Chemistry 225 Semester 01-2012
Homework for Submission #1
Answer Key

Total marks 27

1) Propene is slowly oxidised in a plentiful supply of air at a constant temperature according to the following equation:

$$2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

The rate of reaction with respect to ethane is measured at a certain time as 1.7×10^{-4} M s^{-1}. Calculate:

a) The rate of reaction with respect to oxygen. (2)

From the equation, (rate w.r.t. propene)/2 = (rate w.r.t. oxygen)/9

$\therefore$ rate w.r.t. oxygen = (rate w.r.t. propene) × (9/2) = $1.7 \times 10^{-4} \times (9/2) = 7.65 \times 10^{-4} \otimes 7.7 \times 10^{-4}$ Ms^{-1} to 2 significant figures.

b) The rate of reaction with respect to carbon dioxide. (2)

From the equation, (rate w.r.t. ethane)/2 = (rate w.r.t. carbon dioxide)/6

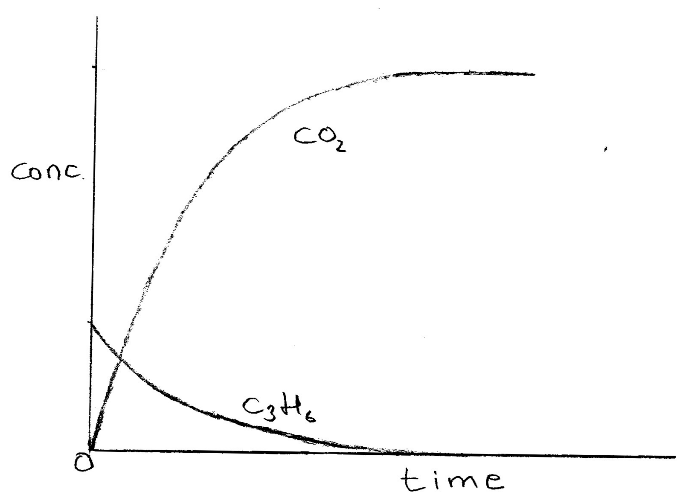
$\therefore$ rate w.r.t. carbon dioxide = (rate w.r.t. to ethane) × (6/2) = $1.7 \times 10^{-4} \times 3 = 5.1 \times 10^{-4}$ Ms^{-1} to 2 significant figures

c) The general rate of reaction. (2)

General rate is given by (rate w.r.t. ethane)/2 or (rate w.r.t. oxygen)/6 or rate w.r.t. (carbon dioxide)/4 or (rate w.r.t. water)/6

$\therefore$ general rate = $(1.7 \times 10^{-4})/2 = 8.5 \times 10^{-5}$ Ms^{-1} to 2 significant figures.

d) Sketch graphs (use one pair of axes) of the concentrations of propene and carbon dioxide against time. (5)



Note that the final concentration of carbon dioxide must be three times the initial concentration of propene because of the stoichiometry of the reaction.

e) If the concentration of propene falls from 0.0140 M to 0.0108 M during the course of 25 s, what is the rate of reaction with respect to propene? (2)

$$Rate = \left(\frac{-\Delta[C_3H_6]}{\Delta t}\right) = \frac{-(0.0108-0.0140)}{25} = 1.28\times10^{-4} = \underline{\underline{1.3\times10^{-4}\,Ms^{-1}}} \text{ to 2 s.f.}$$

(Note that this is an average rate over this interval.)

2) The reaction between nitrogen oxide and hydrogen proceeds according to the following equation at 904°C:

$$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

Experiments to determine the initial rate of reaction at various concentrations of reactants gave results as follows:

Run #	**[NO] / M**	**[H_2] / M**	**Rate / M s^{-1}**
1	0.210	0.122	0.0169
2	0.210	0.244	0.0338
3	0.630	0.122	0.1521

a) Determine the rate law for the reaction, indicating how you arrive at your answer.

Comparing runs 1 and 2, we note that the concentration of NO remains the same whilst the concentration of H_2 doubles. Since the rate also doubles, the reaction must be first order w.r.t. H_2. (3)

Comparing runs 1 and 3, we note that the concentration of H_2 remains the same, whilst that of NO is multiplied by 3. The rate of the reaction is increased by a factor of 0.1521/0.0169 = 9. Since $3^2 = 9$, the order w.r.t. NO must be 2. (3)

Alternatively (just considering the order w.r.t. NO as an example):

$$\frac{R_1}{R_3} = \frac{k([NO]_1{}^m)([H_2]_1)^n}{k([NO]_3{}^m)([H_2]_3{}^n)}$$

Since k and [H_2] remain the same, they cancel, and so:

$$\frac{R_1}{R_3} = \frac{([NO]_1)^m}{([NO]_3)^m} = \left(\frac{[NO]_1}{[NO]_3}\right)^m$$

Putting in values from the table, we have:

$$\frac{0.0169}{0.1521} = \left(\frac{0.210}{0.630}\right)^m$$

$$\therefore \frac{1}{9} = \left(\frac{1}{3}\right)^m$$

$$\therefore m = 2$$

Hence the rate equation is:

$R = k[NO]^2[H_2]$ (2)

b) Calculate the rate constant.

Rearranging the rate equation gives:

$$k = \frac{R}{[NO]^2[H_2]}$$

Selecting one of the runs gives a set of values to solve for *k*. For example, selecting run 1 gives:

$$k = \frac{R}{[NO]^2[H_2]} = \frac{0.0169}{(0.210)^2(0.122)} = 3.141147 = 3.14$$

The units must also be calculated:

$$\frac{R}{[NO]^2[H_2]} = \frac{Ms^{-1}}{M^2M} = M^{-2}s^{-1}$$

Hence the rate constant, k = $\underline{\underline{3.14\ M^{-2}s^{-1}}}$ to 3 sig. figs. (2)

c) Calculate the rate of reaction when [NO] = 0.630 M and $[H_2]$ = 0.244 M. Is this an average or an instantaneous rate? Explain.

Employing the rate equation with the calculated rate constant and the above values of concentrations:

$$R = k[NO]^2[H_2] = 3.141147 \times (0.630)^2 \times (0.244) = 0.3042 \approx \underline{\underline{0.304\,Ms^{-1}}} \text{ to 3 sig. figs.}$$

(2)

This is an instantaneous rate, since the concentrations have fixed values. They can only have fixed values at an instant in time: concentrations change over any interval of time. (2)