10 ~~6~~

You can use the same method for gaseous reactions with pressure**!**

➔ P is directly proportional to Concentration.

Example: The reaction between H_2 and NO occurs according to the equation

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

Six trials of the reaction were carried out. The initial rate of change of pressure for each trial was measured and recorded below.

Trial	Initial Pressure (atm)		Initial Rate
	P_{NO}	P_{H_2}	Δatm min $^{-1}$
I	0.50	2x 0.09	2x 0.025
II	0.50	0.18	0.050
III	0.50	0.27	0.075
IV	0.09 2x	0.80	0.0063 ~4x
V	0.18	0.80	0.025
VI	0.27	0.80	0.056

Based on these results, what is the rate law for this reaction? Justify your answer.

Trials I and II: $2 \times P_{H_2}$ w/ P_{NO} constant doubled the rate, so rxn is 1st order w/ respect to H_2. (NOT P_{H_2}!)

Trials IV and V: $2 \times P_{NO}$ w/ P_{H_2} constant ↑ rate × 4, so rxn is 2nd order w/ respect to NO. (NOT P_{NO}!)

$$\Rightarrow \boxed{\text{rate} = k(P_{NO})^2(P_{H_2}) \quad \text{OR: rate} = k[NO]^2[H_2]}$$

Multiple Choice Practice

1. The reaction below was performed several times, and the following data was gathered. What is the rate law for this reaction?

$$2NO(g) + Br_2(g) \leftrightarrow 2NOBr(g)$$

Trial	$[NO]_{init}$ (*M*)	$[Br_2]_{init}$ (*M*)	Initial Rate of Reaction (*M*/min)
1	0.20 *M*	0.10 *M* 2x	5.20×10^{-3}
2	2x 0.20 *M*	0.20 *M*	1.04×10^{-2}
3	0.40 *M*	0.10 *M*	2.08×10^{-2}

5.2 2x 4x
10.4
20.8

(A) Rate = $k[NO][Br_2]^2$
(B) Rate = $k[NO]^2[Br_2]^2$
(C) Rate = $k[NO][Br_2]$
[✓ (D)] Rate = $k[NO]^2[Br_2]$

1 : 2

2. A chemical reaction A → 2B + C performed with an initial [A] = 5.25 x 10^{-3} is found to have [A] = 4.74 x 10^{-3} M after 2.5 minutes. What is the concentration of B after 2.5 minutes?

(A) 5.10 x 10^{-4} M (B) 9.48 x 10^{-3} M [✓ (C) 1.02 x 10^{-3} M] (D) 4. x 10^{-4} M

$$A \text{ used} = (5.25 - 4.74) \times 10^{-3} = 0.51\,E\text{-}3\,M \times \frac{2 \text{ mol } B}{1 \text{ mol } A} = 1.02\,E\text{-}3\,M\,B$$

3. A(g) + B(g) ↔ 2 C(g)

When the concentration of substance B in the reaction above is doubled, all other factors being held constant, it is found that the rate of the reaction remains unchanged. The most probable explanation for this observation is that

[✓ a)] the order of the reaction with respect to substance B is zero
b) substance B is not involved in the reaction
c) substance B is probably a catalyst, and as such, its effect on the rate of the reaction does not depend on its concentration
d) the reactant with the smallest coefficient in the balanced equation generally has little or no effect on the rate of the reaction

More Practice!!!

1. Consider the reaction represented to the right and the experimental data below. A(g) + B(g) → C(g)

Experiment #	*Initial [A]*	*Initial [B]*	*Initial Rate of Formation of C*
1	0.0030 ↓x2	0.0010 ↓x2	X
2	0.0060	0.0020	?

If the reaction is first order with respect to A and second order with respect to B, what is the initial rate of formation of C in experiment 2 (in terms of X)? Justify your answer.

$$rate_1 = k[O_3][NO] = X$$

$$rate_2 = k(2 \times [O_3])(2 \times [NO]) = 4k[O_3][NO] = \boxed{4X}$$

2. A student wants to determine the rate law for the following reaction: $O_3(g) + NO(g) \rightarrow NO_2(g) + O_2(g)$. When the student runs the reaction with a 0.50 M initial concentration of $O_3(g)$, and a 0.20 M initial concentration of NO(g), they determine the rate of the reaction from these starting conditions is 0.0012 M/s.

What other data the student would need to collect to be able to determine the rate law of the reaction? Be specific!

Repeat experiment twice more:

1) Hold $[O_3] = 0.50 M$, change [NO], measure new rxn rate

2) Hold $[NO] = 0.20 M$, change $[O_3]$, measure new rxn rate