How to Derive A Zero Order Rate Law

When a reaction is zero order, $n = 0$, which means $Rate = -\frac{d[A]}{dt} = K[A]^0$, where d = change, t = time, $[A]$ = concentraton of A, and K = constant.

So... start with $rate = -\frac{d[A]}{dt} = K[A]^0$

then... $\frac{d[A]}{dt} = -K$

$$d[A] = -Kdt$$

$$\int d[A] = \int -Kdt$$

can you say integral?!

$$[A] - [A]_0 = -kt$$

$$[A] = -kt[A]_0$$

initial Concentration

concentration

rate constant/ concentration

$[A] = -Kt[A]_0$ can fit into the form $y = mx + b$ and can furthermore be graphed!

How to Derive a First Order Rate Law

When a rate is first order, $n = 1$, which means Rate $= -\frac{d[A]}{dt} = k[A]^1$, where d = change, t = time, $[A]$ = concentration, & k = constant.

So! If Rate $= -\frac{d[A]}{dt} = k[A]^1$

then... $\frac{d[A]}{dt} = -k[A]$

$$d[A] = -k[A]\,dt$$

$$\frac{1}{[A]}\,d[A] = -k\,dt$$

Think: $\int \frac{1}{x}\,dx$!

remember to take the ln → $\ln[A] - \ln[A]_0 = -kt$

$$\ln[A] = -kt + \ln[A]_0$$

Which also fits into $y = mx + b$ form! ☺ Yay!

How to Derive a Second Order Rate Law

When a chemical reaction is second order, $n=2$, and Rate $= -\frac{d[A]}{dt} = K[A]^2$, where $d=$ change, $[A]=$ concentration, $K=$ constant, and $t=$ time.

So, start with Rate $= -\frac{d[A]}{dt} = K[A]^2$

then... $\frac{d[A]}{dt} = -K[A]^2$

$$d[A] = -K[A]^2\,dt$$

$$\frac{1}{[A]^2}\,d[A] = -K\,dt$$

$$-\frac{1}{[A]} + \frac{1}{[A_0]} = -Kt$$

think initial

$$\frac{1}{[A]} = Kt - \frac{1}{[A_0]}$$

y = mx + b much?! :) wooo!