This chapter is a review of the equilibrium state of a system that can undergo chemical reaction.

Operating reactors are not at chemical equilibrium, so why study this?

Find limits of reactor performance

Find operations or design changes that allow these restrictions to be changed and reactor performance improved.

Thermodynamic system

Variables: temperature, $T$, pressure, $P$, the number of moles of each component, $n_j$, $j = 1, \ldots, n_s$.

Specifying the temperature, pressure, and number of moles of each component then completely specifies the equilibrium state of the system.
The Gibbs energy of the system, $G$, is the convenient energy function of these state variables.

The difference in Gibbs energy between two different states

$$dG = -SdT + VdP + \sum_j \mu_j dn_j$$

$S$ is the system entropy,
$V$ is the system volume and
$\mu_j$ is the chemical potential of component $j$.

Consider a closed system. The $n_j$ can change only by the single chemical reaction,

$$\nu_1 A_1 + \nu_2 A_2 \rightleftharpoons \nu_3 A_3 + \nu_4 A_4 \quad \sum_j \nu_j A_j = 0$$

Reaction extent.

$$dn_j = \nu_j d\varepsilon$$

Gibbs energy.

$$dG = -SdT + VdP + \sum_j (\nu_j \mu_j) d\varepsilon$$

(3.2)

For the closed system, $G$ is only a function of $T$, $P$ and $\varepsilon$. 
Partial derivatives

\[ dG = -SdT + VdP + \sum_j (\nu_j \mu_j) \, d\varepsilon \]

\[ S = -\left( \frac{\partial G}{\partial T} \right)_{P,\varepsilon} \quad (3.3) \]

\[ V = \left( \frac{\partial G}{\partial P} \right)_{T,\varepsilon} \quad (3.4) \]

\[ \sum_j \nu_j \mu_j = \left( \frac{\partial G}{\partial \varepsilon} \right)_{T,P} \quad (3.5) \]

**G versus the reaction extent**

A necessary condition for the Gibbs energy to be a minimum

\[ \sum_j \nu_j \mu_j = 0 \quad (3.6) \]
Other forms: activity, fugacity

\[ \mu_j = G_j^\circ + RT \ln a_j \]

- \( a_j \) is the activity of component \( j \) in the mixture referenced to some standard state.
- \( G_j^\circ \) is the Gibbs energy of component \( j \) in the same standard state.
- The activity and fugacity of component \( j \) are related by
  \[ a_j = \frac{f_j}{f_j^\circ} \]
  - \( f_j \) is the fugacity of component \( j \)
  - \( f_j^\circ \) is the fugacity of component \( j \) in the standard state.

The Standard State

- The standard state is: pure component \( j \) at 1.0 atm pressure and the system temperature.
- \( G_j^\circ \) and \( f_j^\circ \) are therefore not functions of the system pressure or composition.
- \( G_j^\circ \) and \( f_j^\circ \) are strong functions of the system temperature.
Gibbs energy change of reaction

\[ \mu_j = G_j^\circ + RT \ln a_j \]

\[
\sum_j v_j \mu_j = \sum_j v_j G_j^\circ + RT \sum_j v_j \ln a_j \tag{3.8}
\]

The term \( \sum_j v_j G_j^\circ \) is known as the standard Gibbs energy change for the reaction, \( \Delta G^\circ \).

\[
\Delta G^\circ + RT \ln \prod_j a_j^{v_j} = 0 \tag{3.9}
\]

Equilibrium constant

\[ K = e^{-\Delta G^\circ / RT} \tag{3.10} \]

Another condition for chemical equilibrium

\[ K = \prod_j a_j^{v_j} \tag{3.11} \]

\( K \) also is a function of the system temperature, but not a function of the system pressure or composition.
Ideal gas equilibrium

The reaction of isobutane and linear butenes to branched C₈ hydrocarbons is used to synthesize high octane fuel additives.

\[
\text{isobutane} + 1\text{-butene} \rightleftharpoons 2,2,3\text{-trimethylpentane} \\
\text{I} + \text{B} \rightleftharpoons \text{P}
\]

Determine the equilibrium composition for this system at a pressure of 2.5 atm and temperature of 400K. The standard Gibbs energy change for this reaction at 400K is \(-3.72 \text{ kcal/mol} \) [5].

Solution

The fugacity of a component in an ideal-gas mixture is equal to its partial pressure,

\[
f_j = P_j = y_j P
\]  

(3.12)

\(f_j^\circ = 1.0 \text{ atm} \) because the partial pressure of a pure component \( j \) at 1.0 atm total pressure is 1.0 atm. The activity of component \( j \) is then simply

\[
a_j = \frac{P_j}{1 \text{ atm}}
\]  

(3.13)

\[
K = \frac{a_P}{a_Ia_B}
\]  

(3.14)
\[ K = e^{-\Delta G^\circ / RT} \]

\[ K = 108 \]

\[ K = \frac{p_P}{p_I p_B} = \frac{y_P}{y_I y_B} P \]

in which \( P \) is 2.5 atm. Three unknowns, one equation,

\[ \sum_j y_j = 1 \]

Three unknowns, two equations. What went wrong?

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**Ideal-gas equilibrium, revisited**

Additional information. The gas is contained in a closed vessel that is initially charged with an equimolar mixture of isobutane and butene. Let \( n_{j0} \) represent the unknown initial number of moles

\[ n_I = n_{I0} - \varepsilon \]

\[ n_B = n_{B0} - \varepsilon \]

\[ n_P = n_{P0} + \varepsilon \quad (3.16) \]

Summing Equations 3.16 produces

\[ n_T = n_{T0} - \varepsilon \]

in which \( n_T \) is the total number of moles in the vessel.
Moles to mole fractions

The total number of moles decreases with reaction extent because more moles are consumed than produced by the reaction. Dividing both sides of Equations 3.16 by \( n_T \) produces equations for the mole fractions in terms of the reaction extent,

\[
y_I = \frac{n_{I0} - \varepsilon}{n_{T0} - \varepsilon} \quad y_B = \frac{n_{B0} - \varepsilon}{n_{T0} - \varepsilon} \quad y_P = \frac{n_{P0} + \varepsilon}{n_{T0} - \varepsilon}
\]

Dividing top and bottom of the right-hand side of the previous equations by \( n_{T0} \) yields,

\[
y_I = \frac{y_{I0} - \varepsilon'}{1 - \varepsilon'} \quad y_B = \frac{y_{B0} - \varepsilon'}{1 - \varepsilon'} \quad y_P = \frac{y_{P0} + \varepsilon'}{1 - \varepsilon'}
\]

in which \( \varepsilon' = \varepsilon / n_{T0} \) is a dimensionless reaction extent that is scaled by the initial total number of moles.

One equation and one unknown

\[
K = \frac{(y_{P0} + \varepsilon')(1 - \varepsilon')}{(y_{B0} - \varepsilon')(y_{I0} - \varepsilon')}P
\]

\[
(y_{B0} - \varepsilon')(y_{I0} - \varepsilon')KP - (y_{P0} + \varepsilon')(1 - \varepsilon') = 0
\]

Quadratic in \( \varepsilon' \). Using the initial composition, \( y_{P0} = 0, y_{B0} = y_{I0} = 1/2 \) gives

\[
\varepsilon'^2(1 + KP) - \varepsilon'(1 + KP) + (1/4)KP = 0
\]

The two solutions are

\[
\varepsilon' = \frac{1 \pm \sqrt{1/(1 + KP)}}{2} \quad (3.19)
\]
Choosing the solution

The correct solution is chosen by considering the physical constraints that mole fractions must be positive. The negative sign is therefore chosen, and the solution is $\varepsilon' = 0.469$. The equilibrium mole fractions are then computed from Equation 3.18 giving

$$
\begin{align*}
\gamma_I &= 5.73 \times 10^{-2} \\
\gamma_B &= 5.73 \times 10^{-2} \\
\gamma_P &= 0.885
\end{align*}
$$

The equilibrium at 400K favors the product trimethylpentane.

Second derivative of $G$.

Please read the book for this discussion. I will skip over this in lecture.
Let’s calculate directly \( G(T, P, \varepsilon') \) and see what it looks like.

\[
G = \sum_j \mu_j n_j \tag{3.34}
\]

\[
\mu_j = G_j^\circ + RT \left[ \ln y_j + \ln P \right]
\]

\[
G = \sum_j n_j G_j^\circ + RT \sum_j n_j \left[ \ln y_j + \ln P \right] \tag{3.35}
\]

For this single reaction case, \( n_j = n_{j0} + \nu_j \varepsilon \), which gives

\[
\sum_j n_j G_j^\circ = \sum_j n_{j0} G_j^\circ + \varepsilon \Delta G^\circ \tag{3.36}
\]

Modified Gibbs energy

\[
\tilde{G}(T, P, \varepsilon') = \frac{G - \sum_j n_{j0} G_j^\circ}{n_{T0} RT} \tag{3.37}
\]

Substituting Equations 3.35 and 3.36 into Equation 3.37 gives

\[
\tilde{G} = \varepsilon' \frac{\Delta G^\circ}{RT} + \sum_j \frac{n_j}{n_{T0}} \left[ \ln y_j + \ln P \right] \tag{3.38}
\]

Expressing the mole fractions in terms of reaction extent gives

\[
\tilde{G} = -\varepsilon' \ln K + \sum_j (y_{j0} + \nu_j \varepsilon') \left[ \ln \left( \frac{(y_{j0} + \nu_j \varepsilon')}{1 + \nu \varepsilon'} \right) + \ln P \right]
\]
Final expression for the modified Gibbs Energy

\[ \tilde{G} = -\varepsilon' \ln K + (1 + \nu \varepsilon') \ln P + \sum_j (y_j + \nu_j \varepsilon') \ln \left( \frac{y_j + \nu_j \varepsilon'}{1 + \nu \varepsilon'} \right) \]  \hspace{1cm} (3.39)

- \( T \) and \( P \) are known values, so \( \tilde{G} \) is simply a shift of the \( G \) function up or down by a constant and then rescaling by the positive constant \( 1/(nT_0RT) \).
- The shape of the function \( \tilde{G} \) is the same as \( G \).
- The minimum with respect to \( \varepsilon' \) is at the same value of \( \varepsilon' \) for the two functions.

Minimum in \( G \) for an ideal gas

Goal: plot \( \tilde{G} \) for the example and find the minimum with respect to \( \varepsilon' \)

\[ I + B \rightleftharpoons P \]

For this stoichiometry: \( \sum_j \nu_j = \nu = -1 \).
Equimolar starting mixture: \( y_{P0} = 0, y_{I0} = y_{B0} = 0.5 \)

\[ \tilde{G}(T, P, \varepsilon') = -\varepsilon' \ln K(T) + (1 - \varepsilon') \ln P + \varepsilon' \ln(\varepsilon') + 2(0.5 - \varepsilon') \ln(0.5 - \varepsilon') - (1 - \varepsilon') \ln(1 - \varepsilon') \] \hspace{1cm} (1)
Recall that the range of physically significant $\varepsilon'$ values is $0 \leq \varepsilon' \leq 0.5$ and what do we see...

A closer look

Good agreement with the calculated value 0.469
The solution is a minimum, and the minimum is unique.
From Equation 3.39, for an ideal gas, the pressure enters directly in the Gibbs energy with the $\ln P$ term.

- Remake the plot for $P = 2.0$.
- Remake the plot for $P = 1.5$.
- How does the equilibrium composition change.
- Does this agree with Le Chatelier's principle?

For single liquid-phase or solid-phase systems, the effect of pressure on equilibrium is usually small, because the chemical potential of a component in a liquid-phase or solid-phase solution is usually a weak function of pressure.

The temperature effect on the Gibbs energy is contained in the $\ln K(T)$ term.

- This term often gives rise to a large effect of temperature on equilibrium.
- We turn our attention to the evaluation of this important temperature effect in the next section.
We usually calculate the standard Gibbs energy change for the reaction, $\Delta G^\circ$, by using the Gibbs energy of formation of the species. The standard state for the elements are usually the pure elements in their common form at 25°C and 1.0 atm.

$$G^\circ_{H_2O_f} = G^\circ_{H_2O} - G^\circ_{H_2} - \frac{1}{2} G^\circ_{O_2}$$

This gives the Gibbs energy change for the reaction at 25°C

$$\Delta G^\circ_i = \sum_j \nu_{ij} G^\circ_{jf} \quad (3.41)$$

Finding appropriate thermochemical data remains a significant challenge for solving realistic, industrial problems.

- Vendors offer a variety of commercial thermochemical databases to address this need.
- Many companies also maintain their own private thermochemical databases for compounds of special commercial interest to them.
- Design Institute for Physical Property Data (DIPPR) database. A web-based student version of the database provides students with access to data for 2000 common compounds at no charge: http://dippr.byu.edu/students/chemsearch.asp.
Temperature Dependence of the Standard Gibbs Energy

The standard state temperature 25°C is often not the system temperature. To convert to the system temperature, we need the temperature dependence of \( \Delta G^\circ \).

Recall from Equation 3.3 that the change of the Gibbs energy with temperature is the negative of the entropy,

\[
\left( \frac{\partial G}{\partial T} \right)_{p,n_j} = -S, \quad \left( \frac{\partial G^\circ_j}{\partial T} \right)_{p,n_j} = -S^\circ_j
\]

Summing with the stoichiometric coefficients gives

\[
\sum_j \frac{\partial (\nu_j G^\circ_j)}{\partial T} = \sum_j -\nu_j S^\circ_j
\]

Defining the term on the right-hand side to be the standard entropy change of reaction, \( \Delta S^\circ \) gives

\[
\frac{\partial \Delta G^\circ}{\partial T} = -\Delta S^\circ
\]

(3.43)

Let \( H \) denote the enthalpy and recall its connection to the Gibbs energy,

\[
G = H - TS \quad (3.44)
\]

Partial molar properties.

Recall the definition of a partial molar property is

\[
\bar{X}_j = \left( \frac{\partial X}{\partial n_j} \right)_{T,P,n_k}
\]

in which \( X \) is any extensive mixture property (\( U, H, A, G, V, S \), etc.).

\[
G^\circ_j = H^\circ_j - TS^\circ_j
\]

summing with the stoichiometric coefficient yields

\[
\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ
\]

(3.46)

\[
\frac{\partial \Delta G^\circ}{\partial T} = \frac{\Delta G^\circ - \Delta H^\circ}{T}
\]
van ’t Hoff equation

Rearranging this equation and division by $RT$ gives

$$\frac{1}{RT} \frac{\partial \Delta G^\circ}{\partial T} - \frac{\Delta G^\circ}{RT^2} = -\frac{\Delta H^\circ}{RT^2}$$

Using differentiation formulas, the left-hand side can be rewritten as

$$\frac{\partial \left( \frac{\Delta G^\circ}{RT} \right)}{\partial T} = -\frac{\Delta H^\circ}{RT^2}$$

which finally can be expressed in terms of the equilibrium constant

$$\frac{\partial \ln K}{\partial T} = \frac{\Delta H^\circ}{RT^2}$$

(3.47)

One further approximation

$$\int_{T_1}^{T_2} \frac{\partial \ln K}{\partial T} dT = \int_{T_1}^{T_2} \frac{\Delta H^\circ}{RT^2} dT$$

If $\Delta H^\circ$ is approximately constant

$$\ln \left( \frac{K_2}{K_1} \right) = -\frac{\Delta H^\circ}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$$

(3.49)
Consider a multicomponent, multiphase system that is at equilibrium and denote two of the phases as $\alpha$ and $\beta$. $T^k$ and $P^k$ are the temperature and pressure of phase $k$. $n^k_j$ is the number of moles of component $j$ in phase $k$. $\mu^k_j$ is the chemical potential of component $j$ in phase $k$. Phase equilibrium conditions:

$$T^\alpha = T^\beta$$
$$p^\alpha = p^\beta$$
$$\mu^\alpha_j = \mu^\beta_j, \quad j = 1, 2, \ldots, n_s \quad (3.50)$$

$$\mu_j = \mu_j^\circ + RT \ln f_j \quad (3.51)$$

If we express Equation 3.51 for two phases $\alpha$ and $\beta$ and equate their chemical potentials we deduce

$$f_j^\alpha = f_j^\beta, \quad j = 1, 2 \ldots, n_s$$

One can therefore use either the equality of chemical potentials or fugacities as the condition for equilibrium.

Gaseous Solutions

Let $f^G_j$ denote the fugacity of pure component $j$ in the gas phase at the mixture’s temperature and pressure.

The simplest mixing rule is the linear mixing rule

$$f^G_j = f^G_j y_j \quad \text{(ideal mixture)}$$

An ideal gas obeys this mixing rule and the fugacity of pure $j$ at the mixture’s $T$ and $P$ is the system’s pressure, $f^G_j = P$.

$$f^G_j = P y_j \quad \text{(ideal gas)}$$
The simplest mixing rule for liquid (and solid) mixtures is that the fugacity of component \( j \) in the mixture is the fugacity of pure \( j \) at the mixture’s temperature and pressure times the mole fraction of \( j \) in the mixture.

\[
f_L^j = f_L^{j^*} x_j
\]

This approximation is usually valid when the mole fraction of a component is near one.

In a two-component mixture, the Gibbs-Duhem relations imply that if the first component obeys the ideal mixture, then the second component follows Henry’s law

\[
f_L^2 = k_2 x_2
\]

in which \( k_2 \) is the Henry's law constant for the second component. Is \( k_2 = f_L^2 \)?

---

Fugacity pressure dependence.

For condensed phases, the fugacity is generally a weak function of pressure. See the notes for this derivation

\[
f_j \bigg|_{P_2} = f_j \bigg|_{P_1} \exp \left[ \frac{V_j (P_2 - P_1)}{RT} \right]
\]

The exponential term is called the Poynting correction factor. The Poynting correction may be neglected if the pressure does not vary by a large amount.
For gaseous mixtures, we define the fugacity coefficient, $\phi_j$

$$f_j^G = P_j \phi_j$$

The analogous correcting factor for the liquid phase is the activity coefficient, $\gamma_j$.

$$f_j^L = f_j^L x_j \gamma_j$$

These coefficients may be available in several forms. Correlations may exist for systems of interest or phase equilibrium data may be available from which the coefficients can be calculated [2, 3, 6, 4, 1].

Equilibrium Composition for Heterogeneous Reactions

We illustrate the calculation of chemical equilibrium when there are multiple phases as well as a chemical reaction taking place.

Consider the liquid-phase reaction

$$A(l) + B(l) \rightleftharpoons C(l)$$

that occurs in the following three-phase system.

Phase I: nonideal liquid mixture of A and C only. For illustration purposes, assume the activity coefficients are given by the simple Margules equation,

$$\ln \gamma_A = x_C^2 [A_{AC} + 2(A_{CA} - A_{AC})x_A]$$
$$\ln \gamma_C = x_A^2 [A_{CA} + 2(A_{AC} - A_{CA})x_C]$$

Phase II: pure liquid B.
Phase III: ideal-gas mixture of A, B and C.
All three phases are in intimate contact and we have the following data:

\[
A_{AC} = 1.4 \\
A_{CA} = 2.0 \\
P_A^o = 0.65 \text{ atm} \\
P_B^o = 0.50 \text{ atm} \\
P_C^o = 0.50 \text{ atm}
\]

in which \(P_j^o\) is the vapor pressure of component \(j\) at the system temperature.

1. Plot the partial pressures of A and C versus \(x_A\) for a vapor phase that is in equilibrium with only the A–C liquid phase. Compute the Henry’s law constants for A and C from the Margules equation. Sketch the meaning of Henry’s law on the plot and verify your calculation from the plot.

2. Use Henry’s law to calculate the composition of all three phases for \(K = 4.7\). What is the equilibrium pressure?

3. Repeat for \(K = 0.23\).

4. Assume \(K = 1\). Use the Margules equation to calculate the composition of all three phases.

5. Repeat 4 with an ideal mixture assumption and compare the results.
Equate the chemical potential in gas and liquid-phases. Since the gas phase is assumed an ideal-gas mixture:

\[ f^G_A = P_A \text{ gas phase}, \quad f^L_A = f^L_A x_A y_A \text{ liquid phase} \]

The fugacity of pure liquid A at the system temperature and the vapor pressure of A at the system temperature is known; it is simply the vapor pressure, \( P^\circ_A \).

If we neglect Poynting

\[ f^L_A = P^\circ_A, \quad P_A = P^\circ_A x_A y_A \quad (3.56) \]

The analogous expression is valid for \( P_C \).
Part 1.

Henry’s law for component A is

\[ f^L_A = k_A x_A, \quad f^L_A = f^L_A x_A \gamma_A \]

which is valid for \( x_A \) small.

\[ k_A = P^o_A \gamma_A \]

which is also valid for small \( x_A \). Computing \( \gamma_A \) from the Margules equation for \( x_A = 0 \) gives

\[ \gamma_A(0) = e^{A_{AC}} \]

So the Henry’s law constant for component A is

\[ k_A = P^o_A e^{A_{AC}} \]

The analogous expression holds for component C. Substituting in the values gives

\[ k_A = 2.6, \quad k_C = 3.7 \]

The slope of the tangent line to the \( P_A \) curve at \( x_A = 0 \) is equal to \( k_A \).

The negative of the slope of the tangent line to the \( P_C \) curve at \( x_A = 1 \) is equal to \( k_C \).
Part 2.

For $K = 4.7$, one expects a large value of the equilibrium constant to favor the formation of the product, C. We therefore assume that $x_A$ is small and Henry’s law is valid for component A.

The unknowns in the problem: $x_A$ and $x_C$ in the A–C mixture, $y_A$, $y_B$ and $y_C$ in the gas phase, $P$.

We require six equations for a well-posed problem: equate fugacities of each component in the gas and liquid phases, the mole fractions sum to one in the gas and A–C liquid phases. The chemical equilibrium provides the sixth equation.

\[ K = \frac{a_C}{a_A a_B} \]

\[ a^L_A = \frac{f^L_A}{f^o_A} = \frac{k_A x_A}{f^o_A} \]

$f^o_A$ is the fugacity of pure liquid A at the system temperature and 1.0 atm. Again, this value is unknown, but we do know that $P^o_A$ is the fugacity of pure liquid A at the system temperature and the vapor pressure of A at this temperature.

The difference between 0.65 and 1.0 atm is not large, so we assume $f^o_A = P^o_A$.

$x_C$ is assumed near one, so

\[ a^L_C = \frac{f^L_C}{f^o_C} = \frac{f^L_C x_C}{f^o_C} \]

Now $f^L_C$ and $f^o_C$ are the fugacities of pure liquid C at the system temperature and the system pressure and 1.0 atm, respectively.

If the system pressure turns out to be reasonably small, then it is a good assumption to assume these fugacities are equal giving,

\[ a^L_C = x_C \]

Since component B is in a pure liquid phase, the same reasoning leads to

\[ a^L_B = \frac{f^L_B}{f^o_B} = \frac{f^L_B}{f^o_B} = 1 \]
Substituting these activities into the reaction equilibrium condition gives

$$K = \frac{x_C}{x_A k_A P_A^o \cdot 1}$$  \hspace{1cm} (3.58)

Solving Equation 3.58 for $x_A$ yields

$$x_A = \left(1 + \frac{k_A K}{P_A^o}\right)^{-1}$$

$$x_C = \left(1 + \frac{P_A^o}{k_A K}\right)^{-1}$$

Substituting in the provided data gives

$x_A = 0.05$, $x_C = 0.95$

The assumption of Henry’s law for component A is reasonable.

The vapor compositions now are computed from the phase equilibrium conditions.

$$P_A = k_A x_A$$

$$P_B = P_B^o$$

$$P_C = P_C^o x_C$$

Substituting in the provided data gives

$P_A = 0.13$ atm, $P_B = 0.50$ atm, $P_C = 0.48$ atm

The system pressure is therefore $P = 1.11$ atm.

Finally, the vapor-phase concentrations can be computed from the ratios of partial pressures to total pressure,

$y_A = 0.12$, $y_B = 0.45$, $y_C = 0.43$
Part 3.

For $K = 0.23$ one expects the reactants to be favored so Henry’s law is assumed for component C. You are encouraged to work through the preceding development again for this situation. The answers are

\[ x_A = 0.97, \quad x_C = 0.03 \]

\[ y_A = 0.51, \quad y_B = 0.40, \quad y_C = 0.09 \]

\[ P = 1.24 \text{ atm} \]

Again the assumption of Henry’s law is justified and the system pressure is low.

Part 4.

For $K = 1$, we may not use Henry’s law for either A or C. In this case we must solve the reaction equilibrium condition using the Margules equation for the activity coefficients,

\[ K = \frac{x_C y'_C}{x_A y'_A} \]

Using $x_C = 1 - x_A$, we have one equation in one unknown,

\[ K = \frac{(1 - x_A) \exp \left[ x_A^2 (A_{CA} + 2(A_{AC} - A_{CA})(1 - x_A)) \right]}{x_A \exp \left[ (1 - x_A)^2 (A_{AC} + 2(A_{CA} - A_{AC})x_A) \right]} \]

(3.59)
Part 4.

Equation 3.59 can be solved numerically to give $x_A = 0.35$.

$$P_j = P^o_j x_j y_j, \quad j = A, C$$

The solution is

$x_A = 0.35, \quad x_C = 0.65$

$y_A = 0.36, \quad y_B = 0.37, \quad y_C = 0.28$

$P = 1.37$ atm

Part 5.

Finally, if one assumes that the A–C mixture is ideal, the equilibrium condition becomes

$$K = \frac{x_C}{x_A}$$

which can be solved to give $x_A = 1/(1 + K)$. For $K = 1$, the solution is

$x_A = 0.5, \quad x_C = 0.5$

$y_A = 0.30, \quad y_B = 0.47, \quad y_C = 0.23$

$P = 1.08$ atm

The ideal mixture assumption leads to significant error given the strong deviations from ideality shown in Figure 3.4.
Multiple Reactions

We again consider a single-phase system but allow \( n_r \) reactions

\[
\sum_j \nu_{ij} A_j = 0, \quad i = 1, 2, \ldots, n_r
\]

Let \( \varepsilon_i \) be the reaction extent for the \( i \)th reaction

\[
n_j = n_{j0} + \sum_i \nu_{ij} \varepsilon_i \quad (3.60)
\]

We can compute the change in Gibbs energy as before

\[
dG = -SdT + VdP + \sum_j \mu_j dn_j
\]

Using \( dn_j = \sum_i \nu_{ij} d\varepsilon_i \), gives

\[
dG = -SdT + VdP + \sum_j \mu_j \sum_i \nu_{ij} d\varepsilon_i
\]

\[
= -SdT + VdP + \sum_i \left( \sum_j \nu_{ij} \mu_j \right) d\varepsilon_i \quad (3.61)
\]

At constant \( T \) and \( P \), \( G \) is a minimum as a function of the \( n_r \) reaction extents. Necessary conditions are therefore

\[
\left( \frac{\partial G}{\partial \varepsilon_i} \right)_{T,P,\varepsilon_{i\neq i}} = 0, \quad i = 1, 2, \ldots, n_r
\]
Multiple Reactions

\[
\frac{\partial G}{\partial \varepsilon_i} = \sum_j \nu_{ij} \mu_j = 0
\]

Figure 3.6: Gibbs energy versus two reaction extents at constant \( T \) and \( P \).

Evaluating the partial derivatives in Equation 3.61 gives

\[
\sum_j \nu_{ij} \mu_j = 0, \quad i = 1, 2, \ldots, n_r
\]  
\[ (3.62) \]

Defining the standard Gibbs energy change for reaction \( i \), \( \Delta G_i^\circ = \sum_j \nu_{ij} G_j^\circ \) gives

\[
\sum_j \nu_{ij} \mu_j = \Delta G_i^\circ + RT \sum_j \nu_{ij} \ln a_j
\]

Finally, defining the equilibrium constant for reaction \( i \) as

\[
K_i = e^{-\Delta G_i^\circ / RT}
\]  
\[ (3.63) \]

allows one to express the reaction equilibrium condition as

\[
K_i = \prod_j a_j^{\nu_{ij}}, \quad i = 1, 2, \ldots, n_r
\]  
\[ (3.64) \]
In addition to the formation of 2,2,3-trimethylpentane, 2,2,4-trimethylpentane may also form

\[ \text{isobutane} + 1\text{-butene} \rightleftharpoons 2,2,4\text{-trimethylpentane} \]

Recalculate the equilibrium composition for this example given that \( \Delta G^\circ = -4.49 \text{ kcal/mol} \) for this reaction at 400K.

Let \( P_1 \) be 2,2,3 trimethylpentane, and \( P_2 \) be 2,2,4-trimethylpentane. From the Gibbs energy changes, we have

\[ K_1 = 108, \quad K_2 = 284 \]

The total number of moles is then \( n_T = n_{T0} - \varepsilon_1 - \varepsilon_2 \). Forming the mole fractions yields

\[
\begin{align*}
    y_I &= \frac{y_{I0} - \varepsilon'_1 - \varepsilon'_2}{1 - \varepsilon'_1 - \varepsilon'_2} \\
    y_B &= \frac{y_{B0} - \varepsilon'_1 - \varepsilon'_2}{1 - \varepsilon'_1 - \varepsilon'_2} \\
    y_{P1} &= \frac{y_{P10} + \varepsilon'_1}{1 - \varepsilon'_1 - \varepsilon'_2} \\
    y_{P2} &= \frac{y_{P20} + \varepsilon'_2}{1 - \varepsilon'_1 - \varepsilon'_2}
\end{align*}
\]

Applying Equation 3.64 to the two reactions gives

\[
\begin{align*}
    K_1 &= \frac{y_{P1}}{y_I y_B P} \\
    K_2 &= \frac{y_{P2}}{y_I y_B P}
\end{align*}
\]
Substituting in the mole fractions gives two equations for the two unknown reaction extents,

\[ PK_1 (y_{I0} - \varepsilon'_1 - \varepsilon'_2)(y_{B0} - \varepsilon'_1 - \varepsilon'_2) - (y_{P10} + \varepsilon'_1)(1 - \varepsilon'_1 - \varepsilon'_2) = 0 \]

\[ PK_2 (y_{I0} - \varepsilon'_1 - \varepsilon'_2)(y_{B0} - \varepsilon'_1 - \varepsilon'_2) - (y_{P20} + \varepsilon'_2)(1 - \varepsilon'_1 - \varepsilon'_2) = 0 \]

Initial condition: \( y_I = y_B = 0.5, y_{P1} = y_{P2} = 0. \)

Numerical solution

Using the initial guess: \( \varepsilon_1 = 0.469, \varepsilon_2 = 0 \), gives the solution

\( \varepsilon_1 = 0.133, \quad \varepsilon_2 = 0.351 \)

\( y_I = 0.031, \quad y_B = 0.031, \quad y_{P1} = 0.258, \quad y_{P2} = 0.680 \)

Notice we now produce considerably less 2,2,3-trimethylpentane in favor of the 2,2,4 isomer.

It is clear that one cannot allow the system to reach equilibrium and still hope to obtain a high yield of the desired product.
The other main approach to finding the reaction equilibrium is to minimize the Gibbs energy function. We start with

\[ G = \sum_j \mu_j n_j \]  

(3.67)

and express the chemical potential in terms of activity

\[ \mu_j = G_j^\circ + RT \ln a_j \]

We again use Equation 3.60 to track the change in mole numbers due to multiple reactions,

\[ n_j = n_{j0} + \sum_i \nu_{ij} \varepsilon_i \]

Expression for Gibbs energy

Using the two previous equations we have

\[ \mu_j n_j = n_{j0} G_j^\circ + G_j^\circ \sum_i \nu_{ij} \varepsilon_i + \left[ n_{j0} + \sum_i \nu_{ij} \varepsilon_i \right] RT \ln a_j \]  

(3.68)

It is convenient to define the same modified Gibbs energy function that we used in Equation 3.37

\[ \tilde{G}(T, P, \varepsilon_i') = \frac{G - \sum_j n_{j0} G_j^\circ}{n_{T0} RT} \]  

(3.69)

in which \( \varepsilon_i' = \varepsilon_i / n_{T0} \).

If we sum on \( j \) in Equation 3.68 and introduce this expression into Equations 3.67 and 3.69, we obtain

\[ \tilde{G} = \sum_i \varepsilon_i' \frac{\Delta G_i^\circ}{RT} + \sum_j \left[ y_{j0} + \sum_i \nu_{ij} \varepsilon_i' \right] \ln a_j \]
Expression for $\tilde{G}$

\[
\tilde{G} = - \sum_i \varepsilon'_i \ln K_i + \sum_j \left[ y_{j0} + \sum_i \nu_{ij} \varepsilon'_i \right] \ln a_j 
\]  

(3.70)

We minimize this modified Gibbs energy over the physically meaningful values of the $n_r$ extents.

The main restriction on these extents is, again, that they produce nonnegative mole numbers, or, if we wish to use intensive variables, nonnegative mole fractions. We can express these constraints as

\[-y_{j0} - \sum_i \nu_{ij} \varepsilon'_i \leq 0, \quad j = 1, \ldots, n_s\]  

(3.71)

Optimization problem

Our final statement, therefore, for finding the equilibrium composition for multiple reactions is to solve the optimization problem

\[
\min_{\varepsilon'_i} \tilde{G} \n\]

subject to Equation 3.71.
Multiple reactions with optimization

Revisit the two-reaction trimethylpentane example, and find the equilibrium composition by minimizing the Gibbs energy.

\[
a_j = \frac{P}{1 \text{ atm}} y_j \quad \text{(ideal-gas mixture)}
\]

Substituting this relation into Equation 3.70 and rearranging gives

\[
\tilde{G} = - \sum_i \varepsilon'_i \ln K_i + \left( 1 + \sum_i \nu_i \varepsilon'_i \right) \ln P
\]

\[
+ \sum_j \left( y_{j0} + \sum_i \nu_{ij} \varepsilon'_i \right) \ln \left[ \frac{y_{j0} + \sum_i \nu_{ij} \varepsilon'_i}{1 + \sum_i \nu_i \varepsilon'_i} \right]
\]

Constraints

The constraints on the extents are found from Equation 3.71. For this problem they are

\[-y_{f0} + \varepsilon'_1 + \varepsilon'_2 \leq 0 \quad -y_{b0} + \varepsilon'_1 + \varepsilon'_2 \leq 0 \quad -y_{p10} - \varepsilon'_1 \leq 0 \quad -y_{p20} - \varepsilon'_2 \leq 0\]

Substituting in the initial conditions gives the constraints

\[\varepsilon'_1 + \varepsilon'_2 \leq 0.5, \quad 0 \leq \varepsilon'_1, \quad 0 \leq \varepsilon'_2\]
We see that the minimum is unique.
The numerical solution of the optimization problem is

\[ \varepsilon'_1 = 0.133, \quad \varepsilon'_2 = 0.351, \quad \tilde{G} = -2.569 \]

The solution is in good agreement with the extents computed using the algebraic approach, and the Gibbs energy contours depicted in Figure 3.6.
The Gibbs energy is the convenient function for solving reaction equilibrium problems when the temperature and pressure are specified. The fundamental equilibrium condition is that the Gibbs energy is minimized. This fundamental condition leads to several conditions for equilibrium such as:

For a single reaction:

\[ \sum_j v_j \mu_j = 0 \]

\[ K = \prod_j a_j^{v_j} \]

For multiple reactions:

\[ \sum_j v_{ij} \mu_j = 0, \quad i = 1, \ldots, n_r \]

\[ K_i = \prod_j a_j^{v_{ij}}, \quad i = 1, \ldots, n_r \]

in which the equilibrium constant is defined to be

\[ K_i = e^{-\Delta G_i^\circ / RT} \]
You should feel free to use whichever formulation is most convenient for the problem.
The equilibrium “constant” is not so constant, because it depends on temperature via
\[
\frac{\partial \ln K}{\partial T} = \frac{\Delta H^\circ}{RT^2}
\]
or, if the enthalpy change does not vary with temperature,
\[
\ln \left( \frac{K_2}{K_1} \right) = -\frac{\Delta H^\circ}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)
\]

- The conditions for phase equilibrium were presented: equalities of temperature, pressure and chemical potential of each species in all phases.
- The evaluation of chemical potentials of mixtures was discussed, and the following methods and approximations were presented: ideal mixture, Henry’s law, and simple correlations for activity coefficients.
- When more than one reaction is considered, which is the usual situation faced in applications, we require numerical methods to find the equilibrium composition.
- Two approaches to this problem were presented. We either solve a set of nonlinear algebraic equations or solve a nonlinear optimization problem subject to constraints. If optimization software is available, the optimization approach is more powerful and provides more insight.
Notation I

- $a_j$: activity of species $j$
- $a_{jl}$: formula number for element $l$ in species $j$
- $A_j$: $j$th species in the reaction network
- $\overline{C}_{Pj}$: partial molar heat capacity of species $j$
- $E_l$: $l$th element constituting the species in the reaction network
- $f_j$: fugacity of species $j$
- $G$: Gibbs energy
- $\overline{G}_j$: partial molar Gibbs energy of species $j$
- $\Delta G^\circ_i$: standard Gibbs energy change for reaction $i$
- $H$: enthalpy
- $\overline{H}_j$: partial molar enthalpy of species $j$
- $\Delta H^\circ_i$: standard enthalpy change for reaction $i$
- $i$: reaction index, $i = 1, 2, \ldots, n_r$
- $j$: species index, $j = 1, 2, \ldots, n_s$
- $k$: phase index, $k = 1, 2, \ldots, n_p$
- $K$: equilibrium constant
- $K_i$: equilibrium constant for reaction $i$
- $l$: element index, $l = 1, 2, \ldots, n_e$
- $n_j$: moles of species $j$

Notation II

- $n_r$: total number of reactions in reaction network
- $n_s$: total number of species in reaction network
- $P$: pressure
- $P_j$: partial pressure of species $j$
- $R$: gas constant
- $S$: entropy
- $\overline{S}_j$: partial molar entropy of species $j$
- $T$: temperature
- $V$: volume
- $\overline{V}_j$: partial molar volume of species $j$
- $x_j$: mole fraction of liquid-phase species $j$
- $y_j$: mole fraction of gas-phase species $j$
- $z$: compressibility factor of the mixture
- $\gamma_j$: activity coefficient of species $j$ in a mixture
- $\varepsilon$: reaction extent
- $\varepsilon_i$: reaction extent for reaction $i$
- $\mu_j$: chemical potential of species $j$
- $\nu_{ij}$: stoichiometric number for the $j$th species in the $i$th reaction
- $\nu_j$: stoichiometric number for the $j$th species in a single reaction
- $\nu = \sum_j \nu_j$
\[ \nu_i \sum_j \nu_{ij} \]
\[ \phi_j \text{ fugacity coefficient of species } j \text{ in a mixture} \]

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