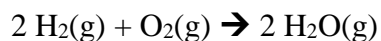


Name: _____

1. (10 pt.) Consider the reaction:



The following data was obtained at 298 K.

| $[\text{H}_2(\text{g})]_0$ (M) | $[\text{O}_2(\text{g})]_0$ (M) | Initial Rate (M/s) | Temp (K) |
|--------------------------------|--------------------------------|--------------------|----------|
| 1.50 | 1.50 | 0.33 | 298 K |
| 1.50 | 0.75 | 0.17 | 298 K |
| 3.00 | 1.50 | 0.65 | 298 K |

What is the rate constant for the reaction at 298 K?

1st, I need the rate law. You can do it the long way, if you want.

Just looking at the 1st two mixtures, when I doubled the O₂ (0.75 to 1.50 M) while keeping H₂ constant, the rate approximately doubled (0.17 to 0.33 M/s). It is 1st order in O₂.

If I look at the 1st and 3rd mixtures, doubling the H₂ (1.50 to 3.00 M) while keeping O₂ constant resulted in the rate doubling (0.33 to 0.65 M/s). It is 1st order in H₂.

So:

$$\text{Rate} = k[\text{H}_2][\text{O}_2]$$

Now, I can just plug in the data to get the rate constant. They are all slightly different, so I'll take the average.

$$0.33 \frac{\text{M}}{\text{s}} = k[1.50 \text{ M}][1.50 \text{ M}]$$

$$k = 0.147 \frac{1}{\text{Ms}}$$

$$0.17 \frac{\text{M}}{\text{s}} = k[0.75 \text{ M}][1.50 \text{ M}]$$

$$k = 0.151 \frac{1}{\text{Ms}}$$

$$0.65 \frac{\text{M}}{\text{s}} = k[3.00 \text{ M}][1.50 \text{ M}]$$

$$k = 0.144 \frac{1}{Ms}$$

$$k_{avg} = 0.147 \frac{1}{Ms}$$