

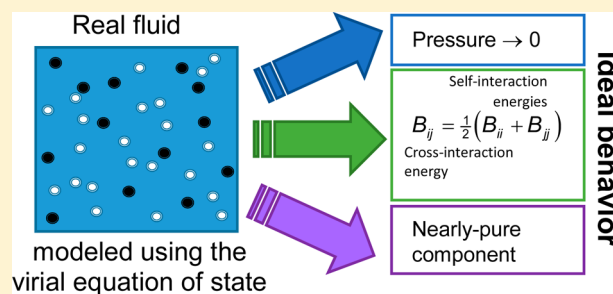
Improving Students' Understanding of the Connections between the Concepts of Real-Gas Mixtures, Gas Ideal-Solutions, and Perfect-Gas Mixtures

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ABSTRACT: In many textbooks of chemical-engineering thermodynamics, a gas mixture obeying the fundamental law $pV_m = RT$ is most often called ideal-gas mixture (in some rare cases, the term perfect-gas mixture can be found). These textbooks also define the fundamental concept of ideal solution which in theory, can be applied indifferently to liquid or gas mixtures but in practice, is nearly always introduced through the instance of liquid solutions. Undergraduate students are thus faced with different theoretical solution models that are all named “ideal” which can be a source of confusion. It is indeed often observed either that the connections between the concepts of perfect gas and ideal solution are missed by undergraduate students, or that they do not make any distinction between the concepts of gaseous ideal-solutions and perfect-gas mixtures. In this article, a simple example is proposed to pedagogically clarify all the misconceptions inherited from the multiple definitions of ideal mixtures and to explicate the connections between the various concepts in terms of molecular interactions. In particular, it is illustrated how an ideal solution (liquid or gas) can be derived from a real mixture (liquid or gas) by imposing constraints on molecular interactions. Similarly, it is shown how a perfect-gas mixture can be derived from a gas ideal-solution. At the same time, this example also highlights how the pressure and the composition influence the ideal character of a mixture.

KEYWORDS: Chemical Engineering, Thermodynamics, Graduate Education/Research, Upper-Division Undergraduate, Physical Chemistry, Analogies/Transfer, Misconceptions/Discrepant Events, Gases, Liquids



INTRODUCTION

In most of reference chemical-engineering thermodynamics textbooks,^{1–3} gases obeying the macroscopic law $pV_m = RT$ (with p , the pressure; V_m , the molar volume; T , the temperature; and R , the gas constant) are often named ideal gases and rarely perfect gases, while liquids obeying Raoult's law are called (liquid) ideal-solutions. Although ideal solutions are often introduced by considering the instance of liquid mixtures, it is sometimes mentioned that the concept of ideal solution is not limited to liquid solutions and is straightforwardly transferable to gaseous solutions. Doing so, it becomes necessary to define a gas ideal-solution (or something equivalent) in addition to the ideal-gas mixture (or perfect-gas mixture), both of them sharing a variety of common features but exhibiting differences also.

According to our teaching experience, the multiplicity of “ideal-mixture” definitions (ideal-gas/perfect-gas mixture, liquid ideal-solution, and gas ideal-solution) can be source of confusion for undergraduate students. As an illustration, let us cite the following sentence stemming from the didactic textbook by O'Connell and Haile⁴ which emphasizes the difficulty to juggle with the terminologies: “Gas-phase ideal solutions differ from ideal-gas mixtures” (page 430).

From a pedagogical viewpoint, it becomes necessary to explicate carefully all the relationships between the various concepts of ideal mixtures to avoid generalized confusion. It is indeed often observed that students do not make a clear distinction between the concepts of ideal-gas/perfect-gas mixture and gas ideal-solution and that few of them realize how the concepts of (liquid or gas) ideal solution and ideal-gas/perfect-gas mixture are interrelated.

The first issue is rather simple to solve since it is solely a matter of terminology: in our opinion, a gas obeying the macroscopic law $pV_m = RT$ should always be called perfect gas instead of ideal gas since, as described below, a perfect-gas mixture is a particular case of a gas ideal-solution. Note that the term *perfect gas* is rarely used in reference textbooks of thermodynamics, with a few exceptions. Among them, let us mention Rowlinson who wrote a comprehensive review entitled “The Perfect Gas”⁵ as well as Atkins and De Paula who stated:⁶ “although the term “ideal gas” is almost universally used in place of “perfect gas”, there are reasons for preferring the latter term. In an ideal system [...] the interactions between molecules

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in a mixture are the same. In a perfect gas, not only are the interactions all the same but they are in fact zero" (page 6).

Note incidentally that according to our bibliographic review, a few isolated authors—essentially from the mechanical-engineering community—define two categories of gases obeying the law $pV_m = RT$ called ideal gas and perfect gas depending on whether the heat capacities are temperature dependent or not.^{7,8} There is no doubt that such a practice is likely to add another degree of confusion in students' minds whereas there is no fundamental reason for introducing different names for so similar mixtures. Consequently, any gas obeying $pV_m = RT$ will be called perfect gas thereafter.

Once terminology issues are solved, the main pedagogical challenge remains: how to clarify the relationships between a perfect-gas mixture, a gas ideal-solution, and a real-gas solution in thermodynamics courses? It is worth noting that this issue is rarely discussed in the open literature dealing with education aspects (whereas articles dealing only with ideal solutions are much more frequent^{9–11}).

This article is intended to propose a pedagogical way to explicate the concepts of ideal mixtures to undergraduate students. It is shown how the aforementioned relationships can be easily highlighted through a practical example. The first sections are devoted to a brief review of the macroscopic and molecular interpretations of a perfect gas and an ideal solution. The example used to connect the various concepts is then discussed. It is worth noting that the proposed approach was successfully implemented during the thermodynamic classes at the ENSIC School (chemical-engineering department of the University of Lorraine, France), as mentioned in the subsequent sections.

MODEL FLUIDS

What is a Perfect Gas?

A perfect gas is macroscopically defined as a theoretical fluid obeying the equation of state $pV_m = RT$. From a molecular viewpoint, a perfect gas is defined as a fluid in which the individual gas particles have no interaction upon one another (neither attraction, nor repulsion) (see page 67 of ref 12). In other words, interaction energies A–A, B–B, and A–B in a perfect-gas mixture of A and B are all set to zero (see page 2 of ref 6). For illustration, Figure 1 compares the intermolecular potentials of a real fluid (continuous line) and a perfect gas (dashed line).

The perfect-gas law predicts that if a fixed quantity of gas is cooled at constant pressure, its molar volume tends to zero as the temperature tends to zero. Since lowering the temperature cannot destroy the molecular volume, the only possible way this condition can exist is for the molecular volume to be zero justifying why perfect-gas molecules are often considered as point masses. Alternatively, Elliott and Lira refer to the absence of repulsive interactions in a perfect-gas mixture to justify the point-mass behavior. They claim "Due to the lack of repulsive forces, ideal gas particles can "pass through" one another. Ideal gas molecules are sometimes called "point masses" to communicate this behavior" (see page 19 of ref 13).

In practice, a perfect gas is a limiting case of real-gas behavior given that any real fluid tends to behave as a perfect gas when its pressure tends to zero (it is recalled that at a fixed temperature and a zero pressure, the molecules of a fluid dispose of an infinite volume and consequently, interactions with other molecules become unlikely).

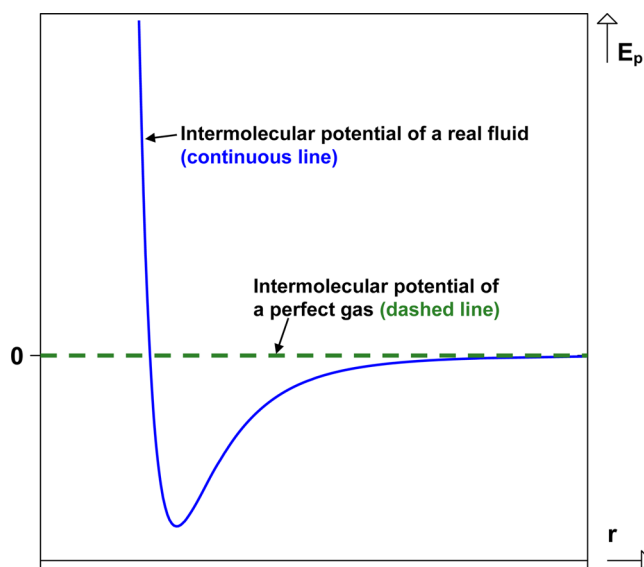


Figure 1. Intermolecular potential energy (E_p) as a function of the intermolecular distance (r) in real-fluid and perfect-gas systems.

The chemical potential and the fugacity of any component i of a perfect-gas mixture are

$$\begin{cases} \bar{g}_i^{\text{PG}}(T, p, \mathbf{y}) = g_{\text{pure } i}^{\text{PG}}(T, p) + RT \ln \gamma_i \\ \tilde{f}_i^{\text{PG}}(T, p, \mathbf{y}) = \gamma_i \cdot f_{\text{pure } i}^{\text{PG}}(T, p) = \gamma_i \cdot p \end{cases} \quad (1)$$

where \bar{g}_i and $g_{\text{pure } i}$ denote the chemical potential of mixed i and pure i , respectively (note that the chemical potential is also frequently denoted by the greek letter μ in chemical-engineering textbooks: $\bar{g}_i \equiv \mu_i$ and $g_{\text{pure } i} \equiv \mu_{\text{pure } i}$); \tilde{f}_i and $f_{\text{pure } i}$ are the fugacities of mixed i and pure i , respectively; the superscript pg stands for perfect gas. T , p , and \mathbf{y} are the temperature, pressure, and mole-fraction vector of component i , respectively.

What is an Ideal Solution?

Several definitions can be proposed for an ideal (liquid or gas) solution depending on the scale used to describe the matter (macroscopic or molecular). Historically, this concept has emerged from the proposal of the Raoult's law in the late 19th century: an ideal liquid solution in equilibrium with a perfect-gas mixture at fixed T and p obeys the relationship:

$$x_i \cdot p_i^{\text{sat}}(T) = \gamma_i \cdot p \quad (2)$$

with x_i and y_i , the mole fractions of component i in the liquid and gas phases, respectively; $p_i^{\text{sat}}(T)$ is the vapor pressure of pure i at fixed T . From a molecular viewpoint, an ideal solution A + B is often described as a mixture the intermolecular potential energy of which is independent of the assignment of the different molecules to the various locations (see page 803 in ref 14). In practice, an ideal solution is obtained "if the molecules of the various components are sufficiently alike from the point of view of the molecular interactions which they exert on one another, and from the point of view of their shapes and sizes" (see page 316 in Prigogine and Defay¹⁵) or in other words, if "the molecules have the same size and [if] the intermolecular forces between pairs of like molecules of each type, as well as between unlike molecules, are all the same" (see page 177 in ref 16). It is thus claimed that the pair interactions between like (A–A, B–B) and unlike (A–B) molecules must

be equal. According to modern theories of mixtures (e.g., the Kirkwood-Buff theory), this characterization of an ideal solution is more a corollary than a definition (in other words, this is a sufficient but not a necessary condition to observe ideal solutions). Strictly speaking, the structural definition of an ideal solution requires that the energy of A–B interactions (E_{A-B}) in the mixture is equal to the mean of the A–A and B–B interactions in the pure fluids (necessary and sufficient condition for the mixture to be ideal) (see page 168 in ref 6, page 802 in ref 14, and ref 17). That is, we must have

$$\langle E_{A-B} \rangle = \frac{\langle E_{A-A} + E_{B-B} \rangle}{2} = \frac{\langle E_{A-A} \rangle + \langle E_{B-B} \rangle}{2} \quad (3)$$

where the $\langle \rangle$ notation denotes the average over all molecular orientations and environments.

In practice, there is no dilemma between both approaches (i.e., equality of all pair interactions on one side; pair interaction between unlike molecules equal to the average of pair interactions between like molecules, on the other side) since ideal solutions are experimentally observed in mixtures containing closely similar molecules (e.g., isomeric mixtures).

Note that the general definition of an ideal solution clearly includes the structural definition of a perfect-gas mixture for which, the pair interactions A–A, B–B, and A–B are all equal to zero. As a consequence and as stated in the introduction, a perfect-gas mixture is a particular case of a gas ideal-solution.

Since the general structural definition does not mention the state of the fluid (temperature, pressure, or composition), an ideal solution must remain ideal regardless of the temperature, pressure, or composition domains investigated.^{3,18}

To conclude this section, let us mention that the chemical potential and fugacity of any component i in a (liquid or gas) ideal-solution are

$$\begin{cases} \bar{g}_{i,\phi}^{\text{id}}(T, p, \mathbf{z}) = g_{\text{pure } i,\phi}(T, p) + RT \ln z_i \\ \bar{f}_{i,\phi}^{\text{id}}(T, p, \mathbf{z}) = z_i f_{\text{pure } i,\phi}(T, p) \end{cases} \quad (4)$$

where the superscript id stands for ideal solution and \mathbf{z} is the mole fraction vector; ϕ denotes the aggregation state common to the solution and the pure component i . It can be immediately observed that eqs 1 and 4 are identical provided $\phi = \text{pg}$ thus highlighting that a perfect-gas mixture is an ideal solution (whereas the converse proposition is not true, as detailed below).

■ GAS IDEAL-SOLUTION VERSUS PERFECT-GAS MIXTURE: CLARIFICATION OF THESE CONCEPTS THROUGH A PRACTICAL EXAMPLE

The application case presented in this section introduces a simple model for the description of real-gas mixtures from which, a model for gas ideal-solution and the model of perfect-gas mixture can be derived by using appropriate assumptions on pair-interaction coefficients thus highlighting the relations between the three kinds of gaseous mixtures (real, ideal, and perfect). This practical example was successfully tested out during the present academic year with ENSIC School students: during the final examination of the chemical-engineering thermodynamics class, we observed that more than 80% of our students were able to elaborate on the differences and similarities between the concepts of real solutions, ideal

solutions, and perfect-gas mixtures whereas one year earlier, this ratio was less than 20%.

In this section, a simple model for real-gas mixtures is considered and makes it possible to:

- illustrate how a gas ideal-solution can be derived from a real-gas mixture from a molecular viewpoint,
- illustrate how a perfect-gas mixture can be derived from either a real-gas mixture or a gas ideal-solution.

A real (nonideal) gaseous N -component mixture is assumed to be well described by the truncated virial equation of state:

$$\begin{cases} V_{\text{m, gas mixture}}(T, p, \mathbf{y}) = \frac{RT}{p} + B_{\text{mixt}}(T, \mathbf{y}) \\ B_{\text{mixt}}(T, \mathbf{y}) = \sum_{i=1}^N \sum_{j=1}^N y_i y_j B_{ij}(T) \end{cases} \quad (5)$$

where $B_{ii}(T)$ denotes the temperature-dependent second virial coefficient of pure i (reflecting the self-interaction energies in pure i); B_{mixt} and $B_{ij} = B_{ji}$ are the mixture second virial coefficient and the so-called mutual second virial coefficient (this latter reflects the cross-interaction energies in a binary mixture of i and j), respectively.

For a binary system (1)+(2), by using the fact that $y_2 = 1 - y_1$, B_{mixt} can be written in the form:

$$\begin{cases} B_{\text{mixt}}(T, \mathbf{y}) = B_{11}y_1 + B_{22}y_2 + y_1y_2(2B_{12} - B_{11} - B_{22}) \\ = B_{11}y_1 + B_{22}y_2 + y_1y_2\delta_{12} \\ \text{Definition: } \delta_{12} = \delta_{21} = 2B_{12} - B_{11} - B_{22} \end{cases} \quad (6)$$

Consequently, the B_{mixt} coefficient aggregates the intermolecular interactions between like pairs i – i and unlike pairs i – j .

The fugacity of components 1 and 2 within the gaseous binary mixture is

$$\begin{cases} \ln \left[\frac{\tilde{f}_{1,\text{gas}}(T, p, \mathbf{y})}{p} \right] \\ = \ln y_1 + \frac{p}{RT} [B_{11}(T) + \delta_{12}y_2^2] \\ \ln \left[\frac{\tilde{f}_{2,\text{gas}}(T, p, \mathbf{y})}{p} \right] \\ = \ln y_2 + \frac{p}{RT} [B_{22}(T) + \delta_{12}y_1^2] \end{cases} \quad (7)$$

For a pure gaseous component, eqs 5 and 7 become:

$$\begin{cases} V_{\text{m, pure } i,\text{gas}}(T, p) = \frac{RT}{p} + B_{ii}(T) \\ \ln \left[\frac{f_{\text{pure } i,\text{gas}}(T, p)}{p} \right] = \frac{B_{ii}(T) \cdot p}{RT} \end{cases} \quad (8)$$

By combining eqs 7 and 8, one obtains:

$$\begin{cases} \frac{\tilde{f}_{1,\text{gas}}(T, p, \mathbf{y})}{f_{\text{pure } 1,\text{gas}}(T, p)} = y_1 \cdot \exp \left(\frac{p \cdot \delta_{12} \cdot y_2^2}{RT} \right) \\ \frac{\tilde{f}_{2,\text{gas}}(T, p, \mathbf{y})}{f_{\text{pure } 2,\text{gas}}(T, p)} = y_2 \cdot \exp \left(\frac{p \cdot \delta_{12} \cdot y_1^2}{RT} \right) \end{cases} \quad (9)$$

When dealing with liquid solutions, the ratio $\frac{\tilde{f}_{i,\text{liq}}(T, p, \mathbf{x})}{f_{\text{pure } i, \text{liq}}(T, p)}$ is usually called activity of component i (in the liquid phase). Symmetrically, the ratio $\frac{\tilde{f}_{i,\text{gas}}(T, p, \mathbf{y})}{f_{\text{pure } i, \text{gas}}(T, p)}$ could be named activity of species i in the gas phase, even though such a practice is rather unusual.

Remark: it can be also noted that the term $\exp\left(\frac{p \cdot \delta_{ij} \gamma_i^2}{RT}\right)$ plays the role of an activity coefficient of i (activity divided by mole fraction) denoted γ_i . Therefore, the expressions of the molar excess Gibbs energy, molar excess entropy, and molar excess enthalpy of the gas phase are

$$\left\{ \begin{array}{l} g_{\text{gas phase}}^E(T, p, \mathbf{y}) = RT(y_1 \ln \gamma_1 + y_2 \ln \gamma_2) \\ \quad = y_1 y_2 \cdot p \cdot \delta_{12}(T) \\ s_{\text{gas phase}}^E(T, p, \mathbf{y}) = -(\partial g_{\text{gas phase}}^E / \partial T)_{p, \mathbf{y}} \\ \quad = -y_1 y_2 \cdot p \cdot (d\delta_{12} / dT) \\ h_{\text{gas phase}}^E(T, p, \mathbf{y}) = g_{\text{gas phase}}^E(T, p, \mathbf{y}) \\ \quad + T \cdot s_{\text{gas phase}}^E(T, p, \mathbf{y}) \\ \quad = y_1 y_2 \cdot p [\delta_{12} - T \cdot (d\delta_{12} / dT)] \end{array} \right. \quad (10)$$

This derivation highlights that the virial equation of state is capable of predicting both entropic effects (reflected by the term $s_{\text{gas phase}}^E$) as well as enthalpic effects (reflected by the term $h_{\text{gas phase}}^E$). It can be observed that the temperature-dependent δ_{12} function plays a central role in the reproduction of such effects.

Component 1 will behave as in a gas ideal-solution provided its activity coefficient, i.e., the term $\exp\left(\frac{p \cdot \delta_{12} \gamma_2^2}{RT}\right)$, is unity. In other words, one out of these three conditions needs to be fulfilled:

- Case 1: the mole fraction of component 2 tends to zero or equivalently, the mole fraction of 1 tends to 1.
- Case 2: the interaction parameter δ_{12} is zero.
- Case 3: the pressure p tends to zero.

Analyses of these three occurrences are proposed in the next sections and supported by Figure 2.

It could be also claimed that the activity coefficient becomes unity at infinite temperature. This instance is addressed below, through a final remark before conclusion.

On the Ideal Behavior of Nearly-Pure Components in Real Fluids

As a well-known result of classical thermodynamics, when the mole fraction x_i of a given mixed component i approaches one, the component adopts a nearly-ideal behavior, even if the mixture is highly non-ideal. In terms of chemical potential and fugacity, one has:

$$\left\{ \begin{array}{l} \tilde{g}_{i,\phi}^{\text{real}}(T, p, \mathbf{x}) \underset{x_i \rightarrow 1}{\sim} \tilde{g}_{i,\phi}^{\text{id}}(T, p, \mathbf{x}) = g_{\text{pure } i, \phi}(T, p) \\ \quad + RT \ln x_i \\ \tilde{f}_{i,\phi}^{\text{real}}(T, p, \mathbf{x}) \underset{x_i \rightarrow 1}{\sim} \tilde{f}_{i,\phi}^{\text{id}}(T, p, \mathbf{x}) = x_i \cdot f_{\text{pure } i, \phi}(T, p) \end{array} \right. \quad (11)$$

This law can be simply related to the structural definition of an ideal solution: when $x_i \rightarrow 1$, each molecule i of a nonideal

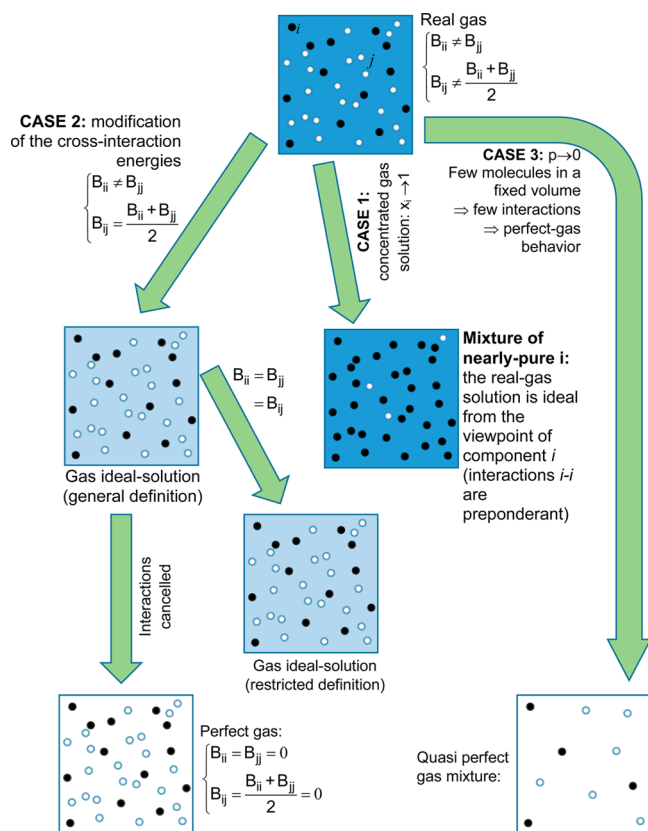


Figure 2. Graphical illustrations of the relationships between gas ideal-solutions, perfect gases and real-gas solutions.

mixture is mainly surrounded by neighbor molecules of the same type (molecules i). Even if molecules j are such that the cross-interaction energy i - j is not the average of the self-interaction energies i - i and j - j , molecules j are so rare that interaction energies i - i are preponderant and govern the expression of the chemical potential of i . Note however that molecules j , the mole fraction of which tends to zero, are essentially surrounded by molecules i and do not exhibit an ideal behavior.

As expected, Case 1 highlights that a nearly-pure component within a real-gas mixture behaves as in an gas-ideal solution (see Figure 2 for a graphical illustration of this instance):

$$\tilde{f}_{i,\text{gas}}^{\text{real}}(T, p, \mathbf{y}) \underset{y_i \rightarrow 1}{\sim} \tilde{f}_{i,\text{gas}}^{\text{id}}(T, p, \mathbf{y}) = y_i \cdot f_{\text{pure } i, \text{gas}}(T, p) \quad (12)$$

For a mixture made up of component 2 infinitely diluted in component 1, eq 9 highlights a nonideal behavior:

$$\frac{\tilde{f}_{2,\text{gas}}(T, p, \mathbf{y})}{f_{\text{pure } 2, \text{gas}}(T, p)} \underset{y_1 \rightarrow 1}{\sim} y_2 \cdot \exp\left(\frac{p \cdot \delta_{12}}{RT}\right)$$

Derivation of a Gas Ideal-Solution and a Perfect-Gas Mixture from a Real-Gas Mixture

Contrary to Case 1, Case 2 makes it possible to derive a general definition of a gas ideal-solution. Indeed, according to eq 6:

$$\delta_{12} = 0 \Leftrightarrow B_{12} = \frac{B_{11} + B_{22}}{2} \quad (13)$$

A gas ideal-solution is thus obtained for any value of T, p , and \mathbf{y} , as long as the mutual virial coefficient is equal to the average

of the second virial coefficients of the pure components. This conclusion fits perfectly the rigorous structural definition of an ideal solution. In such a case, the virial equation for gas-ideal solutions would have the following expression:

$$\begin{cases} V_{m,\text{gas mixture}}^{\text{id}}(T, p, \mathbf{y}) = \frac{RT}{p} + B_{\text{mixt}}^{\text{id}}(T, \mathbf{y}) \\ B_{\text{mixt}}^{\text{id}}(T, \mathbf{y}) = B_{11}(T)y_1 + B_{22}(T)y_2 \end{cases} \quad (14)$$

It is thus observed that the choice of a linear mixing rule for B_{mixt} induces a gas ideal-solution behavior (see Figure 2 for a graphical illustration of this instance).

From this general definition of a gas ideal-solution modeled with the truncated virial equation of state, the case of the perfect-gas mixture can be simply introduced by removing the pair interactions between like molecules $B_{11} = B_{22} = 0$. The pair interaction between unlike molecules is thus also null: $B_{12} = \frac{B_{11} + B_{22}}{2} = 0$. Eventually, it is found that a perfect-gas mixture is a particular instance of the much more general concept of gas ideal-solution.

Derivation of a Perfect-Gas Mixture from a Real-Gas Mixture

Case 3 simply illustrates that any real fluid tends to behave as a perfect gas at low-enough pressures (see Figure 2 for a graphical illustration of this instance; note that the real fluid under low pressure is named quasi-perfect gas). This result was expected as it is well-known that in any real fluid at low pressure and fixed temperature, molecules are dispersed in an infinite molar volume so that molecular interactions vanish and consequently, a perfect gas, which is a particular case of gas ideal-solution, is obtained. Thus, Case 3 makes it possible to generate an ideal behavior of the mixture on a limited pressure range (very low pressures) and consequently, cannot be used to define the gas ideal-solution concept which is supposed to hold at any pressure.

Remark on Real Solutions at High Temperature

The temperature-dependence of the second virial coefficient is well described in the literature. At sufficiently-low temperatures, the second virial coefficient is negative. As temperature increases, the second virial coefficient increases and becomes zero at the so-called Boyle temperature T_B . In such a case, "attractive and repulsive forces between pairs of molecules are approximately balanced".¹⁹ At temperatures above T_B , the second virial coefficient is positive and increases slowly. As a noticeable feature, experimental evidence has been given for some fluids that the second virial coefficient can reach a maximum (the corresponding temperature is named inversion temperature).²⁰ At high temperature, the second virial coefficient approaches zero as described by Beattie and Stockmayer: "At the Boyle temperature and according to theory, at infinite temperature [the second virial coefficient] limit is zero".²¹ As a consequence and following eq 6, it can be claimed that

$$B_{11}(T = +\infty) = B_{22}(T = +\infty) = B_{12}(T = +\infty) = 0 \Rightarrow \delta_{12} = 0 \quad (15)$$

At infinite temperature, a gas mixture obeying the truncated virial model exhibits thus the behavior of a perfect-gas mixture (since all the second virial coefficients approach zero) and consequently, of an ideal solution (since $\delta_{12} = 0$).

CONCLUSION

The idea to write the present article had emerged one year ago, after having asked our students, during the final examination of the chemical-engineering thermodynamics class, to discuss the differences and similarities between a gas ideal-solution and a perfect-gas mixture. Nearly none of them were able to identify the connections between these concepts. A pedagogical way is proposed in this article to clarify these issues. The truncated virial equation of state has been used to generate real-gas mixture properties. It has been shown that this approach makes it possible to intuitively develop the general concept of gas ideal-solution encompassing the particular case of perfect-gas mixtures.

It has been highlighted that a perfect-gas mixture is necessarily a gas ideal-solution whereas a gas ideal-solution is not necessarily a perfect-gas mixture. Consequently, it is proposed to definitively call perfect gas a gas obeying the law: $pV_m = RT$ (instead of ideal gas).

The relationships between real-gas (modeled using the virial equation of state), gas ideal-solution (modeled using the virial equation of state by assuming that the cross-second virial coefficients are related to the pure-component second virial coefficients through: $B_{ij} = 1/2(B_{ii} + B_{jj})$), and perfect-gas behaviors are summed up in Figure 2.

Such an approach was successfully implemented at the ENSIC School (chemical-engineering department of the University of Lorraine), making it possible to avoid confusion with the concepts of real solution, ideal solution, and perfect gases and to understand how these concepts are all interrelated.

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Notes

The authors declare no competing financial interest.

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