Differential Rate Equation

- Gives the rate of the entire reaction
- Determined experimentally

Example

For the reaction, $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$, the following data was obtained. Determine the overall rate for this reaction.

Trial	Initial rate (mol/L·s)	Initial $[NO_2]$ (mol/L)	Initial [CO] (mol/L)
1	0.0050	0.10	0.10
2	0.080	0.40	0.10
3	0.0050	0.10	0.20

Solution:
(1) Take the ratio of the initial rates for the two trials in which only one reactant is changed.

$$\frac{\text{Trial 2 } [NO_2]}{\text{Trial 1 } [NO_2]} = \frac{0.40}{0.10} = 4 \text{ times (quadrupled the concentration)}$$

$$\frac{\text{Trial 2 rate}}{\text{Trial 1 rate}} = \frac{0.080}{0.0050} = 16 \text{ times (rate increases 16 times)}$$

By increasing the concentration four times, the effect on the reaction time is that it is increased by 16. This means that the rate depends on the square of the concentration of NO_2. The reaction is second order with respect to NO_2.

The rate law would be rate = $k[NO_2]^2$

(2) Now take the ratio of the initial rates for trials 1 and 3. These trials have the concentration of CO changing and the concentration of NO_2 constant.

$$\frac{\text{Trial 3 [CO]}}{\text{Trial 1 [CO]}} = \frac{0.20}{0.10} = 2 \text{ times (doubled the concentration)}$$

$$\frac{\text{Trial 3 rate}}{\text{Trial 1 rate}} = \frac{0.0050}{0.0050} = 1 \text{ time (rate does not increase)}$$

By increasing the concentration of CO, the experimental data shows that the reaction rate does not change. It doesn't matter how much CO there is, the rate of reaction does not depend on [CO]. Therefore, the reaction is zero order with respect to CO.

The rate law would be rate = $k[NO_2]^2[CO]^0 = k[NO_2]^2(1) = k[NO_2]^2$