I am able to:
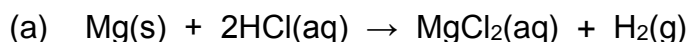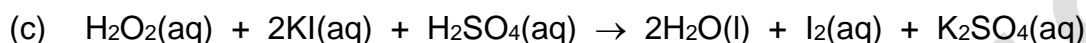
1   Define *rate of reaction* and give the usual units for it.

Change in concentration of reactants/products per unit time   e.g. mol dm^{-3} s^{-1}

2   Describe a method that could be used to measure the rate of each of the following reactions

   (a) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

   measure volume of gas/change in mass (although H_2 will not give much mass change)/change of pH per unit time

   (b) $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$ measure volume of gas/change in mass per unit time

   (c) $H_2O_2(aq) + 2KI(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + I_2(aq) + K_2SO_4(aq)$

   use a colorimeter to measure how the intensity of the colour increases per unit time

3   Explain two different methods which could be used to determine the rate of the reaction between $CaCO_3$ and dilute hydrochloric acid.

   (a)   use a gas syringe/gas burette/measuring cylinder
   Record volume of gas every 10-15 s
   Suitable experimental set-up (see diagram)
   Fixed temperature
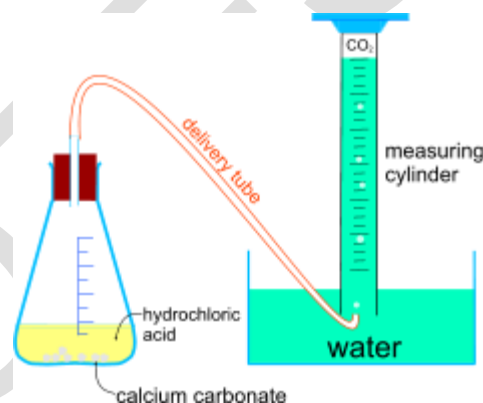
   (b)   use balance
   Record mass every 10-15 s
   Suitable experimental set-up (see diagram)
   Fixed temperature
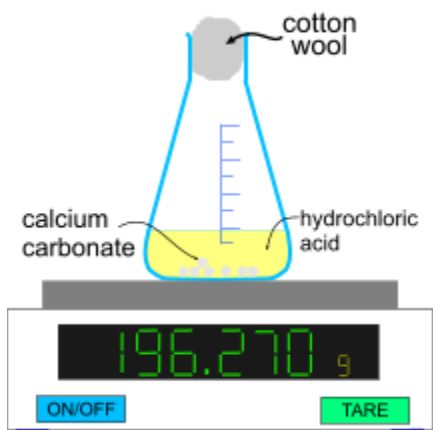
4   Use the graph shown (for the reaction between $CaCO_3$ and HCl)

   (a)   to explain how the rate of reaction varies with time
   Graph steepest at start - Rate initially high –
   highest concentration of HCl – most frequent
   collisions. As HCl used up, lower
   concentration, less frequent collisions –
   gradient of graph decreses.
   From 90s graph flat – reaction has finished –
   one or both reactants have been used up.
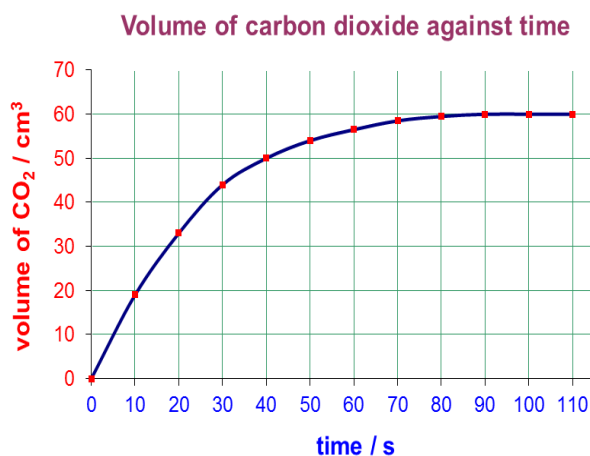
   (b)   to calculate the initial rate of reaction   draw tangent at initial point – rate slightly greater than 2 cm^3 s^{-1}

   (c)   to work out the average rate of reaction for the first 90 s.   60/90 = 0.67 cm^3 s^{-1}

5   Explain the *collision theory*.

For a reaction to occur particles must collide in the correct orientation with $E > E_a$

6   Explain the effects of changing: (a) particle size (b) concentration on the rate of a reaction.

   (a) smaller particles of solid – higher surface area – more particles exposed on the surface – more frequent collisions – faster rate

   (b) more particles in a certain volume – more frequent collisions – faster rate

7   Explain how increasing the pressure would affect the rate of the reaction between hydrogen and iodine:
        $H_2(g) + I_2(g) \rightarrow 2HI(g)$
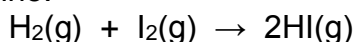
   Rate increases – more particles in a certain volume – particles collide more frequently

8   Sketch graphs showing the effect of different reaction conditions.

   The graph shown is for 25.0 cm^3 of 1 mol dm^{-3} HCl reacting with excess calcium carbonate (medium lumps) at 25 °C.

   Sketch graphs for:

   (a) 25.0 cm^3 of 2 mol dm^{-3} HCl reacting with excess calcium carbonate (medium lumps) at 25 °C. Graph steeper initially, levels off sooner, 2x volume of gas produced

   (b) 50.0 cm^3 of 0.5 mol dm^{-3} HCl reacting with excess calcium carbonate (medium lumps) at 25 °C. Graph less steep initially, levels off later, same volume of gas produced

   (c) 25.0 cm^3 of 1 mol dm^{-3} HCl reacting with excess calcium carbonate (medium lumps) at 35 °C.   Graph steeper initially, levels off sooner, same volume of gas produced

   (d) 25.0 cm^3 of 1 mol dm^{-3} HCl reacting with excess calcium carbonate (powder) at 25 °C. Graph steeper initially, levels off sooner, same volume of gas produced
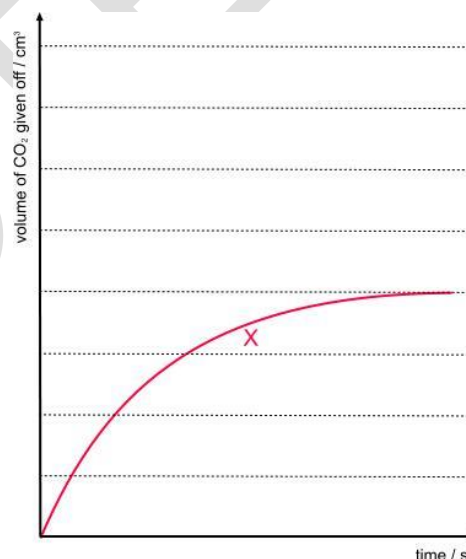
9   State that the average energy of particles in a gas is proportional to the temperature in K.

10   Explain how the average kinetic energy and speed of nitrogen molecules changes as the temperature is increased from 25°C to 50°C
   Increases by a factor of 333/298 – does not double (only for doubling temperature in K)

11   Explain what is meant by *activation energy*
   The minimum amount of energy that colliding particles need for a collision to be successful/result in a reaction

12   Sketch the Maxwell-Boltzmann energy distribution.
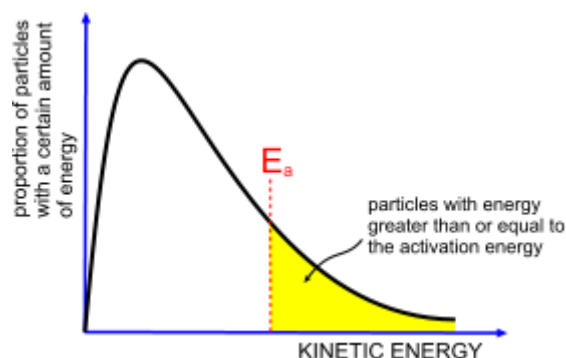
13 Explain why the rate of a chemical reaction increases as the temperature increases.

More particles have E≥E_a therefore greater chance that a collision will results in a reaction, therefore more successful collisions per second. Much smaller factor: particles collide more frequently.

14 Explain what a catalyst is.　　　　A substance that speeds up a reaction but is not used up

15 Explain how a catalyst speeds up a chemical reaction.

Provides an alternative pathway of lower activation energy – greater proportion of particles have E≥E_a

16 Sketch energy profiles for exo- and endothermic reactions with and without catalysts.

| Exothermic | Endothermic |
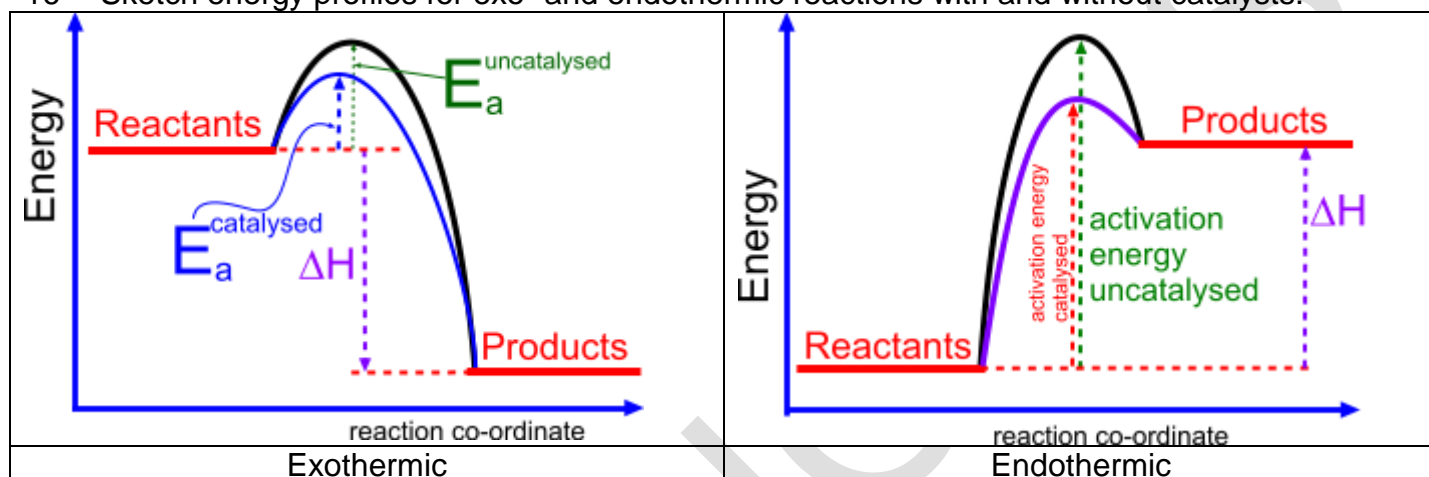
17 Explain the terms: *rate constant*, *order of reaction*, *overall order of reaction*.

- rate constant - constant of proportionality relating the concentrations in the experimentally determined rate equation to the rate of a chemical reaction.
- order of reaction with respect to a particular reactant is the power of the reactants concentration in the experimentally determined rate equation.
- overall order of reaction is the sum of the powers of the concentration terms in the experimentally determined rate equation.

18 Write the rate equation and calculate the rate constant (including units) for the reaction A + B → P + Q

[A] /mol dm^{-3}	[B] /mol dm^{-3}	Initial Rate /mol dm^{-3} s^{-1}
1.0	1.0	2.0×10^{-3}
2.0	1.0	4.0×10^{-3}
2.0	3.0	36×10^{-3}

Rate = $k[A][B]^2$　　k = 2.0×10^{-3} mol^{-2} dm^6 s^{-1}

19 Write the rate equation and calculate the rate constant (including units) for the reaction:
2Z + 3Y → W + X
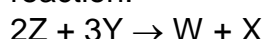
| Expt.	[Z] /mol dm^{-3}	[Y] /mol dm^{-3}	Rate /mol dm^{-3} s^{-1}
1	2.0×10^{-3}	2.0×10^{-3}	0.050
2	8.0×10^{-3}	2.0×10^{-3}	0.800
3	1.0×10^{-2}	6.0×10^{-3}	3.750

Rate = $k[Z]^2[Y]$　　k = 6.25×10^6 mol^{-2} dm^6 s^{-1}

20 Write down possible units for the rate constants for zero, first and second order reactions.

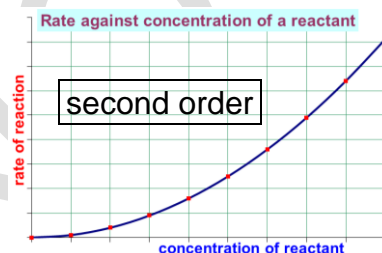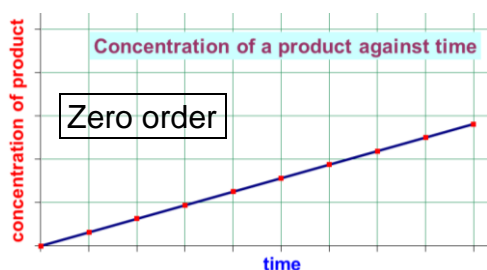| Zero order	conc time^{-1}	mol dm^{-3} s^{-1}
First order	time^{-1}	min^{-1}
Second order	conc^{-1} time^{-1}	mol^{-1} dm^3 h^{-1}

21   Draw graphs of concentration against time for zero, first and second order reactions.
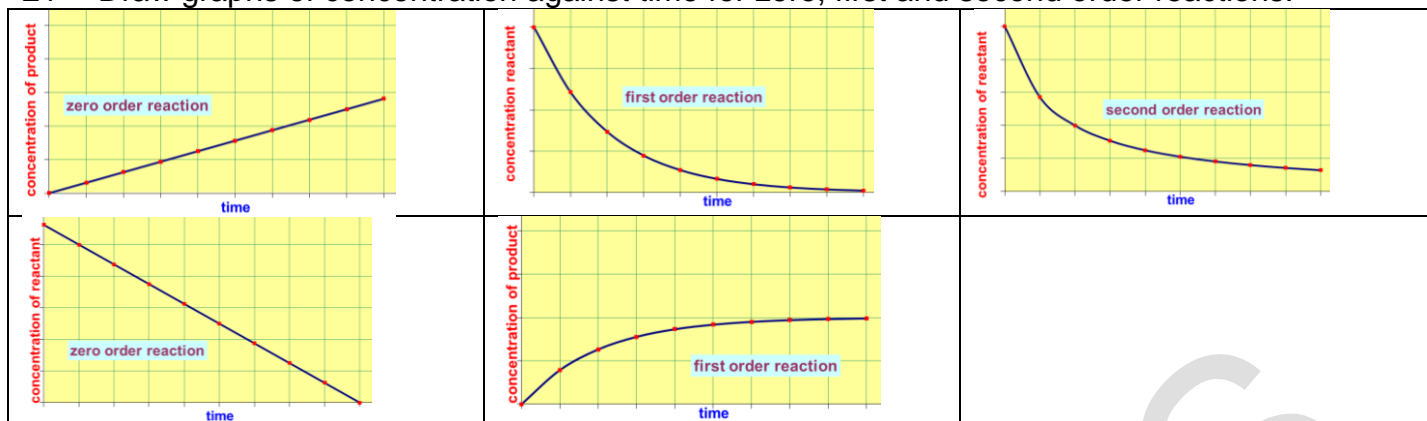
22   Draw graphs of rate against concentration for zero, first and second order reactions.
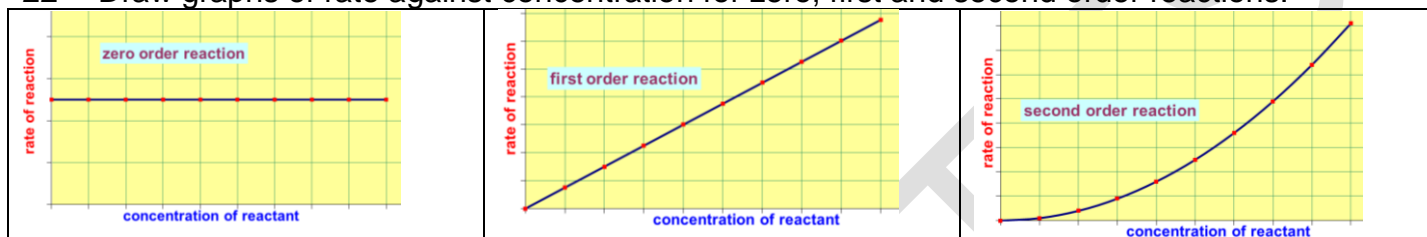
23   Work out whether the graphs represent zero, first or second order reactions.

24   Explain what is meant by the *rate determining step* for a reaction

Slowest step/step with highest activation energy in a reaction mechanism

25   Work out the rate expression for the reaction P + 2Q → R + S given the mechanism:

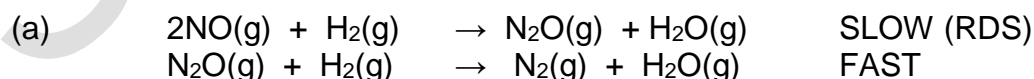Q + P → Y + S   SLOW   Q + Y → R   FAST

Rate = k[Q][P]

26   Deduce whether mechanisms are correct or incorrect
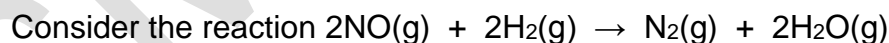
Consider the reaction $2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$
The rate equation is Rate = $k[NO]^2[H_2]$

Explain why each of the following mechanisms is incorrect:

(a) $2NO(g) + H_2(g)$ → $N_2O(g) + H_2O(g)$ SLOW (RDS)
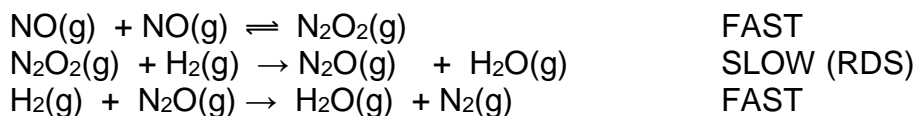 $N_2O(g) + H_2(g)$ → $N_2(g) + H_2O(g)$ FAST

Incorrect – highly unlikely that 3 particles will collide at the same time
Maximum number of particles colliding in 1 step should be 2

(b) $H_2(g) + NO(g) \rightleftharpoons H_2O(g) + N(g)$ FAST
           $H_2(g) + N(g) \rightarrow NH_2(g)$ SLOW (RDS)
           $NH_2(g) + NO(g) \rightarrow H_2O(g) + N_2(g)$ FAST
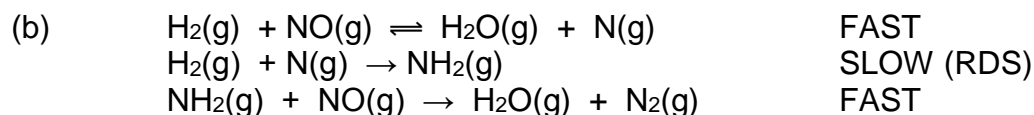
Incorrect – 2 H_2 involves up to and including the RDS therefore should have $[H_2]^2$ is the rate equation
$[NO]^2$ in the rate equation therefore NO should appear twice in the mechanism up to and including the RDS but it only appears once

27 Deduce two possible mechanisms for the reaction in 26

$H_2(g) + NO(g) \rightleftharpoons H_2NO(g)$ FAST
$H_2NO(g) + NO(g) \rightarrow N_2(g) + H_2O_2(g)$ SLOW (RDS)
$H_2(g) + H_2O_2(g) \rightarrow 2H_2O(g)$ FAST

$NO(g) + NO(g) \rightleftharpoons N_2O_2(g)$ FAST
$N_2O_2(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$ SLOW (RDS)
$H_2(g) + N_2O(g) \rightarrow H_2O(g) + N_2(g)$ FAST
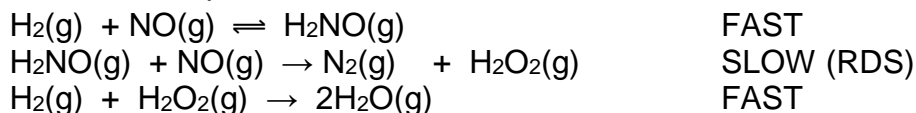
28 Explain what is meant by the term *molecularity* and the difference between molecularity and order.

Molecularity - the number of molecules (or ions) that take part in a particular step (usually the rate determining step) in a mechanism.
Order – an experimentally determined quantity relating rate to concentration – the power of the concentration in the rate equation. Molecularity is deduced from the rate equation, it is not determined experimentally.

29 Explain the difference between a species that acts as a *catalyst* and one that is an *intermediate* in a reaction. Classify X and Y as catalyst or intermediate in the mechanism:

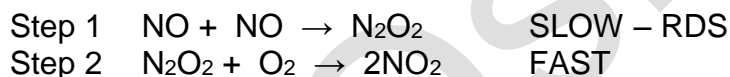Step 1: A + X → Y   FAST   Step 2: Y + B → C + X   SLOW

Catalyst is used up in a step and then produced again
An intermediate is produced in a step and then used up again
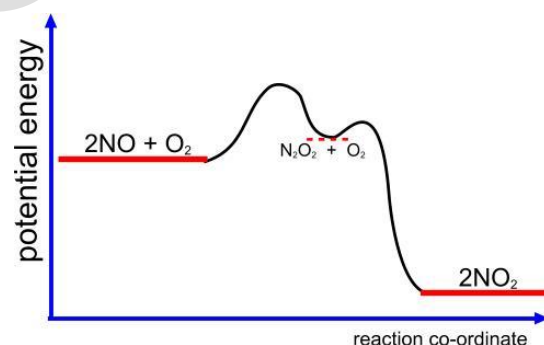X is a catalyst
Y is an intermediate

30 Deduce which is the rate determining step and the mechanism from a potential energy profile.

Step 1 $NO + NO \rightarrow N_2O_2$ SLOW – RDS
Step 2 $N_2O_2 + O_2 \rightarrow 2NO_2$ FAST

Step 1 is the RDS – higher activation energy.

31 Explain why the rate equation cannot be deduced from the stoichiometric equation.

The reaction may not occur in a single step – the mechanism is not known

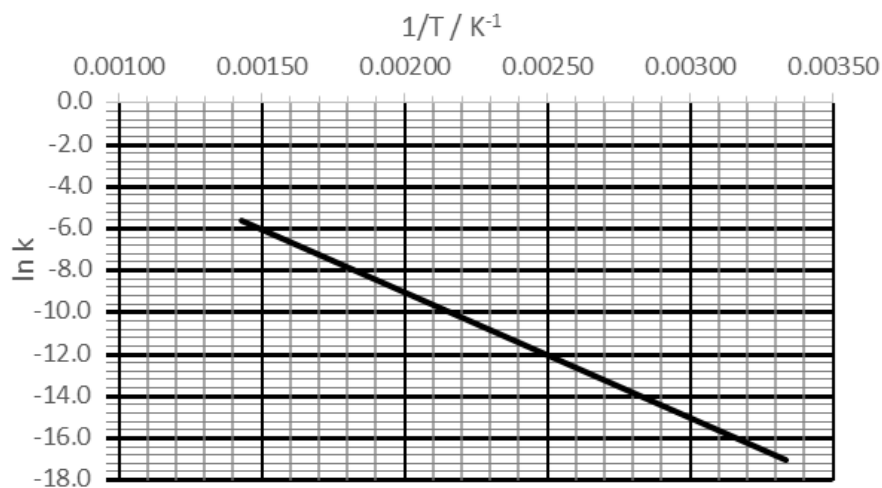32 Explain how the value of the rate constant changes as the temperature increases.

Increases exponentially

33 Explain how the Arrhenius equation can be used to determine the activation energy

- Conduct a series of experiments measuring the rate at a range of temperatures
- Calculate a rate constant (k) for each temperature
- Plot a graph of ln k (y axis) against 1/T (x axis), where T is the absolute temperature (in Kelvin) - this graph should be a straight line.
- Determine the gradient of the graph (this will be a negative value with units K)
- Gradient of the graph is equal to $-E_a/R$, where R is the gas constant.
- Cancel the negative signs and multiply the gradient by R to get the value of E_a in J mol^{-1}.
- Divide the value obtained by 1000 to get the activation energy in kJ mol^{-1}

34   Determine the activation energy from graphical data

50 kJ mol^{-1}

35   Determine the value of the frequency factor using data from the graph and suggest a possible unit for A if the reaction is first order.

20.9 s^{-1} (same units as the rate constant)

36   Explain the relationship between the frequency factor and the complexity of the molecules colliding.

The more complex the molecule the lower the frequency factor – the more complex the molecule, the lower probability that a collision will occur with the correct orientation.