Oxygen and carbon dioxide are involved in the same biological processes in the ocean, but in opposite ways; photosynthesis consumes CO\(_2\) and produces O\(_2\), while respiration and decomposition consume O\(_2\) and produce CO\(_2\). Therefore it should not be surprising that oceanic CO\(_2\) profiles are essentially the opposite of dissolved oxygen profiles (Figure 5.5.1). At the surface, photosynthesis consumes CO\(_2\) so CO\(_2\) levels remain relatively low. In addition, organisms that utilize carbonate in their shells are common near the surface, further reducing the amount of dissolved CO\(_2\).

In deeper water, CO\(_2\) concentration increases as respiration exceeds photosynthesis, and decomposition of organic matter adds additional CO\(_2\) to the water. As with oxygen, there is often more CO\(_2\) at depth because cold bottom water holds more dissolved gases, and high pressures increase solubility. Deep water in the Pacific contains more CO\(_2\) than the Atlantic as the Pacific water is older and has accumulated more CO\(_2\) from the respiration of benthic organisms.
But the behavior of carbon dioxide in the ocean is more complex than the figure above would suggest. When CO₂ gas dissolves in the ocean, it interacts with the water to produce a number of different compounds according to the reaction below:

\[ \text{CO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3 \leftrightarrow \text{H}^+ + \text{HCO}_3^- \leftrightarrow 2\text{H}^+ + \text{CO}_3^{2-} \]

CO₂ reacts with water to produce carbonic acid (H₂CO₃), which then dissociates into bicarbonate (HCO₃⁻) and hydrogen ions (H⁺). The bicarbonate ions can further dissociate into carbonate (CO₃²⁻) and additional hydrogen ions (Figure 5.5.2).

**Figure 5.5.2** The fate of dissolved carbon dioxide in the oceans. Most of the carbon ends up in the form of bicarbonate (PW).

Most of the CO₂ dissolving or produced in the ocean is quickly converted to bicarbonate. Bicarbonate accounts for about 92% of the CO₂ dissolved in the ocean, and carbonate represents around 7%, so only about 1% remains as CO₂, and little gets absorbed back into the air. The rapid conversion of CO₂ into other forms prevents it from reaching equilibrium with the atmosphere, and in this way, water can hold 50-60 times as much CO₂ and its derivatives as the air.

**CO₂ and pH**
The equation above also illustrates carbon dioxide’s role as a buffer, regulating the pH of the ocean. Recall that pH reflects the acidity or basicity of a solution. The pH scale runs from 0-14, with 0 indicating a very strong acid, and 14 representing highly basic conditions. A solution with a pH of 7 is considered neutral, as is the case for pure water. The pH value is calculated as the negative logarithm of the hydrogen ion concentration according to the equation:

\[ \text{pH} = -\log_{10}[H^+] \]

Therefore, a high concentration of $H^+$ ions leads to a low pH and acidic condition, while a low $H^+$ concentration indicates a high pH and basic conditions. It should also be noted that pH is described on a logarithmic scale, so every one point change on the pH scale actually represents an order of magnitude (10 x) change in solution strength. So a pH of 6 is 10 times more acidic than a pH of 7, and a pH of 5 is 100 times (10 x 10) more acidic than a pH of 7.

Carbon dioxide and the other carbon compounds listed above play an important role in buffering the pH of the ocean. Currently, the average pH for the global ocean is about 8.1, meaning seawater is slightly basic. Because most of the inorganic carbon dissolved in the ocean exists in the form of bicarbonate, bicarbonate can respond to disturbances in pH by releasing or incorporating hydrogen ions into the various carbon compounds. If pH rises (low $[H^+]$), bicarbonate may dissociate into carbonate, and release more $H^+$ ions, thus lowering pH. Conversely, if pH gets too low (high $[H^+]$), bicarbonate and carbonate may incorporate some of those $H^+$ ions and produce bicarbonate, carbonic acid, or $CO_2$ to remove $H^+$ ions and raise the pH. By shuttling $H^+$ ions back and forth between the various compounds in this equation, the pH of the ocean is regulated and conditions remain favorable for life.

**CO₂ and Ocean Acidification**

In recent years there has been rising concern about the phenomenon of ocean acidification. As described in the processes above, the addition of CO₂ to seawater lowers the pH of the water. As anthropogenic sources of atmospheric CO₂ have increased since the Industrial Revolution, the oceans have been absorbing an increasing amount of CO₂, and researchers have documented a decline in ocean pH from about 8.2 to 8.1 in the last century. This may not appear to be much of a change, but remember that since pH is on a logarithmic scale, this decline represents a 30% increase in acidity. It should be noted that even at a pH of 8.1 the ocean is not actually acidic; the term “acidification” refers to the fact that the pH is becoming lower, i.e. the water is moving towards more acidic conditions.

Figure 5.5.3 presents data from observation stations in and around the Hawaiian Islands. As atmospheric levels of CO₂ have increased, the CO₂ content of the ocean water has also increased, leading to a reduction in seawater pH. Some models suggest that at the current rate of CO₂ addition to the atmosphere, by 2100 ocean pH may be further reduced to around 7.8, which would represent more than a 120% increase in ocean acidity since the Industrial Revolution.
Why is this important? Declining pH can impact many biological systems. Of particular concern are organisms that secrete calcium carbonate shells or skeletons, such as corals, shellfish, and many planktonic organisms. At lower pH levels, calcium carbonate dissolves, eroding the shells and skeletons of these organisms (Figure 5.5.4).

Not only does a declining pH lead to increased rates of dissolution of calcium carbonate, it also diminishes the amount of free carbonate ions in the water. The relative proportions of the different carbon compounds in seawater is dependent on pH (Figure 5.5.6). As pH declines, the amount of carbonate declines, so there is less available for organisms to incorporate into their shells and skeletons. So ocean acidification both dissolves existing shells and makes it harder for shell formation to occur.
Figure 5.5.6 Proportions of carbon compounds in the ocean at various pH levels. As the ocean pH declines, the proportion of carbonate ions also declines, reducing rates of shell formation (NOAA).

Additional links for more information: