Lesson 1 Rates of reaction

In order for a reaction to occur reactant particles must collide with enough energy to break bonds. These collisions are known as "***Fruitful collisions***".

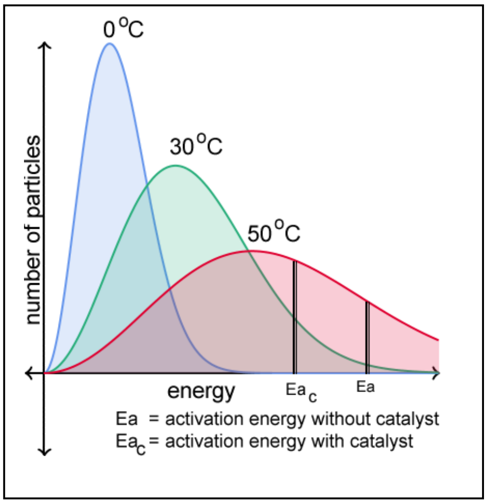
The rate of a chemical reaction depends on

- Surface area – the greater the surface area the more reactant particles that are exposed to undergo collisions.

- Concentration – the more particles per unit volume the greater the chance of fruitful collisions.

- Gas pressure – increasing the pressure of gas by decreasing the volume of the container at constant temperature is the same as increasing the concentration of the gas.

- Temperature – The higher the temperature the greater the <u>average</u> kinetic energy of the reactant particles hence greater number of collisions and greater energy at impact resulting in a greater proportion of these collisions being fruitful.
  On the right is a Maxwell-Boltzmann energy distribution curve at three different temps.
  Note – A molecule in a gas at any temperature could have any amount of Kinetic energy. Molecules of a gas at a high temperature have a higher average kinetic energy than the molecules of a gas at a low temperature, however, in both cases there will be molecules with very little kinetic energy and molecules with high kinetic energy.

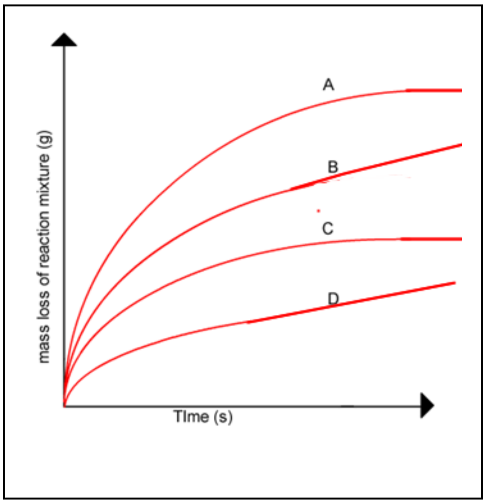


- A catalyst - decreases the activation energy required for reactant particles to react. That means a greater proportion of particles have the required activation energy at the given temperature to undergo a fruitful collision. Catalysts, therefore, increase the proportion of molecular collisions that are successful and <u>**not**</u> the number of collisions. All the factors mentioned above increase the frequency of collisions.

1) A student conducts four experiments by placing solid $CaCO_3$ in a solution of HCl . He is supplied with $CaCO_3$ powder and $CaCO_3$ chips, as well as 0.10 M HCl and 10.0 M HCl.

$2HCl(aq) + CaCO_3(s) \rightarrow CO_2(g) + CaCl_2(aq) + H_2O(l)$

The mass loss of the reaction mixture is measured and plotted on the set of axes on the right.



a) In experiment A and B the student uses $CaCO_3$ powder with 0.10 M HCl solutions.

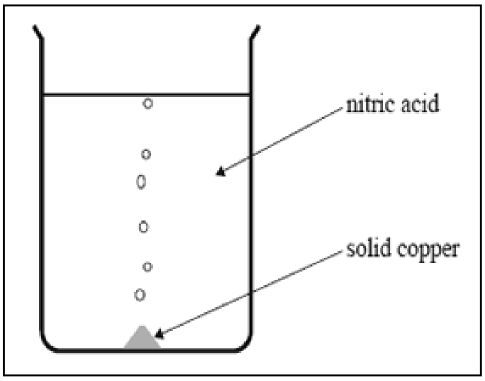
What one factor could have caused the difference in graph A and graph B. Explain

b) In experiment D and C the student used the 0.10 M HCl solution at 25°C. What factor could have caused the difference in graph D and graph C if the type of powder was the same in both experiments.

2) Copper metal reacts with nitric acid to produce NO_2 gas.

Which one of the following will not increase the rate of the reaction shown below? Explain why.

$4HNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$



A. decreasing the size of the solid copper particles

B. increasing the temperature of HNO_3 by 20 °C

C. increasing the concentration of HNO_3

D. allowing NO_2 gas to escape

3) In the above reaction, the number of successful collisions per second is a small fraction of the total number of collisions. The major reason for this is that

A. the nitric acid is ionised in solution.

B. some reactant particles have too much kinetic energy.

C. the kinetic energy of the particles is reduced when they collide with the container's walls.

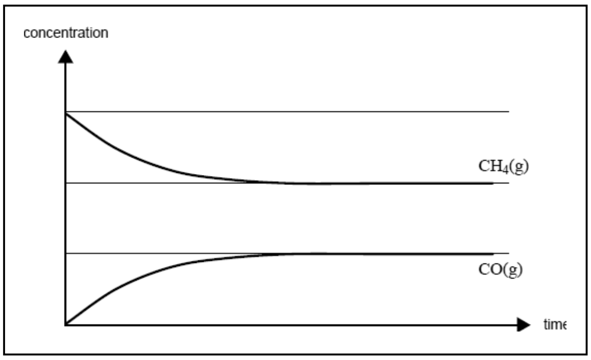
D. not all reactant particles have the minimum kinetic energy required to initiate the reaction.

4) Explain how the number of fruitful collisions per second, in 3) above, can be increased.

5) Carbon monoxide and hydrogen can be produced from the reaction of methane with steam according to the equation

$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$; $\Delta H = +206$ kJ mol^{-1}

Some methane and steam are placed in a closed container and allowed to react at a fixed temperature. The graph on the right shows the change in concentration of methane and carbon monoxide as the reaction progresses.



a) On the graph above, draw a line to show the change in concentration of hydrogen gas as the reaction progresses. Label this line.

b) On the graph above, draw two lines to show how the usage of methane and the production of carbon monoxide would differ over time in the presence of a catalyst. Label these lines.

6) The two statements below give possible explanations for changes that occur when the temperature of a reaction mixture is increased.

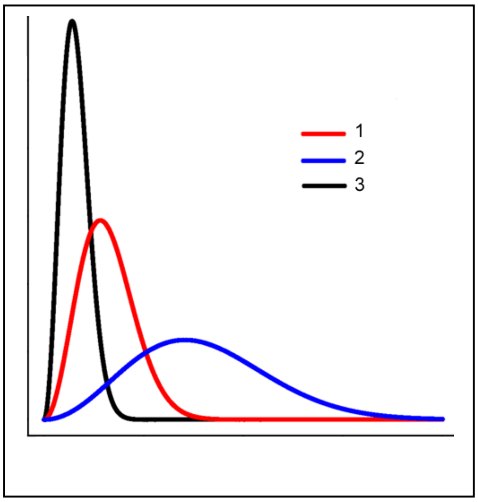
i. At a higher temperature, particles move faster and the reactant particles collide more frequently.

ii. At a higher temperature, more particles have energy greater than the activation energy.

Which statement best explains why the observed reaction rate is greater at higher temperatures and why the other is not significant in determining the rate?

7) Explain how a catalyst increases the rate of a reaction using the Maxwell-Boltzmann distribution curve. Draw a diagram as part of your explanation.

8) Consider the Maxwell-Boltzmann distribution curves shown on the set of axes on the right.

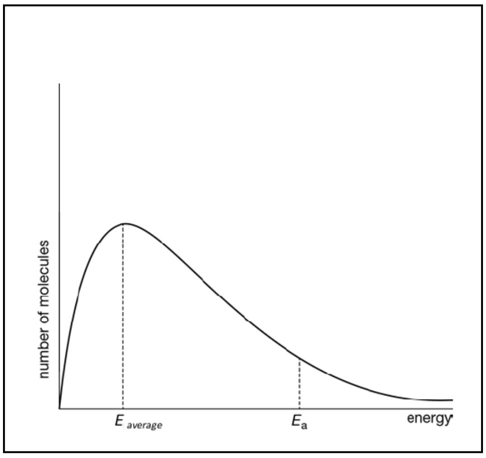


a) Label the axes

b) Place the curves 1, 2 and 3 in order of increasing temperature

c) Consider the statement below.
Indicate if it is true or false and give a reason using the graphs shown on the right.
" All molecules have higher kinetic energy at 100 °C than they do at 50°C"

9) The figure on the right shows the Maxwell-Boltzmann distribution curve of molecular energies of a gas at constant temperature.



i. State how the average kinetic energy changes if the amount of gas is doubled at constant temperature.

ii. State how the number of molecules with the average kinetic energy changes as the temperature is reduced.

iii. State what happens to the number of molecules with energy greater than the activation energy (E_a) changes as the temperature increases without an increase in the number of molecules.

iv. State how the shape of the distribution curve changes as a result of adding a catalyst.

10) A student was asked to draw the shape of the Maxwell-Boltzmann energy distribution curve of a sample of gas molecules at 30 °C given the curve at 10°C.
The student produced the curve in red as shown on the right. Is it accurate? Discuss.