



AP[®] Chemistry 2003 Sample Student Responses

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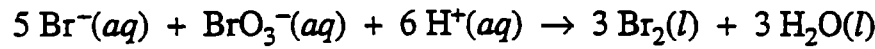
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3A₁



3. In a study of the kinetics of the reaction represented above, the following data were obtained at 298 K.

Experiment	Initial $[\text{Br}^-]$ (mol L ⁻¹)	Initial $[\text{BrO}_3^-]$ (mol L ⁻¹)	Initial $[\text{H}^+]$ (mol L ⁻¹)	Rate of Disappearance of BrO_3^- (mol L ⁻¹ s ⁻¹)
1	0.00100	0.00500	0.100	2.50×10^{-4}
2	0.00200	0.00500	0.100	5.00×10^{-4}
3	0.00100	0.00750	0.100	3.75×10^{-4}
4	0.00100	0.01500	0.200	3.00×10^{-3}

$\times 2$
 $\times 1.5$

(a) From the data given above, determine the order of the reaction for each reactant listed below. Show your reasoning.

(i) Br^- *first order*

(ii) BrO_3^-

(iii) H^+

(b) Write the rate law for the overall reaction.

(c) Determine the value of the specific rate constant for the reaction at 298 K. Include the correct units.

(d) Calculate the value of the standard cell potential, E° , for the reaction using the information in the table below.

Half-reaction	E° (V)
$\text{Br}_2(\text{l}) + 2 e^- \rightarrow 2 \text{Br}^-(\text{aq})$	+1.065
$2 \text{BrO}_3^-(\text{aq}) + 12 \text{H}^+(\text{aq}) + 12 e^- \rightarrow 2 \text{Br}_2(\text{l}) + 6 \text{H}_2\text{O}(\text{l})$	+1.52

(e) Determine the total number of electrons transferred in the overall reaction.

(i) The order for $[\text{Br}^-]$ is first order because from experiment one to experiment two, the concentration of Br^- is doubled from 0.00100 to 0.00200, while the other concentrations remain the same. The doubling in the concentration of Br^- , caused the rate to double from 2.50×10^{-4} to 5.00×10^{-4} . Thus, Br^- is first order in the reaction.

(ii) The order for BrO_3^- is also first order. We know this because from experiment one to experiment three, the concentration

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ADDITIONAL PAGE FOR ANSWERING QUESTION 3.

of BrO_3^- increases by a multiplication of 1.5, (from .00500 to .00750) and the concentrations of Br^- and H^+ remain the same. From experiment one to experiment three, the rate also increased by a multiplication of 1.5 (from 2.5×10^{-4} to 3.75×10^{-4}) thus making $[\text{BrO}_3^-]$ first order in the reaction.

(iii) H^+ in the reaction is _____ order. This one was trickier but from experiment one to experiment four the concentration of H^+ is doubled, but the concentration of BrO_3^- is tripled. Since we already know $[\text{BrO}_3^-]$ is first order, the tripling would cause the rate to triple from 2.50×10^{-4} in the first experiment to 7.5×10^{-4} in the fourth experiment. This means that H^+ must be second order because quadrupling 7.5×10^{-4} gives a rate of .003, which is the correct answer. Therefore while the $[\text{H}^+]$ doubled from experiment one to four, the rate quadrupled making H^+ second order.

$$(b) \text{rate} = k [\text{Br}^-] [\text{BrO}_3^-] [\text{H}^+]^2$$

$$(c) \text{rate} = k [\text{Br}^-] [\text{BrO}_3^-] [\text{H}^+]^2$$

$$2.5 \times 10^{-4} \frac{\text{mol}}{\text{L} \cdot \text{s}} = k (.00100) (.00500) (.100)^2$$

$$2.5 \times 10^{-4} \frac{\text{mol}}{\text{L} \cdot \text{s}} = k 5 \times 10^{-8} \text{M}^3$$

$$k = 5000 \frac{1}{\text{M}^2 \cdot \text{s}}$$

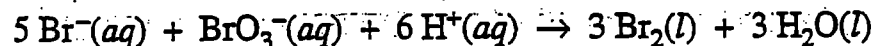
$$(d) E^\circ = -1.065 + 1.52 = .455 \text{ V}$$

$$E^\circ = .455 \text{ V}$$

e) The number of electrons transferred in the overall reaction will be $10e^-$ because the second reaction must be multiplied by 2 in order to clear the fraction, so the $5e^-$ will become $10e^-$.

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3B₁



3. In a study of the kinetics of the reaction represented above, the following data were obtained at 298 K.

Experiment	Initial [Br ⁻] (mol L ⁻¹)	Initial [BrO ₃ ⁻] (mol L ⁻¹)	Initial [H ⁺] (mol L ⁻¹)	Rate of Disappearance of BrO ₃ ⁻ (mol L ⁻¹ s ⁻¹)
1	0.00100	0.00500	0.100	2.50 × 10 ⁻⁴
2	0.00200	0.00500	0.100	5.00 × 10 ⁻⁴
3	0.00100	0.00750	0.100	3.75 × 10 ⁻⁴
4	0.00100	0.01500	0.200	3.00 × 10 ⁻³

- (a) From the data given above, determine the order of the reaction for each reactant listed below. Show your reasoning.
- (i) Br⁻
 - (ii) BrO₃⁻
 - (iii) H⁺
- (b) Write the rate law for the overall reaction.
- (c) Determine the value of the specific rate constant for the reaction at 298 K. Include the correct units.
- (d) Calculate the value of the standard cell potential, E°, for the reaction using the information in the table below.

Half-reaction	E° (V)
Br ₂ (l) + 2 e ⁻ → 2 Br ⁻ (aq)	+1.065
BrO ₃ ⁻ (aq) + 6 H ⁺ (aq) + 5 e ⁻ → $\frac{1}{2}$ Br ₂ (l) + 3 H ₂ O(l)	+1.52

- (e) Determine the total number of electrons transferred in the overall reaction.

(i) first, when the concentration of Br⁻ doubles and everything else is constant, the rate doubles

(ii) first, when the concentration of BrO₃⁻ is multiplied by 1.5 and everything else is constant the rate multiplies by 1.5

(iii) second, between experiments 2 and 4 the rate becomes 8 times more, it doubles because BrO₃⁻ doubles. since that is first order

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that leaves a change of 4 times when $[H^+]$ is doubled, so H^+ is second order

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

$$c) 2.5 \times 10^{-4} = k (0.001 M) (0.005 M) (0.1 M)^2$$

$$2.5 \times 10^{-4} = k 5.00 \times 10^{-8} M^4$$

$$\frac{2.5 \times 10^{-4} \frac{\text{mol}}{L \cdot s}}{5.00 \times 10^{-8} \frac{\text{mol}^4}{L^4 \cdot s}} = k$$

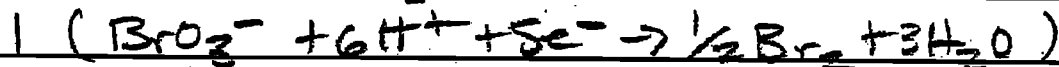
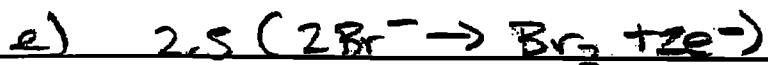
$$5.00 \times 10^3 \frac{\text{mol}^3}{L^3 \cdot s} = k$$

$$5.00 \times 10^3 \frac{L^3}{\text{mol}^3 \cdot s} = k$$

$$d) E^{\circ} = E^{\circ}_{\text{reduced}} - E^{\circ}_{\text{oxidized}}$$

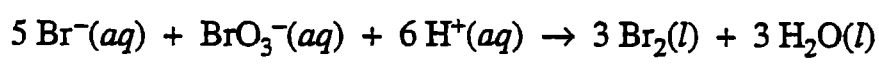
$$1.52 - 1.065$$

$$E^{\circ} = .46 V$$



5 electrons are transferred in each half reaction for a total of 10 electrons transferred

3C,



3. In a study of the kinetics of the reaction represented above, the following data were obtained at 298 K.

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1	0.00100	0.00500	0.100	2.50×10^{-4}
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- (a) From the data given above, determine the order of the reaction for each reactant listed below. Show your reasoning.
- (i) Br^-
 - (ii) BrO_3^-
 - (iii) H^+
- (b) Write the rate law for the overall reaction.
- (c) Determine the value of the specific rate constant for the reaction at 298 K. Include the correct units.
- (d) Calculate the value of the standard cell potential, E° , for the reaction using the information in the table below.

Half-reaction	E° (V)
$\text{Br}_2(l) + 2 e^- \rightarrow 2 \text{Br}^-(aq)$	+1.065
$\text{BrO}_3^-(aq) + 6 \text{H}^+(aq) + 5 e^- \rightarrow \frac{1}{2} \text{Br}_2(l) + 3 \text{H}_2\text{O}(l)$	+1.52

(e) Determine the total number of electrons transferred in the overall reaction.

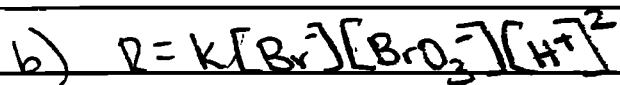
a) $R = k [\text{Br}^-]^x [\text{BrO}_3^-]^y [\text{H}^+]^z$

(i) $\frac{2.50 \times 10^{-4}}{5.00 \times 10^{-4}} = \frac{k(0.001)^x (0.005)^y (0.1)^z}{k(0.002)^x (0.005)^y (0.1)^z}$ $\frac{1}{2} = \left(\frac{1}{2}\right)^x$ $x=1$ first order for Br^-

(ii) $\frac{2.50 \times 10^{-4}}{3.75 \times 10^{-4}} = \frac{k(0.001)^x (0.005)^y (0.1)^z}{k(0.001)^x (0.0075)^y (0.1)^z}$ $\frac{2}{3} = \left(\frac{2}{3}\right)^y$ $y=1$ first order for BrO_3^-

(iii) $\frac{2.50 \times 10^{-4}}{3.00 \times 10^{-3}} = \frac{k(0.001)^x (0.005)^y (0.1)^z}{k(0.001)^x (0.015)^y (0.2)^z}$ $0.083 = \frac{0.33(0.5)^z}{4}$ $\frac{1}{4} = \left(\frac{1}{2}\right)^z$ $z=2$ second order for H^+

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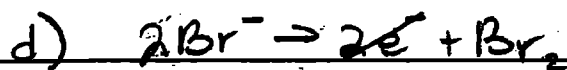


c) $2.50 \times 10^{-4} = k(0.001)(0.005)(.1)^2$

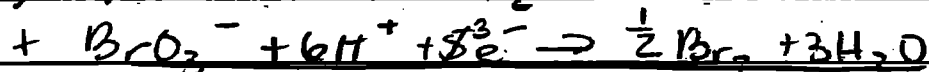
$2.50 \times 10^{-4} = k(5 \times 10^{-3})$

$\boxed{5.00 \times 10^8 \text{ mol}^{-3} \text{ L}^{-1}} = k$

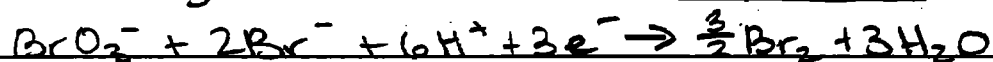
$\frac{\text{mol}}{\text{L s}} = \frac{(\text{mol})^4}{\text{L}^4}$



$E^\circ = -1.065$



$E^\circ = +1.52$



$\boxed{-.455 \text{ V}}$

