Question #61133 – Chemistry – Physical chemistry

Answer:

A **chemical clock** or **oscillating reaction** is a complex mixture of reacting chemical compounds in which the concentration of one or more components exhibits periodic changes, or where sudden property changes occur after a predictable induction time.[1] They are a class of reactions that serve as an example of non-equilibrium thermodynamics, resulting in the establishment of a nonlinear oscillator. The reactions are theoretically important in that they show that chemical reactions do not have to be dominated by equilibrium thermodynamic behavior.

In cases where one of the reagents has a visible color, crossing a concentration threshold can lead to an abrupt color change in a reproducible time lapse. Examples of clock reactions are the Belousov-Zhabotinsky reaction, the Briggs-Rauscher reaction, the Bray-Liebhafsky reaction and the iodine clock reaction. The concentration of products and reactants of oscillatory chemical systems can be approximated in terms of dampedoscillations.

The **iodine clock reaction** (STd3) is a classical chemical clock demonstration experiment to display chemical kinetics in action; it was discovered by Hans Heinrich Landolt in 1886.[1] Two colourless solutions are mixed and at first there is no visible reaction. After a short time delay, the liquid suddenly turns to a shade of dark blue. The iodine clock reaction exists in several variations. In some variations, the solution will repeatedly cycle from colorless to blue and back to colorless, until the reagents are depleted.

Hydrogen peroxide variation[edit]

This reaction starts from a solution of hydrogen peroxide with sulfuric acid. To this is added a solution containing potassium iodide, sodium thiosulfate, and starch. There are two reactions occurring in the solution.

In the first, slow reaction, iodine is produced:

$H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O$

In the second, fast reaction, iodine is reconverted to 2 iodide ions by the thiosulfate:

$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

After some time the solution always changes color to a very dark blue, almost black.

When the solutions are mixed, the second reaction causes the triiodide ion to be consumed much faster than it is generated, and only a small amount of triiodide is present in thedynamic equilibrium. Once the thiosulfate ion has been exhausted, this reaction stops and the blue colour caused by the triiodide – starch complex appears.

Anything that accelerates the first reaction will shorten the time until the solution changes color. Decreasing the pH (increasing H^+ concentration), or increasing the concentration of iodide or hydrogen peroxide will shorten the time. Adding more thiosulfate will have the opposite effect; it will take longer for the blue colour to appear.