Revised August 2009

AP WORKSHEET 15a: ANSWERS

1.

(a) Rate = k $[H_2]^2[Br_2]^0$

$1.92 \times 10\text{-}3\ M^{-1}\ s^{-1}$ $= 1.92 \times 10^{-3} \frac{L}{mol \cdot s}$

(b) 2^{nd} order, (2 + 0 = 2)

2.

Remember that M stands for mol/L

(a) Rate = k $[O_2]\ [NO]^2$

Remember that $s^{-1} = \frac{1}{s}$

(b) $500\ M^{-2}\ s^{-1}$

3.

(a) $(mol\ L^{-1})^{-2}\ s^{-1} = \frac{L^2}{s \cdot mol^2}$

(b) $M^{-1}\ min^{-1} = \frac{L}{min \cdot mol}$

(c) $g\ s^{-1}$ $\frac{g}{s} = k\left[\frac{mol}{L}\right]^0$ $\left[\frac{mol}{L}\right]^0 = 1$

There are, of course, other ways to express these units correctly

So $k = g/s$

4. Start with known concentrations of the two reactants and carry out the reaction. Measure the rate of reaction by collecting the gas, and record the volume produced per unit time. This is the rate

   Repeat the experiment, this time changing the concentration of one of the reactants by a specific, known amount, but leaving the other concentration unchanged. Once again measure the rate of reaction by collecting the gas, and record the volume produced per unit time. This is the rate

   Repeat the experiment once more, this time changing the concentration of the reactant that remained constant in the first repetition by a specific, known amount, but reverting back to the original concentration the reactant that was changed in the first repetition. Once again measure the rate of reaction by collecting the gas, and record the volume produced per unit time. This is the rate

   A comparison of the rate changes with associated concentration changes leads to determination of the orders with respect to each reactant

$\frac{M}{s} \div M^2 \quad \frac{\cancel{M}}{s} \times \frac{1}{\cancel{M}M}$

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AP WORKSHEET 15a: Orders of reaction & rate constants

1. The following data were collected for the reaction below.

$$H_{2(g)} + Br_{2(g)} \rightarrow 2HBr_{(g)}$$

Experiment	[H_2]	[Br_2]	Rate in M s^{-1}
1	0.25 M	0.0012 M	1.20×10^{-4}
2	×2 0.50 M	×1 0.0012 M	×4 4.80×10^{-4}
3	×1 0.50 M	×4 0.0048 M	×1 4.80×10^{-4}

Changing [Br_2] has no affect on rate.

(a) Write the rate equation and calculate the value of the rate constant, k and include units. (6)

$2^n = 4$

rate $= k[H_2]^2[Br_2]^0$. To calculate rate constant, plug it back in: $1.20 \times 10^{-4} = k[0.25]^2$ $\quad k = \frac{1.20 \times 10^{-4}\ M/s}{(.25\ M)^2} = 0.00192$

$\boxed{\text{rate} = 0.00192[H_2]^2}$ $\quad$ units of k: $\frac{M}{s} \cdot \frac{1}{M^2} = \boxed{\frac{1}{sM}}$

(b) What is the overall order of this reaction? (1)

two $\quad (2+0=2)$

2. Nitrogen monoxide can be oxidized to nitrogen dioxide in the reaction below.

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

The following data were collected in a kinetics experiment.

Experiment	Initial [O_2] (M)	Initial [NO] (M)	Rate (M s^{-1})
1	0.20	0.10	1.0
2	×4 0.80	0.10	×4 4.0
3	0.80	×3 0.30	36

(a) Write the rate law. (4)

for [O_2]: $4^? = 4$ $\quad ?=1$ $\quad$ for [NO]: $3^? = 9$ $\quad ?=2$ $\quad$ rate $= k[O_2][NO]^2$

(b) Calculate the rate constant and give its units. (2)

$1.0\frac{M}{s} = k(0.20)(.10)^2$ $\quad$ units: $\frac{\frac{M}{s}}{M^3} = \frac{\cancel{M}}{s} \cdot \frac{1}{\cancel{M}M^2}$

$k =$

$\boxed{\frac{1}{sM^2}}$

3. In each of the following cases where the rate law and units used are described, suggest units for the rate constant, k.(6)

   (a) A third order reaction overall, where the rate is measured in mol L^{-1} s^{-1} and the concentrations of all reactants are measured in units of mol L^{-1}.

$$\frac{mol}{L\,s} = k\left(\frac{mol^3}{L^3}\right) \qquad k = \frac{\cancel{mol}}{\cancel{L}\,s}\cdot\frac{L^{\cancel{3}\,2}}{mol^{\cancel{3}\,2}} = \frac{L^2}{s\,mol^2}$$

   (b) A rate law that has the concentration of two reactants measured in M, each first order and a rate that is measured in mol L^{-1} min^{-1}

$$\frac{mol}{L\,min} = k\,\frac{Mol}{L}\cdot\frac{Mol}{L} \qquad k = \frac{\cancel{mol}}{\cancel{L}\,min}\times\frac{L^{\cancel{2}}}{mol^{\cancel{2}}} = \frac{L}{min\,mol}$$

   (c) A reaction where there are multiple products but only a single reactant that is found to be zero order. The rate is measured in units of g s^{-1}

A single reactant that has no effect on rate? How can that be?

$$\frac{g}{s} = k\left[\frac{M}{L}\right]^0$$

$$k = \frac{g}{s}$$