Chemical kinetics lab raport paper

Profession

FE 106 GENERAL CHEMISTRY EXPERIMENT-3 CHEMICAL KINETICS PREPARED BY BURAK COBAN PURPOSE: In this experiment we will study the rate of decomposition of hydrogen peroxide to form oxygen according to the net equation: $2H_2O_2$ (aq) $2H_2O$(l) + O_2 by measuring the rate at which oxygen evolved, we will investigate how the rate changes with varying initial concentrations of hydrogen peroxide and iodide catalyst. After we will study the affect of changing its concentration the rate oxygen evolution.

At the end of experiment we will summarize our results by attempting to write a rate law for the reaction, showing the defences on the concentrations of H_2O_2 and I

THEORY:

Chemical reactions can be fast (think of any explosion) or slow . It is very important to understand what is affecting the rate of the reaction and what is the mechanism of the reaction with such knowledge, we can often control a reaction to proceed at just the speed we need. we can thus avoid an explosion or speed up a reaction that seems too slow. In this chapter we start out by discussing rates of reactions and the rate law.

The rate law indicates the affect that the concentration of the reactants has on the reaction rate. In general, adding more of a reactant speeds things up (rather like pushing the gas pedal to put more gas into the car engine). But how much faster is the reaction if say the concentration of a reactant is doubled ? the rate law will help us answer such equations. Another way to affect the rate of a reaction is to change the temperature. We refrigeratefoodto slow the rate of bacterial metabolism that can cause food to spoil. If we want food to cook faster, we increase the heat.

We also look at why some reactions seem to need a “ kick start,” although once started, they continue to react. We will examine the activation barrier to reactions and its implications. Rate laws and kinetics experiments also tell us a lot about the mechanism by which a reaction occurs. Understanding the mechanism gives us another way to control the reaction. This knowledge enables us to design better catalyst or to create new compounds (such as new pharmaceuticals) that will be more effective. Differential Rate Laws: In many reactions, the rate of reactions changes as the reaction progresses.

Initially the rate of reaction is relatively large, while at very long times the rate of reaction decreases to zero. In order to characterize the kinetic behaviour of a reaction, it is desirable to determine how the rate of reaction varies as the reaction progresses. A rate law is a mathematical equation that describes the progress of the reaction. There are two forms of a rate law for chemical kinetics: the differential rate law and the integrated rate law. The differential rate law relates the rate of the reaction to the concentrations of the various species in the system.

Differential rate laws can take on many different forms, especially for complicated chemical reaction. However, most chemical reactions obey one of three differential rate laws. Each rate law contains a constant, k, called rate constant. The units for the rate constant depend upon the rate law, because the rate always has units of mole L-1 sec-1 and the concentration always has units of mole L-1. Zero – Order Reaction: For a zero order reaction, the rate of reaction is a constant. When the limiting reactant is completely consumed, the reaction abrupt stops. Differential rate law: R= k The rate constant. k , has units of mole L-1 sec-1. First - Order Reaction: For first order reaction, the rate of reaction is directly proportional to the concentration of ane of the reactants. Differential rate law: R = k[A] The rate constant, k, has units of sec-1. Second – Order Reaction: For a second reaction, the rate of reaction is directly proportional to the square of the concentration of one of the reactants. Differential rate law : R = k [A]2 The rate constant, k, has units of L-1 sec-1.

MATERIALS:

Funnel, Florence flask, Beaker, Pipette ? Thermometer, ring stand, ? Distilled water 0, 2M KI, H2O and H2O2 ? Burette, Rubber stopper, rubber tubes.

PROCEDURE:

Part A: ? 10 ml 0, 2M KI and 15 ml distilled water was taken the flask. ? Flask was swirled until the solution comes to the bath temperature. ? After that 5 ml % 3 H2O2 was added quickly and stopper the flask. ? One of us swirled the flask in the bath, other one observed the change of the volume. ? Other one recorded the time when approximately 2 ml of the gas was evolved. Part B: ? Same experiment was done by using; 10 ml of KI + 10 ml of H2O + 10 ml of H2O2 Part C: ? Same experiment was done by using; 20 ml of KI + 5 ml of H2O + 5 ml of H2O2

DISCUSSION: In this experiment, we discussed the rate of reactions. Reaction rate changed with kinds of reactant. For example in part A we put 10 ml KI and 15 ml distilled water an the other hand; in part B we put 10 ml KI , 10 ml H2O and 10 ml H2O2. After we determined. We saw that part A is slower than part B for this reason we can say rate is changed by nature of reactants. Another important effect is temperature. If temperature is high value reaction finish quickly. Maybe our results were effected temperature

Because we put flask in heat water and rate of reaction is faster than low temperature.

REFERENCES:

GENERAL CHEMISTRY:

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- Chemical kinetics is the study of the speeds, or rates, of the chemical reactions. A small number of factors control how fast a reaction will occur. Investigation of these factors provides clues to the ways in which reactants are transformed into products in chemical reactions. Some of the factors that influence the rate of a reaction are: ? Concentration; ? Pressure (particularly for reactions involving gases); ? Temperature; ? Surface area (for reactions involving solids); ? Catalyst;
- The rate of reaction; R= 1/3*[d[A] / dt] = -1/2 *[d[B] / dt] = 1/4* [d[C] / dt] = d[D] / dt . in this experiment we will investigate concentration of substance according to the rate of reaction. The rate of reactions are effected temperature, pressure of gases, concentrations and volume when one of products appears or one of the reactant is wed up. 4. 50 ml 2 MA 20 ml water added 30 ml 4 MB Initial conc. Of [A] = M= n/V n= 0, 05*2= 0, 1 mol A Initial conc. Of [B] M = n/V n= 0, 03*0, 12 mol B After mixing; V total = 100 ml = 0, 1 L Final conc. Of [A] M= 0, 1 / 0, 1 = 1M A Final conc. Of [B] M = 0, 12 / 0, 1 = 1, 2M B ----------------------- R= k R= k[A] R = k[A]2