Delta G = \Delta H - T\Delta S} \qquad \textbf{(5)}$$

Comparing eqs 4 and 5

$$\Delta G = -T\Delta S_{univ}$$

A constant pressure and constant temperature process will be spontaneous when $\Delta S_{univ} > 0$ or $\Delta G < 0$.

Example: Calculate $\Delta G°$ at 25°C for the following reaction?

$$H_2\,(g) \rightarrow 2\,H\,(g) \qquad \Delta H° = 436\text{ kJ}, \quad \Delta S° = 98.5\text{ J/}\mathbf{K}$$

$$\Delta G° = \Delta H° - T\Delta S° = 436\,kJ - (298\,\boldsymbol{K})(98.5\,\tfrac{J}{\boldsymbol{K}})\left|\frac{1\,kJ}{1000\,J}\right| = 407\,kJ$$

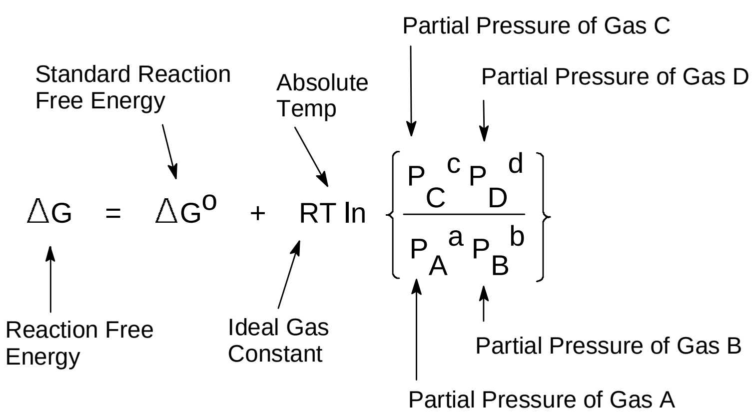
F) Reaction Free Energy

For the generic reaction

$$a\,A\,(g) + b\,B\,(g) \rightarrow c\,C\,(g) + d\,D\,(g)$$

the reaction free energy, ΔG, is defined as

$$\Delta G = \Delta G^{o} + RT\ln\left\{\frac{P_C^{\,c}\,P_D^{\,d}}{P_A^{\,a}\,P_B^{\,b}}\right\}$$



A plot of G for the above reaction versus the extent of reaction yields

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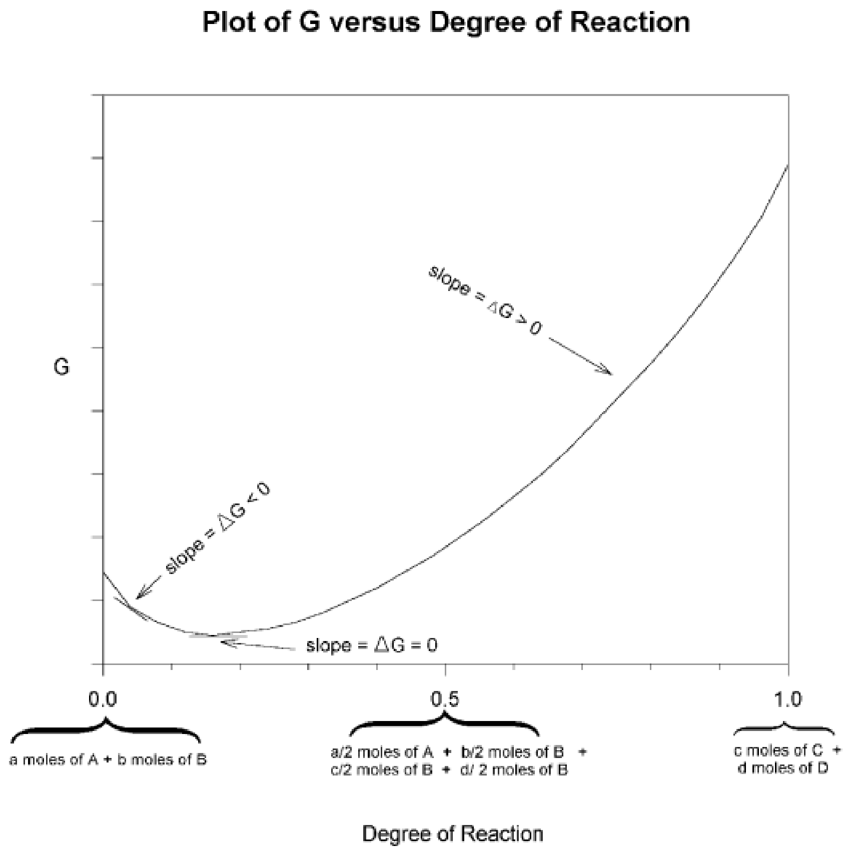

The *degree of reaction* is a measure of the extent to which the reaction has advanced from the left to right.

The slope of the above curve at any point is equal to the reaction free energy, ΔG. When the degree of reaction is 0, only the reactants A and B are present, $\Delta G < 0$, and the reaction will *proceed spontaneously to the right*. When the degree of reaction is 1.0, only the products are present, $\Delta G > 0$, and the reaction will *proceed spontaneously to the left*. When the degree of reaction is 0.16, $\Delta G = 0$ *and the reaction is at equilibrium*.

At equilibrium, ΔG = 0 and

$$\Delta G \quad = \quad 0 \quad = \quad \Delta G^\circ \; + \; RT \ln\left[\frac{(P_C)^c_{eq} \; (P_D)^d_{eq}}{(P_A)^a_{eq} \; (P_B)^b_{eq}}\right] \qquad (6)$$

K_p is defined as

$$K_p = \left[\frac{(P_C)^c_{eq}\ (P_D)^d_{eq}}{(P_A)^a_{eq}\ (P_B)^b_{eq}}\right] \qquad (7)$$

Substituting eq 7 into eq 6 and rearranging

$$\boldsymbol{\Delta G^\circ = -RT\ln(K_p)} \qquad \boldsymbol{or} \qquad \boldsymbol{K_p = \exp\left[\frac{-\Delta G^\circ}{RT}\right]}$$

Relationship between K_c and K_p

For the generic reaction

a A (g) + b B (g) → c C (g) + d D (g)

$$K_p = \left[\frac{(P_C)^c_{eq}\ (P_D)^d_{eq}}{(P_A)^a_{eq}\ (P_B)^b_{eq}}\right]$$

Assuming ideal gas behavior

$$P_A = \left[\frac{n_A}{V}\right]RT = [A]RT; \quad P_B = [B]RT; \quad P_C = [C]RT; \quad P_D = [D]RT$$

Substituting

$$K_p = \left[\frac{(P_C)^c_{eq}\ (P_D)^d_{eq}}{(P_A)^a_{eq}\ (P_B)^b_{eq}}\right] = \left[\frac{([C]RT)^c_{eq}\ ([D]RT)^d_{eq}}{([A]RT)^a_{eq}\ ([B]RT)^b_{eq}}\right] = \left[\frac{[C]^c_{eq}\ [D]^d_{eq}}{[A]^a_{eq}\ [B]^b_{eq}}\right](RT)^{c+d-a-b}$$

and

$$\boldsymbol{K_p = K_c(RT)^{\Delta n}}$$

where $\Delta n = c + d - a - b$.

G) Standard Reaction Free Energy and Standard Absolute Entropy

1) ΔG°

The free energy change for the formation of a compound in its standard state from its elements (most stable form) in their standard states is call the standard free energy of formation, ΔG°_f .

Example:

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g) \qquad \Delta G^\circ_f = -33.2 \text{ kJ}$$

By definition ΔG°_f for elements in their most stable form is zero.

A table of ΔG°_f is found in the text. See example below.

To calculate the standard reaction free energy, ΔG°, from the standard free energies of formation

$$\Delta G^\circ_{reaction} = \sum \nu_{product} \Delta G^\circ_f\,(product) - \sum \nu_{reactant} \Delta G^\circ_f\,(reactant)$$

where ν, called the *mole number*, is the stoichiometric coefficient in the thermochemical equation.

2) S°

The third law of thermodynamics states that the entropy for all pure, perfectly ordered, crystalline substances is zero at absolute zero(0 **K**). Consequently the absolute entropies, S°, of all compounds and elements in their most stable form can be determined. Some of these entropies are list in a table of thermodynamic data found in the text.

Example: Table of Selected Thermodynamic Data

Substance	**ΔG°_f (kJ/mole)**	**S° (J/mole-K)**
CO_2 (g)	- 394.4	213.6
H_2O (l)	- 237.2	70.0
H_2 (g)	0.0	130.5
H (g)	203.3	115

To calculate the change in the standard entropy for a reaction

$$\Delta S^\circ_{reaction} = \sum \nu_{product}\, S^\circ\,(product) - \sum \nu_{reactant}\, S^\circ\,(reactant)$$