

1. For the reaction:



If the value of $-\Delta[\text{Br}^-]/\Delta t = 7.5 \times 10^{-2} \text{ mol L}^{-1} \text{ s}^{-1}$, What will be the value of $\Delta[\text{Br}_2]/\Delta t$ at the same instant (in $\text{mol L}^{-1} \text{ s}^{-1}$)?

- A) 1.25×10^{-2} B) 1.5×10^{-2} C) 4.5×10^{-2} D) 1.25×10^{-1}

SOLUTION

$$-\frac{\Delta[\text{Br}^-]}{5 \times \Delta t} = +\frac{\Delta[\text{Br}_2]}{3 \times \Delta t}, -\frac{\Delta[\text{Br}^-]}{\Delta t} = +\frac{\Delta[\text{Br}_2]}{\Delta t} = -\frac{3 \times \Delta[\text{Br}^-]}{5 \times \Delta t} = \frac{3 \times 7.5 \times 10^{-2}}{5} = 0.045 \text{ mol L}^{-1} \text{ s}^{-1}$$

2. Nitric oxide reacts with chlorine to form nitrosyl chloride:



From the following data, the reaction rate law (rate equation) is:

Experiment	[NO] (M)	[Cl ₂] (M)	Initial rate (M min ⁻¹)
1	0.22	0.064	0.96
2	0.66	0.064	8.64
3	0.22	0.032	0.48

- A) Rate = $k[\text{NO}]^2[\text{Cl}_2]^2$ B) Rate = $k[\text{NO}]^2[\text{Cl}_2]^{1/2}$
 C) Rate = $k[\text{NO}][\text{Cl}_2]^{1/2}$ D) Rate = $k[\text{NO}]^2[\text{Cl}_2]$

SOLUTION

$$\text{rate} = k[\text{NO}]^x[\text{Cl}_2]^y$$

$$\frac{\text{rate}_2}{\text{rate}_1} = \frac{k[\text{NO}]_2^x[\text{Cl}_2]_2^y}{k[\text{NO}]_1^x[\text{Cl}_2]_1^y}, \frac{8.64}{0.96} = \left(\frac{0.66}{0.22}\right)^x, 9 = 3^x, x = 2$$

$$\frac{\text{rate}_3}{\text{rate}_1} = \frac{k[\text{NO}]_3^y[\text{Cl}_2]_3^y}{k[\text{NO}]_1^y[\text{Cl}_2]_1^y}, \frac{0.48}{0.96} = \left(\frac{0.032}{0.064}\right)^y, \frac{1}{2} = \left(\frac{1}{2}\right)^y, y = 1$$

$$\text{rate} = k[\text{NO}]^2[\text{Cl}_2]$$

3. At 25°C, the rate constant for the first order decomposition of a pesticide solution is $6.4 \times 10^{-3} \text{ min}^{-1}$. If the initial concentration of pesticide is $0.0314 \text{ mol L}^{-1}$. What its concentration (in mol L^{-1}) after 62.0 min at 25°C?

- A) 0.011 B) 0.0131 C) 0.0191 D) 0.0211

SOLUTION

$$\ln \frac{[\text{pesticide}]_0}{[\text{pesticide}]_t} = kt, \ln \frac{0.0314}{[\text{pesticide}]_t} = 6.4 \times 10^{-3} \times 62, [\text{pesticide}]_t = 0.0211 \text{ mol L}^{-1}$$

4. A first order reaction has a rate constant of $2.8 \times 10^{-2} \text{ s}^{-1}$. How many seconds will it take for the [reactant] to decrease from 0.88 mol L^{-1} to 0.14 mol L^{-1} ?

- A) 59.5 B) 65.7 C) 74.8 D) 88.6

SOLUTION

$$\ln \frac{[\text{reactant}]_0}{[\text{reactant}]_t} = kt, \ln \frac{0.88}{0.14} = 2.8 \times 10^{-2} \times t, t = 65.65 \text{ min}$$

5. Increasing of which of the following would alter the value of the rate constant (k) for the reaction?



- A) [A] only B) [B] only C) [A] and [B] D) temperature

SOLUTION

Because this is a heterogeneous equilibrium:

$$K_c = [\text{NH}_3][\text{H}_2\text{S}] \text{ and } K_p = P_{\text{NH}_3} \times P_{\text{H}_2\text{S}}$$

$$\begin{aligned} K_p &= K_c (R \times T)^{\Delta n_g} \\ K_p &= 1.2 \times 10^{-4} (0.0821 \times 295)^2 \\ K_p &= 0.0703 \\ K_p &= P_{\text{NH}_3} \times P_{\text{H}_2\text{S}} = 0.0703 \\ P_{\text{NH}_3} &= P_{\text{H}_2\text{S}} = \sqrt{0.0703} \\ P_{\text{NH}_3} &= 0.265 \text{ atm} \end{aligned}$$

OR

$$\begin{aligned} K_c &= [\text{NH}_3][\text{H}_2\text{S}] = 1.2 \times 10^{-4} \\ [\text{NH}_3] &= [\text{H}_2\text{S}] = \sqrt{1.2 \times 10^{-4}} \\ [\text{NH}_3] &= 0.01095 \text{ M} \\ P_{\text{NH}_3} &= \frac{n}{V} \times R \times T = [\text{NH}_3] \times R \times T \\ P_{\text{NH}_3} &= 0.0109 \times 0.0821 \times 295 \\ P_{\text{NH}_3} &= 0.265 \text{ atm} \end{aligned}$$

11. For following reaction at equilibrium:



Which of the following statements could be true? K_p increases with:

- 1) decreasing temperature. 2) increasing temperature.
 3) decreasing total pressure. 4) increasing total pressure.
 A) 1 only B) 1 and 4 C) 2 only D) 2 and 3

SOLUTION

Equilibrium constant for any reaction does not change except if its temperature changes. And for exothermic reactions equilibrium constants increases as temperature decreases.

12. 4.21 moles of S_2Cl_4 gas are introduced in 2.0 L vessel, and at equilibrium 1.25 moles of S_2Cl_4 are found in the reaction vessel, and K_c for the reaction is:



- A) 14.0 B) 17.0 C) 19.5 D) 21.5

SOLUTION

$$[\text{S}_2\text{Cl}_4]_{\text{initial}} = \frac{n_{\text{S}_2\text{Cl}_4}}{V} = \frac{4.21}{2} = 2.105 \text{ M}, [\text{S}_2\text{Cl}_4]_{\text{equilibrium}} = \frac{n_{\text{S}_2\text{Cl}_4}}{V} = \frac{1.25}{2} = 0.625 \text{ M}$$



$$2.105 - X = 0.625 \text{ M}, X = 1.48$$

$$[\text{SCl}_2]_{\text{equilibrium}} = 2X = 2 \times 1.48 = 2.96 \text{ M}$$

$$K_c = \frac{[\text{SCl}_2]^2}{[\text{S}_2\text{Cl}_4]} = \frac{(2.96)^2}{0.625} = 14.02$$

13. If the following equilibrium mixture concentrations are $[\text{SO}_3] = 10.00 \text{ M}$, $[\text{SO}_2] = 0.01 \text{ M}$ and $[\text{O}_2] = 0.10 \text{ M}$ at 700 K:



- A) the reaction mixture is at equilibrium.
 B) the reaction must proceed to the right to reach equilibrium.
 C) the reaction must proceed to the left to reach equilibrium.
 D) there is not enough information to answer.

SOLUTION

$$Q = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} = \frac{(10)^2}{(0.01)^2 \times (0.1)} = 1.00 \times 10^7$$

SOLUTION

$$[\text{OH}^-] = \sqrt{C_b K_b} = \sqrt{0.1 \times 1.8 \times 10^{-5}} = 1.34 \times 10^{-3} \text{ M}, \text{pOH} = -\log [\text{OH}^-]$$
$$\text{pOH} = -\log 1.34 \times 10^{-3} \times 10^{-4} = 2.87, \text{pH} = 11.13$$