

## Kinetics of the Oxidation of $\text{Fe}^{2+}$ to $\text{Fe}^{3+}$

**Part I:  $\text{pH} > 3.5$ .** The data for this set of conditions are an example of the initial rate method for finding a rate law. The key to working with data of this sort is to isolate pairs of experiments in which the concentration of only one species is changing.

- To find the reaction order for  $\text{Fe}^{2+}$  we compare experiments 1 and 2, finding that doubling its concentration ( $1.0 \times 10^{-4} \text{ M} \rightarrow 2.0 \times 10^{-4} \text{ M}$ ) leads to a doubling of the rate ( $2.4 \times 10^{-6} \text{ M/d} \rightarrow 4.8 \times 10^{-6} \text{ M/d}$ ); thus, the rate is first-order in  $\text{Fe}^{2+}$ .
- To find the reaction order for  $\text{O}_2$  we compare experiments 2 and 4, finding that increasing the partial pressure by 50% ( $0.20 \rightarrow 0.30$ ) leads to a 50% increase in the rate ( $4.8 \times 10^{-6} \text{ M/d} \rightarrow 7.2 \times 10^{-6} \text{ M/d}$ ); thus, the rate is first-order in  $\text{O}_2$ .
- To find the reaction order for  $\text{H}_3\text{O}^+$  we compare experiments 1 and 3, finding that increasing its concentration by a factor of 10 ( $1.0 \times 10^{-5} \text{ M} \rightarrow 1.0 \times 10^{-4} \text{ M}$ ) causes the rate to decrease by a factor of 100 ( $2.4 \times 10^{-6} \text{ M/d} \rightarrow 2.4 \times 10^{-8} \text{ M/d}$ ); thus, the rate is inverse second-order in  $\text{H}_3\text{O}^+$ . Yes, sometimes a species can have an inhibitory effect on a reaction's rate.

At this point we know that the rate law is

$$\text{rate} = k_1[\text{Fe}^{2+}][\text{H}_3\text{O}^+]^{-2}P_{\text{O}_2}$$

To find the rate constant we substitute concentrations, partial pressures and rates from each experiment into the rate law and solve for  $k_1$ . Doing this for each experiment produces the same result of  $1.2 \times 10^{-11} \text{ M}^2 \text{ atm}^{-1} \text{ d}^{-1}$ . Pay particular attention to the units and convince yourself that they are correct.

To rewrite the rate law in terms of  $\text{OH}^-$  we note that

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

Solving this for  $[\text{H}_3\text{O}^+]$  and substituting back into our rate law gives

$$\text{rate} = k_1 K_w^{-2}[\text{Fe}^{2+}][\text{OH}^-]^2 P_{\text{O}_2} = k_2[\text{Fe}^{2+}][\text{OH}^-]^2 P_{\text{O}_2}$$

The value for  $k_2$ , therefore, is  $k_1/K_w^2$ , or  $1.2 \times 10^{17} \text{ M}^{-2} \text{ atm}^{-1} \text{ d}^{-1}$ . Again, be sure to convince yourself that the units are correct.

**Part II:  $\text{pH} < 3.5$ .** To show that the data are indeed first-order we can estimate the first several half-lives and observe their pattern. The first three half-lives are approximately

29,000 d each, which is consistent with the observation that the half-life for a first-order process is independent of concentration.

Knowing the half-life allows us to calculate the pseudo-order rate constant,  $k_4$ ; this is

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{29000 \text{ d}} = 2.4 \times 10^{-5} \text{ d}^{-1}$$

This, in turn, allows us to calculate  $k_3$  since

$$k_4 = k_3 P_{\text{O}_2}$$

Thus

$$k_3 = k_4 / P_{\text{O}_2} = 2.9 \times 10^{-5} \text{ d}^{-1} / 0.20 \text{ atm} = 1.2 \times 10^{-4} \text{ d}^{-1} \text{ atm}^{-1}$$

Finally, at a pH of 3.0 and a partial pressure of 0.20 atm a 4.7 mM solution of  $\text{Fe}^{2+}$  will follow the pseudo-order kinetics given by the rate law

$$\text{rate} = k_4 [\text{Fe}^{2+}]$$

For a first-order reaction we know that

$$\ln[\text{Fe}^{2+}]_t = \ln[\text{Fe}^{2+}]_0 - k_4 t$$

Thus

$$\ln(1.2 \text{ mM}) = \ln(4.7 \text{ mM}) - (2.4 \times 10^{-5} \text{ d}^{-1})t$$

$$t = 56,900 \text{ d}$$

Note that we can leave the concentrations in mM if we prefer.

**Part III: Combining the Rate Laws.** To find the pH where the two mechanisms for the oxidation of  $\text{Fe}^{2+}$  have the same rate we first set them equal to each other

$$k_1 [\text{Fe}^{2+}] [\text{H}_3\text{O}^+]^{-2} P_{\text{O}_2} = k_3 [\text{Fe}^{2+}] P_{\text{O}_2}$$

and then simplify, giving

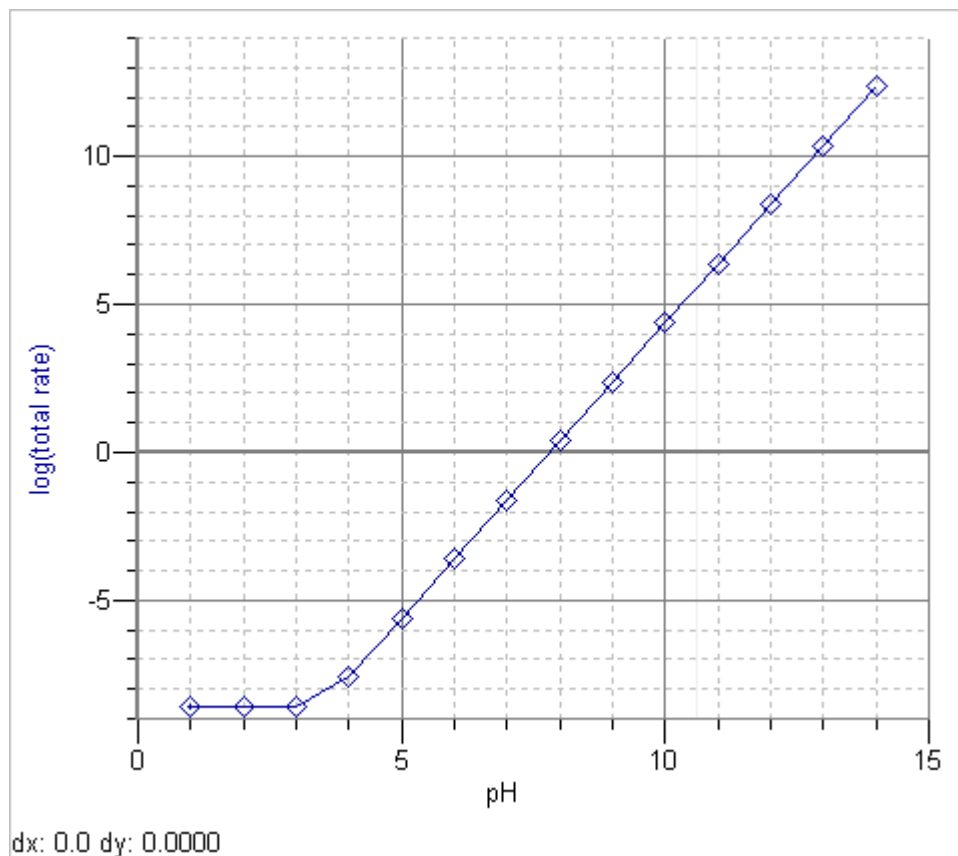
$$k_1 [\text{H}_3\text{O}^+]^{-2} = k_3$$

Next, we solve this for the concentration of  $\text{H}_3\text{O}^+$

$$[\text{H}_3\text{O}^+] = \sqrt{\frac{k_1}{k_3}} = \sqrt{\frac{1.2 \times 10^{-11} \text{ M}^2 \text{ atm}^{-1} \text{ d}^{-1}}{1.2 \times 10^{-4} \text{ atm}^{-1} \text{ d}^{-1}}} = 3.16 \times 10^{-4} \text{ M}$$

which corresponds to a pH of 3.5.

A graph of  $\log(\text{rate})$  vs. pH is shown here



Note that the rate remains constant until the pH is greater than 3 and then it starts to increase quickly. Clearly the rate of oxidation is favored at more basic pH levels. The reason for this is the formation of various  $\text{Fe}^{2+}\text{-OH}^-$  complexes. It turns out that the oxidation of  $\text{Fe}^{2+}$  is much easier when the iron is complexed with hydroxide. There are important environmental implications for this oxidation process. Ground water percolating through gob piles from strip mining is often fairly acidic and somewhat anaerobic (low in  $\text{O}_2$ ), conditions that favor the presence of  $\text{Fe}^{2+}$ . When the water seeps into the open and mixes with other water, the pH increases and, with the presence of atmospheric  $\text{O}_2$ , the  $\text{Fe}^{2+}$  quickly oxidizes to  $\text{Fe}^{3+}$ , which precipitates as  $\text{Fe}(\text{OH})_3$  (iron(III)hydroxide is much less soluble than iron(II)hydroxide). If you have ever encountered standing water that has an orange gunky looking precipitate, then you have seen the result of this process!