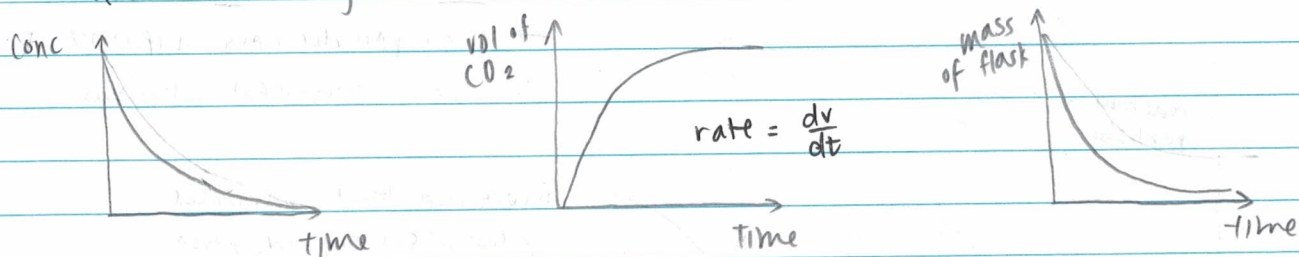


CHEMICAL KINETICS

rates of reaction & collision theory

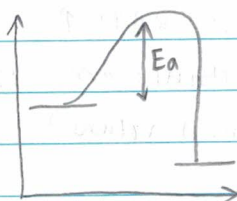
rate of reaction is the INCREASE in concentration of products over unit time
or the DECREASE in conc. of reactants per unit time
it is measured in $\text{mol dm}^{-3} \text{s}^{-1}$

you can measure the rate of reaction by monitoring changes in mass, volume, and colour e.g. $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$



but! for a reaction to occur... **3** conditions.

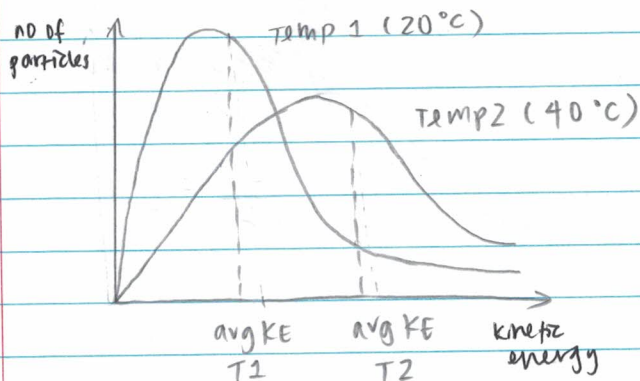
- I. particles must collide
- II. must collide with appropriate geometry & orientation
- III. must collide with sufficient energy [E_a , activation energy]



- E_a is the minimum amount of energy required
- Any factor increases the A. Frequency of collisions
B. energy of collisions
= make the reaction go FASTER

Maxwell-Boltzmann distribution

- moving particles in gas/liquid don't move at the same speed. some more fast, others more slower. ∴ i.e. some have more KE, some have less
- the MBD curve shows the distribution of kinetic energies.



as temperature increases, the area under the curve does NOT change (same no. of particles).

BUT... more particles have HIGHER KE so... the curve BROADENS.

rate expression & reaction mechanism



altering $[A]$ or $[B]$ may affect the rate of reaction BUT have different effects - the mathematical way to determine how $[A]$ and $[B]$ affect rate is:

$$\text{rate} = k[A]^x[B]^y \quad \text{where } k = \text{constant} (\uparrow \text{ as temp } \uparrow)$$

x is the order of reaction with respect to A

y is the order of reaction with respect to B.

$$[\text{overall order} = x + y]$$

order definition

0 rate = constant when $[] \times 2$

1 rate = doubles when $[] \times 2$

2 rate = $4x$ when $[] \times 2$

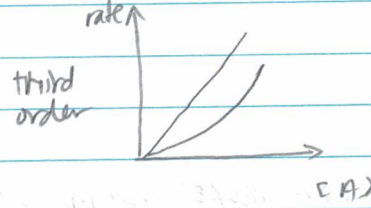
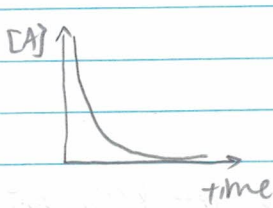
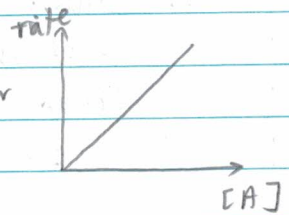
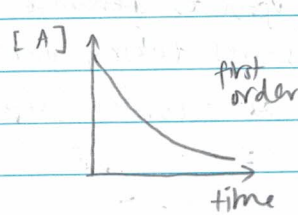
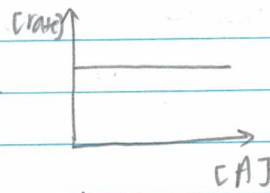
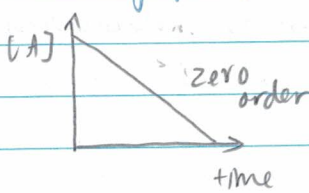
units

s^{-1}

$k[A]$

$dm^3 mol^{-1} s^{-1} k[A][B]$

graphs of rate constant



rate expression through data

	$[A]$	$[B]$	rate of formation	
1	1×10^{-3}	6×10^{-3}	3×10^{-3}	$k[A] = r$
2	2×10^{-3}	6×10^{-3}	6×10^{-3}	$\therefore r = k[A][B]^2$
3	6×10^{-3}	1×10^{-3}	0.5×10^{-3}	$k[B]^2 = r$
4	6×10^{-3}	2×10^{-3}	2.0×10^{-3}	$dm^6 mol^{-2} s^{-1}$

← find k by substituting numbers

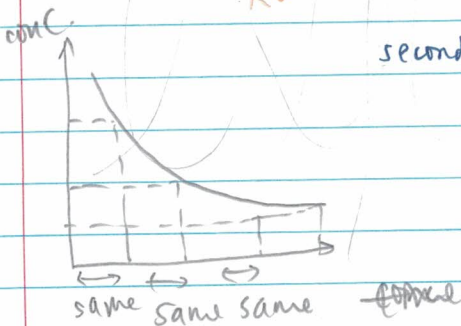
half life

amount of time it takes for reaction concentration to halve.

$$t_{1/2} = \frac{\ln 2}{k}$$

first order: half life is independent of concentration of A. \therefore HALF LIFE IS CONSTANT

second order: half life will double, $t_{1/2}$ is dependant on conc of A.



TOPIC 6 KINETICS (EXAM QUESTIONS)

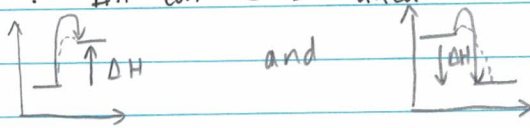
what is important in determining a successful collision?

✓ KE ✓ orientation ✗ collision frequency (rate! not success!)

catalysed and uncatalysed reaction...

? $\Delta H_{cat} < \Delta H_{uncat}$?

? $\Delta H_{cat} = \Delta H_{uncat}$? ← correct! $\Delta H \rightarrow$ change



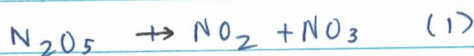
does not change!



no effect, as sol is reacting, NOT solid.

what is rate (if given T, time?)

1 ÷ time (s^{-1})



what is rate exp of (2)?

rate = $k[N_2O_5]$ NOT $[N_2O_5][NO_3]$ as NO_3 is from (1)

how many particles collide at once with $2NO + 2H_2 \rightarrow N_2 + 2H_2O$

4 = N, O, H, H

what is the significance of A in the rate reaction

geometric requirement of reaction

rate ↑ with ↑ °K. this is best explained by... increase in...

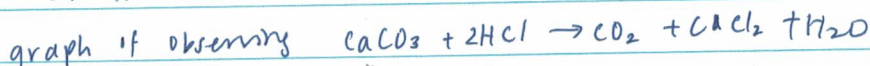
A. avg energy B. no. of collisions **C. no energy $\geq E_a$** D. activation

↳ best = increase in avg no. particles with $\geq E_a$

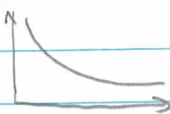
which affect rate?

I. collisions/sec II. energy of collisions III. geometry of collision

⇒ ALL three! RMS III: catalyst!



MASS
LOST



→ "mass" is incorrect.

$r = k[P][Q]$, if volume is $1/4$ of initial, what is new rate?

$k[1][1] = k$, $k[4][4] = 16k$ $k:16k$.

concentration $[NO] = 0.1$ $[H_2] = 0.1$ $r = 2.53 \times 10^6$, what is k ?

→ rate mechanism is $[NO]^2 [H_2]$

DO NOT PUT $[0.1][0.1]$ it is $[0.1]^2 [0.1]$!

activation energy is in $kJ\ mol^{-1}$