The solubility of gases depends on the pressure: an increase in pressure increases solubility, whereas a decrease in pressure decreases solubility. This statement is formalized in Henry’s Law, which states that the solubility of a gas in a liquid is directly proportional to the pressure of that gas above the surface of the solution. This can be expressed in the equation:

\[ C = k \times P_{\text{gas}} \]

---

### Introduction

<table>
<thead>
<tr>
<th>Variable</th>
<th>Represents</th>
</tr>
</thead>
<tbody>
<tr>
<td>( C )</td>
<td>the solubility of a gas in solvent</td>
</tr>
<tr>
<td>( k )</td>
<td>the proportionality constant</td>
</tr>
<tr>
<td>( P_{\text{gas}} )</td>
<td>the partial pressure of the gas above the solution</td>
</tr>
</tbody>
</table>

### Example 1: Using Henry’s Law

The aqueous solubility at 20 degrees Celsius of Ar at 1.00 atm is equivalent to 33.7 mL Ar(g), measured at STP, per liter of water. What is the molarity of Ar in water that is saturated with the air at 1.00 atm and 20 degrees Celsius? Air contains 0.934% Ar by volume. Assume that the volume if the water does not change when it becomes saturated with air.

**Solution**

STP molar volume: (22.414 L = 22,414 mL)

First Determine Molarity

\[
\begin{align*}
    k_{\text{Ar}} &= \frac{C}{P_{\text{Ar}}} \\
    &= \frac{\left( \frac{33.7 \text{ mL Ar}}{1 \text{ L}} \right) \left( \frac{1 \text{ mol Ar}}{22,414 \text{ mL}} \right)}{1 \text{ atm}} \\
    &= 0.00150 \text{ M atm}^{-1}
\end{align*}
\]

Solve for Concentration

\[
\begin{align*}
    C &= k_{\text{Ar}}P_{\text{Ar}} \\
    &= 0.00150 \text{ M atm}^{-1} \times 0.00934 \text{ atm} \\
    &= 1.40 \times 10^{-5} \text{ M Ar}
\end{align*}
\]

The example above illustrates the following:

1. (A) At low pressure, a gas has a low solubility. Decreased pressure allows more gas molecules to be present in the air, with very little being dissolved in solution.
2. (B) At high pressure, a gas has a high solubility. Increased pressure forces the gas molecules into the solution, relieving the pressure that is applied, causing there to be fewer gas molecules in the air and more of it in solution.

Common examples of pressure effects on gas solubility can be demonstrated with carbonated beverages, such as a bottle of soda (above). Once the pressure within the unopened bottle is released, CO₂(g) is released from the solution as bubbles or fizzing.

Deep Sea Divers and "The Bends"

In order for deep sea divers to breathe underwater, they must inhale highly compressed air in deep water, resulting in more nitrogen dissolving in their blood, tissues, and other joints. If the diver returns to the surface too rapidly, the nitrogen gas diffuses out of the blood too quickly and causes pain and possibly death. This condition is known as "the bends."

To prevent the bends, one can return to the surface slowly so that the gases will diffuse slowly and adjust to the partial decrease in pressure or breathe a mixture of compressed helium and oxygen gas, because helium is only one-fifth as soluble in blood than nitrogen. Think of a human body under water as a soda bottle under pressure. Imagine dropping the bottle and trying to open it. In order to prevent the soda from fizzing out, you open the cap slowly to let the pressure decrease. The atmosphere is approximately 78% nitrogen and 21% oxygen, but the body primarily uses the oxygen. Under water, however, the high pressure of water surrounding our bodies causes nitrogen to form in our blood and tissue. And like the bottle of soda, if the body moves or ascends to the water surface the water too quickly, the nitrogen is released too quickly and creates bubbles in the blood.

References

- external link: [http://www.elmhurst.edu/~chm/vchembook/174temppres.html](http://www.elmhurst.edu/~chm/vchembook/174temppres.html)
- More information on the bends can be found here:
  - external link: [http://www.rescuediver.org/med/bends.htm](http://www.rescuediver.org/med/bends.htm)

Contributors

- Michelle Hoang (UCD)