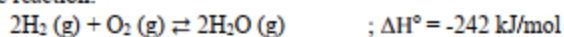


1.

1. (20pts) Given the reverse reaction:



List all below factors that shift reaction to right (forward reaction):

- a) Increase temperature;
- b) Decrease temperature;
- c) Hit há all O_2 ;
- d) Remove H_2 ;
- e) Use all water to make cocktail;
- f) Supply O_2 by photosynthesis in plant;
- g) Maintain H_2O concentration;
- h) Use catalyst;
- i) Steal H_2 form rocket and supply to this reaction;
- j) Increase pressure of the system.

2.

a) The rate of appearance of NO_2 :

$$d[\text{NO}_2]/dt = \frac{1}{2} d[\text{N}_2\text{O}_5]/dt = \frac{1}{2} \times 4,2 \cdot 10^{-7} = 2,1 \cdot 10^{-7} \text{ M/s}$$

b) The rate of appearance of O_2 :

$$d[\text{O}_2]/dt = \frac{1}{2} d[\text{N}_2\text{O}_5]/dt = 2,1 \cdot 10^{-7} \text{ M/s}$$

3)

- Box 1: rate = $k[\text{A}][\text{B}]^2 = k \cdot 5 \cdot 5^2 = 125k$

- Box 2: rate = $k \cdot 7 \cdot 3^2 = 63k$

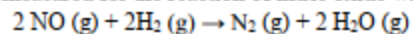
- Box 3: rate = $k \cdot 3 \cdot 7^2 = 147k$

$$\Rightarrow 2 < 1 < 3$$

4.

a) $\text{Rate} = k[\text{NO}]^x [\text{H}_2]^y$
 $\left. \begin{matrix} x = 1 \\ y = 1 \end{matrix} \right\} \text{overall rate order} = 1 + 1 = 2$
 $\Rightarrow \text{Rate} = k[\text{NO}][\text{H}_2]$

4. The following data were measured for the reaction of nitric oxide with hydrogen:



EXPERIMENT NUMBER	[NO] (M)	[H ₂] (M)	INITIAL RATE (M/S)
1	0.10	0.10	1.23×10^{-3}
2	0.10	0.20	2.46×10^{-3}
3	0.20	0.10	4.92×10^{-3}

$2^x = 2$

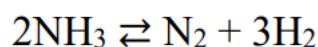
$2^y = 2$

a) Determine the rate law for this reaction by showing your calculation. (6pts)

b) Calculate the rate constant (4pts)

b) $k = \frac{\text{rate}}{[\text{NO}][\text{H}_2]} = \frac{1.23 \times 10^{-3}}{0.10 \times 0.10} = 0.12 \frac{1}{\text{s} \cdot \text{M}}$

5)



Initial	1	0	0
Change	$-2x$	$+x$	$+3x$
Equilibrium	$1 - 2x$	x	$3x$

$$\text{N}_2 \cdot \text{T}_1 / \text{P}_1 = \text{N}_2 \cdot \text{T}_2 / \text{P}_2$$

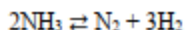
$$\Rightarrow 273 = (1 - 2x + x + 3x) \cdot 819 / 3,3$$

$$\Rightarrow 273 = (1 + 2x) \cdot 819 / 3,3$$

$$\Rightarrow x = 0.05$$

$$K_c = [\text{N}_2] \cdot [\text{H}_2]^2 / [\text{NH}_3]^3 = 2,08 \cdot 10^{-4}$$

6.

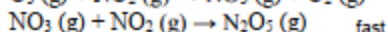
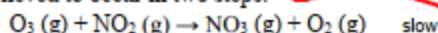


6. Calculate the pH of acetate buffer after mixing 25.00 mL of sodium acetate 0.10M with 50.00 mL of acetic acid, CH_3COOH , 0.20M. The temperature is 25°C. (10pts)

Given: K_a value for acetic acid at 25°C is 1.80×10^{-5} .

$n = 0.20 \times 0.05 = 0.01 \text{ moles}$ | total volume: $0.025 + 0.05 = 0.075 \text{ L}$

7. The reaction is believed to occur in two steps:



The experimental rate law is $\text{rate} = k[\text{O}_3][\text{NO}_2]$. What can you say about the relative rates of the two steps of the mechanism? (5pts)

it is first order w.r.t. to both reactants.

$$\text{pH} = \text{pK}_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$$

$$\frac{[\text{A}^-]}{[\text{HA}]} = \frac{0.01}{0.075} = 0.133$$

$$\text{pK}_a = -\log (K_a)$$

$$= 4.74$$

$$\text{pH} = 4.13$$

7)

- Because the rate law conforms to the molecularity of the first step which must be the rate-determining step. The second step must be faster than the first one.

- The experimental rate law $\text{rate} = k[\text{O}_3][\text{NO}_2]$ suggests that the rate of the reaction depends on the concentrations of O_3 and NO_2 . This implies that the first step of the mechanism, $\text{O}_3 (\text{g}) + \text{NO}_2 (\text{g}) \rightarrow \text{NO}_3 (\text{g}) + \text{O}_2 (\text{g})$, is the rate-determining step. The second step, $\text{NO}_3 (\text{g}) + \text{NO}_2 (\text{g}) \rightarrow \text{N}_2\text{O}_5 (\text{g})$, is faster and does not affect the overall rate of the reaction.

Part II: MULTIPLE CHOICE (30pts)

There may be more than or equal 0 and less than or equal 21 **INCORRECT** answers. But, unfortunately, I forgot how many **incorrect numbers**, so please remind me by giving exact number:

A. 5

B. 7

C. 9

D. 11

E. 13

F. Your answer: _____

List of statements:

- a) The rate of a chemical reaction changes with time.
- b) The rate constant of a reaction generally depends on the concentrations of species.**
- c) The function of buffer is to resist the change in pOH.**
- d) Chemical equilibrium exists when the two opposite reactions occur simultaneously at the same rate.
- e) In the Lowry-Bronsted theory a base is an OH⁻ donor.**
- f) The reaction order is experimentally determined.
- g) Temperature cannot affect the rate of a chemical reaction**
- h) When a catalyst is used in a reaction, it does not affect the final amounts of reactants and products.
- i) The value of the equilibrium constant for the reaction $K_c = 1.26 \times 10^{-12}$ at 500 K implies the product concentrations will be large relative to the reactants at equilibrium.**
- j) Equilibrium is achieved when the reactant and product concentrations become equal.**
- k) The rate law of a chemical reaction bears no relationship with the balancing coefficients of the overall reaction.
- l) Introducing a catalyst can affect the value of the equilibrium constant.**
- m) The rate law expresses how the rate varies with concentration of species.
- n) An electrolyte is a substance that dissolves in water to give an electrically non-conducting solution**
- p) Conjugate acid of Ac⁻ is HAc.
- q) The half-life, $t_{1/2}$, of a reaction is the time it takes for the product concentration to increase to one-half of its final value.**
- r) Hydration is the process that ion is surrounded by water molecules arranged in a specific manner.
- s) The stronger the conjugate base is, the weaker the conjugate acid is.
- t) $\text{pOH} = 14 - \text{pK}_a + \log\left(\frac{[\text{HA}]}{[\text{A}^-]}\right)$**
- u) The units of k for the rate law: $\text{Rate} = k[\text{A}][\text{B}]^2$ is $\text{L}^2 \text{mol}^{-2} \text{s}^{-1}$, when the concentration unit is mol/L.**
- w) Reactions usually occur at faster rates at higher temperatures.