9.5: The Effect of Pressure on Solubility - Henry’s Law

Learning Objectives

- Describe how pressure affects concentration of a solute in a solution.

Pressure has very little effect on the solubility of solids or liquids, but has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the CO_2 molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more CO_2 is added to the space above the container, causing a pressure increase. In this case, more CO_2 molecules are in contact with the water and so more of them dissolve. Thus, the solubility increases as the pressure increases. As with a solid, the CO_2 that is undissolved reaches an equilibrium with the dissolved CO_2, represented by the equation:

$$CO_2\,(g) \rightleftharpoons CO_2\,(aq)$$

At equilibrium, the rate of gaseous CO_2 dissolution is equal to the rate of dissolved CO_2 coming out of the solution.

When carbonated beverages are packaged, they are done so under high CO_2 pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is open, the equilibrium is disrupted because the CO_2 pressure above the liquid decreases. Immediately, bubbles of CO_2 rapidly exit the solution and escape out of the top of the open bottle, see Figure $9.5.1$. The amount of dissolved CO_2 decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more CO_2 comes out of the liquid.

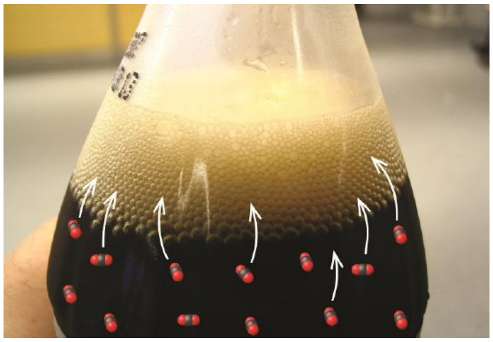

Figure $9.5.1$: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774-1836). **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. Henry's law can be written as follows:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} = k$$

C_1 and P_1 are the concentration and the pressure at an initial set of conditions; C_2 and P_2 are the concentration and pressure at another changed set of conditions; k is a constant at a constant temperature. The solubility of a gas is typically reported in g/L.

Example $9.5.1$

The solubility of a certain gas in water is 0.745 g/L at standard pressure. What is its solubility when the pressure above the solution is raised to 4.50 atm? The temperature is constant at 20^oC.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $C_1 = 0.745\text{ g/L}$

- $P_1 = 1.00\text{ atm}$
- $P_2 = 4.50\text{ atm}$

Unknown

Substitute into Henry's law and solve for C_2.

Step 2: Solve.

$$C_2 = \frac{C_1 \times P_2}{P_1} = \frac{0.745\text{ g/L} \times 4.50\text{ atm}}{1.00\text{ atm}} = 3.35\text{ g/L}$$

Step 3: Think about your result.

The solubility is increased to 4.5 times its original value, according to the direct relationship.

Exercise 9.5.1

Exposing a 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 20.26 kPa (152 torr) resulted in the dissolution of 1.45×10^{-3} g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 101.3 kPa (760 torr).

Answer

7.25×10^{-3} g in 100.0 mL or 0.0725 g/L

Case Study: Decompression Sickness ("The Bends")

Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing an increase of approximately 1 atm for each 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher ambient pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.

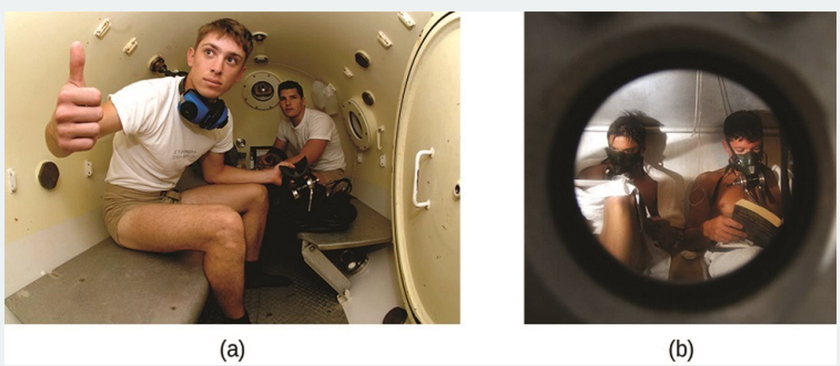

Figure 9.5.4: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.
Two photos are shown. The first shows two people seated in a steel chamber on benches that run length of the chamber on each side. The chamber has a couple of small circular windows and an open hatch-type door. One of the two people is giving a thumbs up gesture. The second image provides a view through a small, circular window. Inside the two people can be seen with masks over their mouths and noses. The people appear to be reading.

As the diver ascends to the surface of the water, the ambient pressure decreases and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that can cause a variety of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompression stops, pausing for several minutes at given depths during the ascent. When these preventive measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (Figure 9.5.4).