

Concept QuickStart – Vapour Pressure of Liquid Solutions

Unit: Unit 1: Solutions

Subject: For CBSE Class 12 Chemistry

Section 1: Understanding the Concept

The concept of vapour pressure is fundamental to understanding the physical properties of solutions. It serves as the gateway to explaining why solutions behave differently from pure solvents, especially during phase transitions like boiling and freezing. By grasping how the presence of a solute affects the tendency of a solvent to evaporate, we can unlock the principles behind several key phenomena in chemistry.

1.1 What Is Vapour Pressure of Liquid Solutions? (Core Idea and Anchor Definition)

At the simplest level, vapour pressure is the pressure exerted by gas molecules that have escaped from a liquid and are bouncing around in the space above it. When a solute is added to a pure liquid, this pressure changes.

Imagine a liquid surface in a closed container. Molecules are constantly jumping off the liquid into the gas phase (evaporation) and crashing back down (condensation). At equilibrium, these rates are balanced, and the resulting gas pressure is the vapour pressure. Now, if you add a non-volatile solute (one that does not evaporate easily), its molecules occupy space at the surface where solvent molecules would normally be. This blocks some of the solvent molecules from escaping. As a result, the total rate of evaporation drops, which in turn lowers the vapour pressure of the solution.

The pressure exerted by gas molecules in equilibrium with a liquid (or solid) at a given temperature; for solutions, Raoult's Law states that vapour pressure equals the mole fraction of solvent times the vapour pressure of pure solvent.

A common point of confusion is that adding a solute might increase the pressure. However, the opposite is true. Remember that non-volatile solute particles replace solvent molecules at the surface, so fewer solvent molecules can evaporate. This means that adding a non-volatile solute *decreases*, not increases, the vapour pressure of the solution.

1.2 Why Vapour Pressure Matters

Understanding vapour pressure is critical because it forms the basis of Raoult's Law, a key principle that predicts the behaviour of ideal solutions. More importantly, this concept is the foundation for explaining all colligative properties, including why adding salt to water raises its boiling point (boiling point elevation) and lowers its freezing point (freezing point depression). A solution boils only when its vapour pressure equals the external atmospheric

pressure; since a solute lowers the vapour pressure, a higher temperature is needed to reach that point.

From an exam perspective, vapour pressure is a foundational concept. Questions on this topic test your understanding of molecular dynamics and are often the first step in solving more complex problems related to colligative properties.

1.3 Why This Concept Exists

Chemists developed the concept of vapour pressure in solutions to solve a fundamental problem: explaining why solutions don't boil at the same temperature as pure solvents. Before this concept was understood, the observation that saltwater boils above 100°C was a mystery. Vapour pressure provides the logical explanation: a solution has a lower vapour pressure than a pure solvent at the same temperature, so it must be heated to a higher temperature for its vapour pressure to equal the atmospheric pressure and begin boiling.

This has direct practical consequences, such as explaining why adding salt to pasta water raises its boiling point, allowing food to cook slightly faster. Historically, it was the French chemist François-Marie Raoult who discovered that the reduction in a solution's vapour pressure is directly proportional to the amount (mole fraction) of solute added, a discovery that became the cornerstone of solution theory.

1.4 Analogies and Mental Image

To make this abstract concept more concrete, we can use a powerful analogy. Think of vapour pressure as **people leaving a crowded party**. In a closed room, people are constantly mingling (the liquid phase). Some decide to step outside for fresh air (evaporation). At equilibrium, the number of people leaving is balanced by the number of people returning. Now, imagine a large, non-moving obstacle is placed in the doorway, representing a solute particle. This obstacle blocks some people from leaving, reducing the flow rate out of the room. The "pressure" of people outside (the vapour pressure) drops because fewer are escaping.

The key components of this analogy map directly to the chemical concepts:

- **The Party** = The bulk liquid solvent.
- **People Leaving** = Solvent molecules evaporating from the surface.
- **The Obstacle** = A non-volatile solute particle occupying surface area.

Another way to think about it is the **evaporation from a wet cloth**. A cloth wet with pure water dries quickly because water molecules can freely escape from the surface. However, a cloth soaked in salt water dries much more slowly because the salt ions occupy surface "spots," reducing the rate of evaporation and thus lowering the vapour pressure.

Picture this... you are looking at a closed beaker of pure water. At the molecular level, water molecules at the surface are constantly vibrating, and some gain enough energy to escape into the gas phase above. At the same time, water molecules in the gas phase are colliding with the surface and condensing back into the liquid. An equilibrium is established where the rates of evaporation and condensation are equal, creating a steady vapour pressure. Now, dissolve some salt into the water. The surface is no longer composed entirely of water molecules; it is now dotted with non-volatile sodium and chloride ions. These ions act as barriers, physically blocking some of the surface area. This reduces the number of "escape routes" for water molecules, causing the rate of evaporation to decrease. To re-establish equilibrium, the pressure in the gas phase must also decrease until the rate of condensation once again matches the new, lower rate of evaporation.

This is what the vapour pressure of a solution looks like in your mind's eye.

1.5 Everyday Context and Applications

A direct, observable example is the difference between boiling pure water and boiling a salt solution. At sea level, pure water boils at exactly 100°C. If you add a significant amount of salt, you will observe that the solution needs to be heated to a higher temperature, perhaps 101-102°C, to achieve a vigorous boil. This visibly higher boiling temperature is a direct consequence of the salt lowering the water's vapour pressure. The solution must get hotter to make its vapour pressure equal to the surrounding atmospheric pressure.

A major real-world application of this principle is found in industrial desalination. Technologies like reverse osmosis fundamentally rely on the properties of solutions. In saltwater, the dissolved salt lowers the water's potential to move. To separate fresh water from the salt, extremely high pressure must be applied to the saltwater side of a semipermeable membrane. This pressure must be great enough to overcome the natural osmotic pressure, which itself is a direct result of the vapour pressure lowering caused by the solute.

Finally, it can be counterintuitive that adding "more stuff" (a solute) to a liquid *lowers* its pressure. You might think that adding more particles would make the system more active and increase its pressure. But actually, adding a non-volatile solute *decreases* the vapour pressure because it is generated only by the volatile solvent, not the non-volatile solute. The solute particles get in the way, occupying surface space and interfering with the solvent's ability to escape into the vapour phase. The more solute you add, the more interference there is, and the lower the vapour pressure becomes.

After understanding the concept of vapour pressure intuitively, it is important to see how this topic is formally presented in the official textbook curriculum.

Section 2: What the Textbook Says (NCERT)

This section distills the core principles, definitions, and examples related to the vapour pressure of liquid solutions as they are presented in the NCERT textbook for Class 12. Its purpose is to align your conceptual understanding with the official curriculum's language, formulas, and focus areas, ensuring you are prepared for board examinations.

2.1 NCERT Key Statements

Based on the NCERT textbook, the following key principles govern the vapour pressure of liquid solutions:

- **Raoult's Law for Volatile Liquids:** For a solution containing volatile components, the partial vapour pressure of each component (p_1) is directly proportional to its mole fraction (x_1) in the solution. This is expressed as $p_1 \propto x_1$, which leads to the formula $p_1 = p_1^0 x_1$, where p_1^0 is the vapour pressure of the pure component.
- **Total Vapour Pressure (Dalton's Law):** The total vapour pressure over the solution is the sum of the partial vapour pressures of all volatile components. For a binary solution, $p_{\text{total}} = p_1 + p_2$.
- **Linear Relationship:** The total vapour pressure over a solution of two volatile liquids varies linearly with the mole fraction of either component.
- **Composition of Vapour Phase:** At equilibrium, the vapour phase is always richer in the component that is more volatile (i.e., the one with the higher pure vapour pressure).
- **Effect of Non-Volatile Solute:** When a non-volatile solute is added to a solvent, the vapour pressure of the solution decreases. This lowering occurs because the solute molecules reduce the fraction of the surface area covered by solvent molecules, thereby reducing the rate of evaporation.
- **Raoult's Law and Henry's Law:** Raoult's Law can be considered a special case of Henry's Law where the proportionality constant, K_H , becomes equal to the vapour pressure of the pure component, p_1^0 .

2.2 NCERT Examples and Distinctions

The NCERT textbook uses the example of mixing chloroform (CHCl_3) and dichloromethane (CH_2Cl_2) to illustrate Raoult's Law for a solution of two volatile liquids. This example is important because it demonstrates the complete calculation process: finding the mole fractions of the components, calculating their partial pressures, summing them to find the total vapour pressure of the solution, and finally, determining the mole fractions of each component in the vapour phase.

The textbook also emphasizes several crucial distinctions that are key to understanding the topic fully:

- The difference between solutions containing two **volatile components** (where both contribute to the total vapour pressure) and solutions with a **non-volatile solute** (where only the solvent contributes).
- The distinction between the composition of the **liquid phase**, described by mole fraction x , and the composition of the **vapour phase**, described by mole fraction y .
- The relationship and comparison between **Raoult's Law**, which applies to ideal solutions of volatile liquids, and **Henry's Law**, which describes the solubility of gases in liquids.

After covering the core concepts and the official textbook facts, the focus now shifts to clarifying common points of confusion and providing tools to help you remember these principles for your exams.

Section 3: Clarity and Memory

This final section is designed as a toolkit for effective exam preparation. Having built a conceptual and textbook-based understanding, we will now reinforce that knowledge by addressing common pitfalls, clarifying key rules, and offering memorable ways to retain the essential principles of vapour pressure.

3.1 Key Clarity Lines

Here are several definitive statements to clarify potential points of confusion and provide you with exam-safe language:

1. Raoult's Law for a non-volatile solute is given by $P_{\text{solution}} = \chi_{\text{solvent}} \times P^{\circ}_{\text{solvent}}$. This means the solution's vapour pressure is simply the pure solvent's vapour pressure reduced by a factor equal to the solvent's mole fraction.
2. The vapour pressure of a solution containing a non-volatile solute is *always lower* than the vapour pressure of the pure solvent at the same temperature.
3. The lowering of vapour pressure is a colligative property; it depends on the *amount* (number of moles) of solute particles, not on their chemical identity.
4. Boiling point elevation and freezing point depression are direct *consequences* of this vapour pressure reduction. A liquid boils when its vapour pressure equals atmospheric pressure; since the solution's vapour pressure is lower, it requires a higher temperature to boil.
5. Raoult's Law, in its pure form, strictly applies only to ideal solutions, where solute-solvent interactions are identical to solvent-solvent and solute-solute interactions.

3.2 How to Remember Vapour Pressure

1. Mnemonic Use the mnemonic **R-O-L** to remember the three core ideas of vapour pressure lowering.

- **Raoult's Law:** This is the governing formula that quantifies the effect.
- **Only solvent evaporates:** The non-volatile solute stays behind and blocks the surface.
- **Lowers the pressure:** The solution's vapour pressure is always lower than the pure solvent's.

2. Memorable Phrase To capture the core logic in a single sentence, remember: "**Raoult's Rule: Less solvent, less evaporation, less vapor pressure.**" This phrase connects the cause (adding solute reduces the relative amount of solvent at the surface) to the effect (fewer molecules can evaporate, leading to lower pressure).

3. Physical Gesture Use your hands to create a physical memory anchor.

- **Pure Liquid:** Hold one hand flat, palm up. With your other hand, point all five fingers up from the flat hand and wiggle them quickly. This represents many solvent molecules escaping freely from the surface.
- **Solution:** Now, on your flat palm, curl three fingers down on your other hand to represent solute particles blocking the surface. Point only the remaining two fingers up and wiggle them slowly. This shows fewer molecules escaping at a lower rate. This physical action reinforces how a solute reduces evaporation.

4. Extreme Association Create a strong mental link for exams: getting Raoult's Law right is the key to mastering all colligative properties. Think of it as a logical chain reaction: **Solute → blocking evaporation → lower vapor pressure → higher boiling point → lower freezing point.** If you understand the first link in this chain (vapour pressure lowering), you can correctly solve problems related to all the subsequent effects. Master this one concept, and you master a major portion of the Solutions unit.

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