Opinion

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A Brief Note on Chemical Kinetics

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Description

Chemical kinetics, often known as reaction kinetics, is a branch of physical chemistry that studies the rates at which chemical reactions occur. Thermodynamics, on the other hand, is concerned with the direction in which a process happens but not with the rate of the process.

Chemical kinetics is the study of how experimental variables impact the rate of a chemical reaction, producing information on the reaction's mechanism and transition stages, as well as the development of mathematical models that can characterise the features of a chemical reaction.

Factors Affecting Reaction Rate

The rate of reaction varies based on the components involved. Acid/base reactions, salt production, and ion exchange are all examples of quick reactions. The reactions tend to be slower when covalent bonds are formed between molecules and when big molecules are created. The nature and strength of bonds in reactant molecules have a big impact on how quickly they turn into products.

Physical State

The physical state of a reactant (solid, liquid, or gas) is also a significant determinant in the rate of change. Thermal motion brings reactants into touch when they are in the same phase, as in an aqueous solution. The reaction is limited to the interface between the reactants when they are in distinct stages. Only at their point of contact, or, in the case of a liquid and a gas, at the liquid's surface, can they react. To complete the reaction, vigorous shaking and stirring may be required [1].

This indicates that the finer a solid or liquid reactant is divided, the more surface area per unit volume it has and the more contact it has with the other reactant, resulting in a faster reaction. To take an analogy, when starting a fire, one utilises wood chips and little branches rather than massive logs right away. On water reactions are the exception to the norm that homogeneous reactions occur faster than heterogeneous reactions in organic chemistry [2].

Surface Area of Solid State

Only the particles at the surface of a material can be engaged in a reaction. Because more particles are present at the surface when a solid is crushed into tiny pieces, the frequency of collisions between these and reactant particles increases, and reaction occurs more quickly.

Sherbet (powder), for example, is made out of very fine malic acid (a weak organic acid) powder and sodium hydrogen carbonate. When these molecules come into touch with saliva in the mouth, they immediately breakdown and react, producing carbon dioxide and causing the fizzy sensation [3].

In addition, fireworks makers alter the surface area of solid reactants to manipulate the pace at which the fuels in pyrotechnics are oxidised, resulting in a variety of effects. Finely split aluminium enclosed in a shell, for example, explodes forcefully. The reaction is slower when larger bits of aluminium are utilised, and sparks appear when pieces of burning metal are evacuated.

Concentration

Collisions of reactant species cause the reactions. The rate at which molecules or ions collide is determined by their concentrations. The molecules are more likely to clash and react with one another as they get more crowded [4].

As a result, increasing the concentrations of the reactants usually leads to an increase in the reaction rate, while decreasing the concentrations usually has the opposite effect. In pure oxygen, for example, combustion occurs more quickly than in air (21 percent oxygen).

Pressure

In a gaseous process, increasing the pressure increases the number of collisions between reactants, which speeds up the reaction. This is due to the fact that a gas's activity is proportional to its partial pressure. This is comparable to the impact of increasing a solution's concentration. In addition to the simple mass-action impact, the rate coefficients might change as a result of pressure. When an inert gas is added to a mixture of high-temperature gas-phase processes, the rate coefficients and products change; variations on this effect are known as fall-off and chemical activation. These phenomena arise when exothermic or endothermic reactions occur faster than heat transmission, resulting in non-thermal energy distributions in the reacting molecules. Increasing the pressure reduces this impact by increasing the heat transfer rate between the interacting molecules and the rest of the system [5].

Chemical reactions are described using mathematical models. Chemical engineers and chemists use kinetics to better comprehend and characterise chemical processes like food degradation, microorganism growth, stratospheric ozone decomposition, and biological system chemistry [6]. These models can also be used to improve product production, separate products more effectively and eliminate dangerous by-products in chemical reactors [7]. Kinetic models can be used to determine the temperature and pressure at which the largest yield of heavy hydrocarbons into gasoline will occur when catalytic cracking of heavy hydrocarbons into gasoline and light gas is performed, for example. Chemical Kinetics is regularly evaluated and investigated using specialist software to solve ordinary differential equations and fit curves [8]. In general, a reaction's free energy change (G) defines whether or not a chemical change will occur, but kinetics describes how fast the reaction occurs.

A reaction can be very exothermic and have a very positive entropy change, but if it is too slow, it will not happen in practice. In most cases, if a reactant may make two products, the thermodynamically more stable one will develop, unless the reaction is said to be under kinetic reaction control [9]. When estimating the product ratio for two reactants that are fast interconverting and each leading to a different product, the Curtin–Hammett principle is used. From free-energy relationships, it is possible to predict reaction rate constants for a reaction. Thermodynamics controls the extent to which reactions occur, whereas chemical kinetics is concerned with the rate of a chemical reaction [10].

Chemical equilibrium is reached in a reversible process when the forward and reverse reaction rates are identical (the principle of dynamic equilibrium) and the reactant and product concentrations do not change. The Haber–Bosch process for mixing nitrogen and hydrogen to make ammonia is an example of this. Component concentrations can oscillate for a long period before reaching equilibrium, as demonstrated by chemical clock reactions like the Belousov-Zhabotinsky reaction.

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