SCH4U1	# Rate Law
Energy Changes & Rates of Reaction	

Each reaction has its own equation which gives its rate as a function of ______________________ ____________________. This is called the "_________________________".

Rate laws __________ be determined experimentally. They can't be deduced from the equation alone. To determine rate law, we will measure the rate at different starting concentrations.

Example: $NH_4^+(aq) + NO_2^- \rightarrow N_2(g) + 2H_2O(l)$

Experiment Number	Initial NH_4^+ Concentration (M)	Initial NO_2^- Concentration (M)	Observed Initial Rate (M/s)
1	0.0100	0.200	5.4×10^{-7}
2	0.0200	0.200	10.8×10^{-7}
3	0.0400	0.200	21.5×10^{-7}
4	0.0600	0.200	32.3×10^{-7}
5	0.200	0.0202	10.8×10^{-7}
6	0.200	0.0404	21.6×10^{-7}
7	0.200	0.0606	32.4×10^{-7}
8	0.200	0.0808	43.3×10^{-7}

Compare Experiment number 1 and 2. What do you notice?

Compare Experiment number 5 and 6. What do you notice?

Exponents tell the order of the reaction with respect to each reactant. This reaction is First-order in $[NH_4^+]$, First-order in $[NO_2^-]$ The overall reaction order can be found by adding the exponents on the reactants in the rate law. This reaction is second-order overall.

The general rate expression:

k is ____________ for each reaction, and varies with _________________. A large k indicates a ________ reaction, while a small k indicates a ________ reaction.

Let's try another example. Find the rate expression from the following data:

$$A + B + 2C \rightarrow D + 2E$$

TABLE 5.1.4 RESULTS OF THE A–B–C EXPERIMENTS

Experiment	Initial [A] (mol dm^{-3})	Initial [B] (mol dm^{-3})	Initial [C] (mol dm^{-3})	Initial rate of production of D (mol dm^{-3} s^{-1})
1	0.05000	0.05000	0.05000	3.125×10^{-6}
2	0.1000	0.05000	0.05000	6.250×10^{-6}
3	0.05000	0.1000	0.05000	1.250×10^{-5}
4	0.05000	0.05000	0.1000	3.125×10^{-6}

Start with A. (Hint, look at 1,2)

Next, B. (Hint, look at 1,3)

Lastly, C. (Hint, 1,4)

Units of Rate Constants

Depending on the rate of reaction with respect to the various reactants, rate expressions may be made up of varying combinations of concentration terms. It is important that the units of the rate of reaction are always mol dm^{-3} s^{-1}, so to maintain this the units of the rate constant, k, will vary from reaction to reaction.

Find the units for k if the rate expression is Rate = k [A]

Find the units for k if the rate expression is Rate = k $[A]^2$

1 State the order of reaction with respect to a reactant, R, if:

a the rate of reaction doubles when [R] doubles

b the rate of reaction is unchanged when [R] triples

c the rate of reaction reduces by a factor of four when [R] is halved.

2 Calculate the average rate of production of carbon dioxide in mol dm^{-3} s^{-1} if 0.480 g of carbon dioxide is produced in a 2.00 dm^3 flask in 1.06 minutes, according to the reaction:

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

3 The rate expression for the reaction
$5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$
is Rate = $k[Br^-][BrO_3^-][H^+]^2$

Assuming that all variables except concentration are kept constant, predict the effect on the rate of reaction if:

a $[Br^-]$ doubles (other reactant concentrations unchanged)

b $[BrO_3^-]$ halves (other reactant concentrations unchanged)

c $[H^+]$ triples (other reactant concentrations unchanged)

4 The reaction system $2H_2(g) + 2NO(g) \rightarrow 2H_2O(g) + N_2(g)$ was studied at 800°C, yielding the results listed in the table below.

Experiment	Initial $[H_2]$ (mol dm^{-3})	Initial [NO] (mol dm^{-3})	Initial rate of reaction (mol dm^{-3} s^{-1})
1	0.0030	0.0030	2.25×10^{-3}
2	0.0030	0.0060	9.00×10^{-3}
3	0.0060	0.0030	4.50×10^{-3}

From these results, deduce:

a the order of reaction with respect to H_2

b the order of reaction with respect to NO

c the rate expression

d the overall reaction order

e the value of the rate constant, with units.