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Chemistry 142 Practice Exam #1

Question 1. For H_2A, $K_{a1} = 2.0\times10^{-3}$ and $K_{a2} = 3.3\times10^{-8}$. For a 1.00*M* solution of H_2A calculate: $[H_2A]$, $[H^+]$, $[HA^-]$, and $[A^{2-}]$.

Question 2. A two step mechanism has been suggested for the reaction of nitric oxide, NO, and bromine, Br_2.

(1) $NO(g) + Br_2(g) \rightarrow NOBr_2(g)$

(2) $NOBr_2(g) + NO(g) \rightarrow 2\ NOBr(g)$

a) What is the overall reaction?

b) What is the role of $NOBr_2$ in this reaction

c) What is the predicted rate law if the first step is much slower than the first step?

d) The experimentally determined rate law is $R = k[NO]^2[Br_2]$. What can you conclude about this mechanism and the rate-determining step?

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Question 3. The following reaction has been studied at 540 K.

$$CO(g) + NO_2(g) \rightarrow CO_2(g) + NO(g).$$

The initial rates and initial concentrations are given in the table.

[CO] mol/L	[NO_2] mol/L	Initial Rate mol/Lhour
5.1×10^{-4}	0.35×10^{-4}	3.4×10^{-8}
5.1×10^{-4}	0.70×10^{-4}	6.8×10^{-8}
5.1×10^{-4}	0.18×10^{-4}	1.7×10^{-8}
1.0×10^{-3}	0.35×10^{-4}	6.8×10^{-8}
1.5×10^{-3}	0.35×10^{-4}	10.2×10^{-8}

Use the information to:
- Write out the rate expression
- Determine the order with respect to each reactant
- Calculate the value (including the units) of the rate constant and
- Write out the complete differential rate law.

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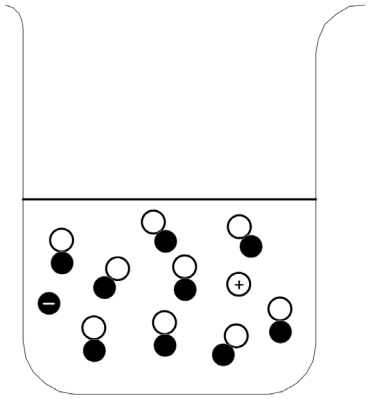
Question 4. The decomposition of SO_2Cl_2 is first order in SO_2Cl_2. The half life for the reaction is 4.1 hours at a certain temperature. If the initial amount of SO_2Cl_2 is 1.6×10^{-3} moles of SO_2Cl_2 in a flask, how many hours elapse until 2.00×10^{-4} moles remain? [Remember – the volume of the flask will not change.]

$$SO_2Cl_2(g) \rightarrow SO_2(g) + CO_2(g)$$

Question 5. What is the K_a of the monoprotic acid indicated by the diagram below?



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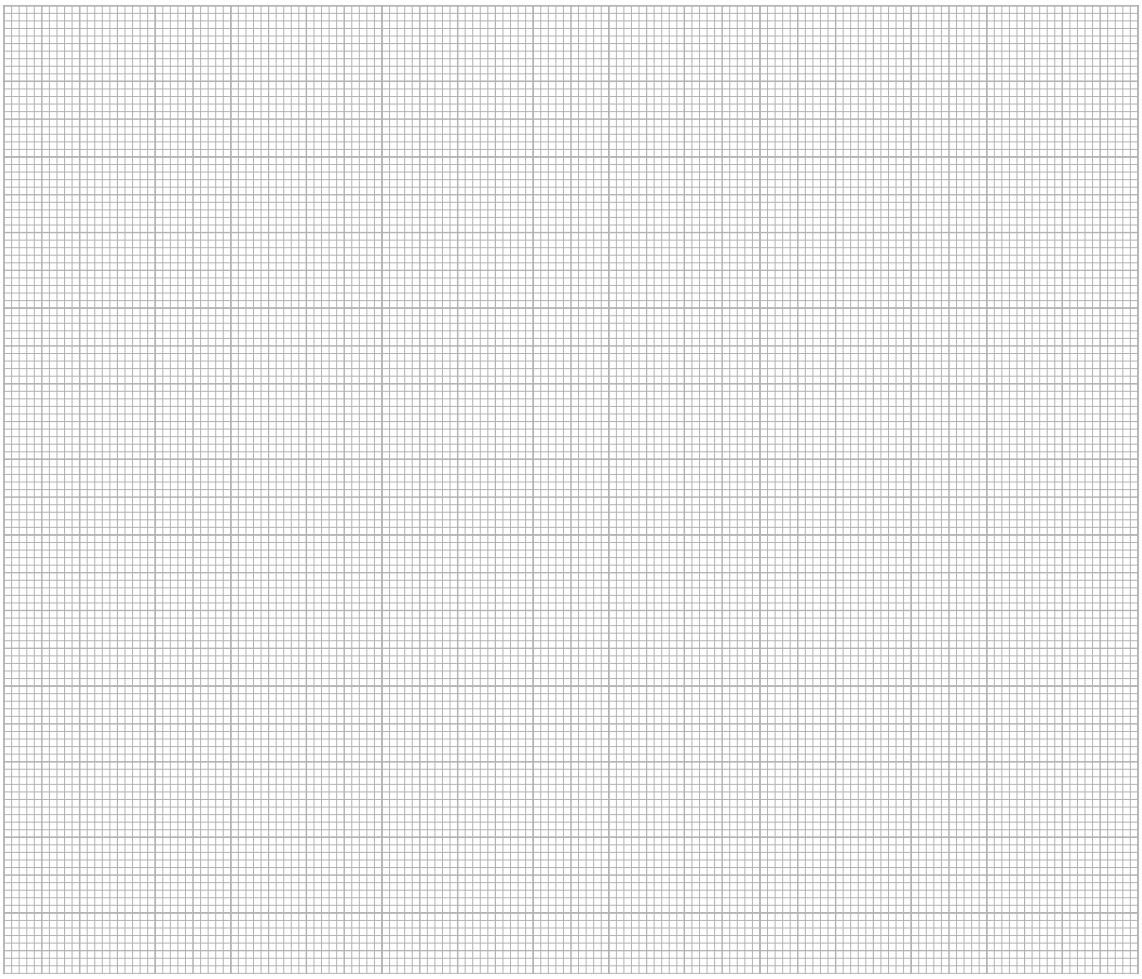
Question 6. The reaction shown in the equation is second order overall. The value of *k*, the specific rate constant, was measured at several temperatures. Determine the activation energy for this reaction by plotting your data. Estimate the linear fit with a ruler.

$$C_2H_5Br(aq) + OH^-(aq) \rightarrow C_2H_5OH(aq) + Br^-(aq)$$

°C	k, L mol^{-1}s^{-1}
25.1	8.76×10^{-5}
30.0	1.63×10^{-4}
34.9	2.80×10^{-4}
40.2	4.97×10^{-4}
45.2	8.63×10^{-4}
49.8	1.44×10^{-3}



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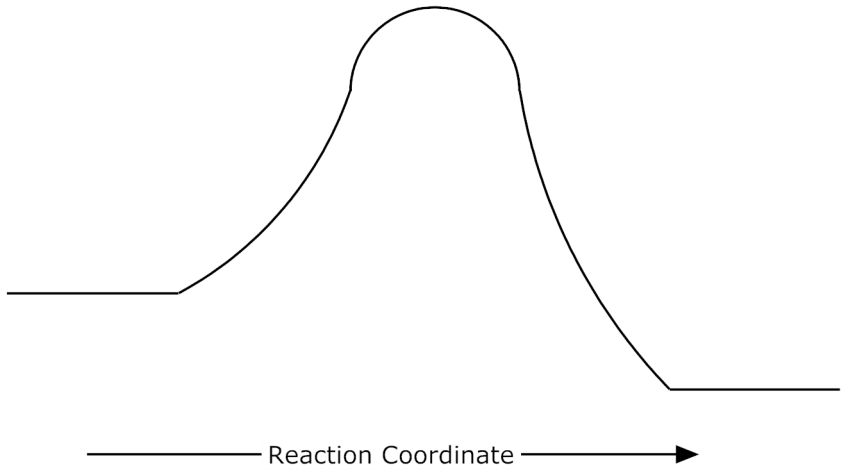
Question 7. Label the diagram below. Show reactants, products, transition state, and illustrate graphically, the activation energy. In this reaction an exothermic or endothermic reaction? Explain your reasoning.



Question 8. While spending the entire day in Room 30-252, Stacey began wondering what the pH of a 0.300*M* solution of oxalic acid is. Help her out and calculate the molarity of all species as well as the pH for a 0.300*M* solution of $H_2C_2O_4$. $K_{a1} = 5.9\times10^{-2}$, $K_{a2} = 6.4\times10^{-5}$.

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Question 9. Acetone is a useful solvent and is used in a variety of industrial processes. It decomposes via a first-order process to form methane and ketene. At 600 °C the rate constants is 8.7 x $10^{-3}s^{-1}$.

a) What is the half-life of the reaction?

e) How much time is required for 34% of a sample to decompose?

Question 10. What is the expected pH of a 0.200 *M* NH_4NO_3 solution? K_b for ammonia is 1.8×10^{-5}.

Question 11. A 1.00×10^{-3} *M* solution of quinine, a malaria drug, has a pH of 9.75. What is the value of K_b for quinine?

Question 12. Will the following salts alter the pH a of distilled water solution? If so, will they be acidic, basic or neutral? Be sure to offer an explanation for each.

a) $NaNO_3$

f) Na_3PO_4

g) $NH_4C_2H_3O_2$

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Kinetics

Concepts

1. How reaction rate depends on concentration, particle size, and temperature
2. The meaning of reaction rate in terms of changing concentrations over time
3. How the rate can be expressed in terms of reactant or product concentrations
4. The distinction between average and instantaneous rate and why the instantaneous rate changes during the reaction
5. The interpretation of reaction rate in terms of reactant and product concentrations
6. The experimental basis of the rate law and the information needed to determine it—initial rate data, reaction orders, and rate constant
7. The importance of reaction order in determining the rate
8. How reaction order is determined from initial rates at different concentrations
9. How integrated rate laws show the dependence of concentration on time
10. The meaning of half-life and why it is constant for a first-order reaction
11. Activation energy and the effect of temperature on the rate constant (Arrhenius equation)
12. Why concentrations are multiplied in the rate law
13. How temperature affects rate by influencing collision energy and, thus, the fraction of collisions exceeding the activation energy
14. Why molecular orientation and complexity influence the number of effective collisions and the rate
15. How the transition state represents the momentary species between reactants and products and how the activation energy is needed to form it
16. How an elementary step represents a single molecular event and its molecularity equals the number of colliding particles
17. How a reaction mechanism consists of several elementary steps, with the slowest step determining the overall rate
18. The criteria for a valid mechanism
19. How a catalyst speeds a reaction by lowering the activation energy
20. The distinction between homogeneous and heterogeneous catalysis

Required Skills

1. Calculating instantaneous rate from the slope of a tangent to a concentration vs. time plot
2. Expressing rate in terms of changes in concentration over time and calculating the instantaneous rate
3. Determining reaction order from a known rate law
4. Determining reaction order from changes in initial rate with concentration
5. Calculating the rate constant and its units
6. Using an integrated rate law to find concentration at a given time or the time to reach a given concentration
7. Determining reaction order graphically with a rearranged integrated rate law
8. Determining the half-life of a first-order reaction
9. Using the Arrhenius equation to calculate the activation energy or the rate constant
10. Using reaction energy diagrams to depict the energy changes during a reaction
11. Postulating a transition state for a simple reaction
12. Determining the molecularity of an elementary step
13. Constructing a mechanism with either a slow or a fast initial step

Equations

1. $\text{Rate} = -\frac{\Delta[\text{A}]}{\Delta t}$
2. $\text{Rate} = k[\text{A}]^m[\text{B}]^n\ldots$
3. $\ln\frac{[\text{A}]_0}{[\text{A}]_t} = kt$
4. $\frac{1}{[\text{A}]_t} - \frac{1}{[\text{A}]_0} = kt$
5. $k = Ae^{-E_a/RT}$
6. $\ln\frac{k_2}{k_1} = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$

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Acid/Base Equilibria

Concepts

1. The Arrhenius and Bronsted-Lowry definitions of acids and bases.
2. How the strength of an acid or base relates to the extent of ionization in water.
3. How relative acid/base strength is related to K_a/K_b.
4. Why water is a weak electrolyte and how its autoionization is related to $K_{w.}$
5. How the relative concentrations of OH^- and H_3O^+ are related to the acidity of a solution.
6. How water acts as both an acid and base.
7. How to write the conjugate acid of a base and the conjugate base of an acid.
8. The relationship between conjugate acid/base strength and the acid/base strength.
9. How percent dissociation of a weak acid increases as the concentration decreases.
10. The relationship between K_a, K_b, and K_w.
11. How a polyporotic acid dissociates in two or more steps.
12. How amines and weak acid ions act as bases in aqueous solution.
13. Why aqueous soluitons of small, highly charged metal ions yield acidic solutions.
14. The various combinations of cations and anions that lead to acidic, basic, or neutral solutions.

Required Skills

1. Classifying strong or weak acids/bases from their formulas.
2. Using K_w to calculate $[H_3O^+]$ and $[OH^-]$
3. Interconverting between pH, pOH, $[H_3O^+]$ and $[OH^-]$
4. Identifying conjugate acid/base pairs.
5. Using relative acid strengths to predict the net direction of a reaction.
6. Calculate K_a from the pH..
7. Calculate pH from K_a/K_b and initial acid/base concentration.
8. Calculate the percent dissociation of a weak acid.
9. Finding K_b of A^- from K_a of HA and K_w.
10. Predicting the relative acidity of salt solutions.