Experiment No.: 11

Title: Investigation of the Order of the Reaction of Iodine with Acetone

Aim: To determine the order of the reaction of Iodine with Acetone by using titration method.

Introduction:

Chemical kinetics is the investigation of the rates at which chemical reactions occur, it is very important as it enables the exploration of reaction mechanism. Even the simplest chemical may consist of a complex sequence of events. Thus, it is customary practice to simplify matters by devising experiment in which, effectively, the concentration of one species only is changing.

In this experiment, reactants of different volumes are given, iodine stored in KI(aq) solution is first mixed with given volume of sulphuric acid, acetone and distilled water into a flask. Students are to be worked in pairs. Our group carried out experiment 2 (20cm^3 of Acetone) and only experiment 1, 2 and 3 were underwent in our class.

Group(Experiment):	1	2	3	4	5
Flask A:					
Volume of I_2 solution(aq)/cm^3	50.0	50.0	50.0	50.0	50.0
Flask B:					
Volume of H_2SO_4 soln/cm^3	25.0	25.0	25.0	25.0	25.0
Volume of Acetone soln/cm^3	25.0	20.0	15.0	12.5	6.25
Volume of distilled H_2O/cm^3	0	5.0	10.0	12.5	18.75

The flask of mixture is placed in a thermostat. At a fix time interval, portions of mixture are added to a flask containing $NaHCO_3$ solution. Since iodine reacted with acetone to give I^- ions, the solution is then titrated against sodium thiosulphate solution, the amount of iodine molecules remained after reaction with acetone are reacted with sodium thiosulphate to give I^- ions. At a fixed time interval, the procedure is repeated, thus the reaction rate can be determined by monitoring the volume of titre in each titration. In the experiment, the $NaHCO_3$ is used as a quenching material to quench the reaction of iodine with acetone.

$$CH_3COCH_3 + I_2 \rightarrow CH_3COCH_2I + H^+ + I^-$$

$$I_2\,(aq) + 2S_2O_3^{\,2-}\,(aq) \rightarrow 2I^-\,(aq) + S_4O_6^{\,2-}\,(aq)$$

Requirements:

- 0.02M I_2 dissolved in aqueous KI solution

1.0M acetone dissolved in water

- 1.0M sulphuric acid
- 0.5M sodium bicarbonate solution
- Very dilute sodium thiosulphate
- 25 cm3 pipette
- 10 cm3 pipette
- At least 5 clean, dry conical flasks
- Water trough
- Burette
- Measuring cylinder
- 100cm3 beakers

Procedures: (for pair 2)

1. 25.0cm^3 of sulphuric acid and 20.0 cm^3 of acetone solution were mixed together with 5.0 cm^3 of distilled water into a clean by using 25cm^3 pipette and 10 cm^3 pipette, dry conical flask, labeled flask B. The flask was stoppered and placed in a thermostat at 25℃ (water bath).
2. 50.0 cm^3 of iodine in KI solution was added in another clean, dry conical flask by using a 25cm3 pipette, labeled flask B.
3. The burette was filled with sodium thiosulphate solution.
4. 10 cm^3 portions of the NaHCO^3 solution was added into each of three conical flask.
5. The contents of flasks A and B were mixed thoroughly and a stopclock was started at once.
6. The flask containing the reaction mixture was reclamped in the thermostat.
7. The temperature of the mixture was checked carefully, and after about 5 minutes, 10 cm^3 of the reaction mixture was pipetted into one of the flask containing the NaHCO^3 solution, time at which the act was done was noted.
8. The contents of the flask were mixed thoroughly and the mixture was then titrated against sodium thiosulphate solution.
9. When the mixture was nearly colourless, two or three drops of a fresh starch solution were added and then the titration was continues until one drop of the sodium thiosulphate solution discharged the blue starch – iodine complex colour.
10. After about 10, 15, 20 and 30 minutes, further 10 cm3 portions of the reaction mixture were withdrew from the thermostat and the above procedure was carried out each time. All times and titration data were recorded.

Data and Analysis:

Graph1 (please refer to the graph paper attached)

The table below shows the volume of sodium thiosulphate solution used against time. (data for graph 1)

Time started/ min	6:00	12:30	18:00	21:00	24:00	27:00
Final reading/ cm^3	39.55	29.50	43.90	28.95	42.65	32.60
Initial reading/ cm^3	23.15	14.20	29.50	14.75	28.95	19.70
Titre/cm^3	16.40	15.30	14.40	14.20	13.70	12.90

Temperature readings remain constant during the experiment: 22.5℃.

Analysis:

From the table, it was noticed that the amount of sodium thiosulphate solution used is directly proportional to the concentration of the remaining iodine. The slope of graph 1 equals to the negative value of the rate of reaction, it implies that iodine concentration drops at a uniform rate. Therefore, rate of change of the iodine concentration is independent of iodine concentration. The reaction of iodine with acetone is **zero order** with respect to I_2.

Graph2 (please refer to the graph paper attached)

By combining the results of Group 1, 2 and 3 class results are obtained. The data is used to plot Graph II: rate of reaction against the volume of acetone added.

Class Results

Group	1	2	3
Slope/ $cm^3 min^{-1}$	0.279	0.200	0.180
		(0.139)	
Volume of acetone/ cm^3	25.0	20.0	15.0

Analysis:

Data marked with ( ) are abandoned because of its great deviation form the estimated value. With the rate of reaction increases with increasing volume of acetone together with a straight line passing through origin, the reaction of iodine with acetone is in **first order** with respect to acetone.

Volume of mixture left = 47 cm^3

Observations:

- Iodine solution was reddish brown in color while the other reactants are colourless.
- Gaseous bubbles were evolved when the reaction mixture is mixed with $NaHCO_3$. due to formation of carbon dioxide gas.
  $2NaHCO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$
- During the titration process, the colour of the mixture containing iodine changed from reddish brown to straw yellow gradually. Once the mixture reached the straw yellow colour, 2-3 drops of starch solution were added into the mixture. The colour turned to dark blue, the solution became colourless after several more drops of sodium thiosulphate were added.

Precautions:

- For mixing the content of flasks A and B, the solution should be transferred alternately between two flasks for 2 – 3 times. The solution is to swirled gently to allow the content inside to mix throughly.

Conclusion:

For a straight line passing through origin obtained in Graph II and a linear decreasing graph is in Graph I, the reaction of iodine with acetone was to be found in **first order** with respect to acetone and **zero order** with respect to I_2.

Discussion:

- The experiment cannot be conducted with more than one variable amounts reactions. For reaction more than one reactant, only the concentration of the reactant under investigated (acetone) changes, and the other reactants remain unchanged.
- **Quenching:**

  What is quenching?

  The analysis of the changes in concentration of reactants or products with time is direct measure of the rate of a reaction. Yet, since the process of analysis takes time, quenching is necessary to slow down the rate of reaction abruptly and assumed to have stopped.

  The methods of quenching of sample mixture include:

  - Rapid cooling by ice
  - Removing the catalyst

- Removing one of the reactants by adding another reagent which can use up the reaction rapidly
- Dilution with a large volume of water

In this experiment, The addition of $NaHCO_3$ is to neutralize the H_2SO_4 , with a view to removing the catalyst to lower the rate of reaction for upcoming titration process. Moreover, ice cubes were added to the reaction mixture to lower the temperature and concentration, thus minimize the reaction rate. The volume of $NaHCO_3$ added is unimportant. It should only be added in excess in order to remove all the catalyst H_2SO_4.

Chemical kinetics:

Chemical kinetics, also known as reaction kinetics, is the study of rates of chemical processes. Chemical kinetics includes investigations of how different experimental conditions can influence the speed of a chemical reaction and yield information about the reaction's mechanism and transition states, as well as the construction of mathematical models that can describe the characteristics of a chemical reaction. In 1864, Peter Waage and Cato Guldberg pioneered the development of chemical kinetics by formulating the law of mass action, which states that the speed of a chemical reaction is proportional to the quantity of the reacting substances.

Rate of reaction

Chemical kinetics deals with the experimental determination of reaction rates from which rate laws and rate constants are derived. Relatively simple rate laws exist for zero order reactions (for which reaction rates are independent of concentration), first order reactions, and second order reactions, and can be derived for others. In consecutive reactions the rate-determining step often determines the kinetics. In consecutive first order reactions, a steady state approximation can simplify the rate law. The activation energy for a reaction is experimentally determined through the Arrhenius equation and the Eyring equation. The main factors that influence the reaction rate include: the physical state of the reactants, the concentrations of the reactants, the temperature at which the reaction occurs, and whether or not any catalysts are present in the reaction.

Factors affecting reaction rate:

Nature of the Reactants

Depending upon what substances are reacting, the time varies. Acid reactions, the formation of salts, and ion exchange are fast reactions. When covalent bond formation takes place between the molecules and when large molecules are formed, the reactions tend to be very slow.

Physical State

The physical state (solid, liquid, or gas) of a reactant is also an important factor of the rate of change. When reactants are in the same phase, as in aqueous solution, thermal motion brings them into contact. However, when they are in different phases, the reaction is limited to the interface between the reactants. Reaction can only occur at their area of contact, in the case of a liquid and a gas, at the surface of the liquid. Vigorous shaking and stirring may be needed to bring the reaction to completion. This means that the more finely divided a solid or liquid reactant, the greater its surface area per unit volume, and the more contact it makes with the other reactant, thus the faster the reaction. To make an analogy, for example, when one starts a fire, one uses wood chips and small branches—one doesn't start with large logs right away. In organic chemistry On water reactions are the exception to the rule that homogeneous reactions take place faster than heterogeneous reactions.

Concentration

Concentration plays a very important role in reactions according to the collision theory of chemical reactions, because molecules must collide in order to react together. As the concentration of the reactants increases, the frequency of the molecules colliding increases, striking each other more frequently by being in closer contact at any given point in time. Think of two reactants being in a closed container. All the molecules contained within are colliding constantly. By increasing the amount of one or more of the reactants it causes these collisions to happen more often, increasing the reaction rate (Figure 1.1).

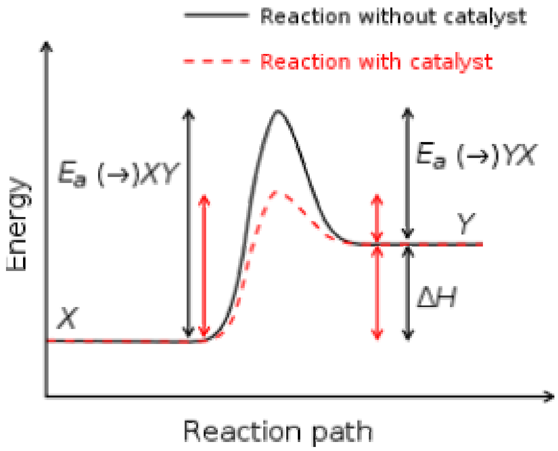
Temperature

Temperature usually has a major effect on the rate of a chemical reaction. Molecules at a higher temperature have more thermal energy. Although collision frequency is greater at higher temperatures, this alone contributes only a very small proportion to the increase in rate of reaction. Much more important is the fact that the proportion of reactant molecules with sufficient energy to react (energy greater than activation energy: $E > E_a$) is significantly higher and is explained in detail by the Maxwell-Boltzmann distribution of molecular energies.

The 'rule of thumb' that the rate of chemical reactions double for every 10 °C temperature rise is a common misconception. This may have been generalized from the special case of biological systems, where the Q10 (temperature coefficient) is often between 1.5 and 2.5.

A reaction's kinetics can also be studied with a temperature jump approach. This involves using a sharp rise in temperature and observing the relaxation rate of an equilibrium process.

Catalysts



Generic potential energy diagram showing the effect of a catalyst in an hypothetical exothermic chemical reaction. The presence of the catalyst opens a different reaction pathway (shown in red) with a lower activation energy. The final result and the overall thermodynamics are the same.

A catalyst is a substance that accelerates the rate of a chemical reaction but remains chemically unchanged afterwards. The catalyst increases rate reaction by providing a different reaction mechanism to occur with a lower activation energy. In autocatalysis a reaction product is itself a catalyst for that reaction leading to positive feedback. Proteins that act as catalysts in biochemical reactions are called enzymes. Michaelis-Menten kinetics describe the rate of enzyme mediated reactions.

In certain organic molecules specific substituents can have an influence on reaction rate in neighbouring group participation.

Agitating or mixing a solution will also accelerate the rate of a chemical reaction, as this gives the particles greater kinetic energy, increasing the number of collisions between reactants and therefore the possibility of successful collisions.

Increasing the pressure in a gaseous reaction will increase the number of collisions between reactants, increasing the rate of reaction. This is because the activity of a gas is directly proportional to the partial pressure of the gas. This is similar to the effect of increasing the concentration of a solution. A catalyst does not affect the position of the equilibria, as the catalyst speeds up the backward and forward reactions equally.

References:

- New Way Chemistry for Hong Kong A-Level Book 2 Ch.13-14 by Y.C. Wong & C.T.Wong
- http://en.wikipedia.org/wiki/Chemical_kinetics