ANSWER KEY.

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5.2 Introduction to Rate Law

AP Chemistry | ANSWER KEY

Rate = $k[A]^m[B]^n$, with rate in M/s. Show all work, including the units of k.

1. The rate law for a reaction is rate = $k[NO]^2[O_2]$.

(a) What is the order in NO? In O_2?

(b) What is the overall order?

(c) What are the units of k?

(a) Second order in NO; first order in O_2 (no exponent written means 1).

(b) 2 + 1 = 3, third order overall.

(c) M/s = $k(M^3)$, so k is in $M^{-2}\ s^{-1}$.

2. A reaction between A and B is zero order in A and first order in B.

(a) Write the rate law.

(b) What are the units of k?

(c) What happens to the rate if [A] is doubled?

(a) rate = k[B] (the $[A]^0$ term equals 1)

(b) First order overall: M/s = k(M), so k is in s^{-1}.

(c) No change: $2^0 = 1$.

3. The rate law is rate = $k[A][B]^2$. By what factor does the rate change when:

(a) [A] is tripled?

(b) [B] is cut in half?

(c) [A] and [B] are both doubled?

(d) The temperature is raised. Does k change? Explain.

(a) $3^1 = 3$, the rate triples.

(b) $(½)^2 = ¼$, the rate is one fourth as fast.

(c) $2^1 \times 2^2 = 2 \times 4 = 8$, the rate is 8 times faster.

(d) Yes. k gets larger at a higher temperature, so the reaction is faster.

4. For $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$, a student writes rate = $k[NO]^2[O_2]$ "because the coefficients are 2 and 1." Explain what is wrong with the student's reasoning.

Orders come from experimental data, not from the coefficients in the balanced equation. The rate law might turn out to be this one, but the student can only know that by measuring how the rate changes when each concentration changes.

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5. Use the data for A + B → C.

Trial	[A] (M)	[B] (M)	Initial rate (M/s)
1	0.15	0.10	3.0×10^{-4}
2	0.30	0.10	6.0×10^{-4}
3	0.15	0.30	2.7×10^{-3}

(a) Find the order in A and the order in B. Show which trials you used.

(b) Write the rate law.

(c) Calculate k with units.

(d) Find the rate when [A] = 0.20 M and [B] = 0.20 M.

(a) Trials 1 and 2: [A] × 2, rate × 2 (6.0/3.0), so $2^m = 2$ and $m = 1$. Trials 1 and 3: [B] × 3, rate × 9 (2.7×10^{-3} / 3.0×10^{-4}), so $3^n = 9$ and $n = 2$.

(b) rate = $k[A][B]^2$

(c) k = (3.0×10^{-4} M/s) / ((0.15 M)(0.10 M)^2) = (3.0×10^{-4}) / (1.5×10^{-3}) = 0.20 M^{-2} s^{-1}

(d) rate = (0.20 M^{-2} s^{-1})(0.20 M)(0.20 M)^2 = 1.6×10^{-3} M/s

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6. Use the data for 2X + Y → Z.

Trial	[X] (M)	[Y] (M)	Initial rate (M/s)
1	0.020	0.050	4.8×10^{-5}
2	0.040	0.050	1.9×10^{-4}
3	0.020	0.100	4.8×10^{-5}

(a) Write the rate law. Justify each order with the data.

(b) Calculate k with units.

(c) The coefficient of Y is 1, but its order is not 1. Explain how this is possible.

(a) Trials 1 and 2: [X] × 2, rate × 4 (1.9×10^{-4} / 4.8×10^{-5} ≈ 4), so second order in X. Trials 1 and 3: [Y] × 2, rate × 1, so zero order in Y. rate = $k[X]^2$

(b) k = (4.8×10^{-5} M/s) / (0.020 M)^2 = (4.8×10^{-5}) / (4.0×10^{-4}) = 0.12 M^{-1} s^{-1}

(c) Orders come from experimental data, not from coefficients. The data show that changing [Y] does not change the rate.

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7. Name a property you could measure over time to follow each reaction. Explain your choice.

(a) Purple MnO_4^-(aq) fades to a colorless product.

(b) $2NO_2(g) \rightarrow 2NO(g) + O_2(g)$ in a sealed, rigid container.

(a) Absorbance (color) with a spectrophotometer. Absorbance drops as [MnO_4^-] drops ($A = \varepsilon bc$).

(b) Gas pressure. 2 mol of gas become 3 mol, so the pressure rises as the reaction goes ($P \propto n$).

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