South Pasadena • AP Chemistry

Name ______________________________________

Period ___ Date ___/___/___

15 • Chemical Kinetics

RATE LAWS

Big Ideas (This is Section 15.3 in the textbook, pages 694 – 700):

- One factor that speeds up a reaction is CONCENTRATION.
- Higher concentrations of reactants means there will be **more collisions** between the particles that need to collide to cause a reaction.
- **How much** the rate of a reaction increases gives clues to the step-by-step **mechanism** of a reaction.
- A Rate Law has a **standard format**. For a reaction such as A + B → C + D, the rate law is:

  $$\text{Rate} = k[A]^x[B]^y$$

  (x = the order of the reaction with respect to A, y = the order with respect to B and *k* = the rate constant)

1. Consider the initial rate data for the reaction, $A + B \rightarrow C$:

	[A]	**[B]**	**Rate (mol·L^{-1}·s^{-1})**
Expt 1	0.10	0.10	2.0
Expt 2	0.10	0.20	4.0
Expt 3	0.20	0.20	32.0

   a. What is the order of reactant A?
   b. What is the order of reactant B?
   c. Write the rate equation for the reaction.
   d. Calculate the value of the specific rate constant, *k*.

2. Consider the initial rate data for the reaction, $3X + 2Y \rightarrow Z$:

	[X]	**[Y]**	**Rate (mol·L^{-1}·s^{-1})**
Expt 1	0.10	0.20	4.0
Expt 2	0.10	0.10	1.0
Expt 3	0.20	0.20	4.0
Expt 4	0.25	0.25	

   a. What is the order of reactant X?
   b. What is the order of reactant Y?
   c. What is the overall order of the reaction?
   d. Write the rate equation for the reaction.
   e. Calculate the value of the specific rate constant, *k*.

   f. Calculate the rate when the [X] and [Y] = 0.25 mol·L^{-1}.

3. Consider the reaction: $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

The following data were obtained from three experiments using the method of initial rates:

	Initial [NO] mol L^{-1}	Initial [O_2] mol L^{-1}	Initial rate NO mol $L^{-1}s^{-1}$
Experiment 1	0.010	0.010	2.5×10^{-5}
Experiment 2	0.020	0.010	1.0×10^{-4}
Experiment 3	0.010	0.020	5.0×10^{-5}

a. Determine the order of the reaction for each reactant.
b. Write the rate equation for the reaction.
c. Calculate the rate constant.

d. Calculate the rate (in mol $L^{-1}s^{-1}$) at the instant when [NO] = 0.015 mol L^{-1} and [O_2] = 0.0050 mol L^{-1}

e. At the instant when NO is reacting at the rate 1.0×10^{-4} mol $L^{-1}s^{-1}$, what is the rate at which O_2 is reacting and NO_2 is forming?

4. The reaction of ^{t}butyl-bromide $(CH_3)_3CBr$ with water is represented by the equation:

$$(CH_3)_3CBr + H_2O \rightarrow (CH_3)_3COH + HBr$$

The following data were obtained from three experiments using the method of initial rates:

	Initial [$(CH_3)_3CBr$] mol L^{-1}	Initial [H_2O] mol L^{-1}	Initial rate mol $L^{-1}min^{-1}$
Experiment 1	5.0×10^{-2}	2.0×10^{-2}	2.0×10^{-6}
Experiment 2	5.0×10^{-2}	4.0×10^{-2}	2.0×10^{-6}
Experiment 3	1.0×10^{-1}	4.0×10^{-2}	4.0×10^{-6}

a. What is the order with respect to $(CH_3)_3CBr$?
b. What is the order with respect to H_2O?
c. What is the overall order of the reaction?
d. Write the rate equation.
e. Calculate the rate constant, k, for the reaction.

5. Some problems can be solved "by inspection", but you can also treat them **mathematically**.

Consider the initial rate data for the reaction, $A + B \rightarrow C$:

	[A]	[B]	Rate ($mol \cdot L^{-1} \cdot s^{-1}$)
Expt 1	0.10	0.10	3.0×10^{-5}
Expt 2	0.10	0.20	6.0×10^{-5}
Expt 3	0.20	0.30	3.6×10^{-4}

Use the rate law to write a reaction for Experiment 2 / Experiment 1

$$\text{Rate} = k\ [A]^x\ [B]^y$$

Experiment 2:

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Experiment 1:

Repeat the procedure with Experiment 3 / Experiment 1

$$\text{Rate} = k\ [A]^x\ [B]^y$$

Experiment 3:

——————————

Experiment 1:

a. What is the order of reactant A?
b. What is the order of reactant B?
c. Write the rate equation for the reaction.
d. Calculate the value of the specific rate constant, *k*.

6. The reaction $2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$ was studied at 904 °C, and the data in the table were collected.

	Initial [NO] $mol\ L^{-1}$	Initial [H_2] $mol\ L^{-1}$	Initial rate N_2 $mol\ L^{-1}s^{-1}$
Experiment 1	0.420	0.122	0.136
Experiment 2	0.210	0.122	0.0339
Experiment 3	0.105	0.488	0.0339

a. Determine the order of the reaction for each reactant.
b. Write the rate equation for the reaction.
c. Calculate the rate constant at 904 °C.
d. Find the rate of appearance of N_2 at the instant when [NO] = 0.350 M and [H_2] = 0.205 M.

FRQ from the 2010 AP Exam
You should be able to do the entire question, but for this worksheet, simply do parts (b) and (c).

$$8\ H^+(aq) + 4\ Cl^-(aq) + MnO_4^-(aq) \rightarrow 2\ Cl_2(g) + Mn^{3+}(aq) + 4\ H_2O(l)$$

3. $Cl_2(g)$ can be generated in the laboratory by reacting potassium permanganate with an acidified solution of sodium chloride. The net-ionic equation for the reaction is given above.

   (a) A 25.00 mL sample of 0.250 M NaCl reacts completely with excess $KMnO_4(aq)$. The $Cl_2(g)$ produced is dried and stored in a sealed container. At 22°C the pressure of the $Cl_2(g)$ in the container is 0.950 atm.

   (i) Calculate the number of moles of $Cl^-(aq)$ present before any reaction occurs.

   (ii) Calculate the volume, in L, of the $Cl_2(g)$ in the sealed container.

   An initial-rate study was performed on the reaction system. Data for the experiment are given in the table below.

Trial	$[Cl^-]$	$[MnO_4^-]$	$[H^+]$	Rate of Disappearance of MnO_4^- in $M\ s^{-1}$
1	0.0104	0.00400	3.00	2.25×10^{-8}
2	0.0312	0.00400	3.00	2.03×10^{-7}
3	0.0312	0.00200	3.00	1.02×10^{-7}

   (b) Using the information in the table, determine the order of the reaction with respect to each of the following. Justify your answers.

   (i) Cl^-

   (ii) MnO_4^-

   (c) The reaction is known to be third order with respect to H^+. Using this information and your answers to part (b) above, complete both of the following:

   (i) Write the rate law for the reaction.

   (ii) Calculate the value of the rate constant, k, for the reaction, including appropriate units.

   (d) Is it likely that the reaction occurs in a single elementary step? Justify your answer.