The iodine clock reaction

Science, Chemistry

Chemical Kinetics: The Iodine-Clock Reaction: S2O82âˆ’(aq) + 2 Iâˆ’(aq) â†’ I2(aq) + 2 SO42âˆ’(aq) To measure the rate of this reaction we must measure the rate of concentration change of one of the reactants or products. Here, it is convenient to carry out a clock reaction involving the product I2. To do this, you will include (to the reacting S2O82âˆ’ and Iâˆ’) i) a small (but accurately known) amount of sodium thiosulfate, Na2S2O3, and ii) some starch indicator. The added Na2S2O3 does not interfere with the rate of above reaction, but it does consume the I2 as soon as it is formed (see below): 2 S2O32âˆ’(aq) + I2(aq) â†’ S4O62âˆ’(aq) + 2 Iâˆ’(aq) This reaction is much faster than the previous, so the conversion of I2 back to Iâˆ’ is essentially instantaneous. In a typical experiment, known concentrations of S2O82- and Iâˆ’ are mixed with a small amount of S2O32- and starch. As reaction the first proceeds, S2O82- and Iâˆ’ react to produce I2, but nothing appears to be happening because the second reaction is consuming the I2 as quickly as it is formed. The important moment is when the S2O32- runs out, for then the I2 is no longer consumed (see the second reaction), and the dark blue starchâˆ’I2 complex forms. The stoichiometry of the second reaction indicates that one I2 molecule must have been generated for every 2 ions of S2O32 âˆ’ initially present in the solution. [I2] generated(from first reaction) = [I2] consumed(in second reaction) = {1/2[S2O32-] initially present} Note that the change in concentration of I2 (before the blue complex forms) is fixed by amount of S2O32- included in the initial reaction mixture. 1 [S O2 âˆ’] d [I2] 2 2 3 rate = = dt t Where t = time for blue color to appear. In this experiment you will: A. Determine Rate Law and Rate Constant {k is constant at constant temperature} B. Vary temperature and estimate EA {Using the Arrhenius Equation} A. Determining the rate constant (k) and the rate law for the iodine-clock reaction: We know [S2O82-] and [I-], because we know initial volume and concentration of each solution and the final volume for each solution, (Remember M1V1 = M2V2) e. g for [S2O82-] in trial 1: [S2O82-] = (0. 100 M)(0. 00100L)/(0. 0100L) = 0. 0100 M and we know the general rate law for the iodine-clock reaction: rate = k[S2O82-]Î± [I-]Î² So we will: 1. Experimentally determine rate of the reaction (stopwatch, color change, [S2O32-]). 1 [S O2 â ˆ’] d [I2] 2 2 3 rate = = dt t 2. Empirically determine Î± and Î². log(rate) = log(k) + Î±log[S2O82-] + Î²log[I-] a. If temperature is constant, k is constant. If [I-] is constant, log[I-] is constant log(rate) = Î±log[S2O82-] + C y = m x + b -4. 50 -5. 00 -5. 50 -6. 00 -2. 20 (Where C = log(k) + Î²log[I-]) y = 1. 1704x - 3. 3137 R2 = 0. 995 -2. 00 -1. 80 -1. 60 log[S2O8=] -1. 40 For trials 1 -5, you should plot log(rate) vs Î±log[S2O82-] slope (rounded to the nearest integer) = Î± b. If temperature is constant, k is constant. If [S2O82-] is constant, log[S2O82-] log(rate) = Î²log[I-] + C y = m x +b (Where C= log(k) + Î±log[S2O82-]) For trials 6 - 10, you should plot log(rate) vs Î²log[I-]: -4. 50 -5. 00 -5. 50 -6. 00 -1. 80 y = 0. 9707x - 3. 9334 R2 = 0. 9835 -1. 60 -1. 40 log[I-] -1. 20 -1. 00 slope (rounded to the nearest integer) = Î² After Î± and Î² have been determined: 3. Solve for k @ constant temperature (°K) for trials 1 - 10 using the rates and Î± and Î² that you have determined from trials 1 - 10. (Find mean k, don't forget units units for k: M-1s-1) 4. Determine the rate law for the iodine-clock reaction (use mean k from #3: k) rate = k[S2O82-]Î± [I-]Î² B. Varying the temperature, measure rate of reaction, calculate k and estimate EA 5. Vary the temperature and solve for k (@ each temp). 6. Using k and T, construct an Arrhenius plot and solve for EA. Arrhenius Equation: ln k = âˆ’ Ea 1 ( )+ A R T k = rate constant EA = activation energy T = temperature (°K) R = Ideal Gas Constant (8. 314 J-mol-1-°K-1) A = reaction specific constant For trials 8, 11 - 14 you should plot ln(k) vs 1/T ln k vs 1/T -3 -3. 5 -4 -4. 5 -5 -5. 5 -6 -6. 5 -7 -7. 5 0. 0031 y = -6227. 7x + 15. 576 R = 0. 9965 2 0. 0032 0. 0033 0. 0034 1/T 0. 0035 0. 0036 0. 0037 slope = -EA/R In this example : -EA = (R)(-6227. 7 °K) EA = (8. 314 J-mol-1-°K-1)(6227. 7 °K) EA = 51777 J/mol or 51. 777 kJ/mol