CHEMISTRY

1. (4)
When volume is reduced to $\frac{1}{4}$, concentrations become four times

2. (2)
$$\frac{-dN_2}{dt} = \frac{-1}{3}\frac{dH_2}{dt} = \frac{1}{2}\frac{dNH_3}{dt}$$
$$\frac{dH_2}{dt} = \frac{3}{2} \times 0.001 = 0.0015 \text{ mol hr}^{-1}.$$

3. (4)
Rate = $K(N_2O_5)$
Hence $2.4 \times 10^{-5} = 3.0 \times 10^{-5}(N_2O_5)$
or $(N_2O_5) = 0.8$ mol L^{-1}

4. (1)
Slower reaction rate indicates higher energy of activation.

5. (2)
When the temperature is increased, heat energy is supplied which increases the kinetic energy of the reacting molecules. this will increase the number of collisions and ultimately the rate of reaction will be enhanced.

6. (4)
$$k = Ae^{-\frac{E^0}{RT}} \quad \underset{y}{\log K} = \underset{c}{\log A} - \underset{mx}{E^o / RT}$$
$\therefore \log k$ vs $\frac{1}{T}$.

7. (2)
Rate of formation of NH_3 (r_f) = 2 × rate of disappearance of nitrogen (r_d)
$r_{f(NH_3)} = \frac{2}{3}$ × rate of disappearance of Hydrogen (H_2)
$$\therefore -\frac{d[N_2]}{dt} = -\frac{1}{3}\frac{d[H_2]}{dt} = \frac{1}{2}\frac{d[NH_3]}{dt}$$

8. (2)
Rate of formation of B = $\frac{2}{3}\times$ rate of disappearance of A
$$\therefore +\frac{d[B]}{dt} = -\frac{2}{3}\frac{d[A]}{dt}$$

9. (4)
In the given rate law
Rate = $[A]^1[B]^2$
Power of [A] = 1 and [B] = 2
So overall order = 1 + 2 = 3
With respect to [A] order = 1
With respect to [B] order = 2

10. (4)
For first of order reaction Rate = k[A]
Rate constant $(k) = \frac{\text{Rate}}{[A]}$
$$= \frac{7.5 \times 10^{-4} \text{molL}^{-1}\text{s}^{-1}}{0.2\text{mol.L}^{-1}} = 3.75 \times 10^{-3} \text{ s}^{-1}$$