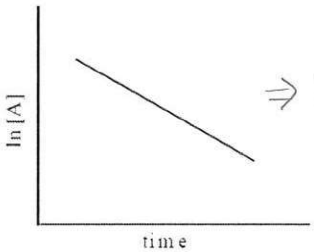

⇒ 1st order!
units of k = $\frac{1}{\text{time}}$

6. Which of the following statements best describe this graph?

× I. The graph represents a reactant that is second order.
✓ II. The absolute value of the slope for the graph is equal to the rate constant, k.
× III. The units for the rate constant, *k* will be M^{-1} time^{-1}

a) I only
[✓ b)] II only
c) I and II only
d) II and III only
e) I, II and III

7. As the reaction A + 2B → C proceeds at constant temperature, the reaction rate

a) remains the same since there is no catalyst present
b) remains the same since the temperature is constant
c) increases because the rate constant is a large number
d) increases because the rate of effective collisions increases over time
[✓ e)] decreases because the concentrations of the reactants decrease as the reaction progresses

8. For the reaction whose rate law is given below, a plot of which of the following is a straight line?

rate = $k[\mathrm{X}]$ ⇒ 1st order

a) [X] vs time
b) 1/[X] vs time
c) ln[X] vs 1/time
[✓ d)] ln[X] vs time

9. What are the potential units for the rate constant for the reaction below?

$\mathrm{A + B \rightarrow C + D}$ rate = $k[\mathrm{A}][B]^2$] 3rd order overall

a) s^{-1}
b) $s^{-1}\,M^{-1}$
[✓ c)] $s^{-1}\,M^{-2}$
d) $s^{-1}\,M^{-3}$

10. $2\,\mathrm{H_2O \rightarrow 2\,H_2 + O_2}$

Which of the following is true of the relative rates of disappearance of the reactants and appearance of the products?

a) O_2 appears at twice the rate that H_2O disappears.
b) H_2 appears at half the rate that H_2O disappears.
c) H_2 appears at twice the rate that H_2O disappears.
[✓ d)] O_2 appears at half the rate that H_2O disappears.

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Free Response Practice!

The reaction between bromate ions and bromide ions in acidic aqueous solution is given by the equation:

$$BrO_3^-(aq) + 5\,Br^-(aq) + 6\,H^+(aq) \rightarrow 3\,Br_2(\ell) + 3\,H_2O(\ell)$$

The table below gives the results of four experiments.

Experiment	Initial $[BrO_3^-]$	Initial $[Br^-]$	Initial $[H^+]$	Measured initial rate (mol/L•s)	
1	0.10	0.10	0.10	8.0×10^{-4}	
2	2x 0.20	0.10	0.10	2x 1.6×10^{-3}	4x
3	0.20	2x 0.20	0.10	2x 3.2×10^{-3}	
4	0.10	0.10	2x 0.20	3.2×10^{-3}	

a. Using the data above, determine the order for all three reactants and the overall reaction order. Justify!

b. Write the rate law for the reaction.

c. What is the value of k, and what are the units of k?

d. For experiment 4, calculate the amount of BrO_3^- remaining when 75% of the H^+ has reacted.

e. What is the rate of formation of Br_2 in experiment 1?

a.) experiments 1+2: 2x $[BrO_3^-]$ while $[Br^-]$ and $[H^+]$ constant ↑ rate x 2, thus rxn is 1st order w/ respect to BrO_3^-.

expt. 2+3: 2x $[Br^-]$ while $[BrO_3^-]$ and $[H^+]$ constant ↑ rate x 2, thus rxn is 1st order w/ respect to Br^-.

expt. 1+4: 2x $[H^+]$ while $[Br^-]$ and $[BrO_3^-]$ constant ↑ rate x 4, thus rxn is 2nd order w/ respect to H^+.

⇒ overall reaction order $= 1+1+2 = \boxed{4^{th} \text{ order overall}}$

b.) $\text{rate} = k[BrO_3^-][Br^-][H^+]^2$

c.) $$k = \frac{\text{rate}}{[BrO_3^-][Br^-][H^+]^2} = \frac{8.0E\text{-}4 \text{ M/s}}{(0.10\text{ M})^4} = 8.0\text{ M}^{-3}\text{s}^{-1} \left(\frac{1}{M^3 s}\right)$$

d.) $$0.20\text{ M} \times 0.75 = 0.15 \frac{mol}{L} H^+ \times \frac{1 \text{ mol } BrO_3^-}{6 \text{ mol } H^+} = 0.025 \text{ M } BrO_3^- \text{ reacted}$$

$$\Rightarrow 0.10 - 0.025 = 0.075 \text{ M } BrO_3^- \text{ remaining} = \boxed{0.08 \text{ M } BrO_3^- \text{ remaining}}$$

mmm, sig figs

e.) $$\text{rate} = \frac{1}{3}\frac{\Delta[Br_2]}{\Delta t} = 8.0E\text{-}4 \text{ M/s}$$

$$\Rightarrow \frac{\Delta[Br_2]}{\Delta t} = 3 \times 8.0E\text{-}4 \text{ M/s} = \boxed{0.0024 \text{ M/s}}$$