

For a reaction:



$[A_0] = 4.0 \text{ M}$ ,  
and  $k = 7.4 \times 10^{-3} \text{ M}^{-1} \text{ s}^{-1}$ ,  
calculate  $[A]$  at  $t = 299 \text{ s}$

Select one:

a.  
0.30 M

b.  
0.63 M

c.  
0.84 M

d.  
0.40 M

e.  
0.99 M

The rate constant for the decomposition of a certain substance is  $2.80 \times 10^{-3} \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$  at  $30 \text{ }^\circ\text{C}$  and  $1.38 \times 10^{-2} \text{ L}\cdot\text{mol}^{-1}\cdot\text{s}^{-1}$  at  $50 \text{ }^\circ\text{C}$ . calculate the activation energy for that reaction.

Select one:

a.

$$2.52 \times 10^5 \text{ J}\cdot\text{mol}^{-1}$$

b.

$$3.67 \times 10^4 \text{ J}\cdot\text{mol}^{-1}$$

c.

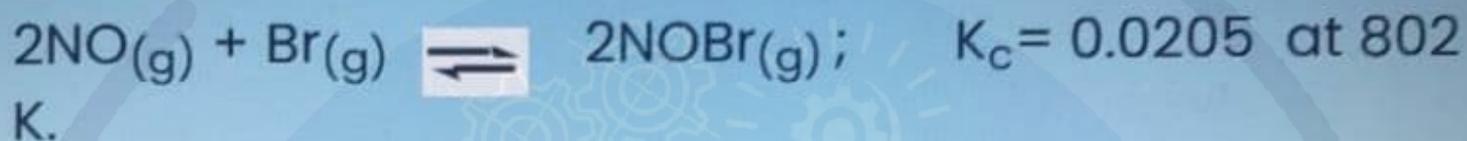
$$3.46 \times 10^5 \text{ J}\cdot\text{mol}^{-1}$$

d.

$$6.49 \times 10^4 \text{ J}\cdot\text{mol}^{-1}$$

e.

$$1.59 \times 10^5 \text{ J}\cdot\text{mol}^{-1}$$

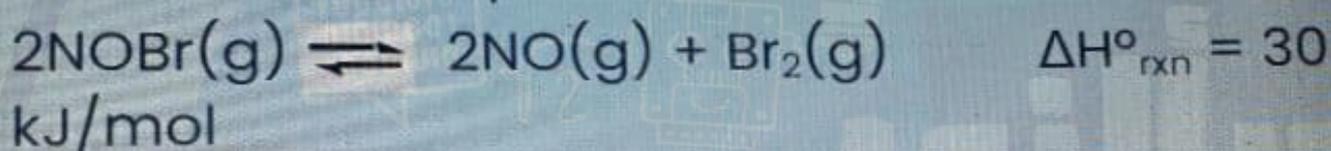


What is  $K_p$ ?

Select one:

- a.  $1.14 \times 10^{-5}$
- b.  $2.75 \times 10^{-6}$
- c.  $5.80 \times 10^{-4}$
- d.  $1.36 \times 10^{-4}$
- e.  $3.11 \times 10^{-4}$

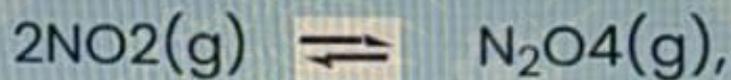
For the following reaction at equilibrium, which choice gives a change that will shift the position of equilibrium to favor the formation of more products?



Select one:

- a.  
Remove  $\text{Br}_2$ .
- b.  
Remove  $\text{NOBr}$  selectively
- c.  
Lower the temperature

For the following reaction which occurred at 100 °C.



If the initial concentration for  $[\text{NO}_2] = 0.78 \text{ M}$ , at equilibrium, the concentration of  $\text{N}_2\text{O}_4$  was found to be  $0.16 \text{ M}$ , What is the value of  $K_c$  at 100 °C?

Select one:

- a. 1.0
- b. 0.76
- c. 2.7
- d. 3.2
- e. 1.0

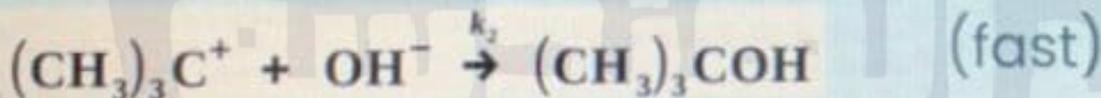
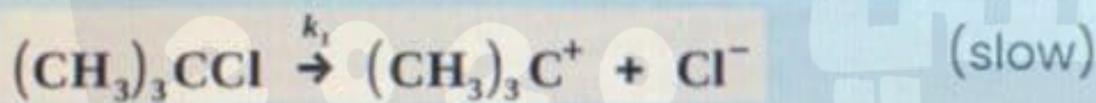
Which of the following statements **must be true**?

Select one:

- a. Transition state or the activated complex is defined as the low-energy state of a reaction.
- b. In the equation,  $\text{Rate} = k[A]^m[B]^n$ , the overall order of the reaction is defined as  $q = m - n$ .
- c. The catalyst raises the activation energy of the reaction by providing an alternative path that avoids the slow step of the uncatalysed reaction.
- d. The integrated rate law for a zero-order reaction is expressed as  $[A] = kt + [A]_0$ .
- e. The rate law can be expressed either in terms of the disappearance of reactants or the appearance of products

In basic solution,  $(\text{CH}_3)_3\text{CCl}$  reacts according to the equation below.

What is the rate law for the reaction based on the following accepted mechanism for the reaction?



Select one:

a.  $k_1 [(\text{CH}_3)_3\text{CCl}]$

b.

Rate constants usually:

are independent of temperature.

Select one:

- a.  
increase with temperature.
- b.  
decrease with time.

The reaction shown below has an equilibrium constant  $K_c$  been assigned to  $3.07 \times 10^{-4}$  at  $38^\circ\text{C}$ . Decide the status of the reaction mixture based on the following compositions:

$[\text{NOBr}] = 1.13 \times 10^{-2} \text{ M}$ ,  $[\text{NO}] = 1.58 \times 10^{-2} \text{ M}$ ,  $[\text{Br}_2] = 1.29 \times 10^{-3} \text{ M}$ .



Select one:

- a.  $Q_c = K_c$  and the mixture is at equilibrium
- b.  $Q_c = 2.52 \times 10^{-3}$  and the reaction will shift to the left
- c.  $Q_c = 6.33 \times 10^{-4}$  and the reaction will shift to the left
- d.

Which of the statements concerning the relative rate of reaction is correct for the decomposition of dinitrogen pentaoxide?



Select one:

a.

$$-\frac{d[\text{N}_2\text{O}_5]}{dt} = \frac{1}{2} \frac{d[\text{O}_2]}{dt}$$

b.

$$-\frac{d[\text{N}_2\text{O}_5]}{dt} = \frac{1}{2} \frac{d[\text{NO}_2]}{dt}$$

c.

1) A  $\rightarrow$  product

$$[A]_0 = 4 \text{ M}$$

$$k = 7.4 \times 10^{-3} \text{ M}^{-1} \text{ s}^{-1}$$

$[A]_t$ ?

$$\text{At } t = 299$$

This rxn is second order, we use

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

$$= 7.4 \times 10^{-3} \times 299 + \frac{1}{4}$$

$$\frac{1}{[A]_t} = 2.463 \text{ M}$$

$$[A]_t = 0.406 \text{ M} \quad \text{d)}$$

2)  $k_1 = 2.8 \times 10^{-3} \text{ L} \cdot \text{mol}^{-1} \cdot \text{s}^{-1}$

$$\text{At } T_1 = 50^\circ \text{C} + 273 = 323 \text{ K}$$

$$k_2 = 1.38 \times 10^{-2} \text{ L} \cdot \text{mol}^{-1} \cdot \text{s}^{-1}$$

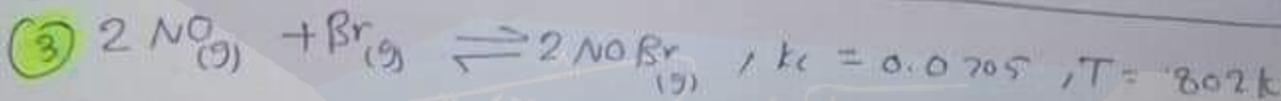
$$T_2 = 50^\circ \text{C} + 273 = 323 \text{ K}$$

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$\ln \left( \frac{2.8 \times 10^{-3}}{1.38 \times 10^{-2}} \right) = \frac{E_a}{8.314 \text{ J/K} \cdot \text{mol}} \left( \frac{1}{323} - \frac{1}{303} \right)$$

$$-1.595 = \frac{E_a}{8.314} \left( -2.04 \times 10^{-4} \right)$$
$$-1.595 = E_a - 2.457 \times 10^{-5}$$

$$E_a = 6.49 \times 10^4 \text{ J} \cdot \text{mol}^{-1} \quad \text{d)}$$



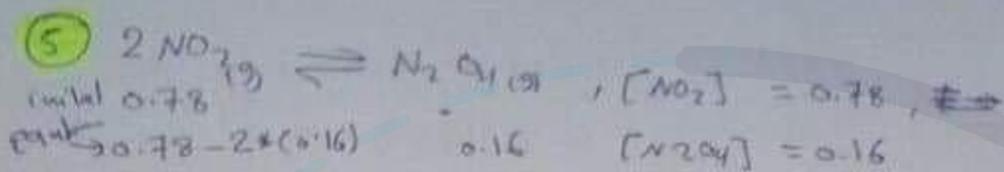
$$k_p = k_c (RT)^{\Delta n}$$

$$= 0.0205 (0.0821 (802))$$

$$k_p = 3.113 \times 10^{-4} \quad \text{e)}$$

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(4) Remove  $Br_2$ , give more space to form the  $2NO$



$$K_c = \frac{[N_2O_4]}{[NO_2]^2} = \frac{0.16}{(0.78 - 2 \times 0.16)^2}$$

$$K_c = \frac{0.16}{(0.78 - 2 \times 0.16)^2} = 0.756$$

(6) The rate law can be expressed either in terms of the disappearance of ~~reactants~~ reactants or the appearance of products.

(7) We write the mechanism for slow step as follows:

$$\text{rate law} = k [(CH_3)_3CCl] \quad (A)$$

(8) (a) Relation between Temperature and  $K_c$  as (9)  
 Increase with temperature.

(a)  $K_c = 3.07 \times 10^{-4}$ , find  $Q_c$

$$Q_c > K_c$$

$$Q_c = \frac{[NO]^2 [Br_2]}{[NOBr]^2} = \frac{(1.58 \times 10^{-3})^2 (1.79 \times 10^{-3})}{(1.13 \times 10^{-2})^2}$$

(b)  $2.52 \times 10^{-3}$  and the reaction will shift to the left.

$$Q_c = 2.52 \times 10^{-3}$$

(10) (b)  $-\frac{1}{2} \frac{d[N_2O_5]}{dt} = \frac{1}{2} \frac{d[NO_2]}{dt}$

$$-d[N_2O_5] = \frac{1}{2} \frac{d[NO_2]}{dt}$$