12.3 Determining the Form of the Rate Law

The order is not the same as the coefficients.
If the concentration becomes half and the rate (slope of the tangent at these points) also becomes one half then it is 1^{st} order for that reactant.
Or Rate = ΔAmol/L.s= k [A]

Method of Initial Rates

Differential Rate Laws give data as [**conc**] and initial **rate**
Study information on page 559, Table 12.4

Rate = $k[NH_4^+]^n[NO_2^-]^m$

For NO_2^-, Exp. 1 and 2, the concentration doubles and the rate doubles while NH_4^+ remains same.

$$\frac{Rate2}{Rate1} = \frac{2.7x10^4\, mol/L.s}{1.35x10^4\, mol/L.s} = \frac{k(0.100 mol/L)^n(0.010 mol/L)m}{k(0.100 mol/L)^n(0.0050 mol/L)^m} = 2 = \frac{(0.010 mol/L)^m}{(0.0050 mol/L)^m}$$

or 2.0 = $(2.0)^m$, m=1, The reaction is first order in the reactant NO_2^-
NH_4^+ is also 1^{st} order. Therefore the rate law is: Rate = $k[NH_4^+][NO_2^-]$

To solve for more complicated exponents:

Log each side:
Ex. 6 = $(2.0)^n$, log both sides gives: 0.778 = (0.301)n
Rearrange and solve for n, n= 2.58

The overall order of the reaction is the sum of exponents: n + m = 1 + 1 = 2
Now solve for the rate constant k: (Using any set of data)
From Exp.1: Rate = $1.35x10^{-7}$mol/L·s = k[0.100][0.0050], $k = 2.7x10^{-4}$ L/mol·s

In Class: Sample Ex. 12.1

Solving; Rate = $k[BrO_3^-][Br^-][H^+]^2$, $k = 8.0L^3/mol^3 \cdot s$, overall reaction order=4

Practice Problems:
Set 1: P. 567 #25-30, 71, 79