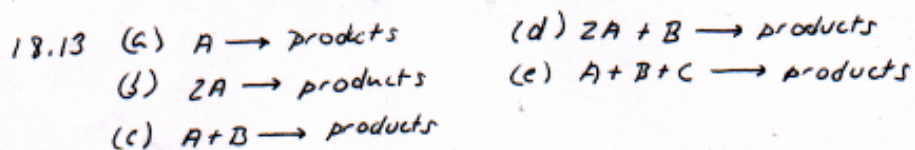


18.12 (a) rate = $k[A][B]$ (d) rate = $k[A][B][C]$
 (b) rate = $k[A]$ (e) rate = $k[A][B]^2$
 (c) rate = $k[A]^2$



termolecular reactions are the least likely to happen.

18.15 false, it depends on mechanism.

18.19 rate of $[O_2]$ & rate of $[NO]^2$ so rate = $k[NO]^2[O_2]$

$$k = \frac{\text{rate}}{[NO]^2[O_2]} = \frac{7.11 \times 10^{-3} \text{ M s}^{-1}}{(1.00 \times 10^{-2} \text{ M})^2 (1.00 \times 10^{-2} \text{ M})} = 7.11 \times 10^3 \text{ M}^{-2} \text{ s}^{-1}$$

18.20 using rate = $7.11 \times 10^3 \text{ M}^{-2} \text{ s}^{-1} [NO]^2 [O_2]$.

TRIAL	Rate (M s^{-1})
1	8.9×10^{-4}
2	1.8×10^{-3}
3	8.52×10^{-2}
4	1.92×10^{-1}

18.41 (a) For reaction 1:
 $k_1(300\text{K}) = (10^3 \text{ s}^{-1}) \exp\left[-\frac{1.00 \times 10^5 \text{ J mol}^{-1}}{(8.314 \text{ J mol}^{-1} \text{ K}^{-1})(300\text{K})}\right] = 3.88 \times 10^{-5} \text{ s}^{-1}$
 $k_1(1000\text{K}) = 5.98 \times 10^2 \text{ s}^{-1}$

For reaction 2:

$$k_2(300\text{K}) = 7.04 \times 10^{-6} \text{ s}^{-1}$$

$$k_2(1000\text{K}) = 1.80 \times 10^3 \text{ s}^{-1}$$

(b) Reaction 1 is faster at 300K.

Reaction 2 is faster at 1000K.

Reaction 1 has a smaller activation energy (which increases rate) but lower "A" (which decreases rate). At lower T, small E_a is more important. At higher T, "A" is more important.