For a reaction:

$A \rightarrow$ products

$[A_o] = 4.0$ M,

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

calculate [A] at t= 299 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 $2.80 \times 10^{-3}\ L \cdot mol^{-1} \cdot s^{-1}$ at 30 °C and $1.38 \times 10^{-2}\ L \cdot mol^{-1} \cdot s^{-1}$ at 50 °C. calculate the activation energy for that reaction.

Select one:

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

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

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

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

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

$2NO_{(g)} + Br_{(g)} \rightleftharpoons 2NOBr_{(g)}$; $K_c = 0.0205$ at 802 K.

What is K_p?

Select one:

- a. 1.14×10^{-5}
- b. 2.75×10^{-6}
- c. 5.80×10^{-4}
- d. 1.36×10^{-4}
- e. 3.11×10^{-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?

$2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g) \qquad \Delta H°_{rxn} = 30$ kJ/mol

Select one:

- a. Remove Br_2
- b. Remove NOBr selectively
- c. Lower the temperature

For the following reaction was occurred at 100 °C.

$$2NO2(g) \rightleftharpoons N_2O4(g),$$

If the initial concentration for $[NO_2] = 0.78$ M, at equilibrium, the concentration of N_2O_4 was found to be 0.16 *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, 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, $(CH_3)_3CCl$ reacts according to the equation below.

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

$$(CH_3)_3CCl \xrightarrow{k_1} (CH_3)_3C^+ + Cl^- \quad \text{(slow)}$$

$$(CH_3)_3C^+ + OH^- \xrightarrow{k_2} (CH_3)_3COH \quad \text{(fast)}$$

Select one:

- a. $k_1[(CH_3)_3CCl]$
- 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 Kc been assigned to 3.07×10^{-4} at 38 °C. Decide the status of the reaction mixture based on the following compositions:

$[NOBr] = 1.13 \times 10^{-2}$ M, $[NO] = 1.58 \times 10^{-2}$ M, $[Br_2] = 1.29 \times 10^{-3}$ M.

$$2NOBr(g) \rightleftharpoons 2NO(g) + Br_2(g)$$

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?

$$2N_2O_5(g) \rightleftharpoons 4NO_2(g) + O_2(g)$$

Select one:

a.

$$-\frac{d[N_2O_5]}{dt} = \frac{1}{2}\frac{d[O_2]}{dt}$$

b.

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

c.

(1) $A \longrightarrow$ product

$[A_0] = 4M$

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

$[A]_t$?!

At $t = 299$

This Rxn is second order, we use

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

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

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

$$[A]_t = 0.406 M \quad (d)$$

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

At $T_1 = 30 c^\circ + 273 = 303 k$

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

$T_2 = 50 c^\circ + 273$

$= 323 k$

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

$$\ln \frac{(2.8 \times 10^{-3})}{(1.38 \times 10^{-2})} = \frac{Ea}{8.314 J/k.mol}\left(\frac{1}{323} - \frac{1}{303}\right)$$

~~[illegible]~~

$-1.595 = Ea \dfrac{-2.04 \times 10^{-4}}{8.314}$

$-1.595 = Ea \; -2.457 \times 10^{-5}$

$Ea = 6.49 \times 10^4 J.mol^{-1}$ (d)

(3) $2NO_{(g)} + Br_{(g)} \rightleftharpoons 2NOBr_{(g)}$, $k_c = 0.0205$, $T = 802 k$

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

$= 0.0205 (0.0821 (802))^{2-3}$

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

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

(5) $2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$, $[NO_2] = 0.78$,

initial 0.78

equil $\rightarrow 0.78 - 2*(0.16)$ 0.16 $[N_2O_4] = 0.16$

$$k_c = \frac{[N_2O_4]}{[NO_2]^2} = \text{[illegible]}$$

$$k_c = \frac{0.16}{(0.78 - 2(0.16))^2} = 0.756$$ (b)

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

(7) we can the mechanism for slow step as follow.

rate law $= k[(CH_3)_3CCl]$ (a)

(8) (a) Relation between Temperature and k as (a)

increase with temperature.

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

$$= \frac{[NO]^2[Br_2]}{[NOBr]^2} = \frac{(1.58\times10^{-3})^2 (1.79\times10^{-3})}{(1.13\times10^{-2})^2}$$

$Q_c > k_c$

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

(b) 2.52×10^{-3} and the Reaction will shift to the left

(10) (b) $$-\frac{1}{2}\frac{d[N_2O_5]}{dt} = \frac{1}{4/2}\frac{d[NO_2]}{dt}$$

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