Rate of Reaction

To reach a speed of 5652 km/h and an altitude of 48.7 km, each of the Space Shuttle's solid rocket boosters burns 4536 kg of fuel…

To reach a speed of 5652 km/h and an altitude of 48.7 km, each of the Space Shuttle's solid rocket boosters burns 4536 kg of fuel…

…per second…

…for 123 seconds.

aluminium + ammonium perchlorate

↓

nitrogen + aluminium oxide + aluminium chloride + water

$$10A\mathit{l} + 6NH_4C\mathit{l}O_4$$

$$\downarrow$$

$$3N_2 + 4A\mathit{l}_2O_3 + 2A\mathit{l}C\mathit{l}_3 + 12H_2O$$

- What has been *oxidised* and what has been *reduced*?

Rate of Reaction

- Duration: 1 min. 28 sec.

Rate of Reaction

Rate of Reaction

- Fireflies are actually a type of beetle. They glow due to a phenomenon known as *bioluminescence*.

Rate of Reaction

Rate of Reaction

- Fireflies flash more *frequently* on *warm* evenings compared to *cool* evenings.

- Scientists have discovered that fireflies flash once every *5 seconds at 21 °C* and once every *11 seconds at 12 °C*.

Rate of Reaction

What do I need to know about *rate of reaction*?

Rate of Reaction

Learning Outcomes

Candidates should be able to:

a) Describe the effect of concentration, pressure, particle size and temperature on the speeds of reactions and explain these effects in terms of collisions between reacting particles.

b) Define the term catalyst and describe the effect of catalysts (including enzymes) on the speeds of reactions.

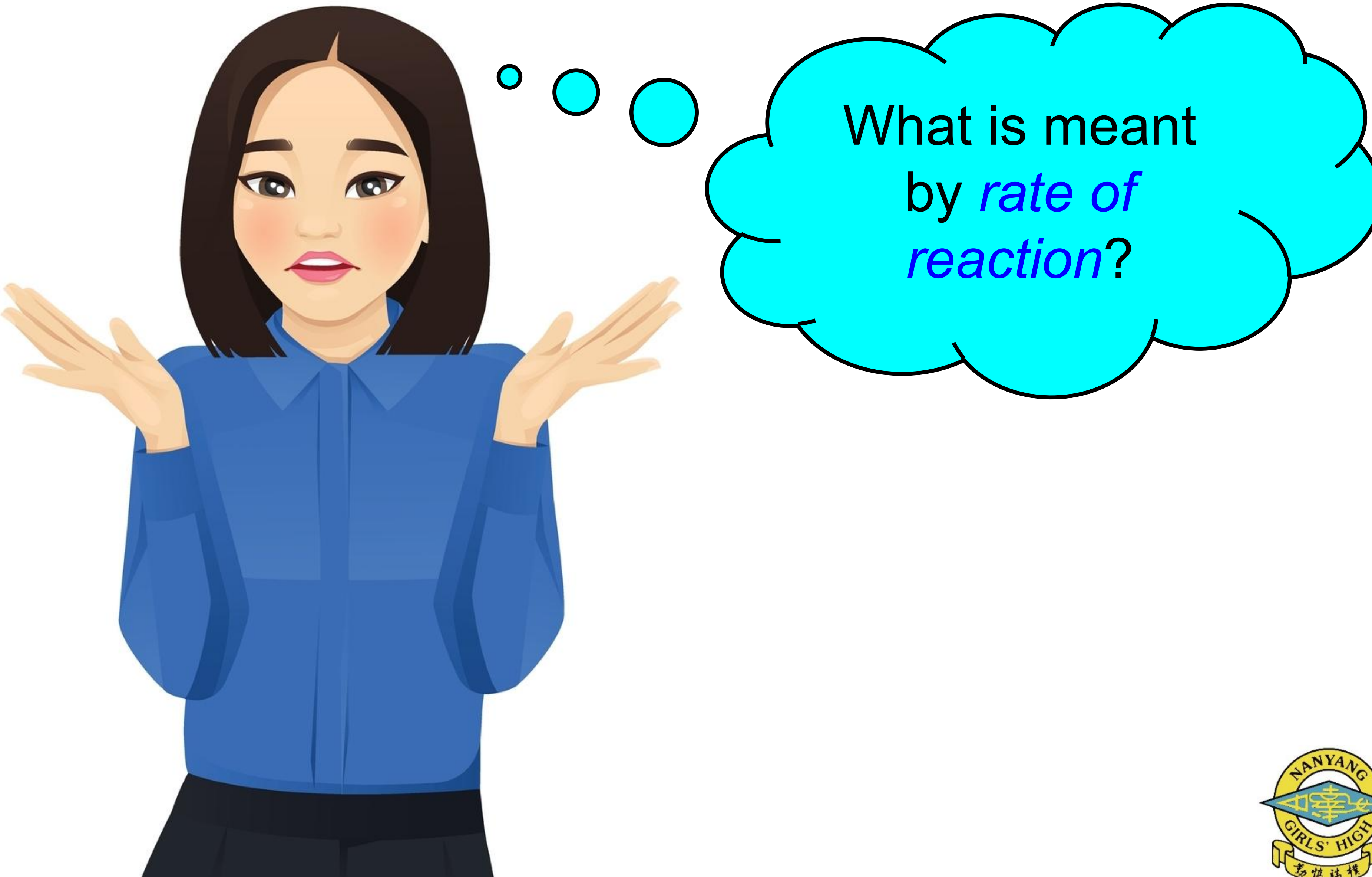
c) Explain how pathways with lower activation energies account for the increase in speeds of reactions.

d) State that some compounds act as catalysts in a range of industrial processes and that enzymes are biological catalysts.

e) Suggest a suitable method for investigating the effect of a given variable on the speed of a reaction.

f) Interpret data obtained from experiments concerned with speed of reaction.

Rate of Reaction



Rate of Reaction

- It is human nature to find things interesting if they are *fast*.

- Guinness World Record For Parallel Parking.

- Duration: 29 sec.

- How would you determine the average speed of this car?

- speed = distance ÷ time

Rate of Reaction

It's about…  …time.

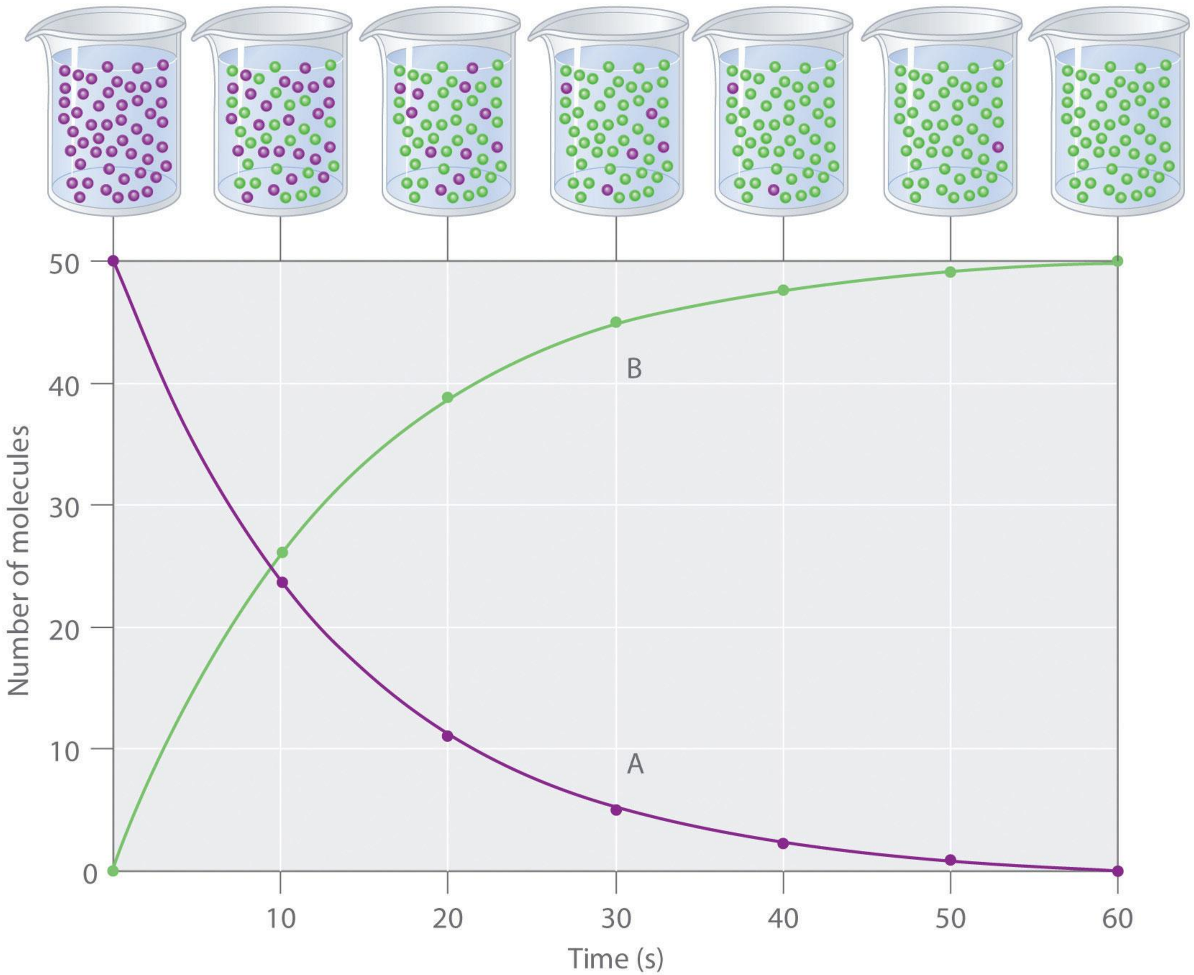
So talk about *frequency*, not *number*.

Rate of Reaction

$$\text{rate of reaction} = \frac{\text{amount of reactant used up}}{\text{time taken}}$$

$$\text{rate of reaction} = \frac{\text{amount of product formed}}{\text{time taken}}$$

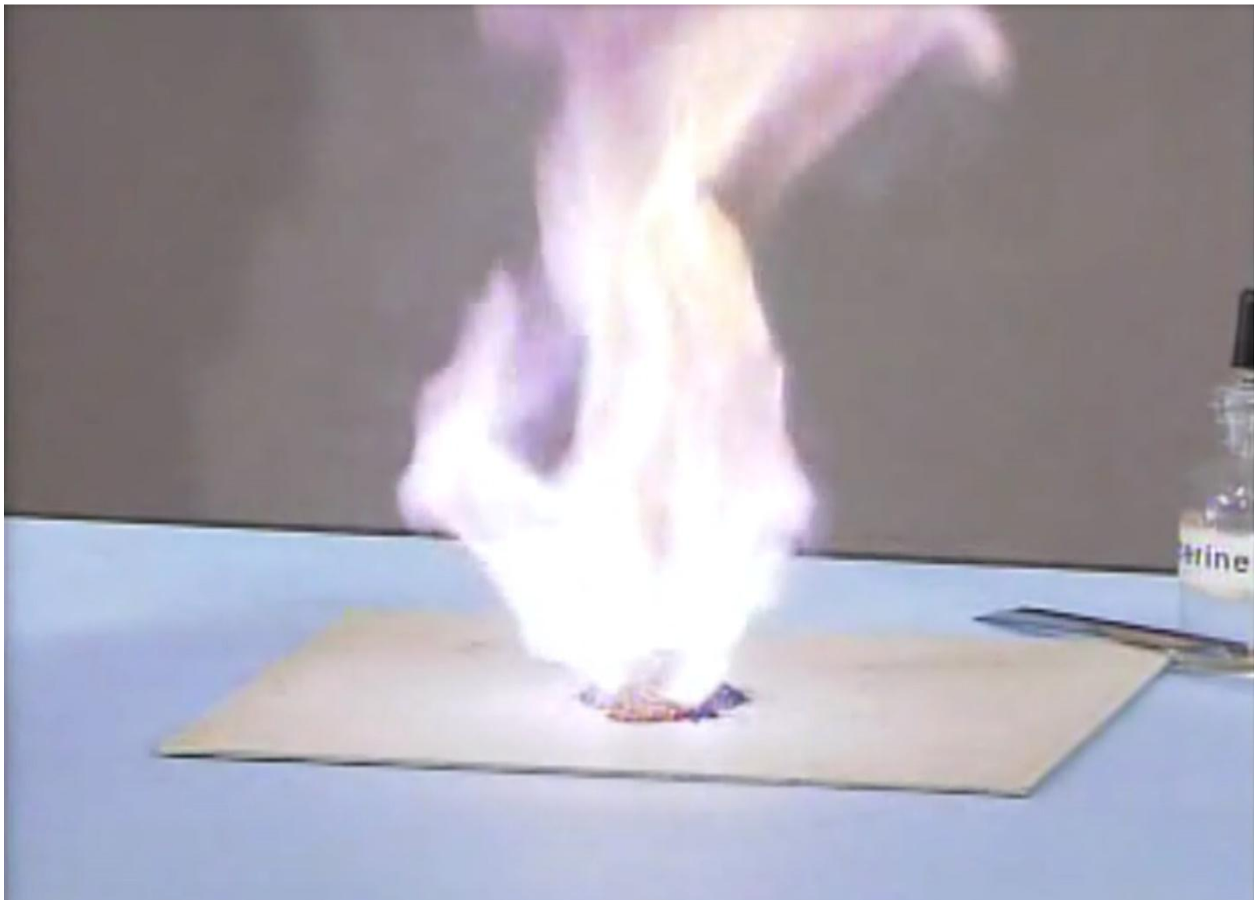
Rate of Reaction



- The *steeper* the gradient, the *faster* the reaction.

Rate of Reaction

- Some chemical reactions are *very fast*, for example, the reaction between potassium manganate(VII) (formula, $KMnO_4$) and glycerol (formula, $C_3H_8O_3$).



- Duration: 1 min. 17 sec.

Rate of Reaction

- Some chemical reactions are *very fast*, for example, the reaction between potassium zinc (Zn) and sulfur (S):

$$Zn(s) + S(s) \rightarrow ZnS(s)$$

- Duration: 44.5 sec.

Rate of Reaction

Rate of Reaction

- While other chemical reactions are very slow, for example, the rusting of iron.

Rate of Reaction

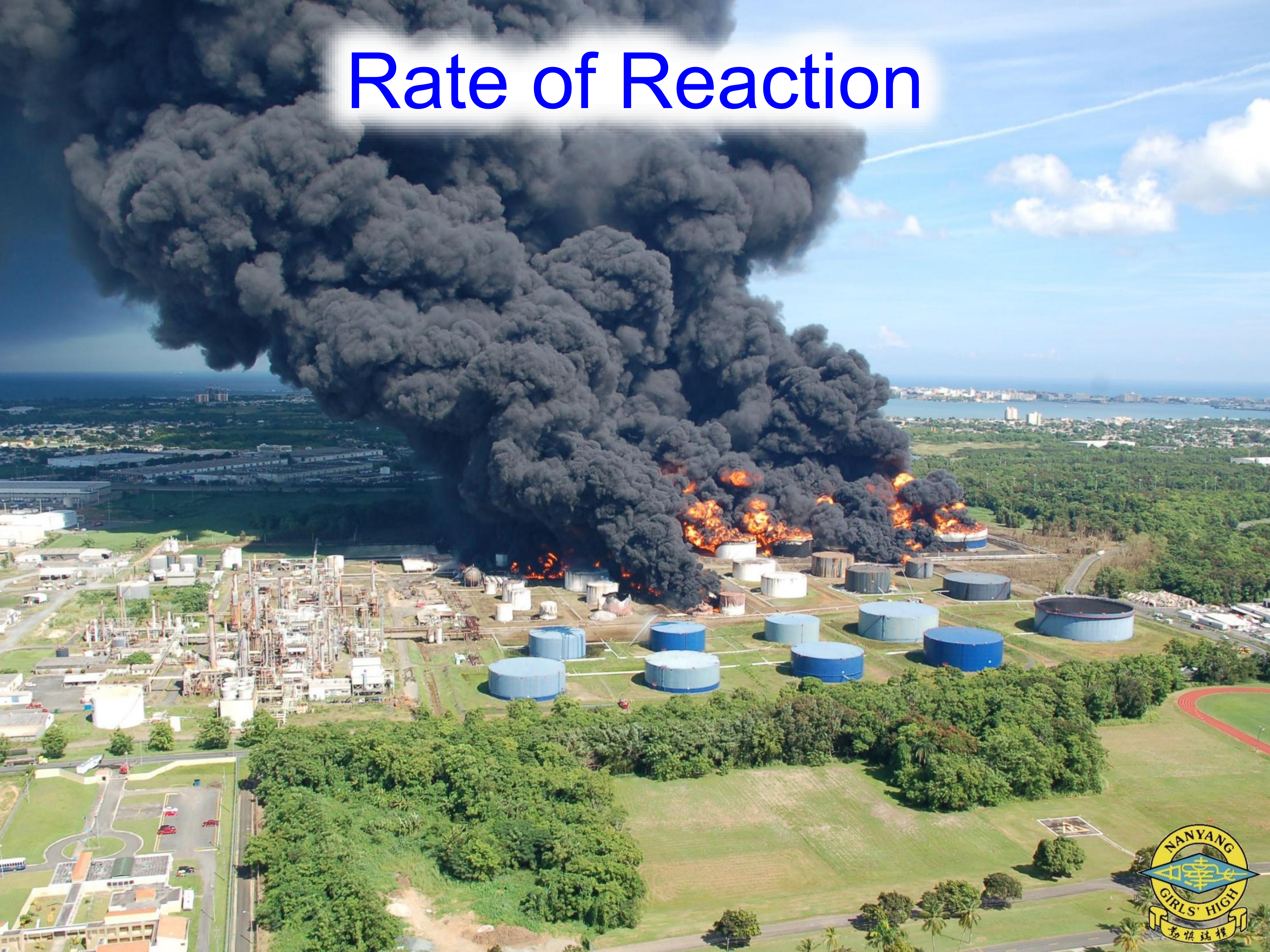
Why is it important to *control* the rate of a chemical reaction?

Rate of Reaction

Rate of Reaction

- Consider why it is important to control the rate of a chemical reaction that takes place on an industrial scale.

Rate of Reaction



NANYANG GIRLS' HIGH

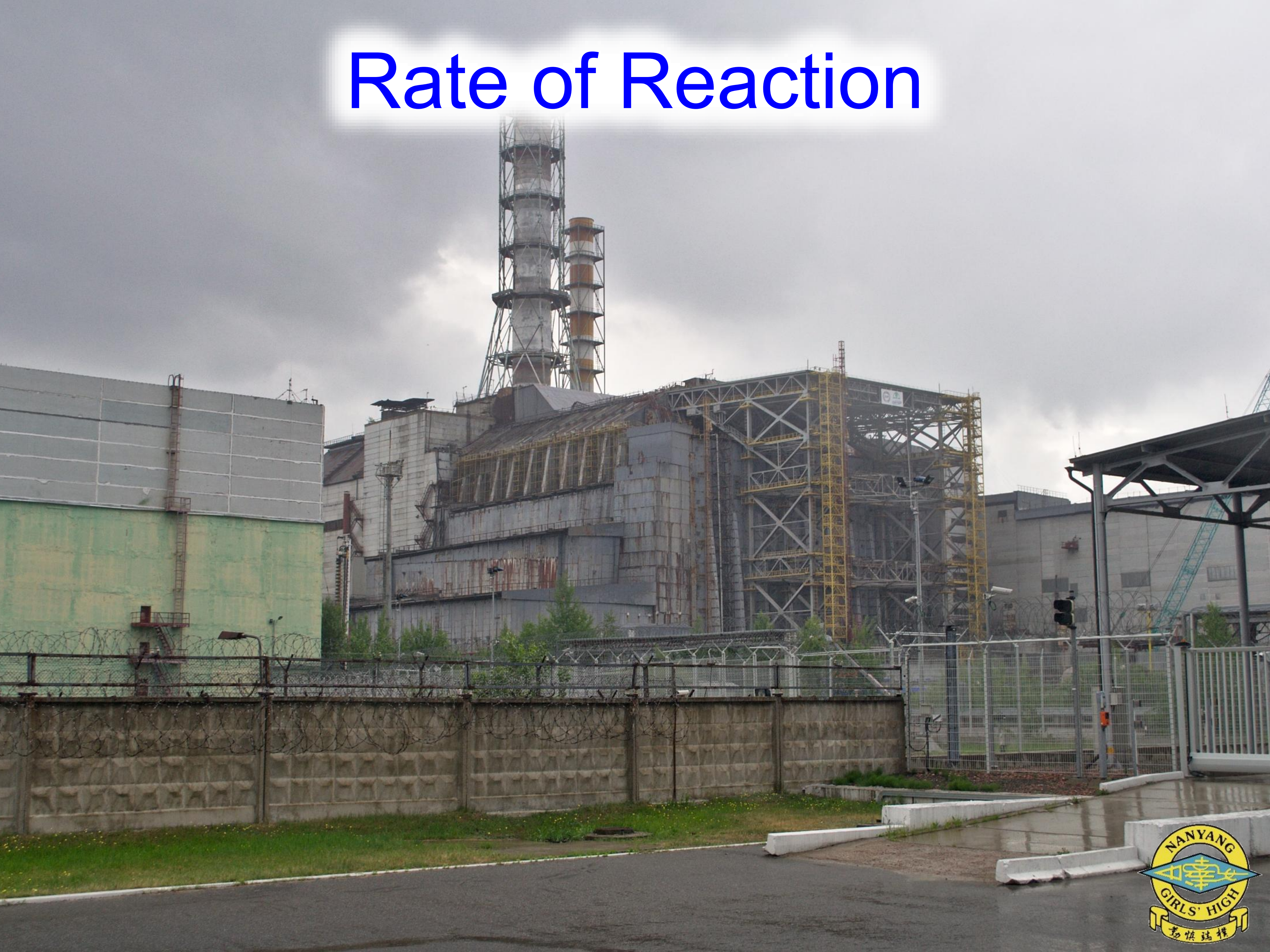
Rate of Reaction

- At a factory making nylon in Flixborough, a small town in England, a cigarette ignited a cloud of highly flammable cyclohexane vapour, which had escaped from a leaking pipe. A tiny amount of energy from the cigarette was enough to start the reaction that released the vast amount of energy stored in the cyclohexane. This destroyed 100 houses and killed 29 people.

cyclohexane + oxygen → carbon dioxide + water

$$C_6H_{12}(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$$

Rate of Reaction



Rate of Reaction

- The causes of the Chernobyl incident, which released a cloud of radioactive material into the atmosphere, are misunderstood. The cause was chemical rather than nuclear. A series of operator errors produced steam inside the reactor. Zirconium metal reacted with the steam to produce hydrogen which ultimately exploded and set fire to the graphite rods that were used to absorb neutrons in the reactor. Radioactive material was then released from the damaged reactor. Essentially, this was a case of a high temperature increasing the rate of a chemical reaction.

Rate of Reaction

Rate of Reaction

- At Bhopal, in India, an insecticide called carbamyl was being manufactured. The process involved several poisonous gases and one of these, methyl isocyanate, was the cause of the disaster. One theory is that a reaction between methyl isocyanate and water caused a build-up of pressure. This normally slow reaction was made much faster by sodium hydroxide (from a cleaning system) acting as a catalyst. The build-up of pressure caused toxic gases to be released into the surrounding area, killing over 2000 people.

Rate of Reaction

Rate of Reaction

Rate of Reaction

- On 13th August 2015 a warehouse complex at Tianjin in China was destroyed by two massive explosions.

- Seismologists were able to measure the strength of the two explosions. The first explosion was equivalent to *3 tonnes* of TNT detonating. The second explosion was equivalent to *21 tonnes* of TNT detonating, and the explosion could be *observed from space*.

Rate of Reaction

- Initial reports stated that the explosions killed dozens of people, injured hundreds more, and devastated a large area of the city.
- It is believed that the warehouse complex was used to store hazardous chemicals such as *calcium carbide* (CaC_2), which is used in the manufacture of plastics, and *potassium nitrate* KNO_3, which is used in the manufacture of fertilisers and explosives.

Rate of Reaction

- First Explosion: Could have been caused by the civil defence force spraying water onto the *calcium carbide* (CaC_2). This reaction would have produced *ethyne* (C_2H_2) and calcium hydroxide.

$$CaC_2(s) + 2H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

- Ethyne is a *highly flammable* gas that releases a very large amount of energy when it burns.

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

$$\Delta H = -1300 \text{ kJ/mol of } C_2H_2(g)$$

Rate of Reaction

- Second Explosion: This was most probably triggered by the first explosion.

- *Potassium nitrate* (KNO_3) is an excellent oxidising agent. Above its melting point of 334 °C, potassium nitrate rapidly decomposes into potassium nitrite (KNO_2) and *oxygen*.

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

- Heat from the first explosion would have caused the potassium nitrate to decompose, releasing enormous quantities of oxygen which would have reacted explosively with other chemicals.

Rate of Reaction

- In addition, it is also believed that the warehouse complex was used to store 700 tonnes of *sodium cyanide* (NaCN) which is used in the mining of precious metals. Sodium cyanide is a *highly toxic chemical*. Once it enters the body, it is rapidly converted into hydrogen cyanide (HCN) which *inhibits cellular respiration*.

Rate of Reaction

- Duration: 1 minute
- Video downloaded from www.theguardian.co.uk
- Warning: Video contains some expletives.

Rate of Reaction

- For industrial reactions, it is often important for the reaction to take place as quickly as possible. By manufacturing the chemical in a shorter space of time, the company will:

i) Meet the needs of its customers, especially if the chemical is in high demand.

ii) Ensure that all equipment is used to its full capacity.

iii) Increase its profit margins / make more money.

- Some chemical reactions are very exothermic. On an industrial scale, it may be necessary to slow down these reactions to reduce the energy that is released and hence reduce the chance of an industrial accident.

Rate of Reaction

It's very important to *monitor* and *control* the rate of a reaction.

Rate of Reaction

How can I *measure* the rate of a chemical reaction?

- Experimental Design:

What could the *dependent variables* be?

Rate of Reaction

- Some observable property of the reaction must be measured against *time*.

Rate of Reaction

- Measuring how these variables *change against time* will enable a Chemist to determine the *rate of a chemical reaction*:

1) Change in temperature.

2) Formation of a precipitate.

3) Change in colour.

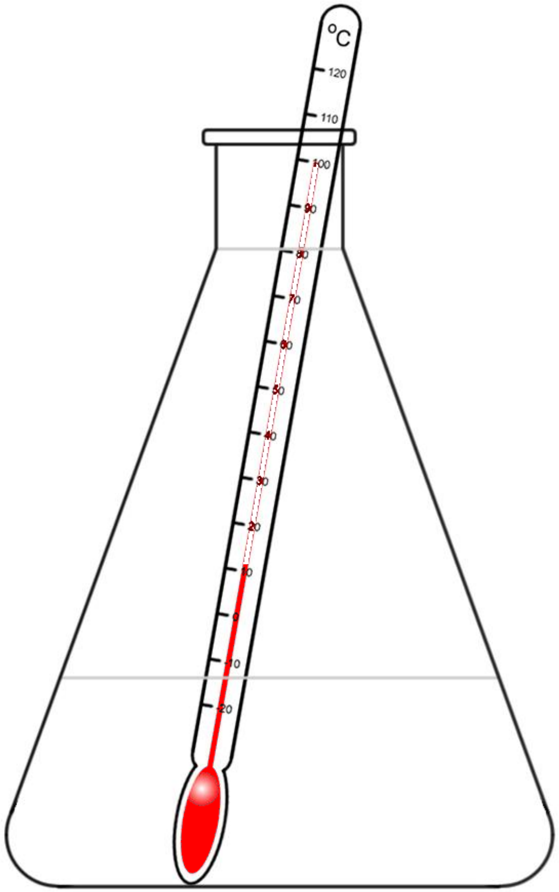
4) Change in pH value.

5) Change in volume / pressure (for gases).

6) Change in mass.

Rate of Reaction

1) Change in *temperature* against time.

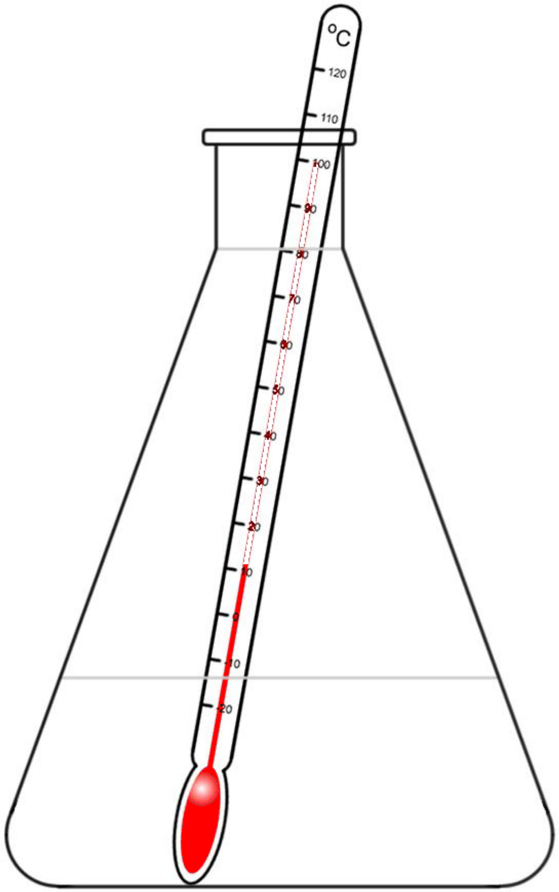


- For *exothermic* reactions, where there is an *increase* in the temperature of the surroundings. The *enthalpy change* for these reactions is *negative*, $-\Delta H$.

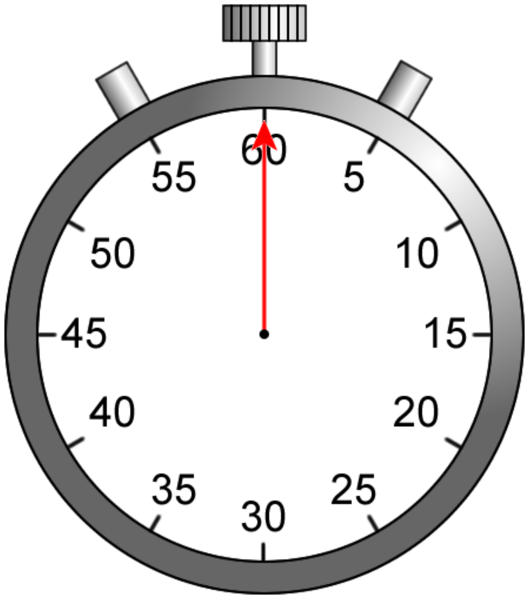
- For *endothermic* reactions, where there is a *decrease* in the temperature of the surroundings. The *enthalpy change* for these reactions is *positive*, $+\Delta H$.

Rate of Reaction

1) Change in *temperature* against time.

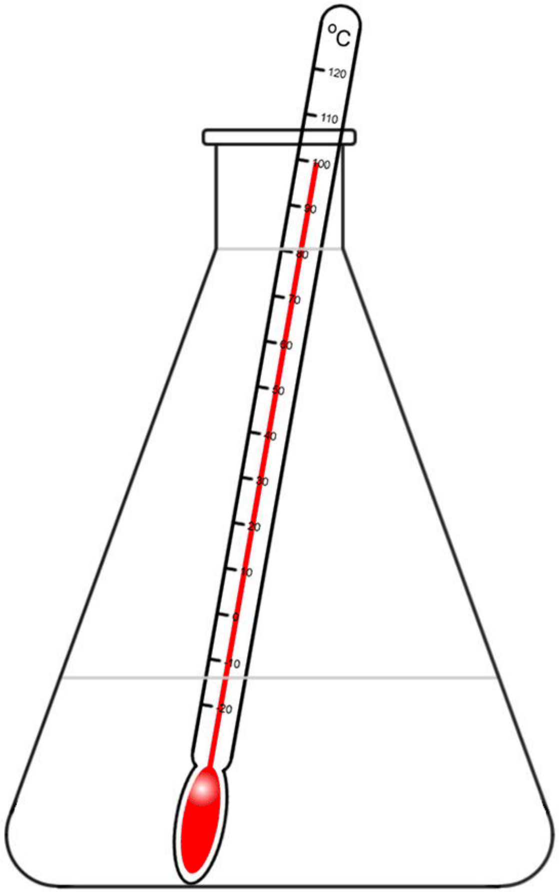


- For *exothermic* and *endothermic* reactions.

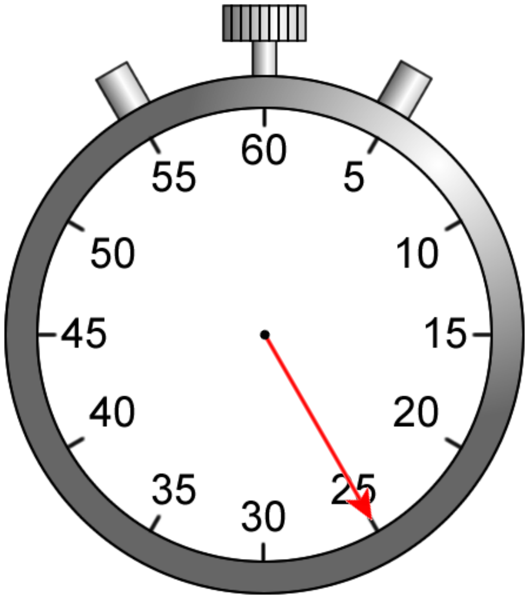


Rate of Reaction

1) Change in *temperature* against time.

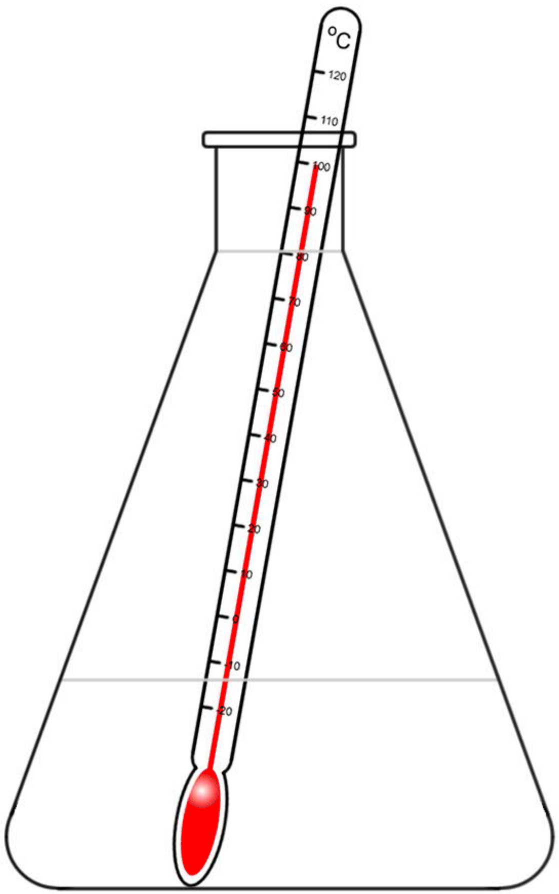


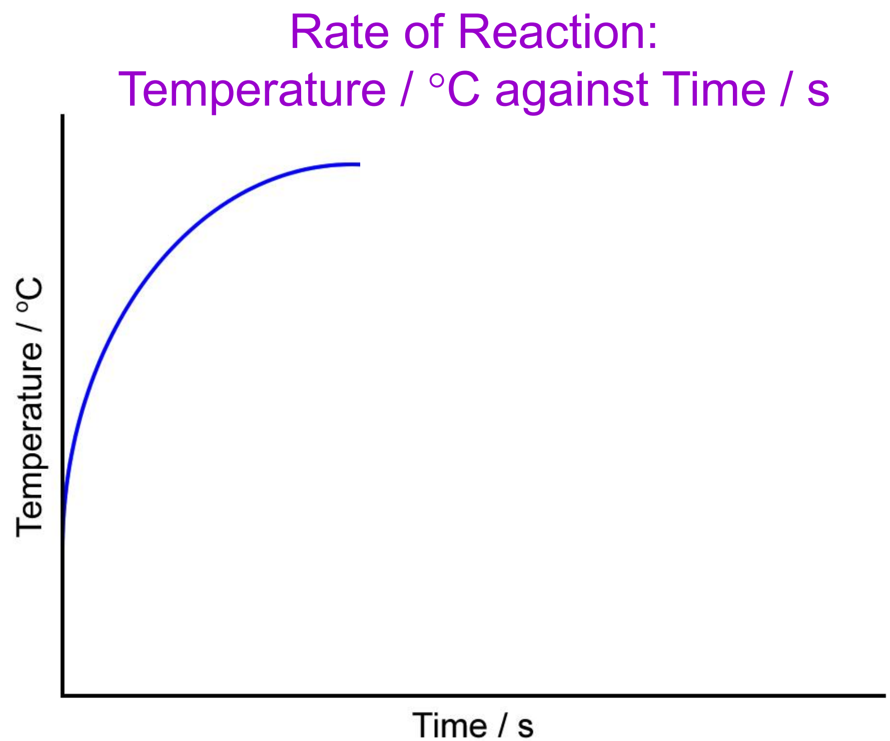
- For *exothermic* and *endothermic* reactions.



Rate of Reaction

1) Change in *temperature* against time.

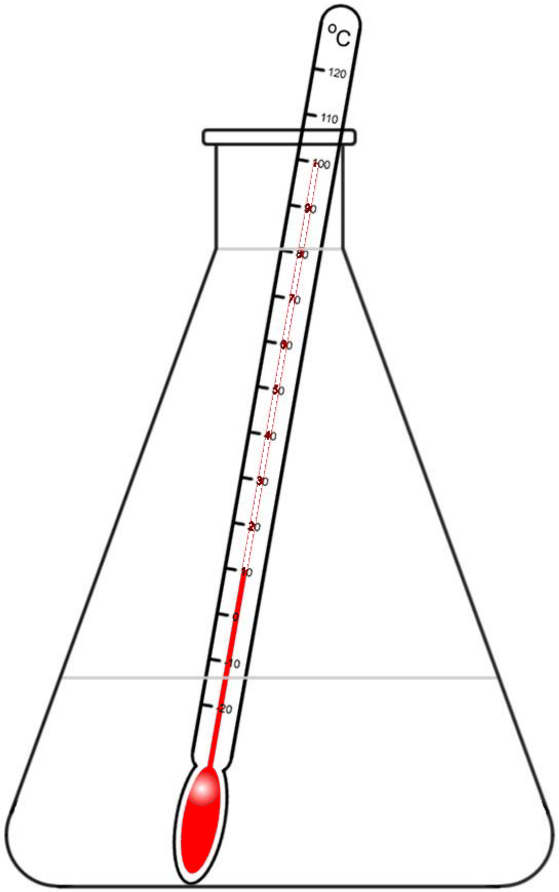


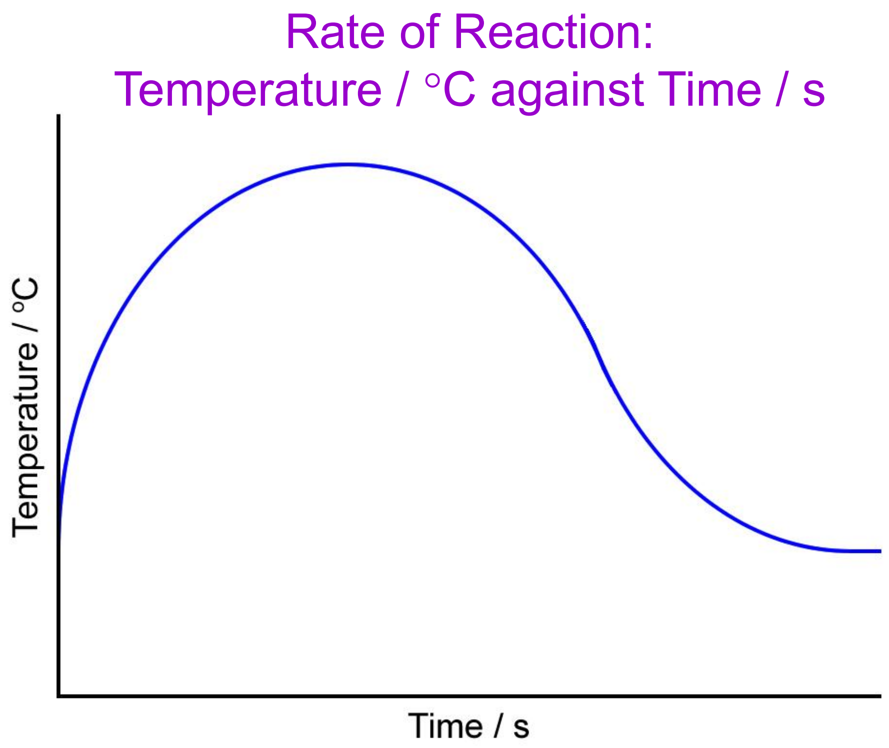


- For an *exothermic* reaction, the temperature will *increase* to a maximum…

Rate of Reaction

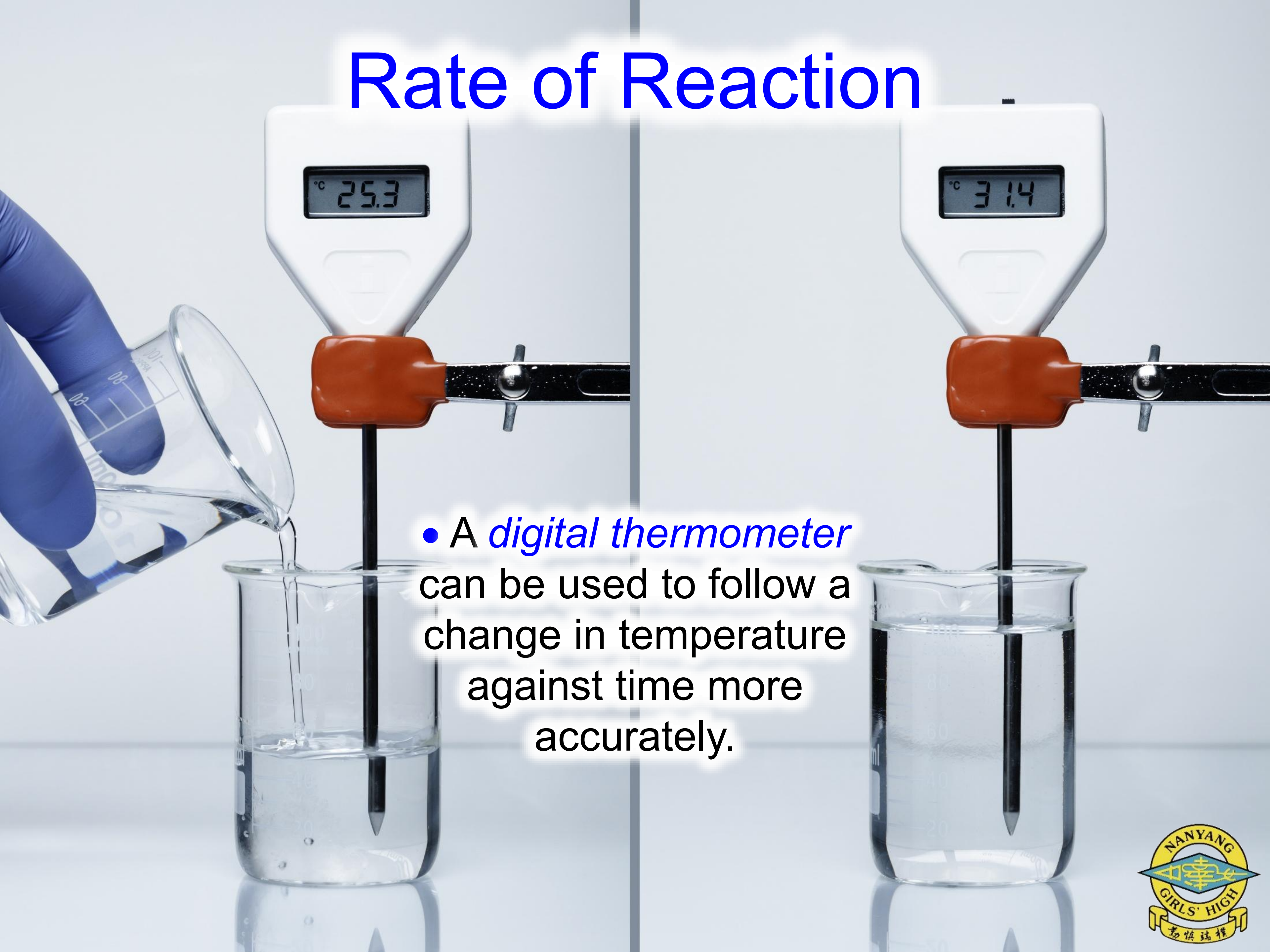
1) Change in *temperature* against time.





…before cooling down and returning to room temperature once more.

Rate of Reaction



- A *digital thermometer* can be used to follow a change in temperature against time more accurately.

Rate of Reaction

1) Change in *temperature* against time.

- Examples:

methane + oxygen → carbon dioxide + water

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad -\Delta H$

zinc + copper(II) sulfate → zinc sulfate + copper

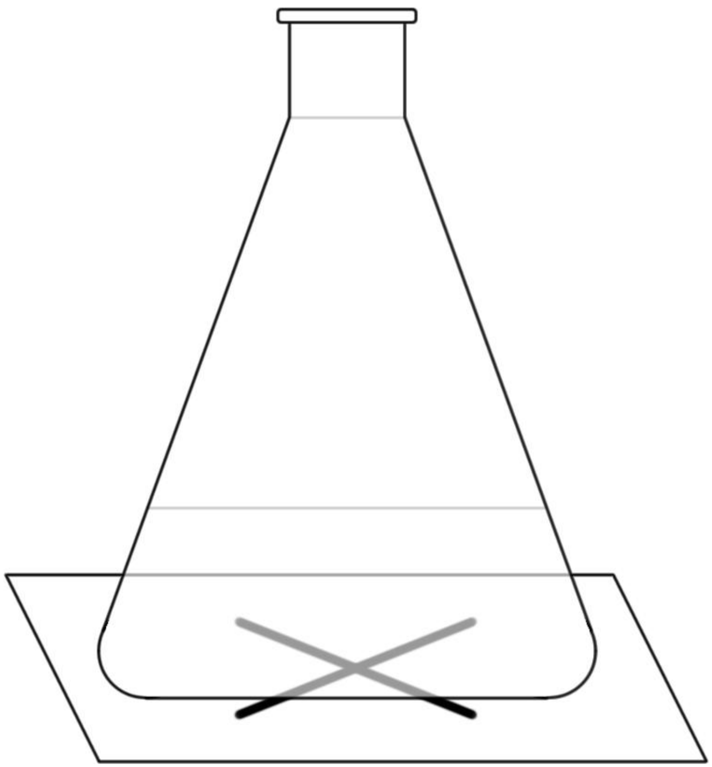
$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s) \quad -\Delta H$

carbon dioxide + water → glucose + oxygen

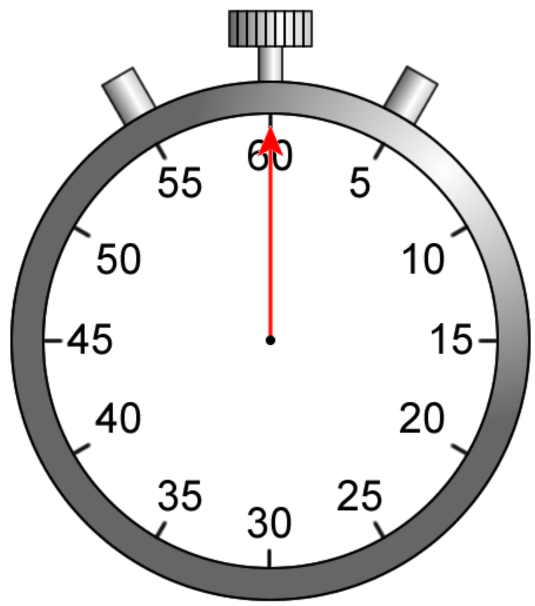
$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g) \quad +\Delta H$

Rate of Reaction

2) Formation of a *precipitate* against time.

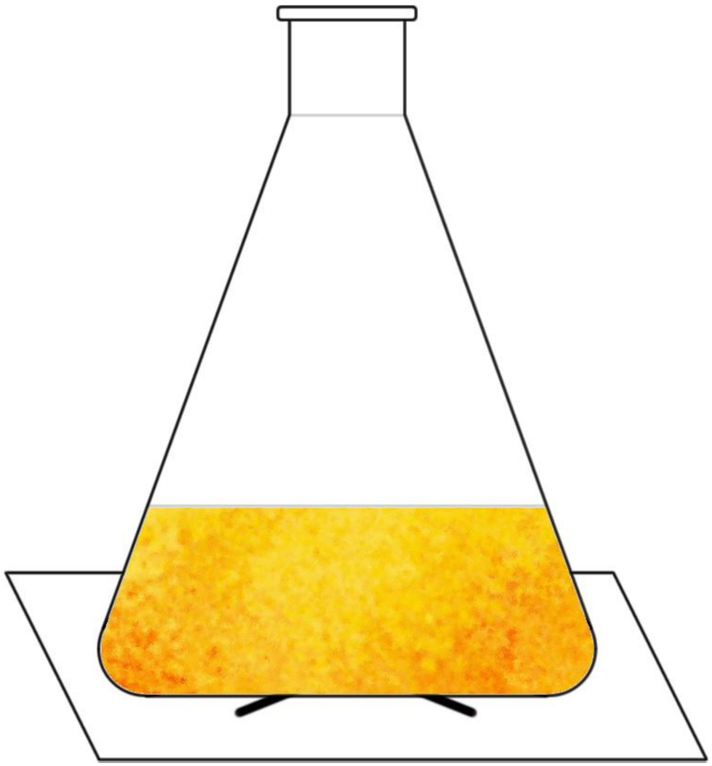


- For reactions that result in the formation of an *insoluble* product.

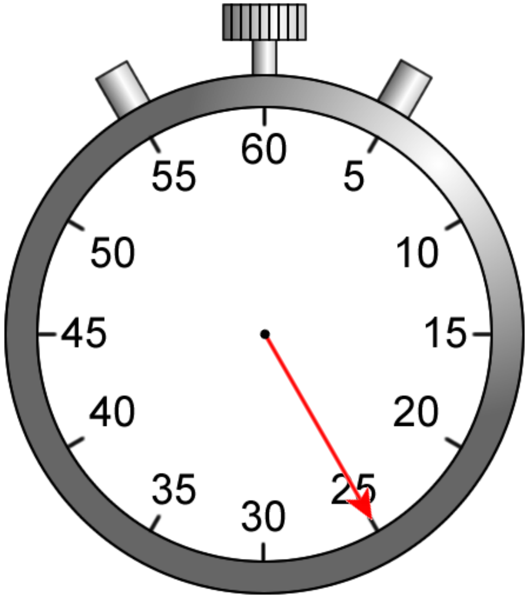


Rate of Reaction

2) Formation of a *precipitate* against time.

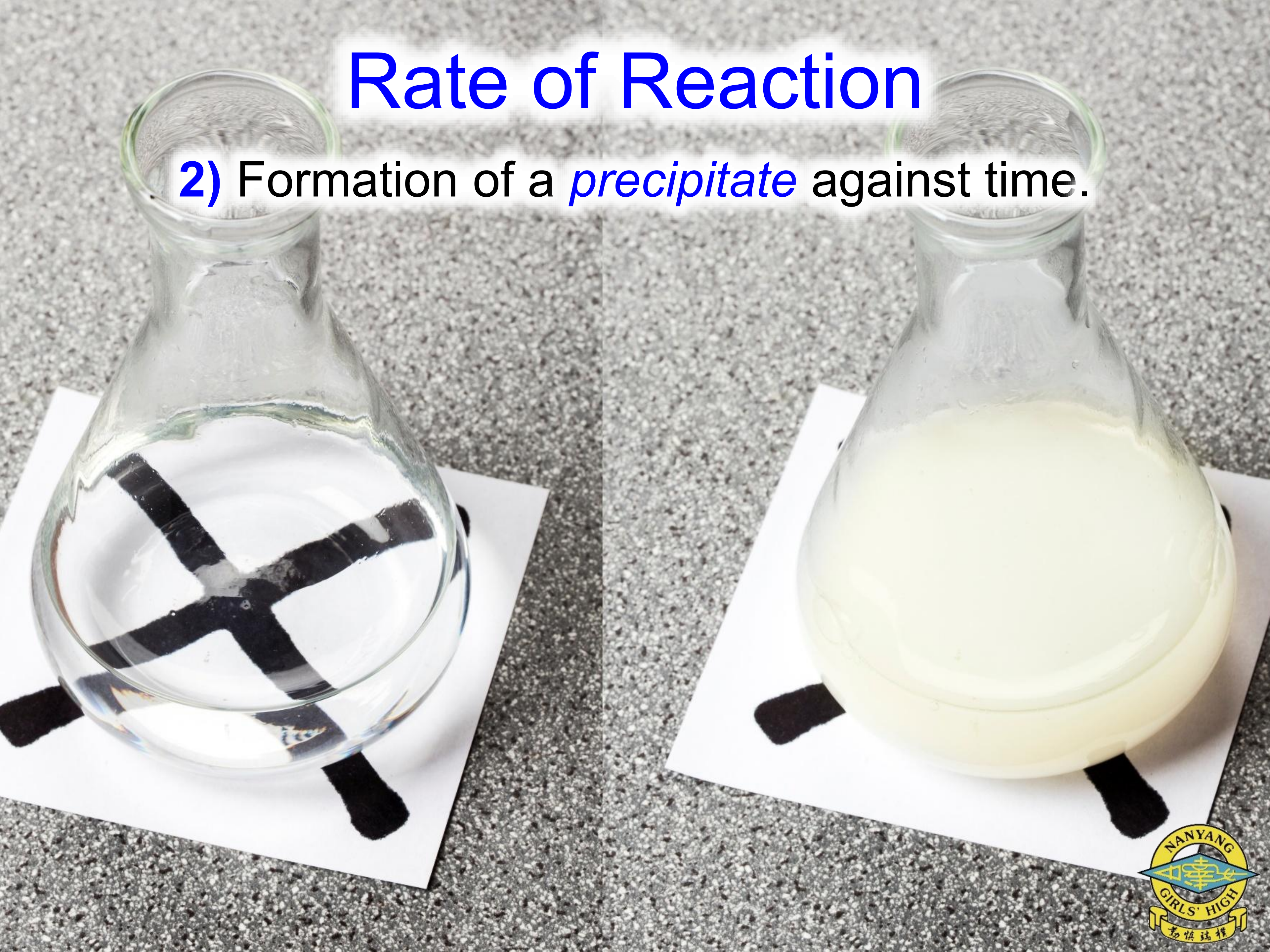


- For reactions that result in the formation of an *insoluble* product.



Rate of Reaction

2) Formation of a *precipitate* against time.



Rate of Reaction

2) Formation of a *precipitate* against time.

- Example:

sodium thiosulfate + hydrochloric acid

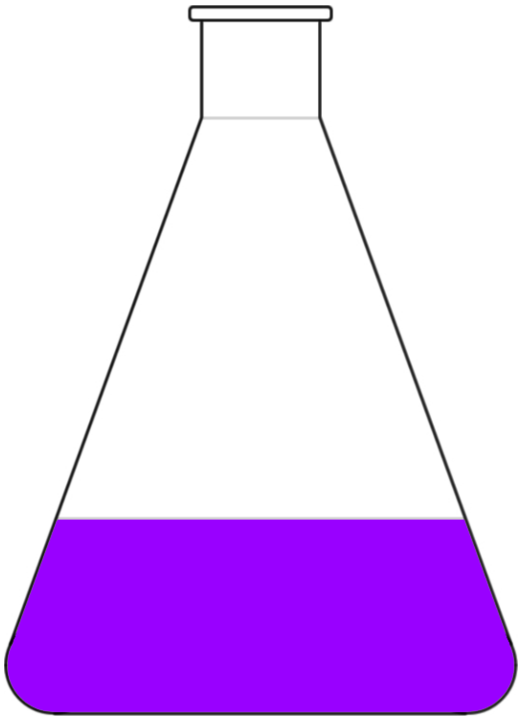
↓

sodium chloride + sulfur + sulfur dioxide + water

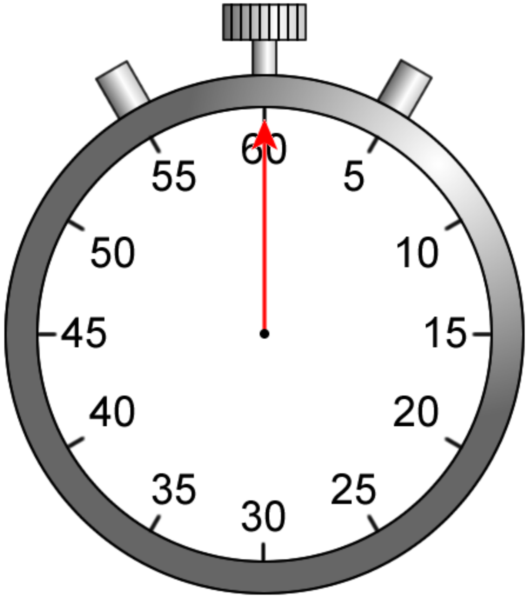
$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + S(s) + SO_2(g) + H_2O(l)$$

Rate of Reaction

3) Change in *colour* against time.



- Mainly for reactions of *transition metal* compounds and *Group 17* elements.



Rate of Reaction

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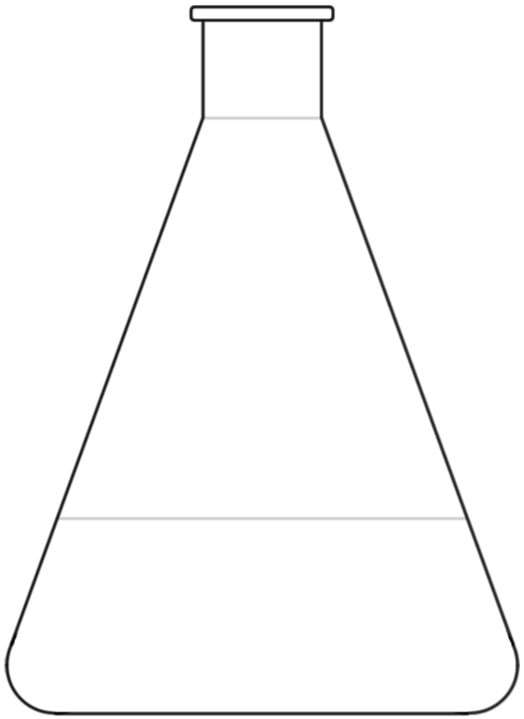
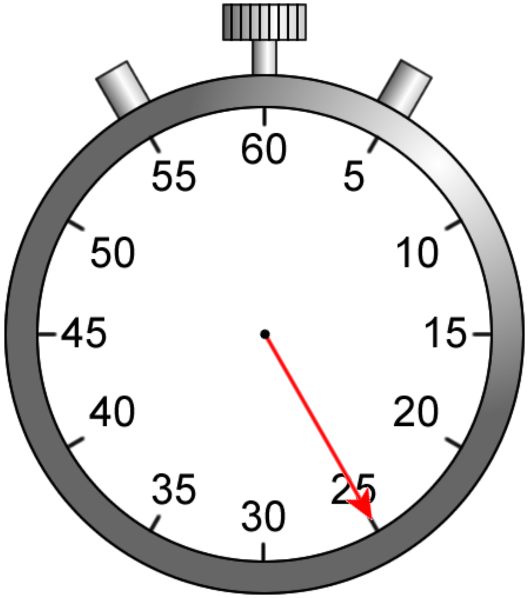
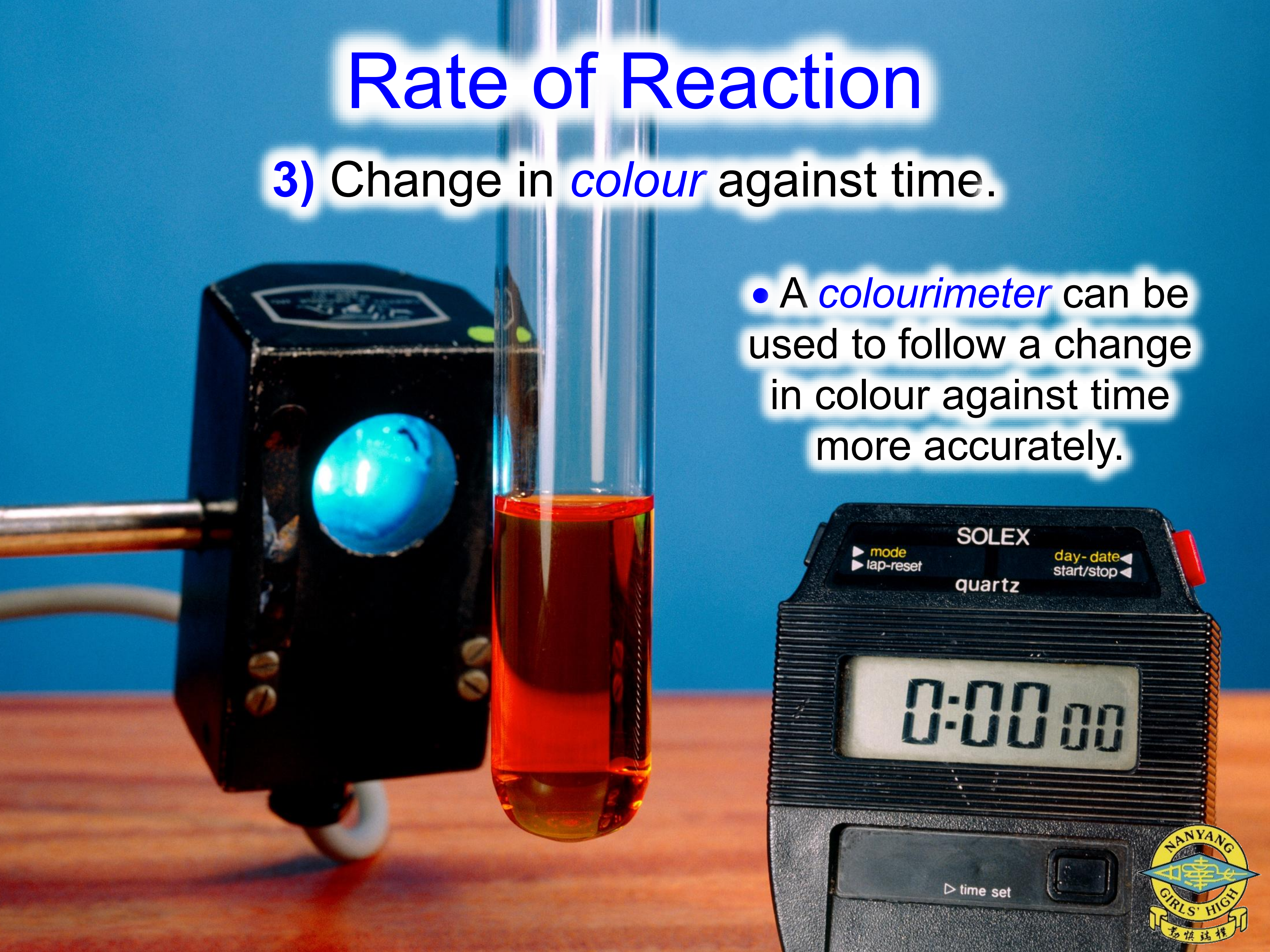




Rate of Reaction

3) Change in *colour* against time.

- A *colourimeter* can be used to follow a change in colour against time more accurately.



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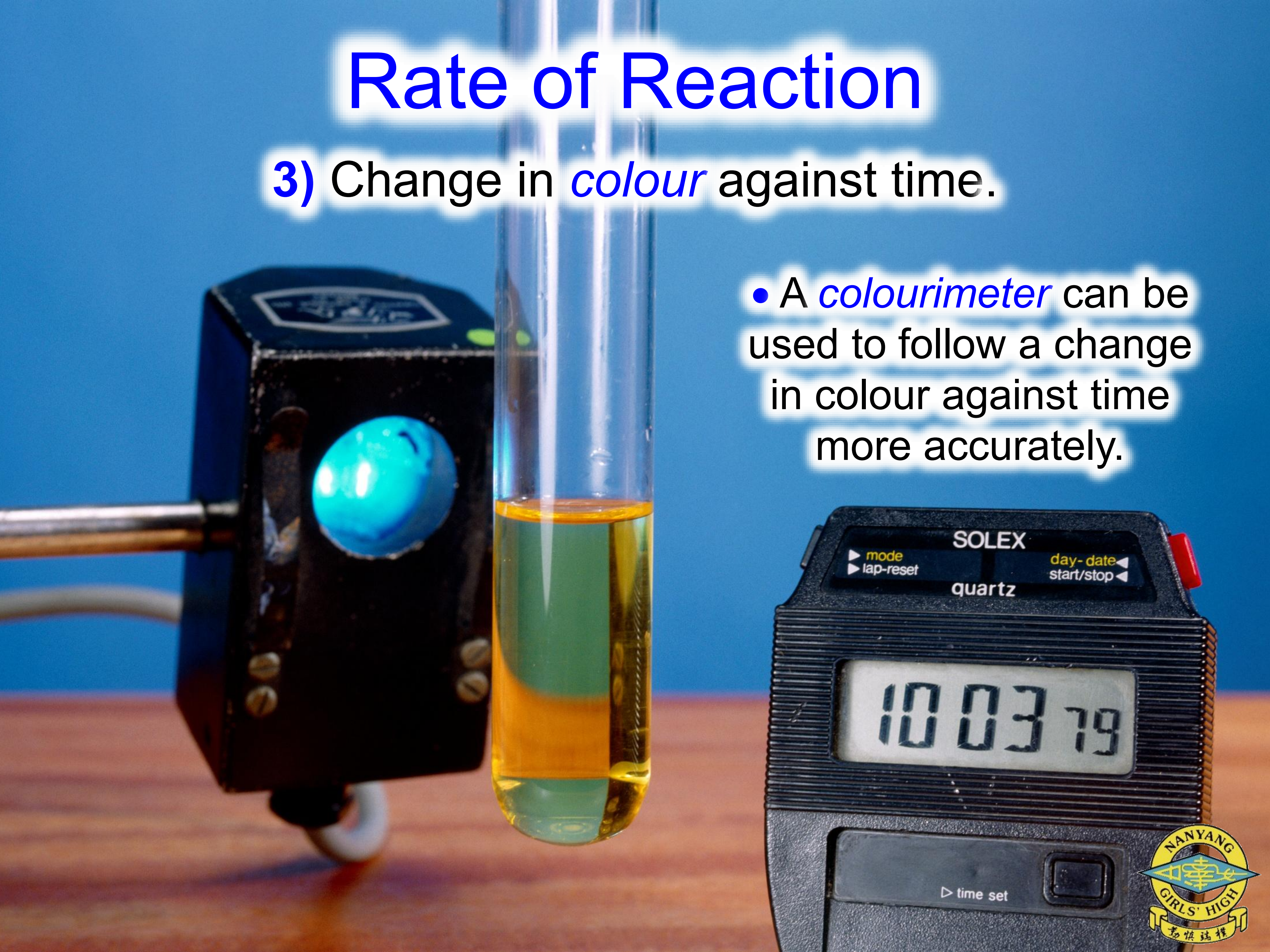

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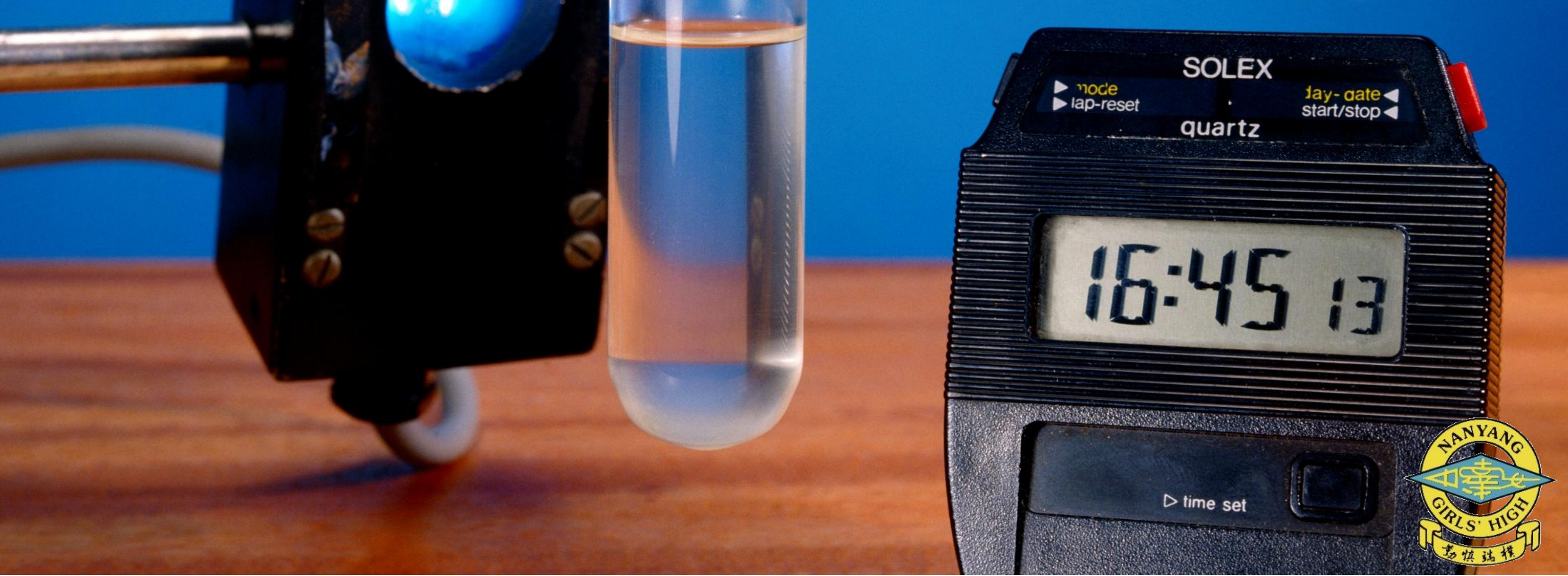


Rate of Reaction

3) Change in *colour* against time.

- Example:

manganate(VII) ions (purple)

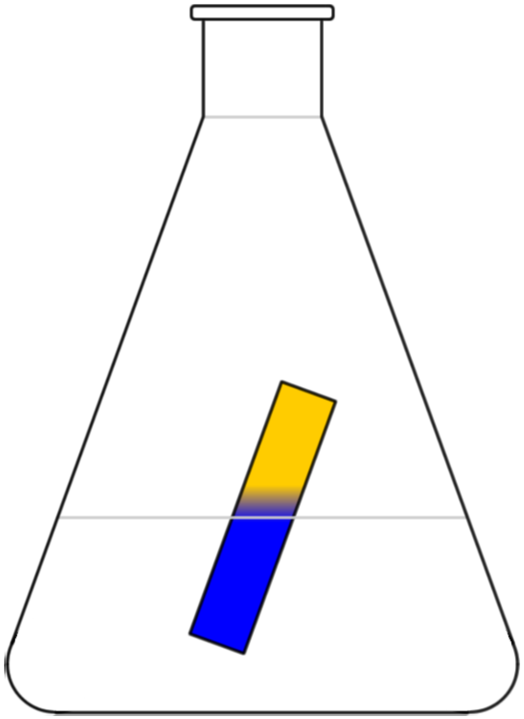
↓

manganese(II) ions (colourless)

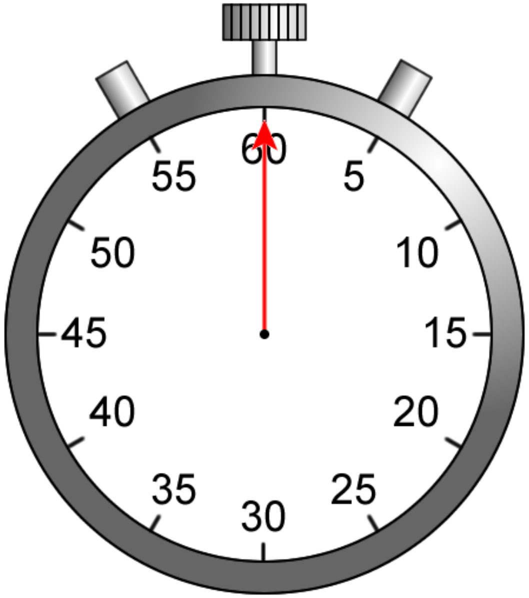
$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

Rate of Reaction

4) Change in *pH* against time.



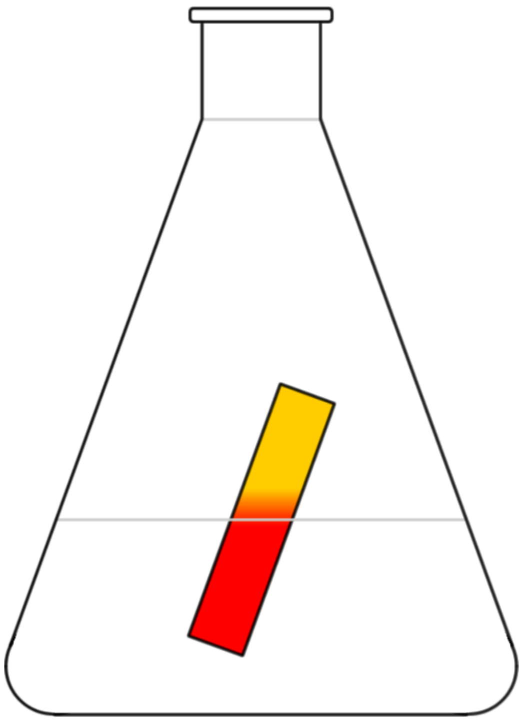
- For reactions involving *acids* and / or alkalis.

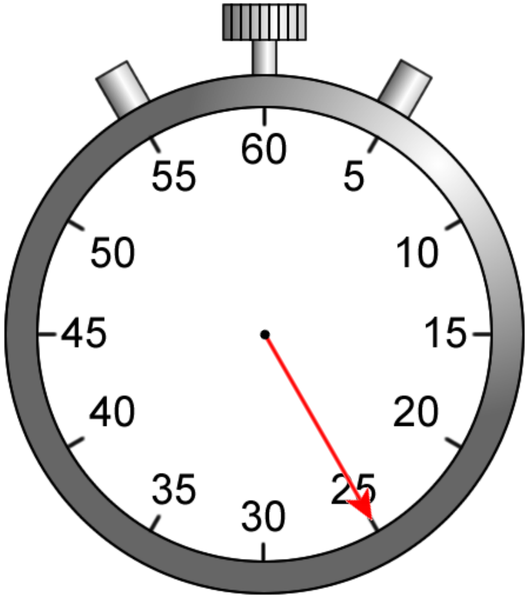


Rate of Reaction

4) Change in *pH* against time.

- For reactions involving *acids* and / or alkalis.





Rate of Reaction

4) Change in *pH* against time.

- Example:

sulfuric acid + sodium hydroxide → sodium sulfate + water

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

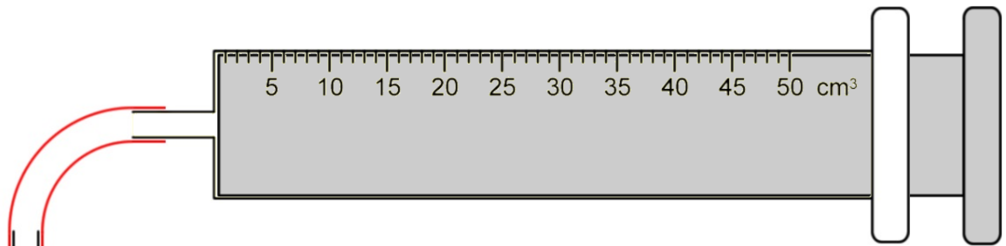
Rate of Reaction

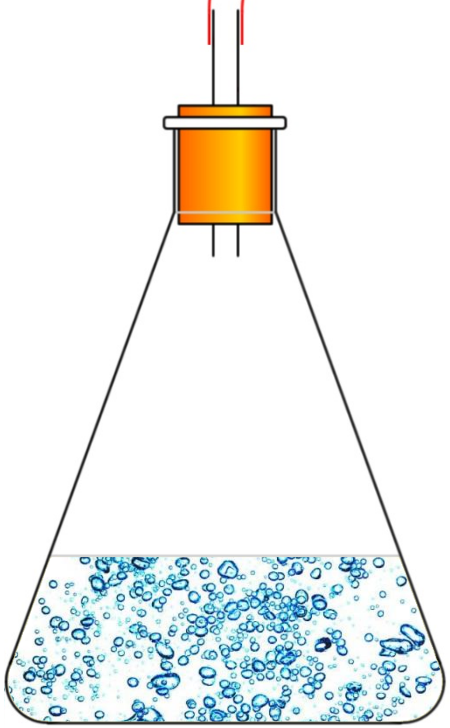
4) Change in *pH* against time.

• A *digital pH meter* can be used to follow a change in pH against time more accurately.

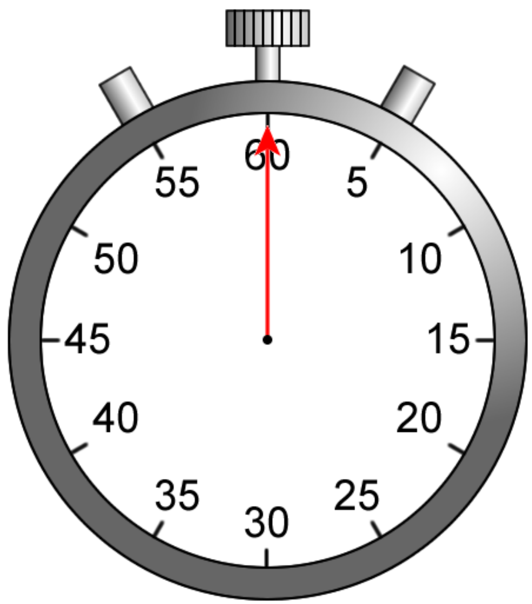
Rate of Reaction

5) Change in *volume* or *pressure* against time.



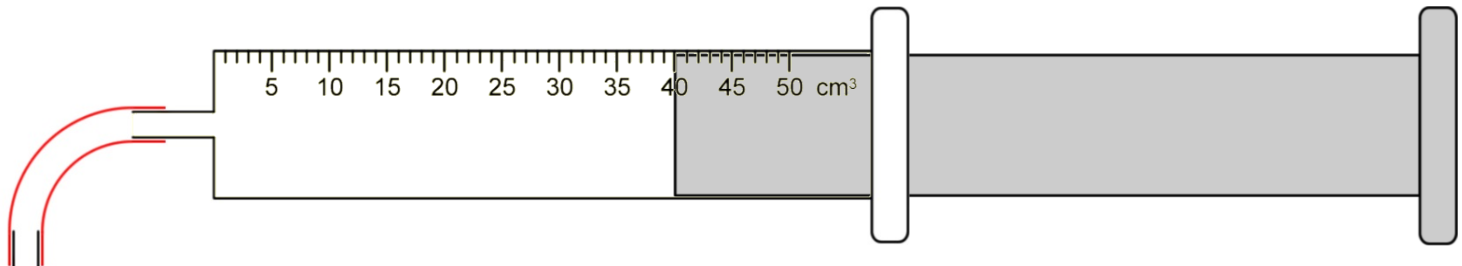


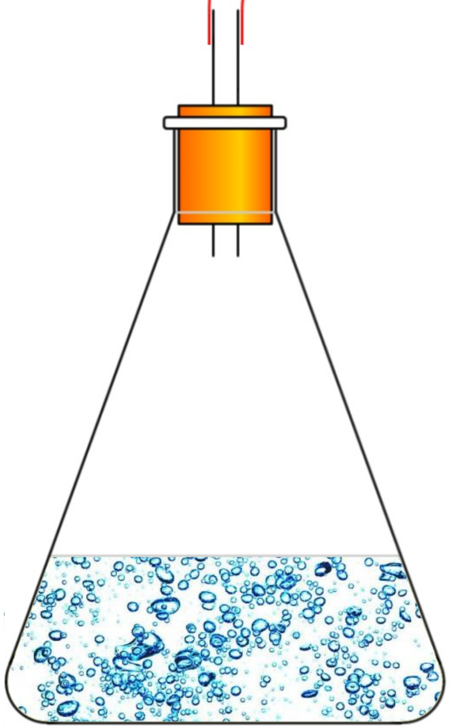
- For reactions involving *gases*.



Rate of Reaction

5) Change in *volume* or *pressure* against time.





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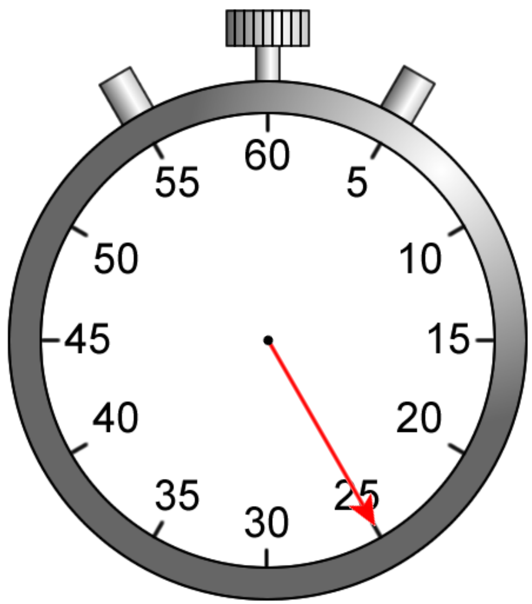
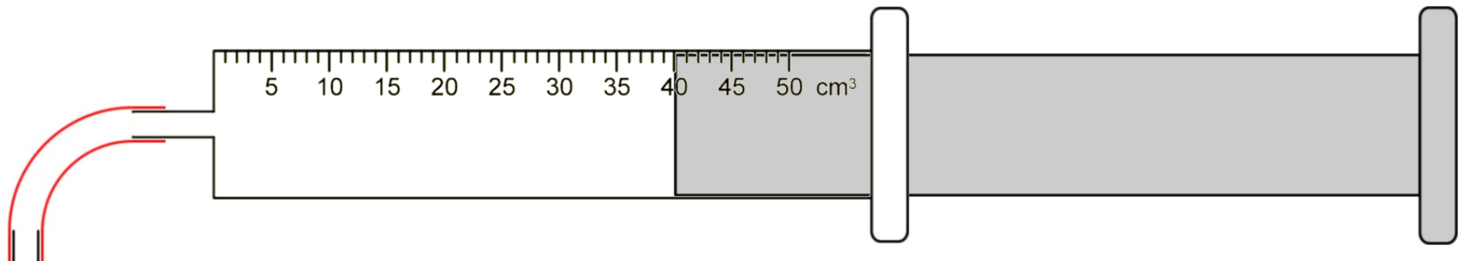
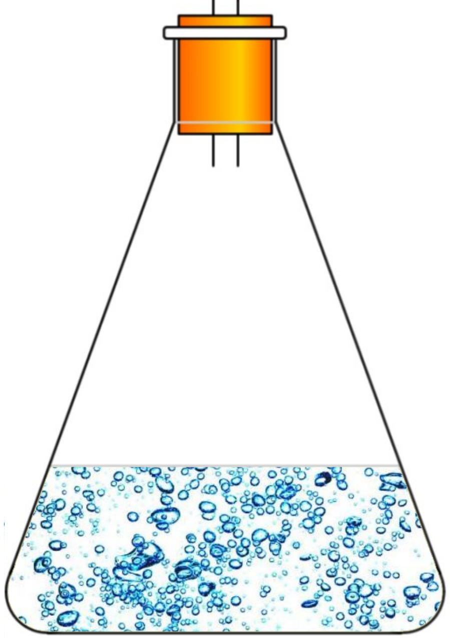
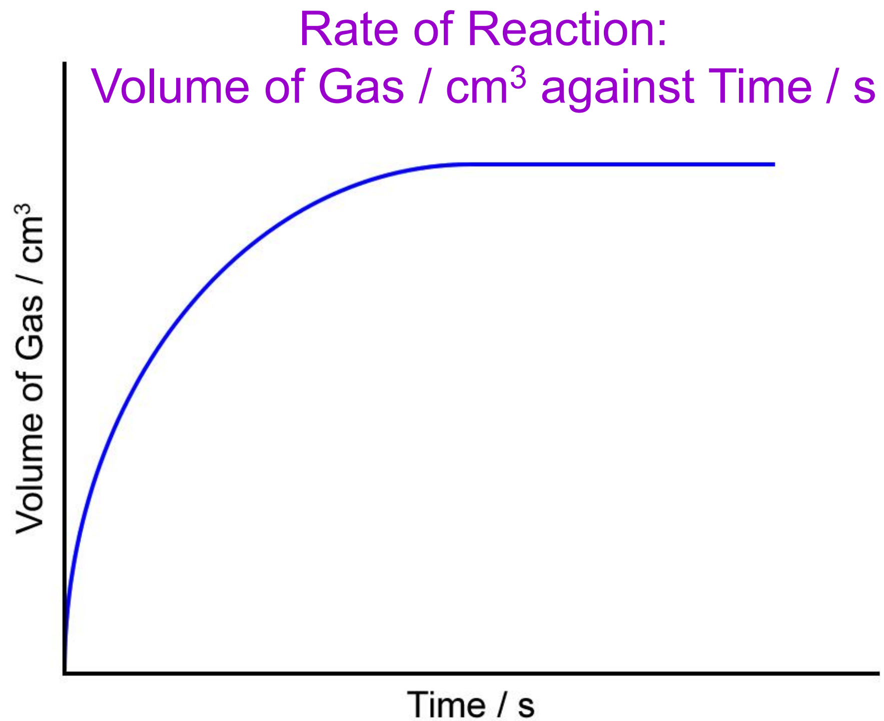


Rate of Reaction

5) Change in *volume* or *pressure* against time.



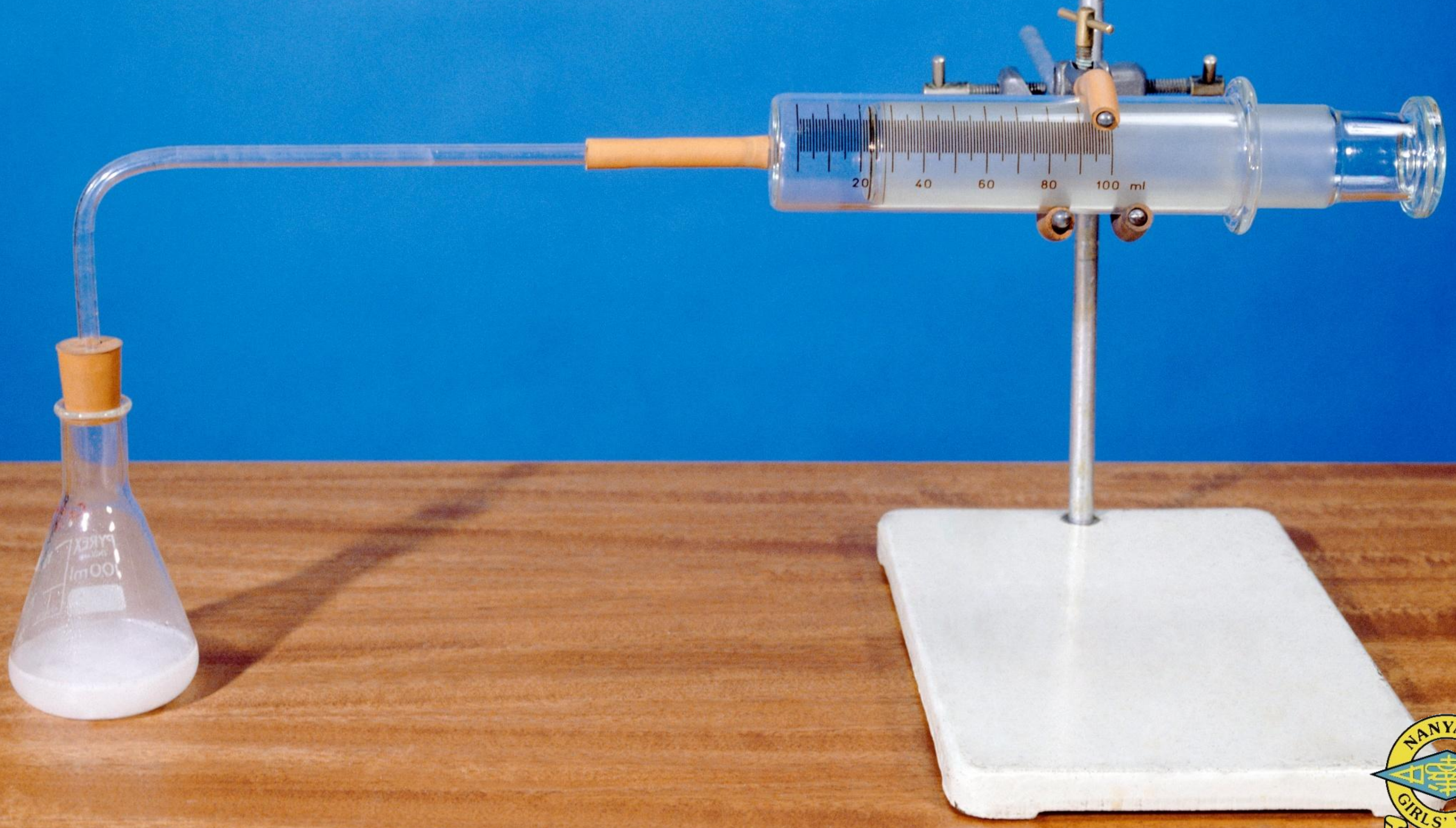




- When the volume reaches a constant value, then the reaction has finished.

Rate of Reaction

5) Change in *volume* or *pressure* against time.



Rate of Reaction

5) Change in *volume* or *pressure* against time.

- Examples:

magnesium + nitric acid → magnesium nitrate + hydrogen

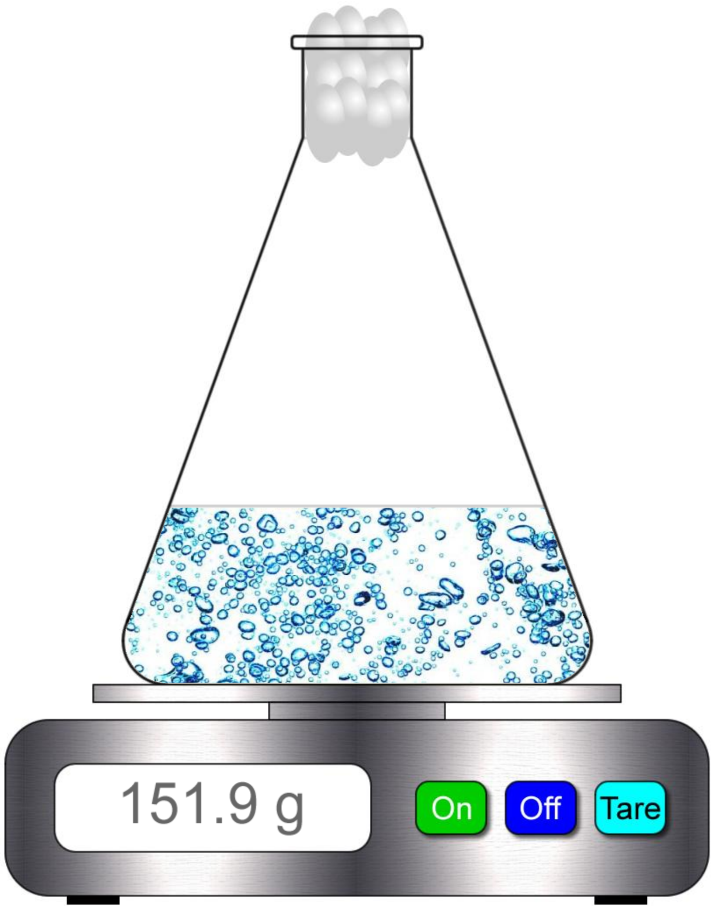
$$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$

hydrogen peroxide → oxygen + water

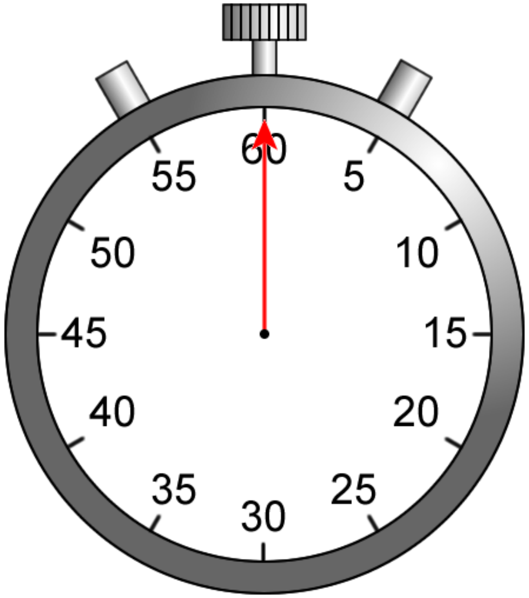
$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$$

Rate of Reaction

6) Change in *mass* against time.



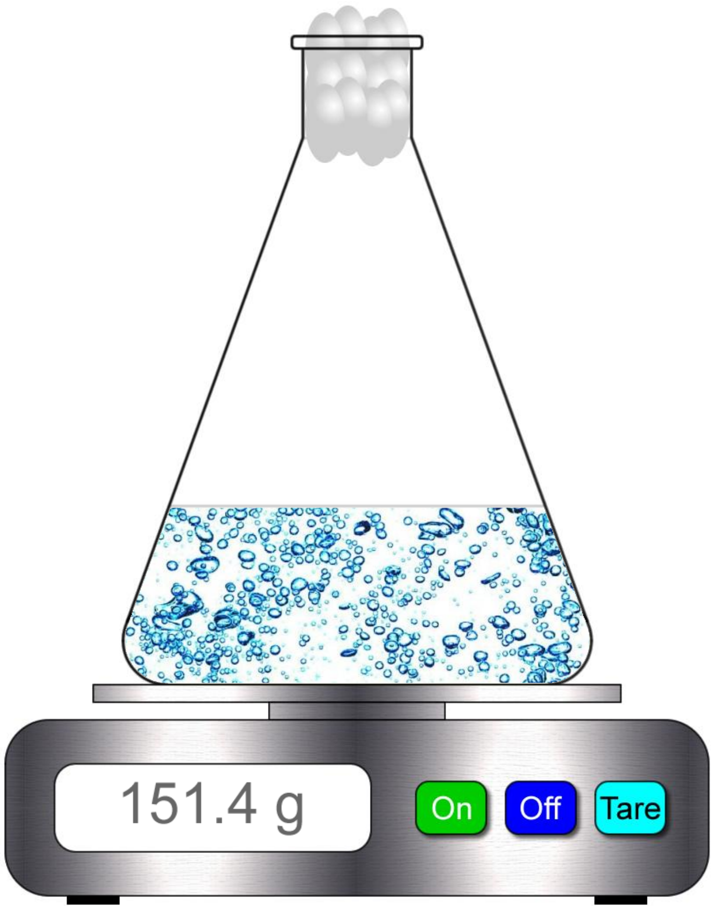
- For reactions in which a dense *gaseous product* is allowed to escape.



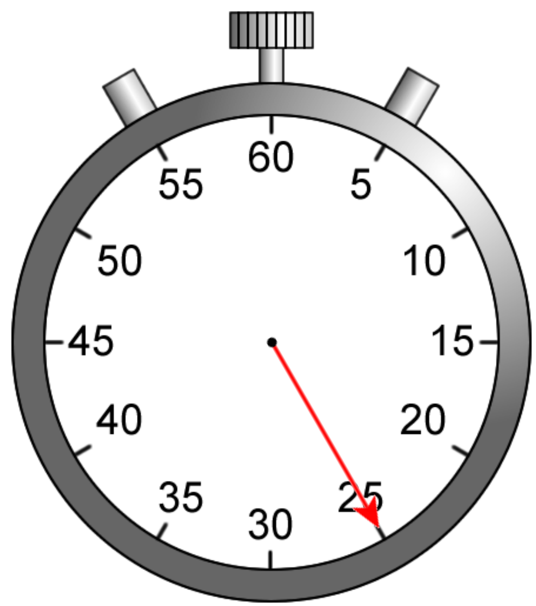
Note: Cotton wool is placed in the mouth of the flask to allow gases to escape, while preventing any of the solution from escaping.

Rate of Reaction

6) Change in *mass* against time.



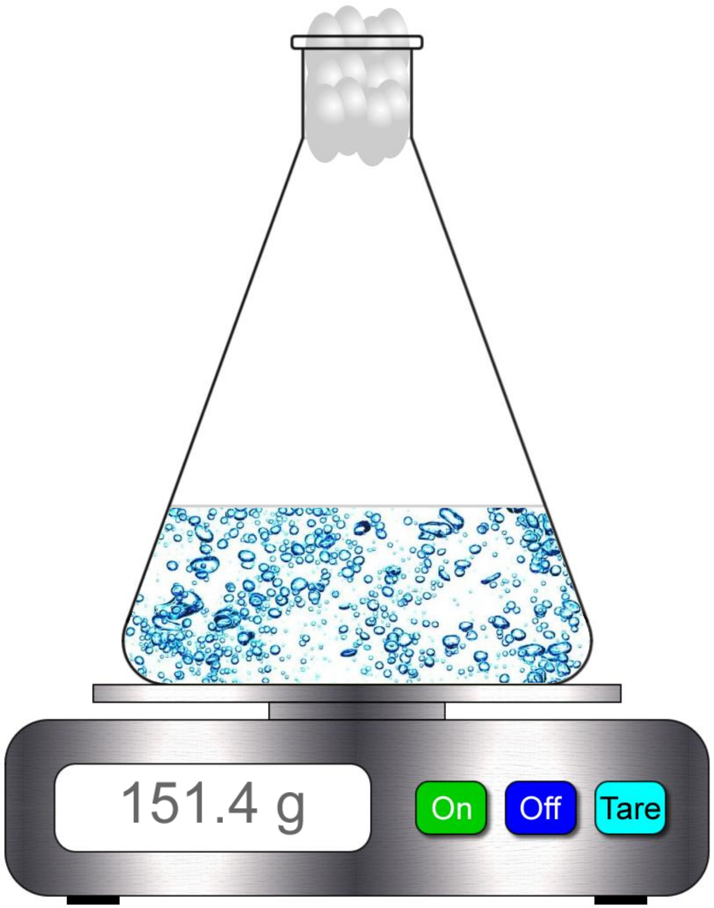
- For reactions in which a dense *gaseous product* is allowed to escape.



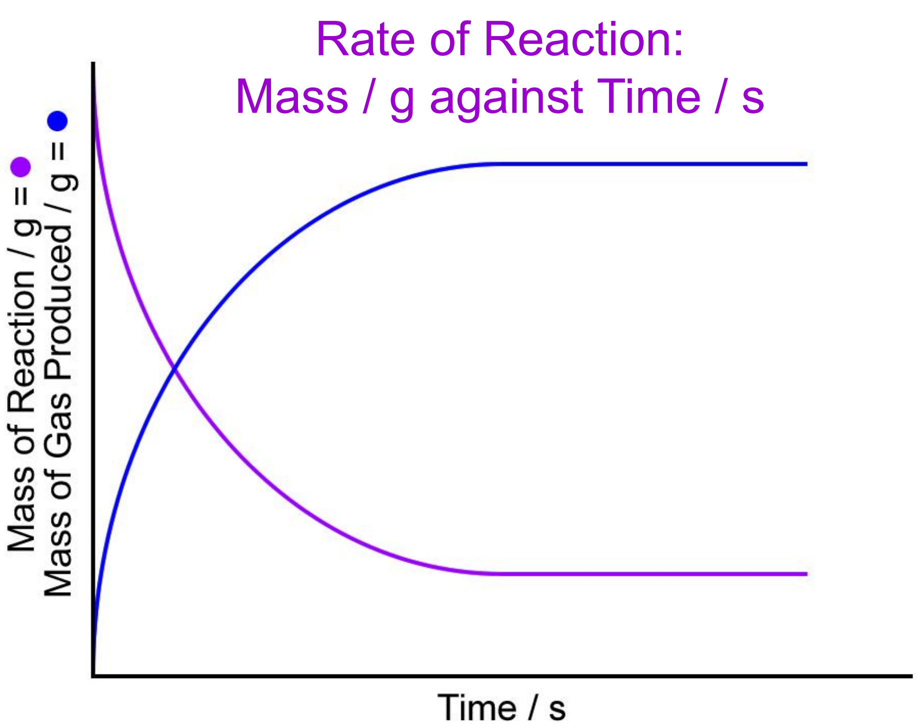
Note: Cotton wool is placed in the mouth of the flask to allow gases to escape, while preventing any of the solution from escaping.

Rate of Reaction

6) Change in *mass* against time.



Note: Cotton wool is placed in the mouth of the flask to allow gases to escape, while preventing any of the solution from escaping.



- When the mass reaches a constant value, then the reaction has finished.

Rate of Reaction

6) Change in *mass* against time.

Rate of Reaction

6) Change in *mass* against time.

- Example:

calcium carbonate + hydrochloric acid

↓

calcium chloride + water + carbon dioxide

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

Note: This method of determining the rate of a chemical reaction is *not suitable* if the gas is *very soluble in water*, *e.g.* $NH_3(g)$, $HCl(g)$ and $SO_2(g)$.

Rate of Reaction

6) Change in *mass* against time.

- Example:

calcium carbonate + hydrochloric acid

↓

calcium chloride + water + carbon dioxide

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

Note: A *significant* change in mass is only observed for gases that have a *high relative molecular mass*, *e.g.* CO_2(g), M_r = 44.0. The change in mass is *insignificant* for gases that have a *low relative molecular mass*, *e.g.* H_2(g), M_r = 2.00.

Rate of Reaction

What variables could you measure to determine the rate of the following reaction?

sodium carbonate + nitric acid

↓

sodium nitrate + water + carbon dioxide

$$Na_2CO_3(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$$

$$\Delta H = -50\ \text{kJ}$$

Rate of Reaction

What variables could you measure to determine the rate of the following reaction?

sodium carbonate + nitric acid

$\downarrow$

sodium nitrate + water + carbon dioxide

$$Na_2CO_3(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$$

$$\Delta H = -50\ kJ$$

- Change in temperature against time.
- Change in volume of gas against time.
- Change in mass against time.
- Change in pH against time.

Rate of Reaction

What must happen in order for two chemicals to *react* with each other?

Rate of Reaction

Firstly, a few important terms…

- **Exothermic**: An *exothermic* process *releases energy*, usually in the form of heat and light, *into* the surroundings.

- **Endothermic**: An *endothermic* process *absorbs energy*, usually in the form of heat and light, *from* the surroundings.

- **Activation Energy**: The *minimum* amount of energy that must be provided / supplied in order for a chemical reaction to take place.

Rate of Reaction

Activation Energy

- Activation energy is the minimum amount of energy that must be supplied to a chemical system in order for a reaction to take place.

Rate of Reaction

Activation Energy

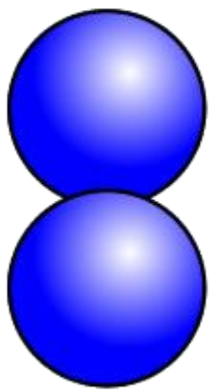
- For example, the phosphorus and sulfur in a match head will not react with the oxygen in the air until the match is struck against the sandpaper on the side of the match box.

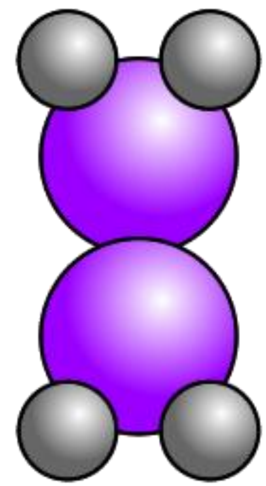
Rate of Reaction

Activation Energy

- Friction between the match head and sandpaper generates heat energy which is used to break chemical bonds in the molecules of phosphorus, sulfur and oxygen, thus allowing them to react.

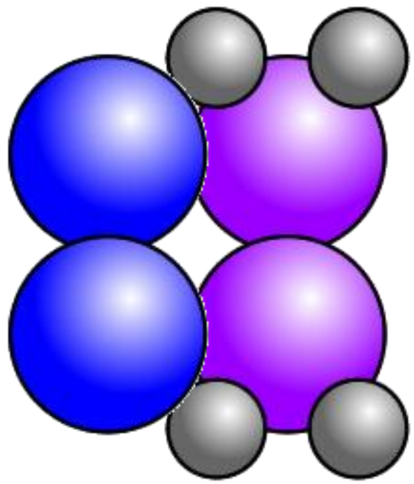
Rate of Reaction





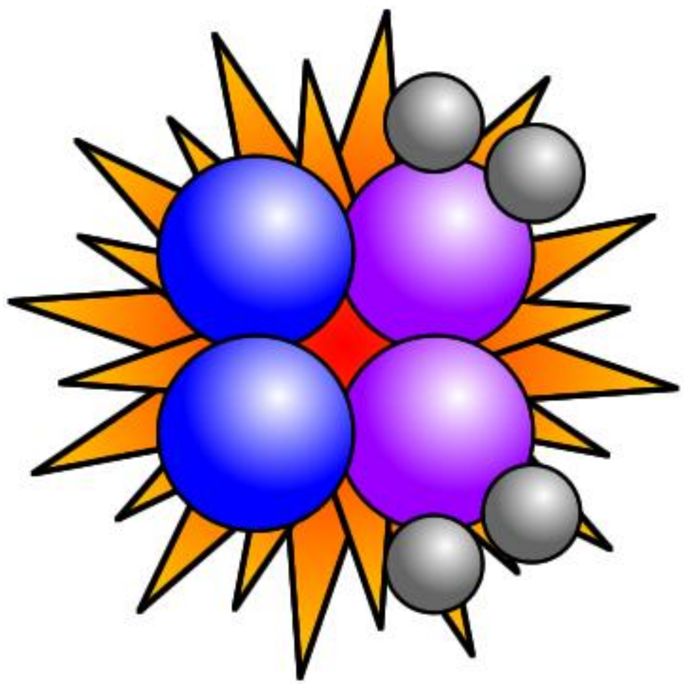
1) Particles of the two reactants are in a *constant state* of *random motion* (kinetic particle theory).

Rate of Reaction



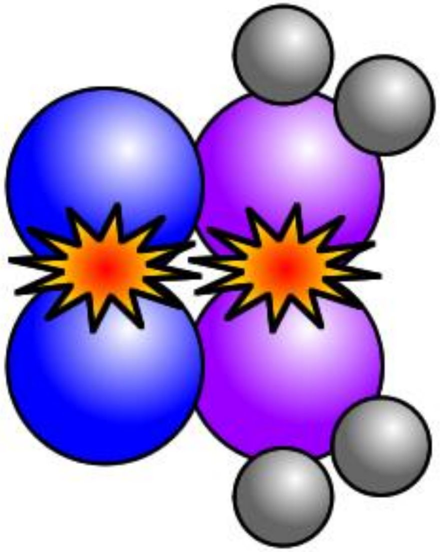
2) Particles of the two reactants eventually *collide*.

Rate of Reaction



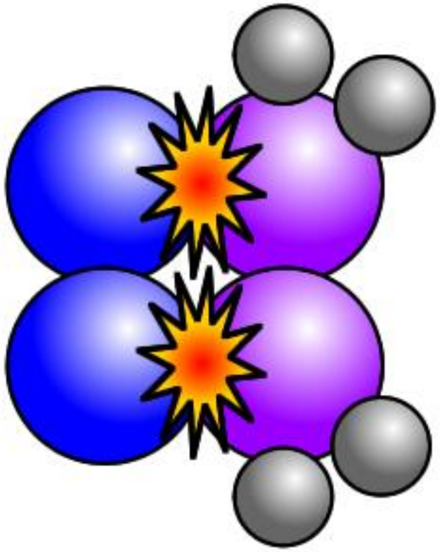
3) The energy of the collision may be less than or greater than the *activation energy* for the reaction.

Rate of Reaction



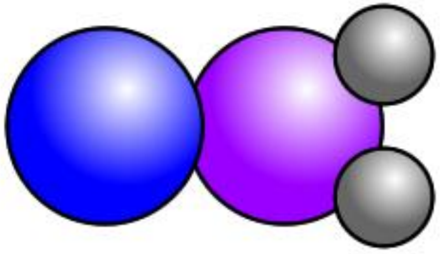
4) If the energy of the collision is *equal to* or *greater* than the activation energy for the reaction, then chemical bonds will be broken (*endothermic*).

Rate of Reaction

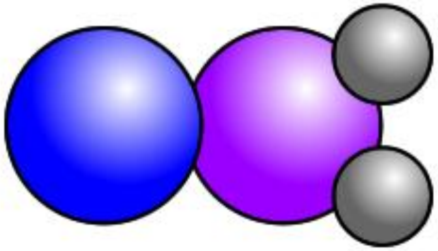


5) And new chemical bonds will be formed (*exothermic*).

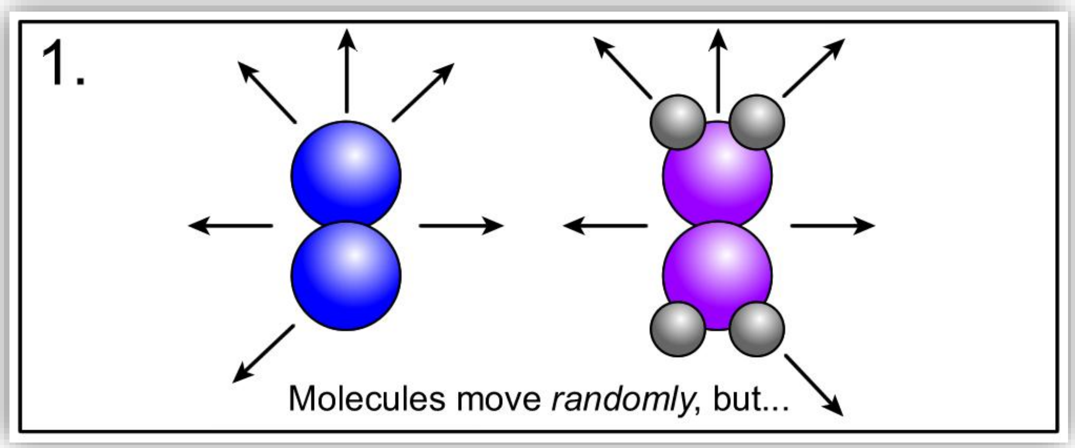
Rate of Reaction



6) Resulting in the formation of new reaction products.



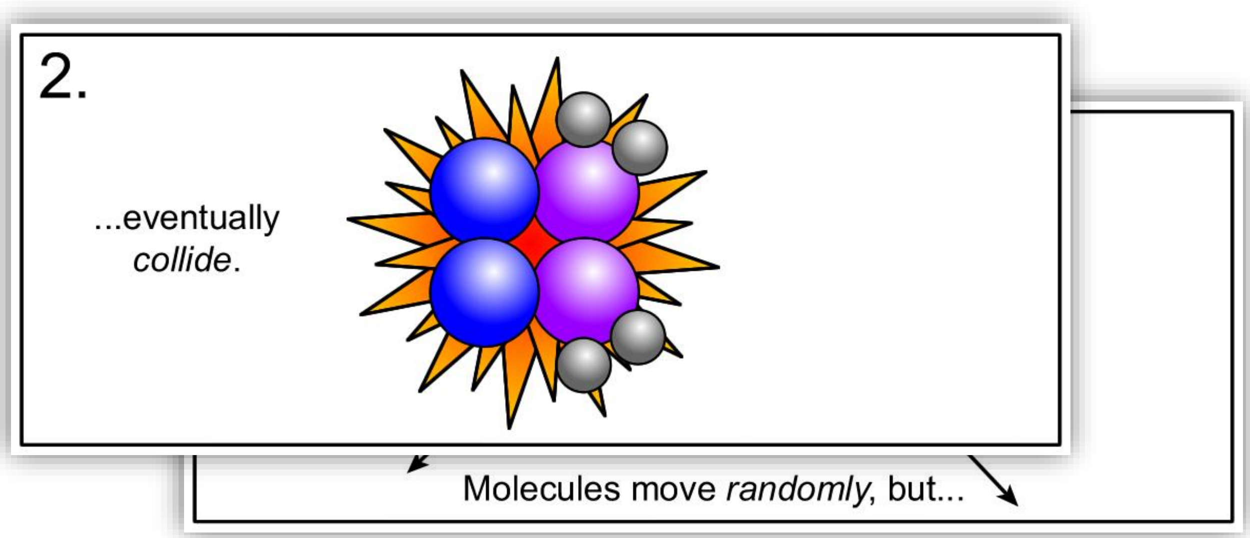
Rate of Reaction



- This is known as…

  collision theory.

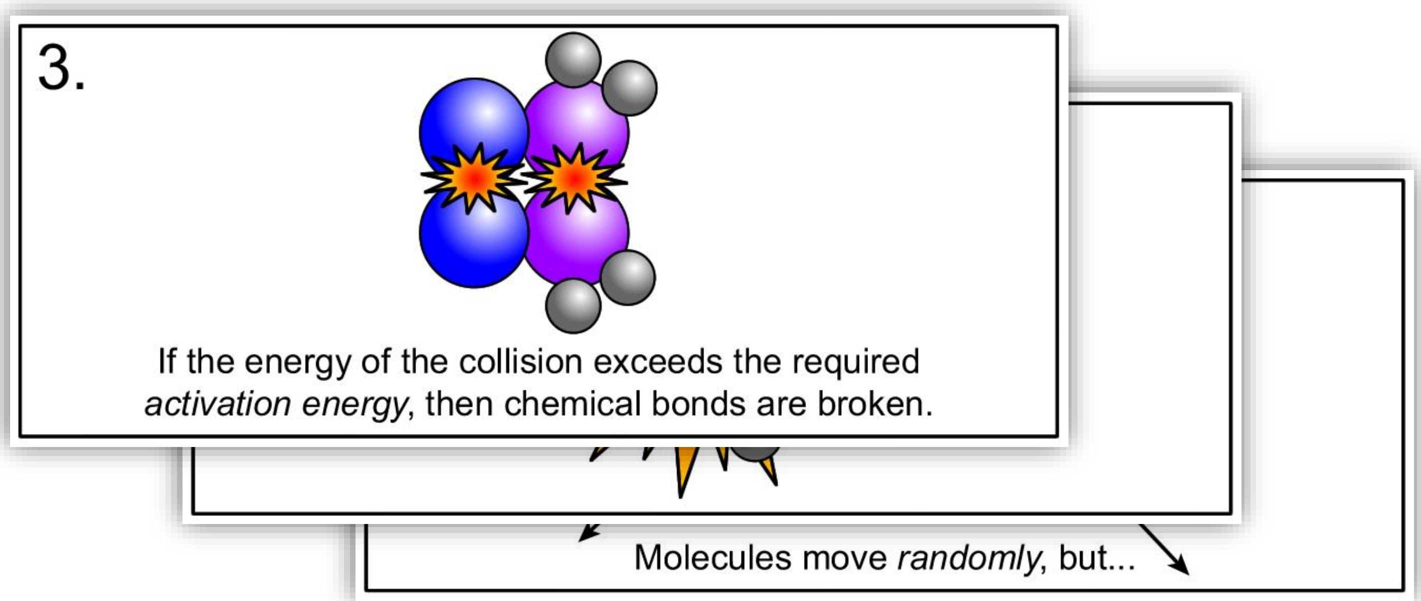
Rate of Reaction



- This is known as…

  collision theory.

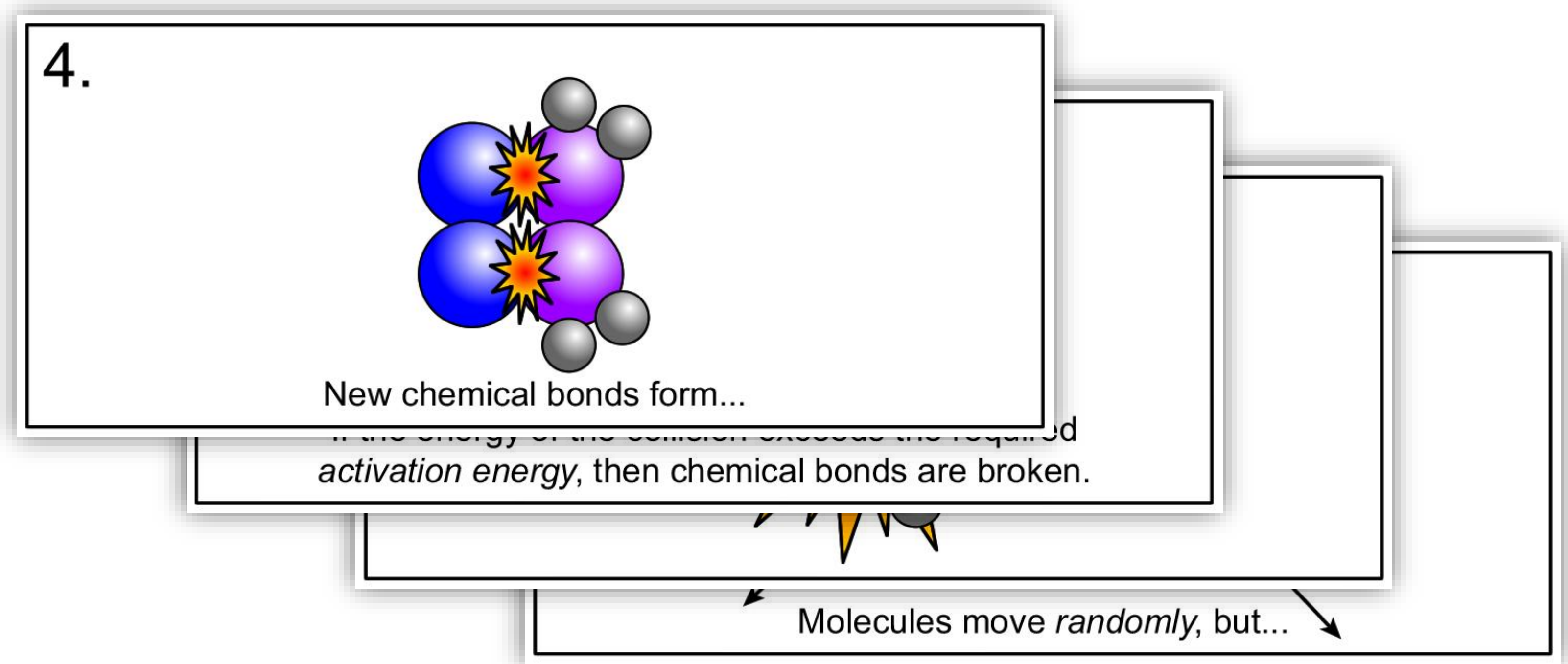
Rate of Reaction



- This is known as…

collision theory.

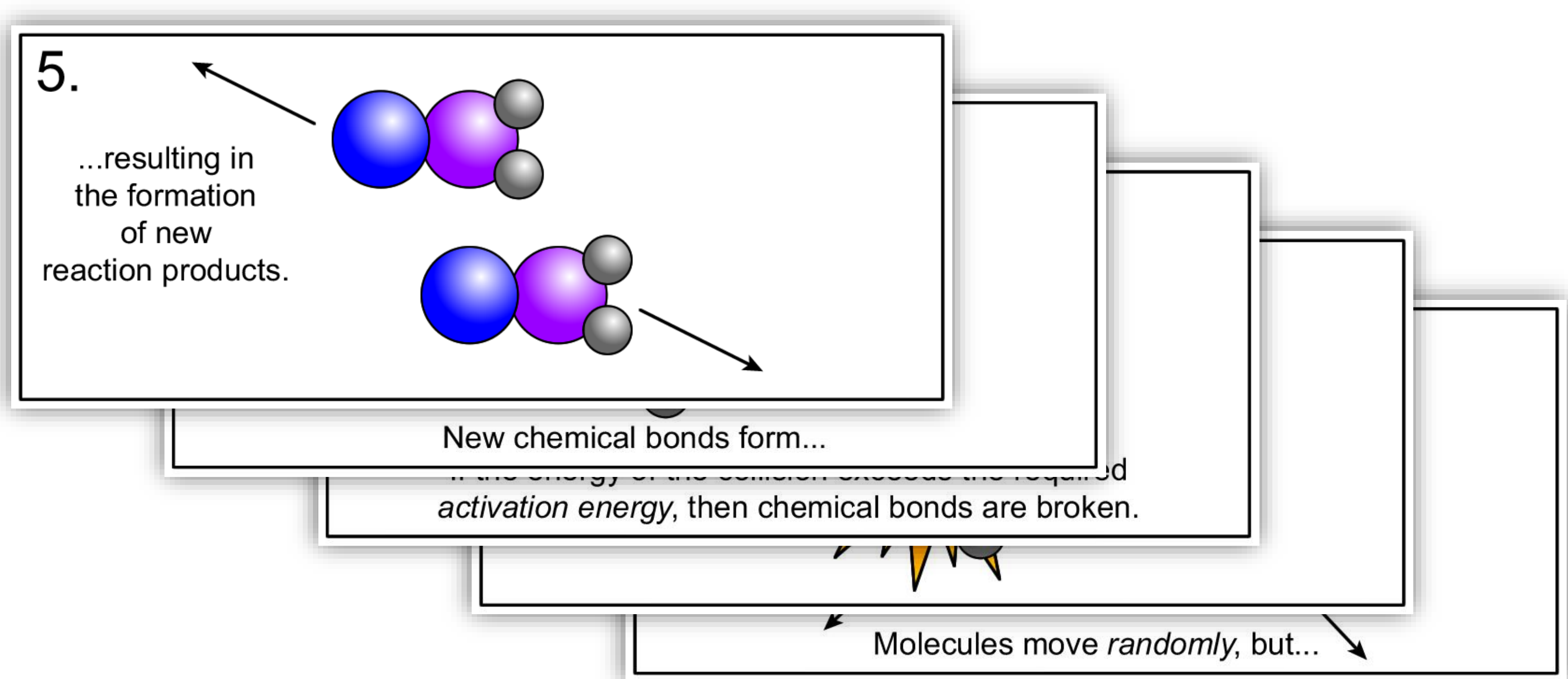
Rate of Reaction



- This is known as…

collision theory.

Rate of Reaction



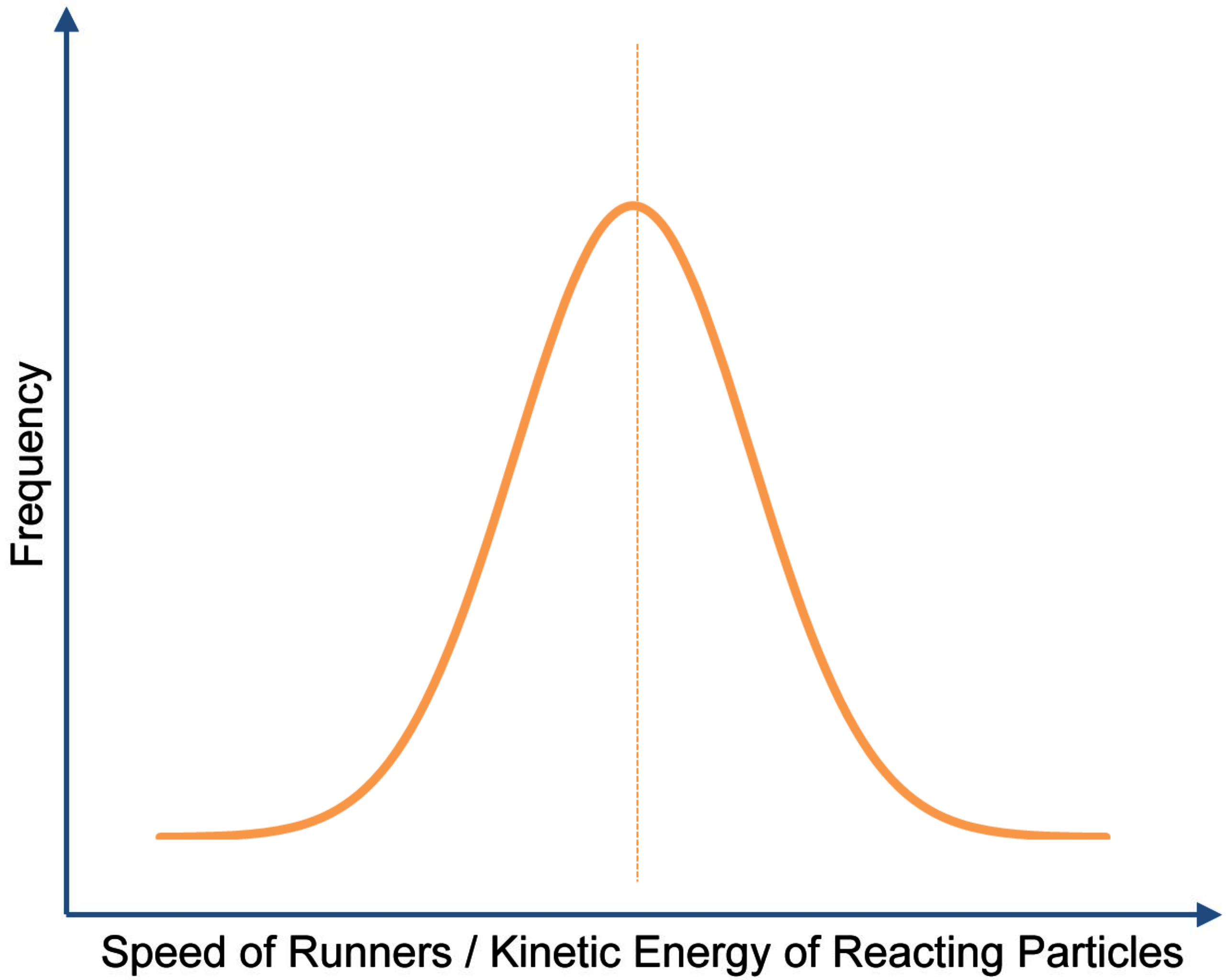
- This is known as…

collision theory.

Rate of Reaction

- **Note:** The particles in a reaction do not all react at once.
- People running a marathon run at different speeds and cross the finishing line at different times.
- It is the same for the particles in a reaction. The particles move at different speeds, with different amounts of kinetic energy, and will collide and react at different times.

Rate of Reaction



Rate of Reaction

• Duration: 18.5 sec.

• For particles to react, they must collide with a minimum amount of energy. Some particles move faster, and posses more kinetic energy, than others.

Rate of Reaction

What *variables* affect the rate of a chemical reaction?

- Experimental Design:

What could the *independent variables* be?

Rate of Reaction

- Any variable that increases the *frequency* with which particles of the reactants collide and / or increases the *energy* with which they collide will increase the rate of a chemical reaction.

Rate of Reaction

- Changing these variables will change the rate of a chemical reaction:

1) Temperature.

2) Concentration of a solution.

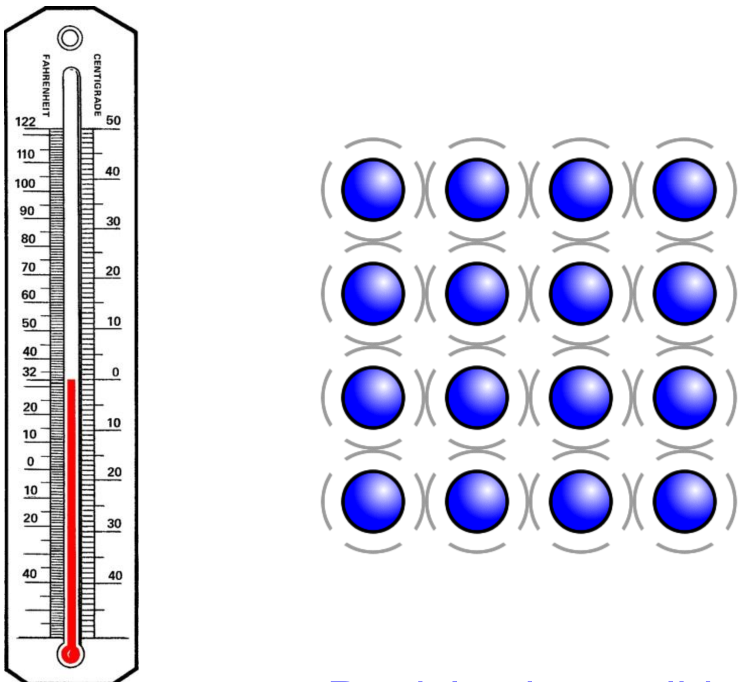
3) Pressure of a gas.

4) Surface area of a solid.

5) Addition of a catalyst or enzyme.

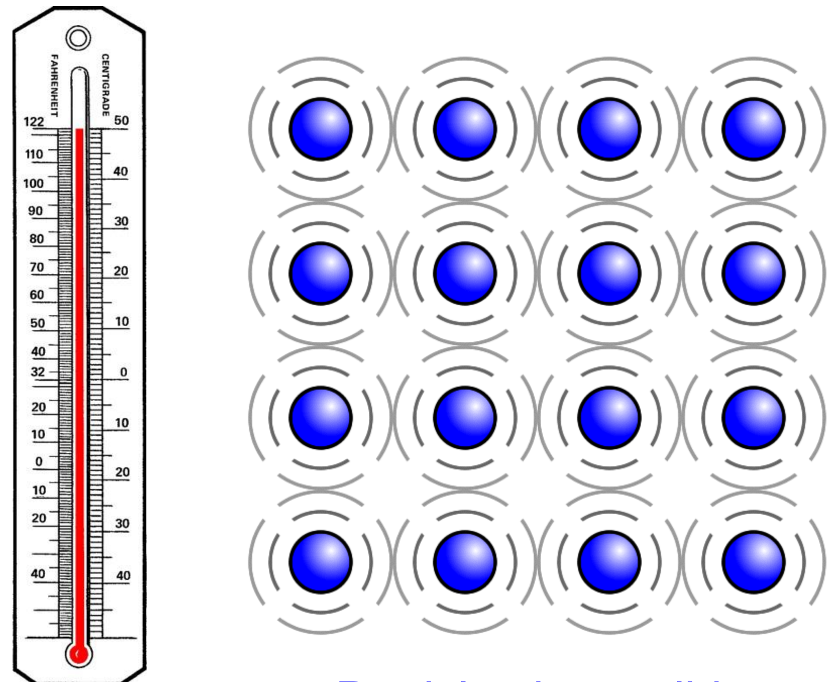
Rate of Reaction

1) Temperature.



- Particles in a solid at *low* temperature.

Relatively low average kinetic energy.



- Particles in a solid at *high* temperature.

Relatively high average kinetic energy.

Rate of Reaction

1) Temperature.

Increasing the temperature will increase the rate of reaction.

- Increasing the temperature of a reaction mixture increases the *average kinetic energy* of the particles in the mixture. The particles collide *more frequently*, and each collision is *more energetic*.

- Increasing the *frequency* of the collisions between the particles increases the chance of a reaction taking place.

- Increasing the *energy* of the collisions between the particles means that more collisions will exceed the reaction's activation energy.

- Increasing the *frequency of effective collisions* between the particles increases the rate of the reaction.

Rate of Reaction

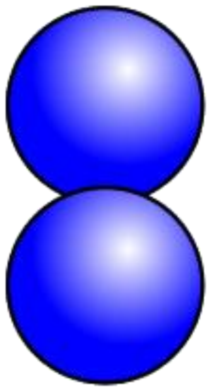
Rate of Reaction

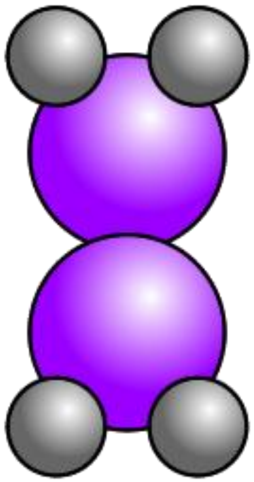
- As a general rule, increasing the temperature of a chemical reaction by *10 °C* will *double the rate* of the reaction.

- Remember the fireflies? They flashed *once every 11 seconds at 12 °C*, but flashed *once every 5 seconds at 21 °C*.

Rate of Reaction

- Low Temperature – Slow Reaction

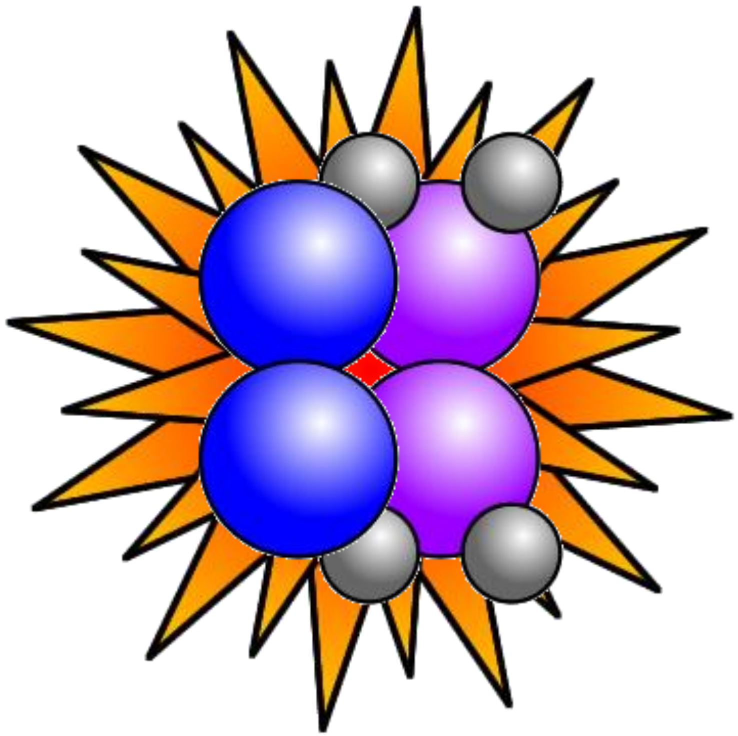




- Frequency of collisions is *low*. Only a *small number* of collisions produce energy > activation energy. The reaction is *slow*.

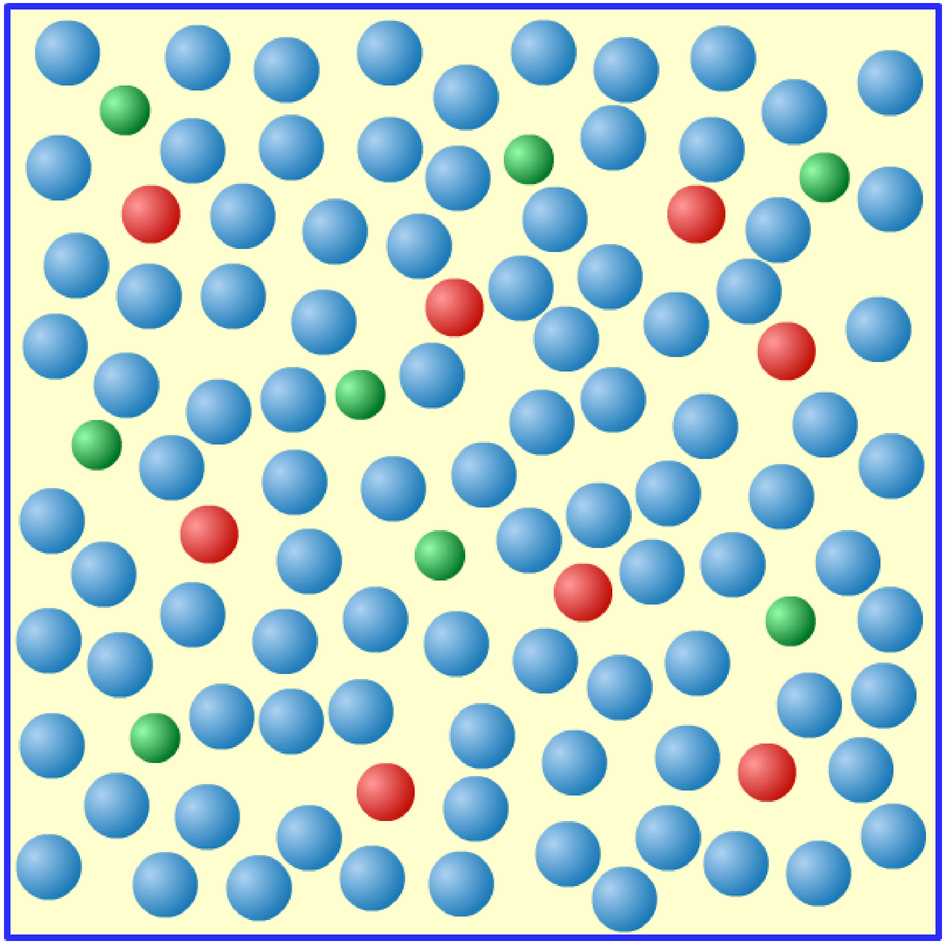
Rate of Reaction

- High Temperature – Fast Reaction



- Frequency of collisions is *high*. *Many* collisions produce energy > activation energy. The reaction is *fast*.

Rate of Reaction



number of collisions between reactants = 0

increase temperature

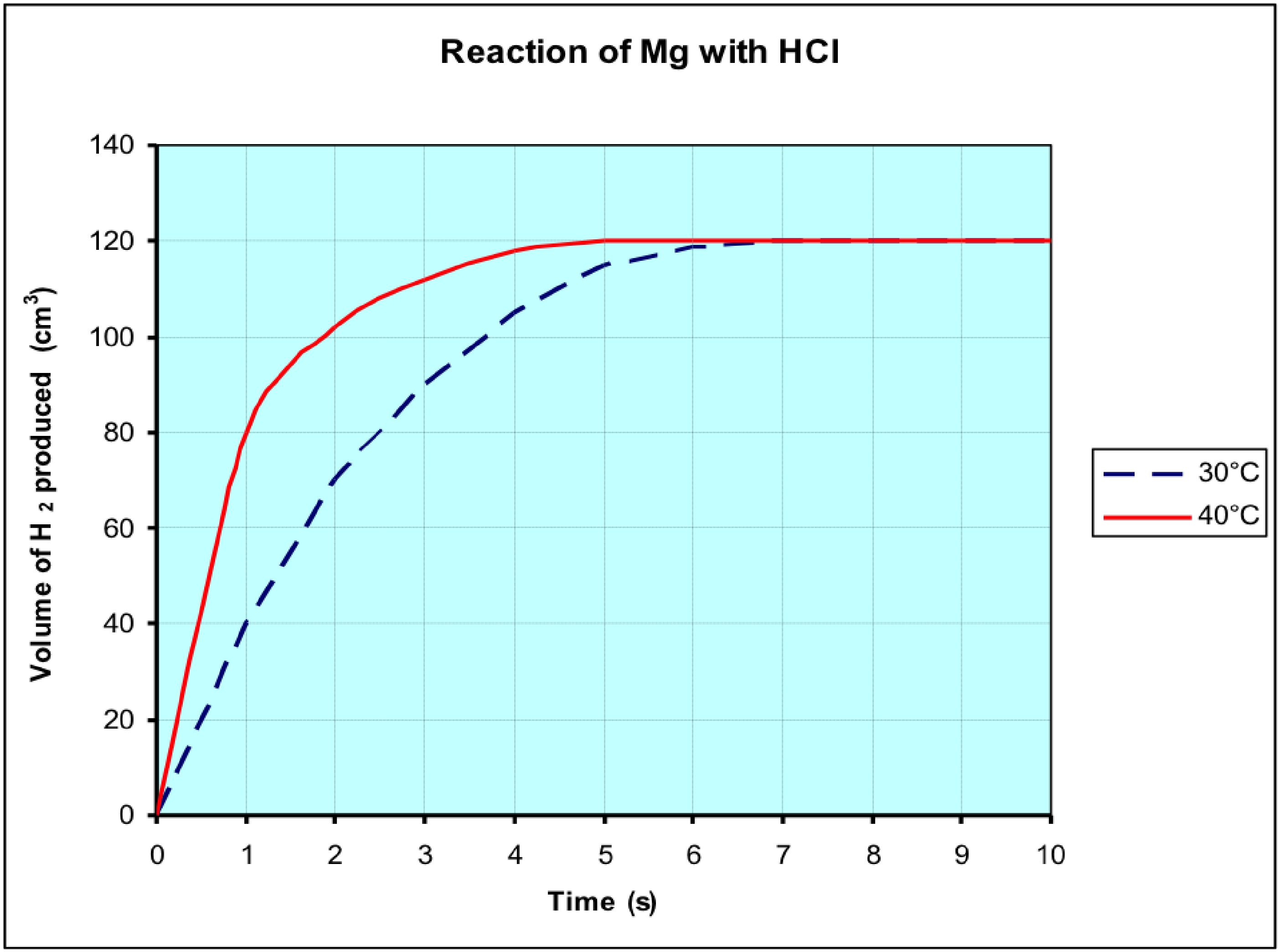
decrease temperature

40°C

0.0 s

Timer

Rate of Reaction



Rate of Reaction

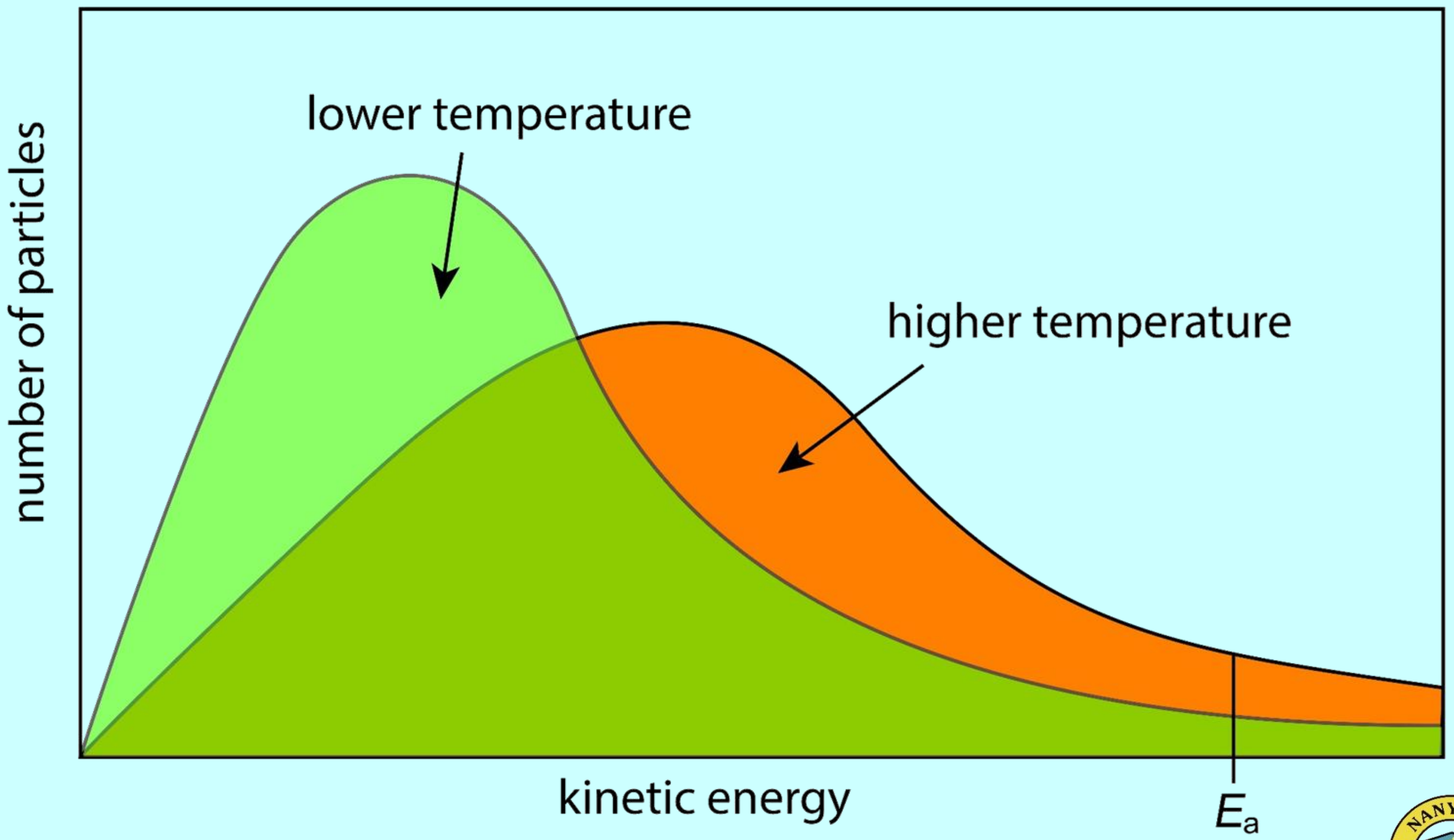
Note: A change in *temperature* will affect the rate of *all* reactions involving *solids*, *solutions*, *gases* or any combination of the three.

Rate of Reaction

Please tell me more about how *temperature* affects the rate of reaction!

Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies



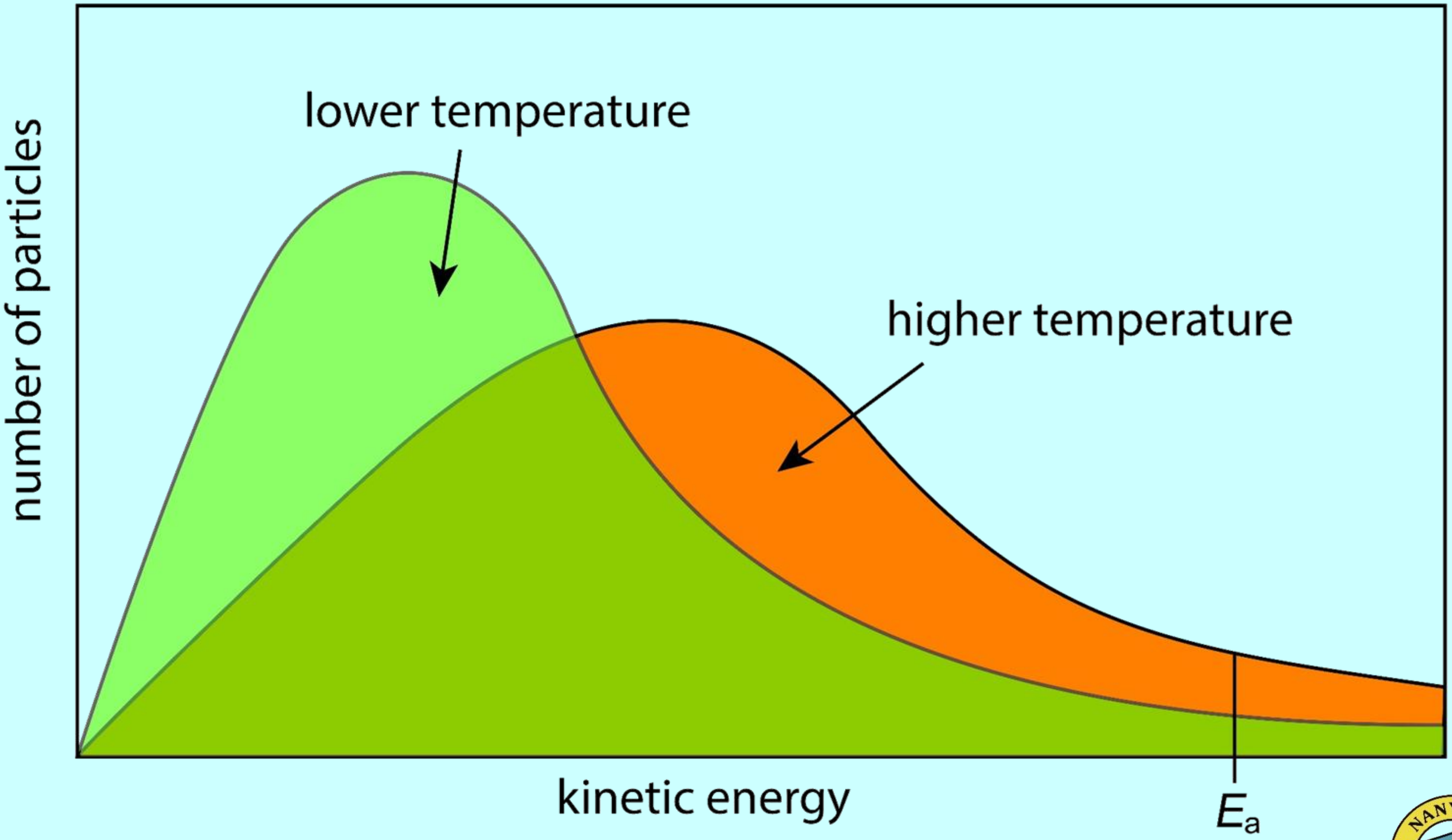
Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies

- Note: The graphs show that the *higher* the *temperature* of the reaction, the *greater* the average *kinetic energy* possessed by the reacting particles.
- Note: The *area* under each graph is directly proportional to the *number of particles* in the reaction mixture.
- Note: The line labelled E_a on the graph represents the *activation energy* of the reaction.

Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies



Rate of Reaction

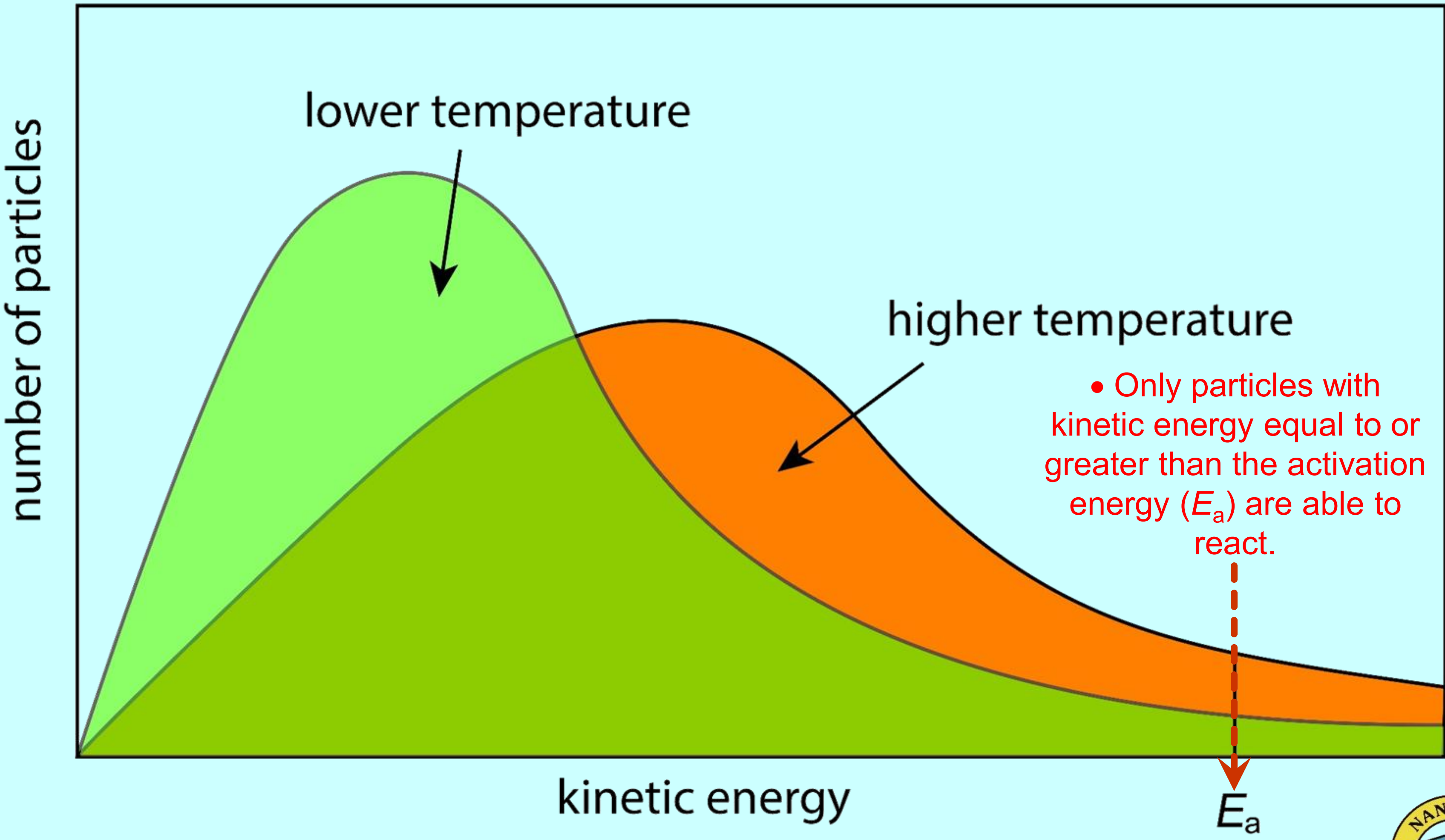
Maxwell-Boltzmann Distribution of Molecular Energies

• Note: The higher the temperature, the greater the proportion of reacting particles that possess kinetic energy equal to or greater than the *activation energy* of the reaction.

• Consequence: An *increase in temperature* increases the average kinetic energy of the reacting particles. More particles posses kinetic energy equal to or greater than the reaction’s activation energy. This *increases the frequency of effective collisions* between the reacting particles, causing the reaction to occur at a *faster rate*.

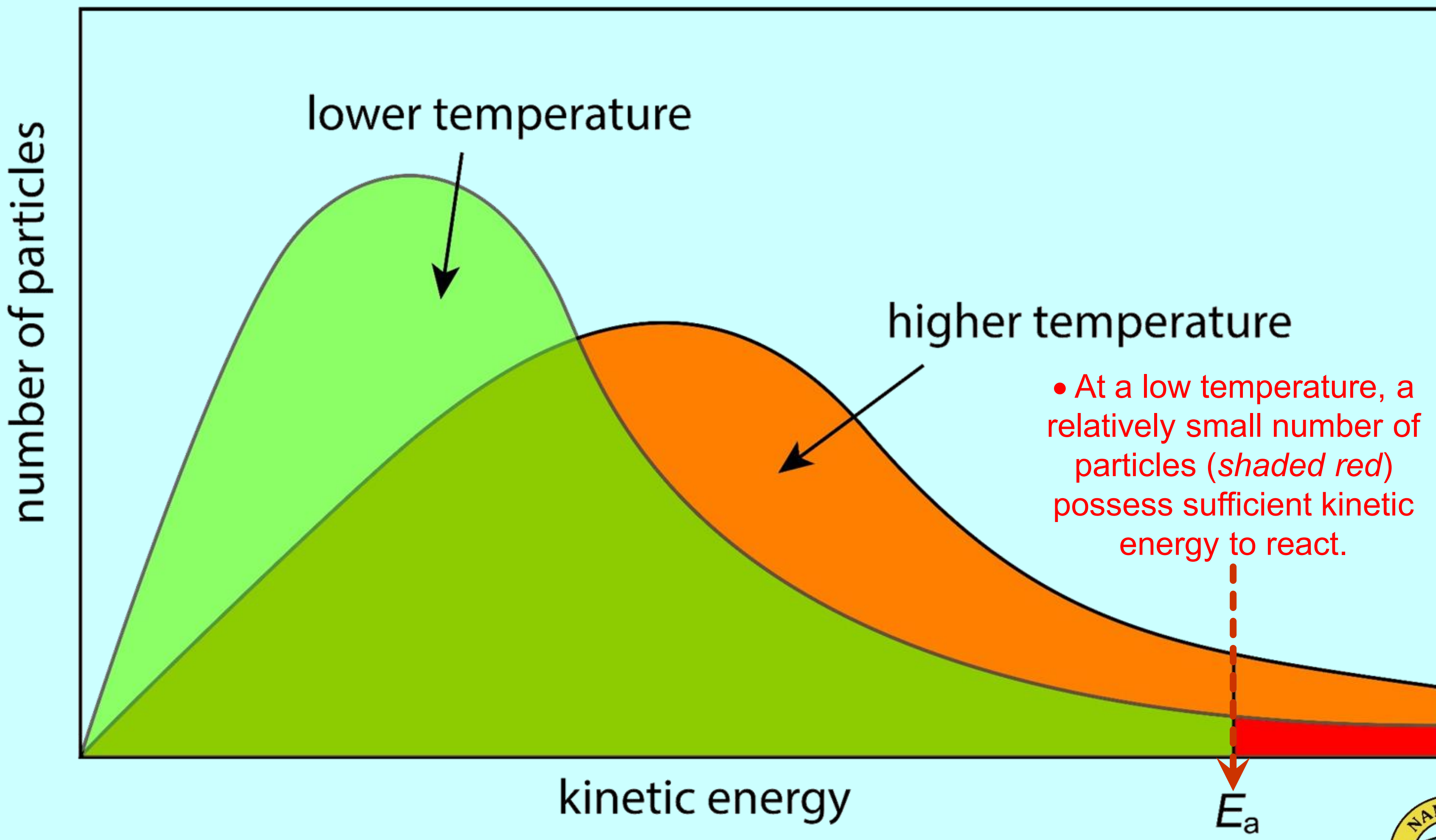
Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies



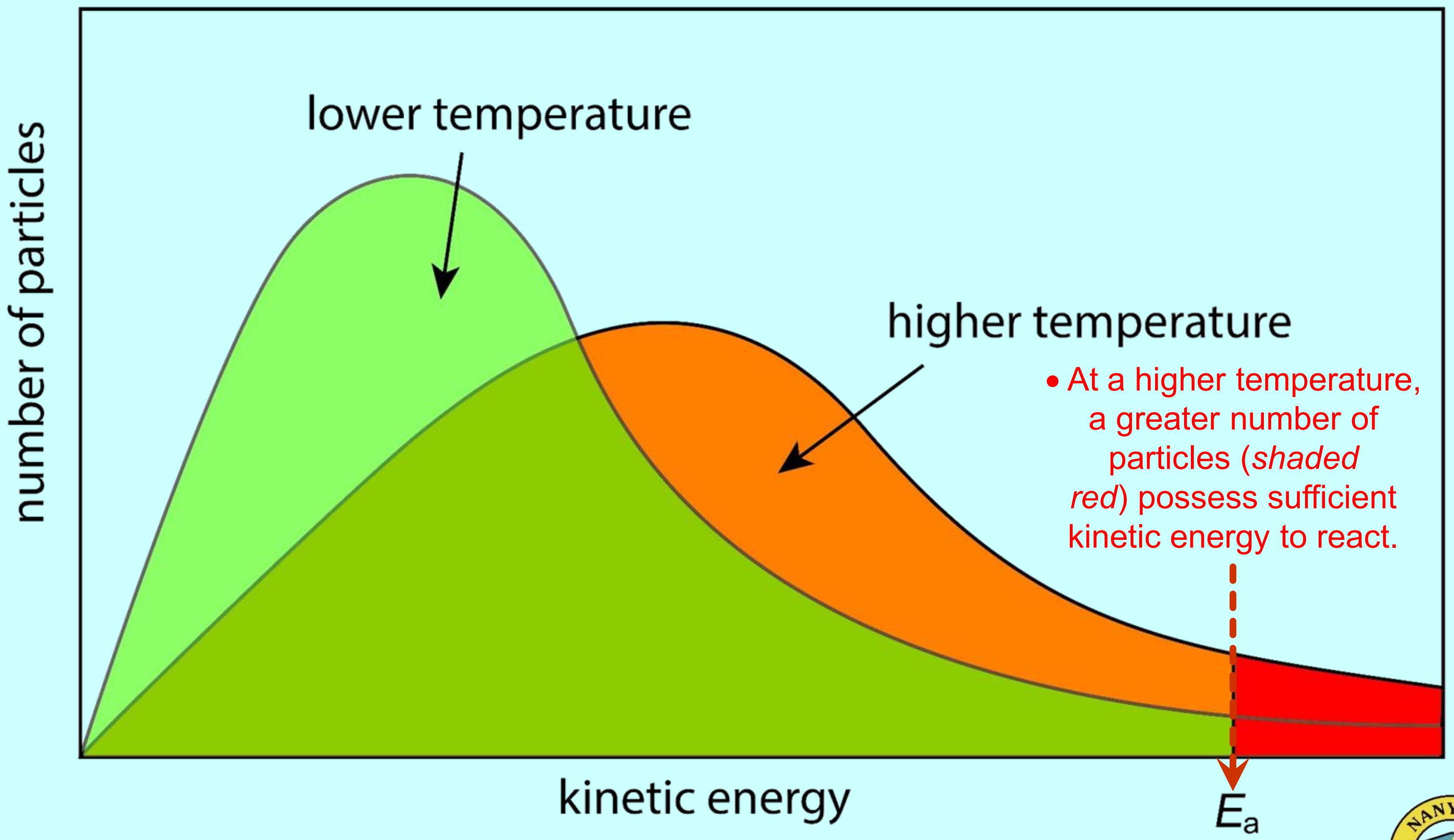
Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies



Rate of Reaction

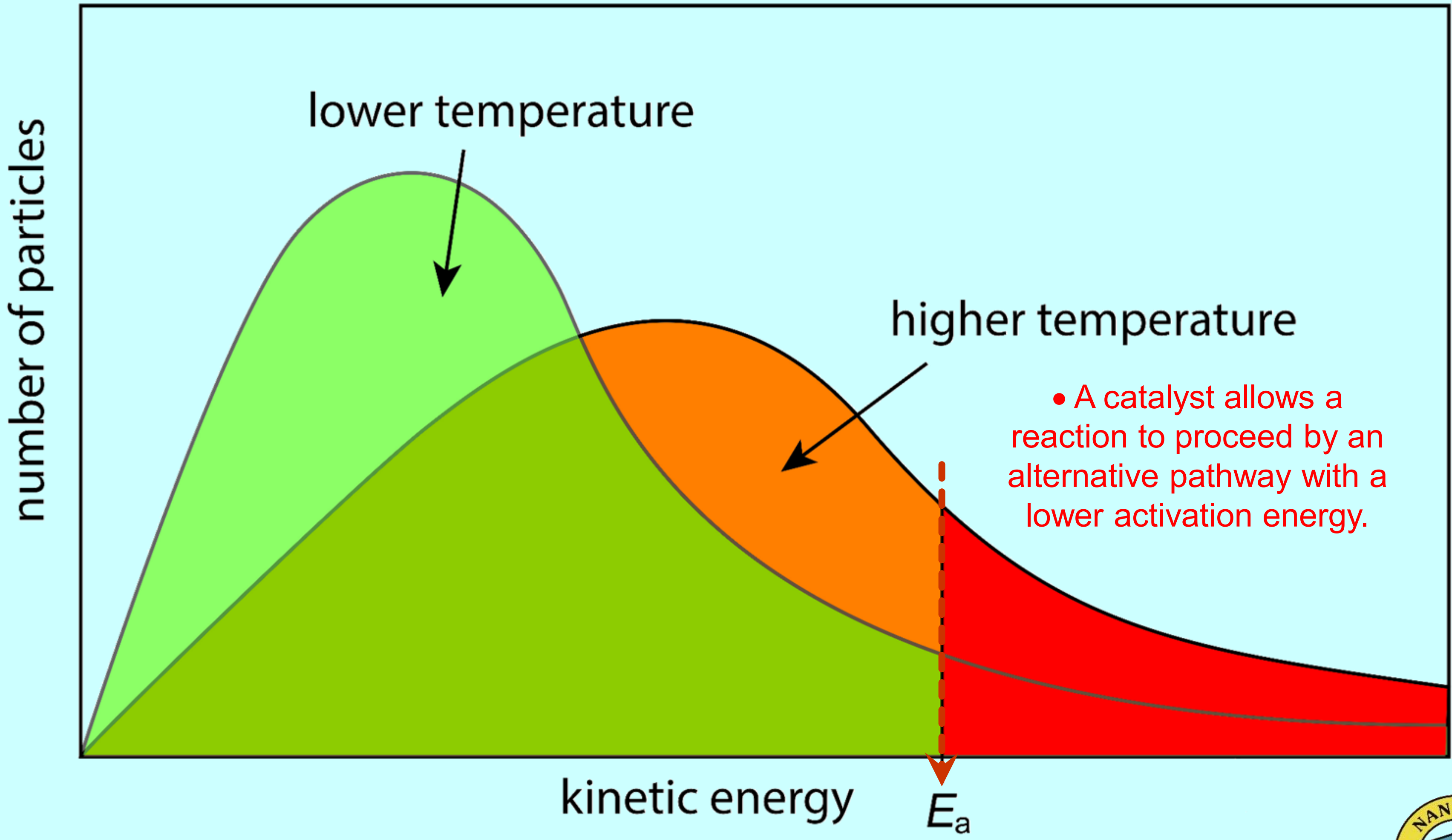
Maxwell-Boltzmann Distribution of Molecular Energies



- At a higher temperature, a greater number of particles (*shaded red*) possess sufficient kinetic energy to react.

Rate of Reaction

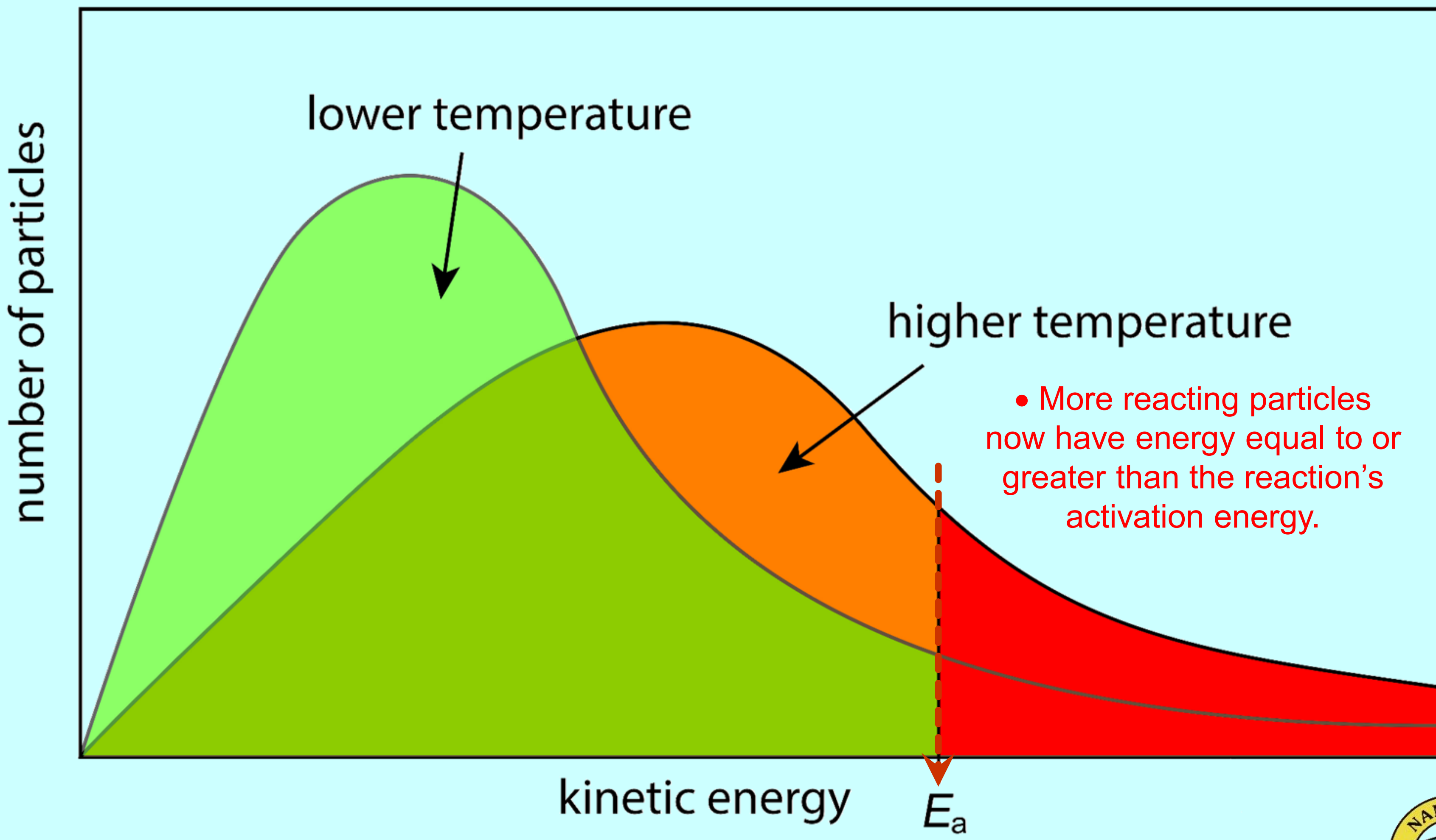
Maxwell-Boltzmann Distribution of Molecular Energies



- A catalyst allows a reaction to proceed by an alternative pathway with a lower activation energy.

Rate of Reaction

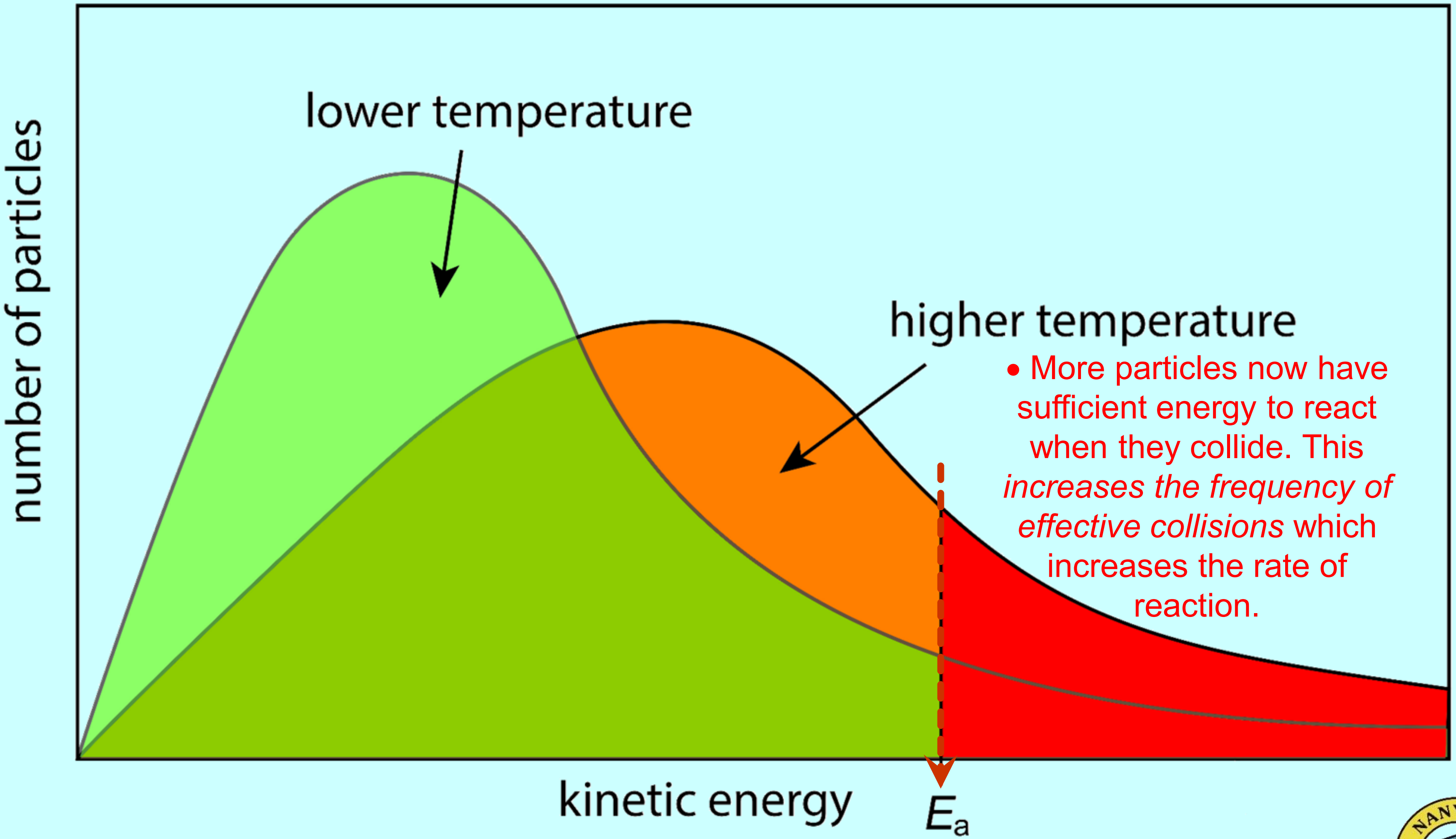
Maxwell-Boltzmann Distribution of Molecular Energies



- More reacting particles now have energy equal to or greater than the reaction's activation energy.

Rate of Reaction

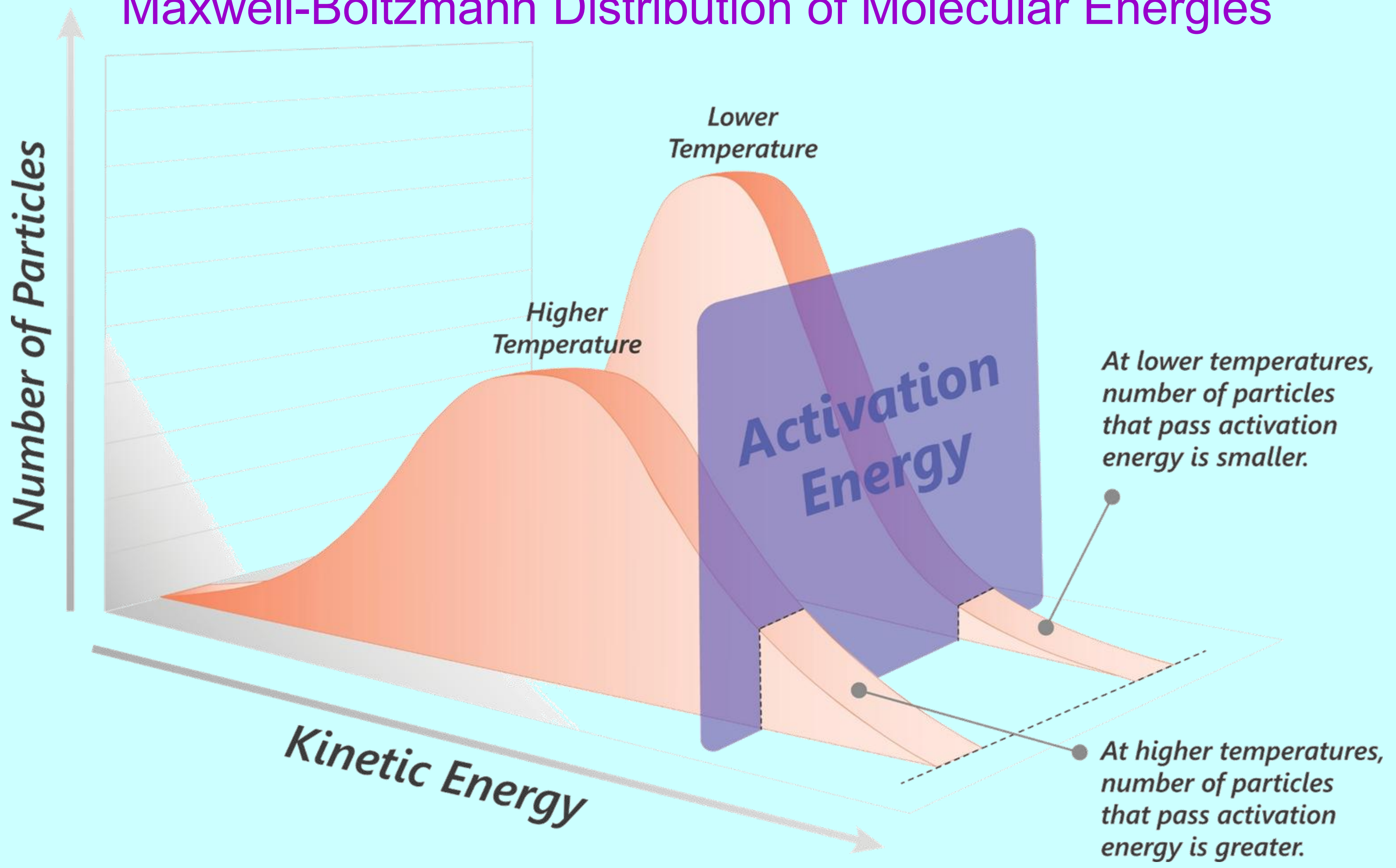
Maxwell-Boltzmann Distribution of Molecular Energies



- More particles now have sufficient energy to react when they collide. This *increases the frequency of effective collisions* which increases the rate of reaction.

Rate of Reaction

Maxwell-Boltzmann Distribution of Molecular Energies



Rate of Reaction

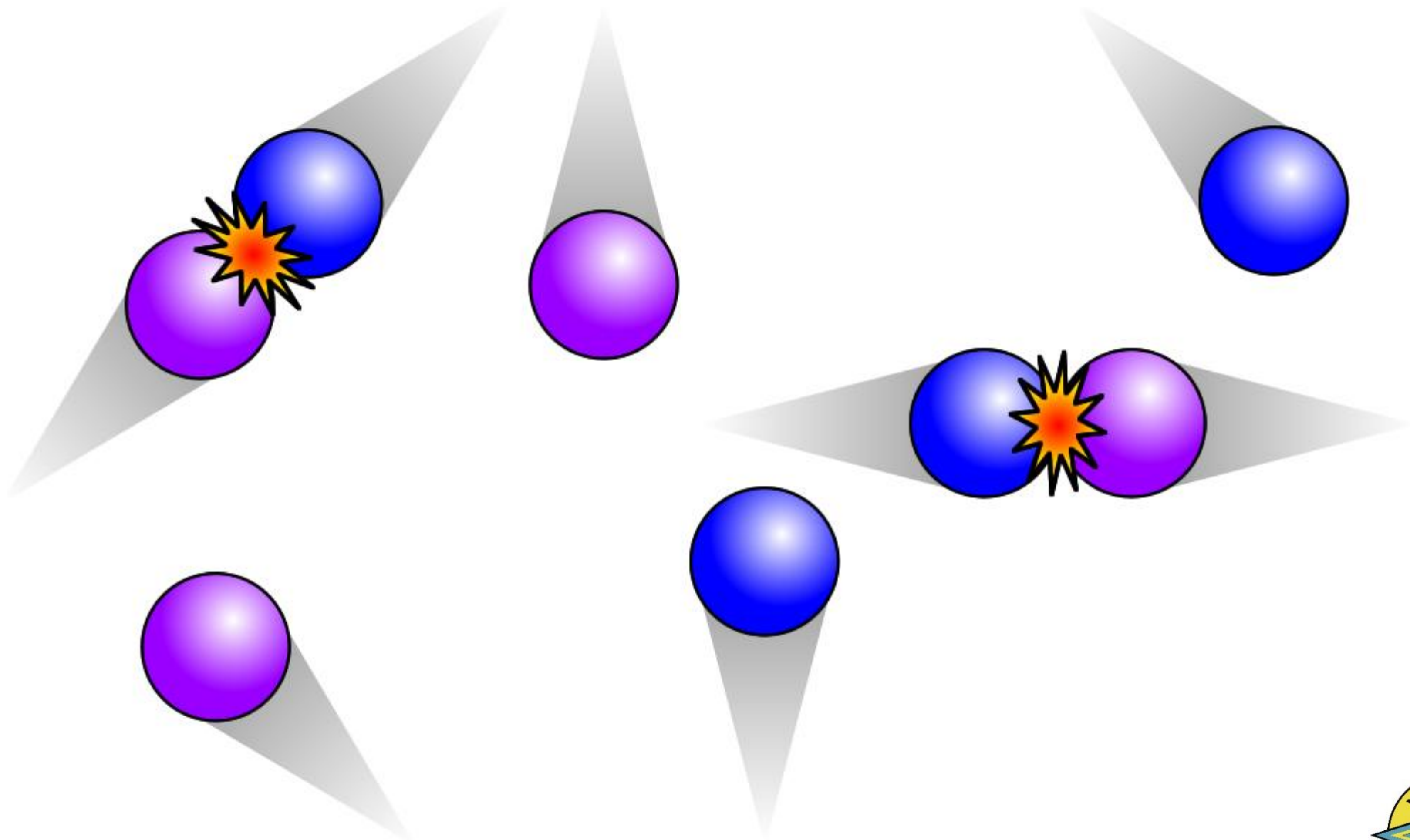
2) Concentration (of a solution).

Increasing the concentration of a solution will increase the rate of reaction.

- The more concentrated a solution is, *the more particles it contains per unit volume*.
- Because there are more particles per unit volume, the particles are *closer together* and therefore collide *more frequently*.
- Increasing the *frequency of effective collisions* between the particles increases the rate of the reaction.
- **NOTE**: A change in concentration does **NOT** affect the *energy* of the collisions between the particles.

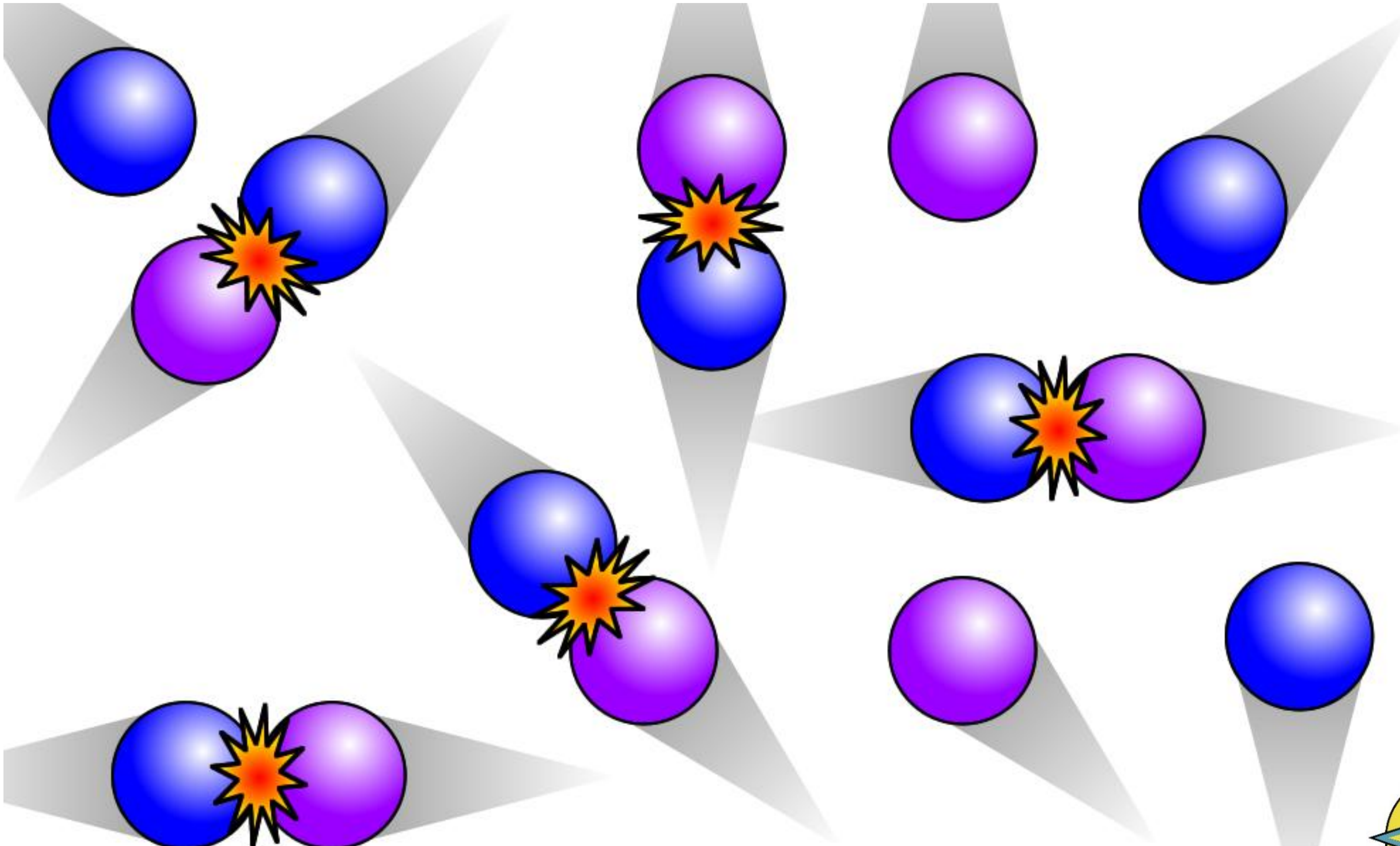
Rate of Reaction

- Low Concentration – Slow Reaction

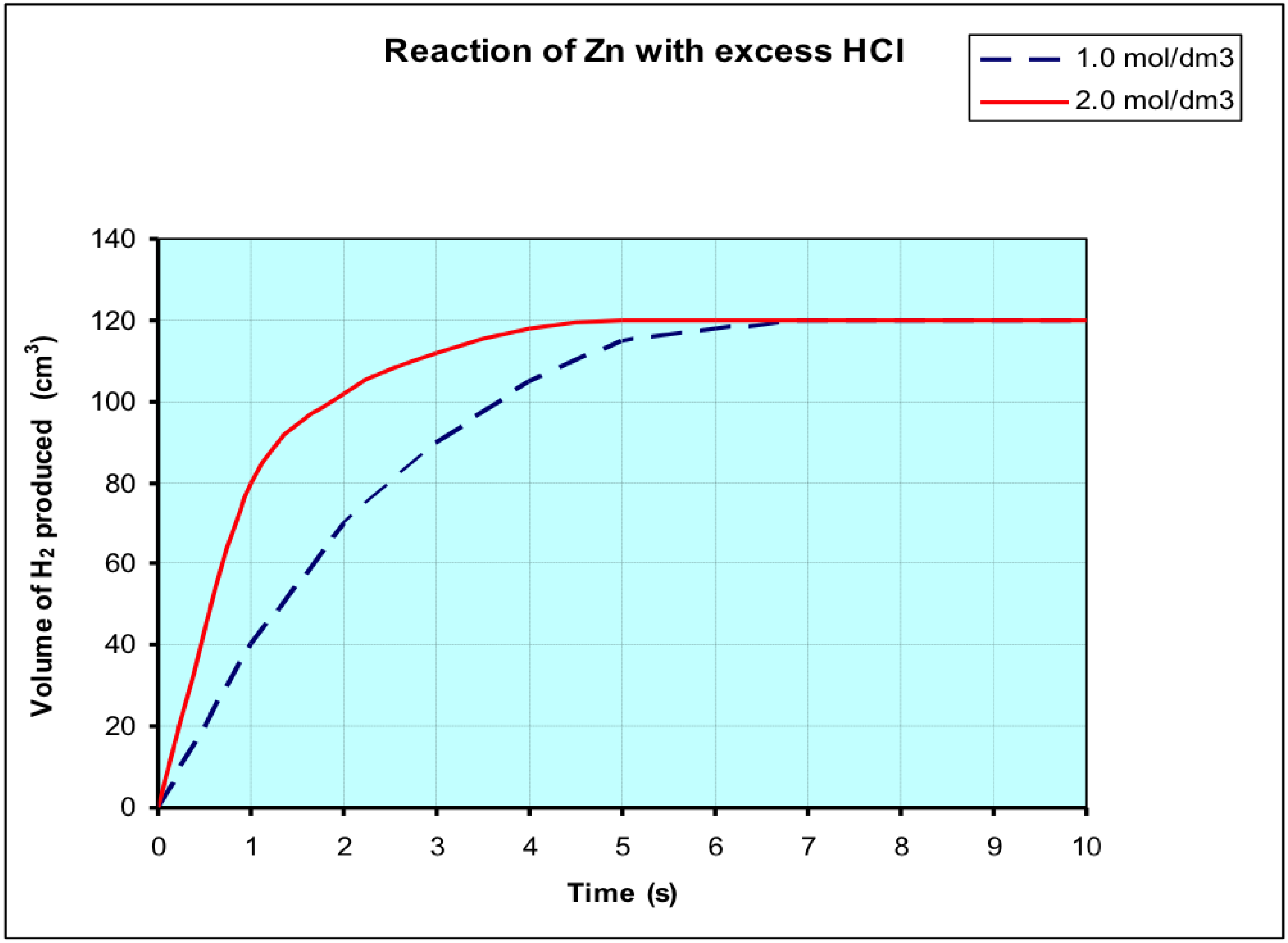


Rate of Reaction

- High Concentration – Fast Reaction



Rate of Reaction



Rate of Reaction

3) Pressure (of a gas).

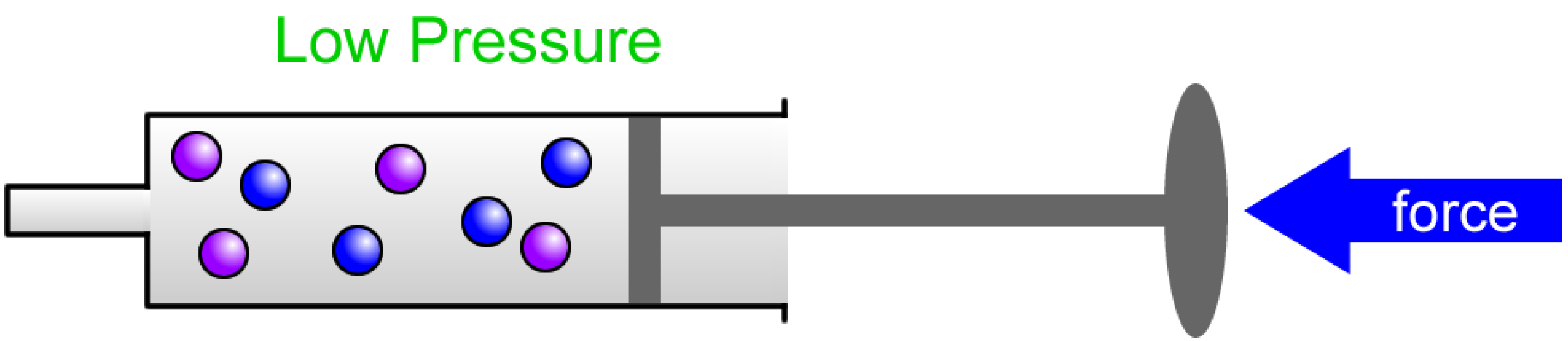
Increasing the pressure of a gas will increase the rate of reaction.

- Increasing the pressure of a gas forces the particles *closer together* (kinetic particle theory – a gas can be compressed).
- There are *more particles per unit volume*. Because the particles are *closer together*, they collide *more frequently*.
- Increasing the *frequency of effective collisions* between the particles increases the rate of the reaction.
- **NOTE**: A change in pressure does **NOT** affect the *energy* of the collisions between the particles.

Rate of Reaction

3) Pressure (of a gas).

Increasing the pressure of a gas will increase the rate of reaction.



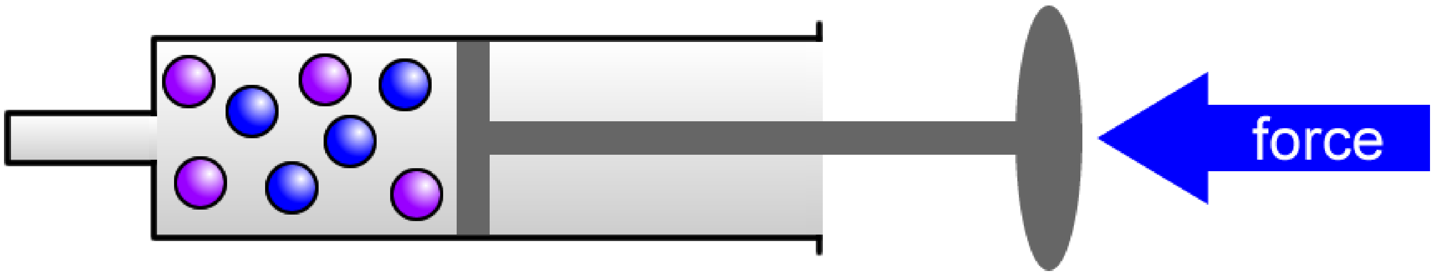
- At a high pressure, there are more particles per unit volume. Because the particles are closer together, they collide *more frequently*. This *increases the frequency of effective collisions*, between the particles which *increases the rate of the reaction*.

Rate of Reaction

3) Pressure (of a gas).

Increasing the pressure of a gas will increase the rate of reaction.

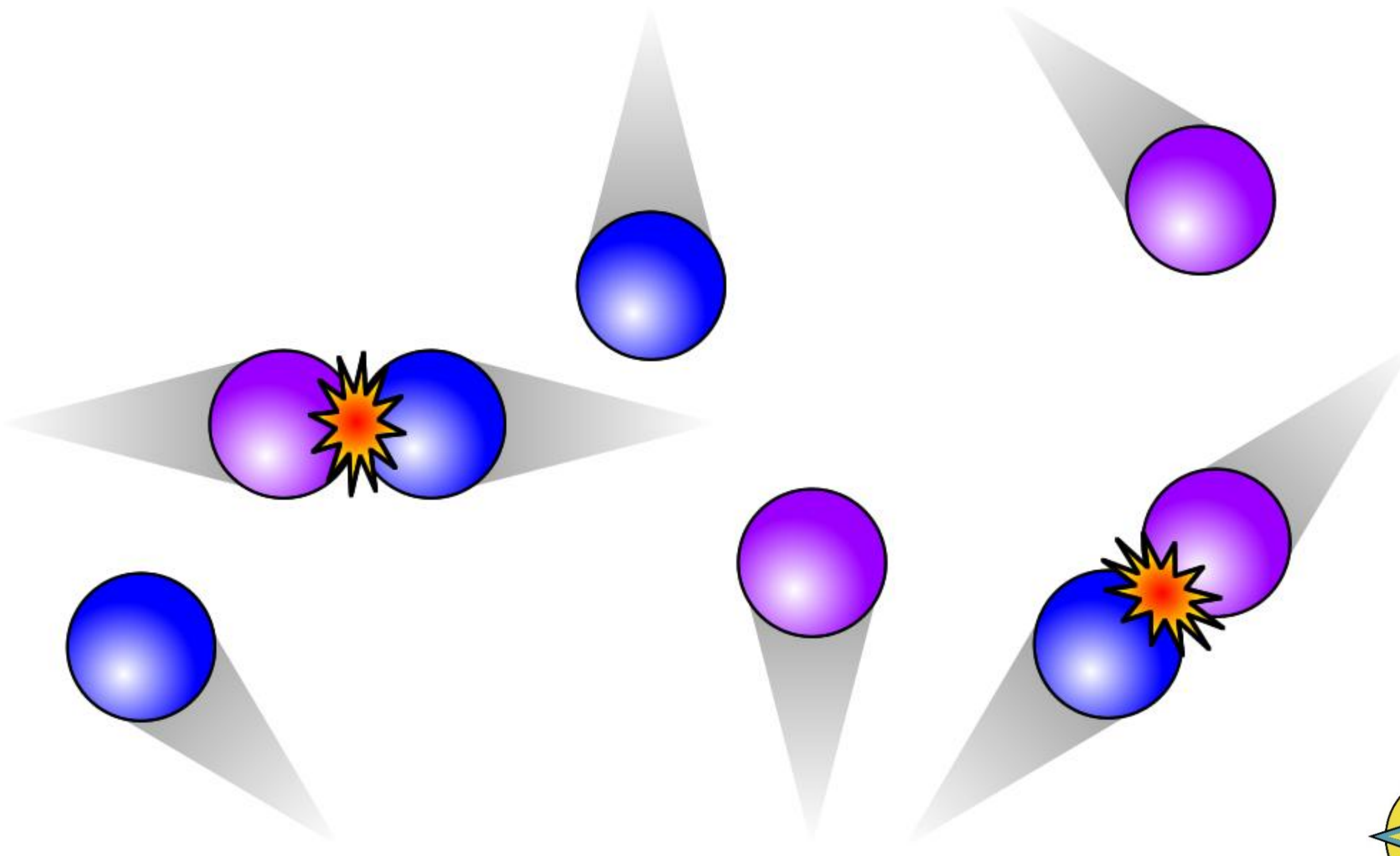
High Pressure



- At a high pressure, there are more particles per unit volume. Because the particles are closer together, they collide *more frequently*. This *increases the frequency of effective collisions*, between the particles which *increases the rate of the reaction*.

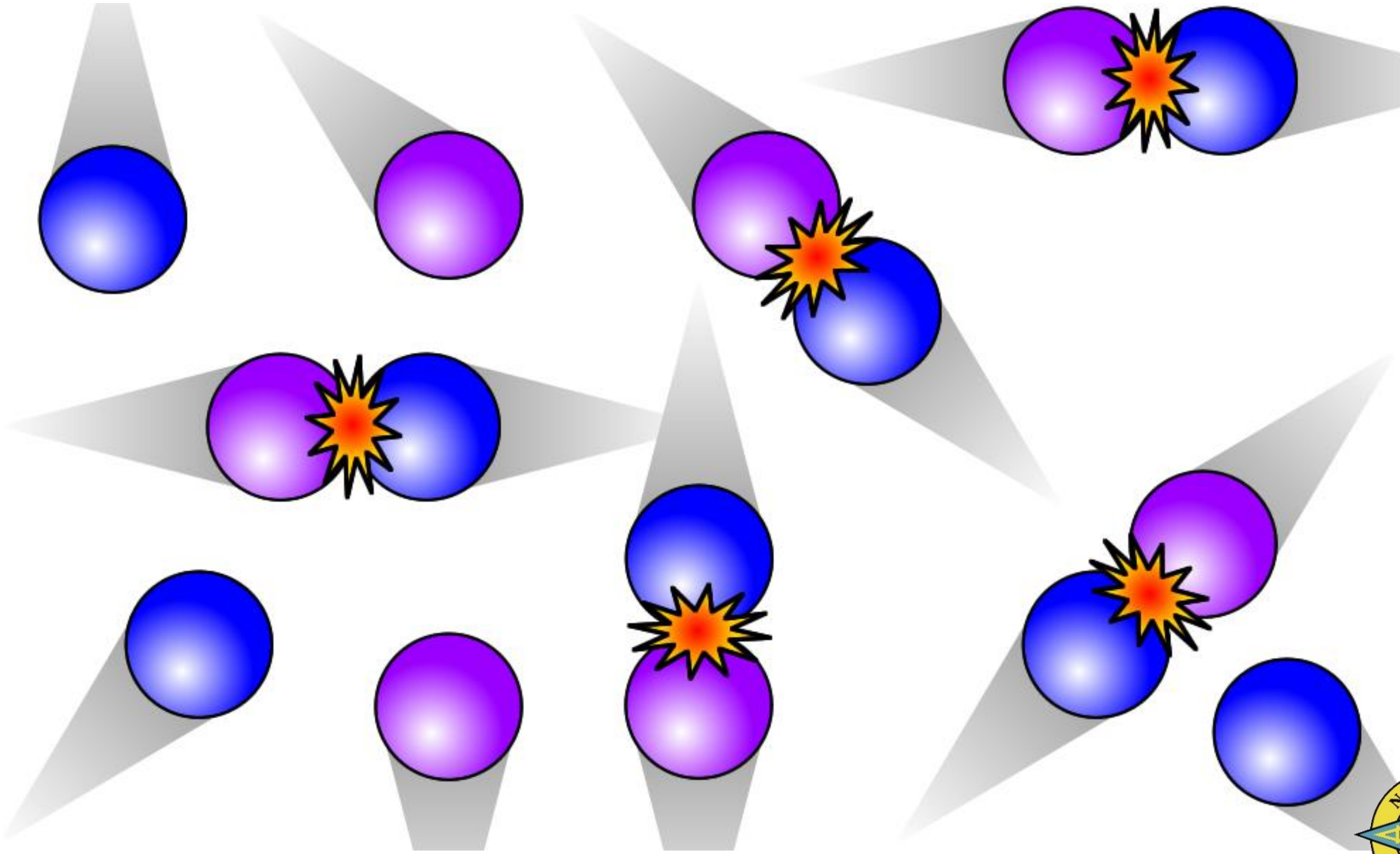
Rate of Reaction

- Low Pressure – Slow Reaction



Rate of Reaction

- High Pressure – Fast Reaction



Rate of Reaction

4) Surface area (of a solid).

Increasing the surface area to volume ratio of a solid will increase the rate of reaction.

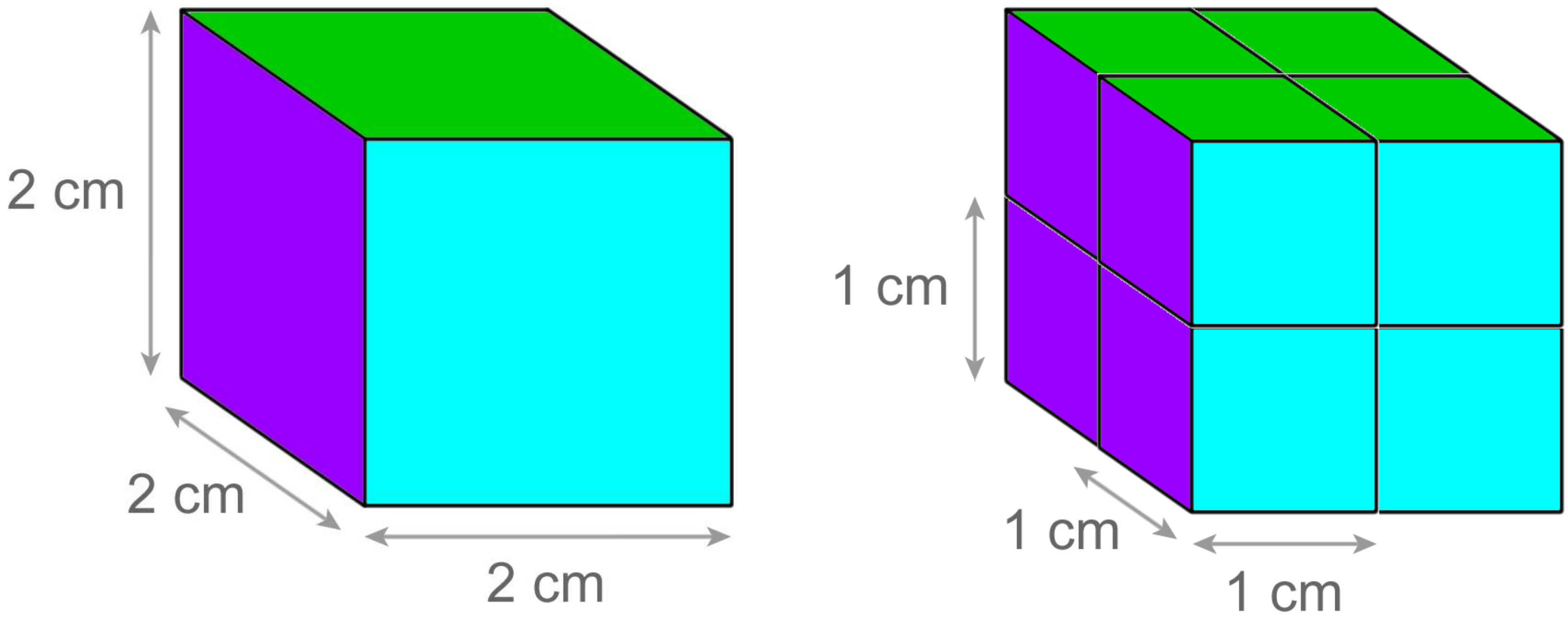
- Increasing the surface area to volume ratio of a solid reagent increases the area over which collisions between the two reacting chemicals can take place.

- This increases the *frequency* of the collisions between particles of the two reagents.

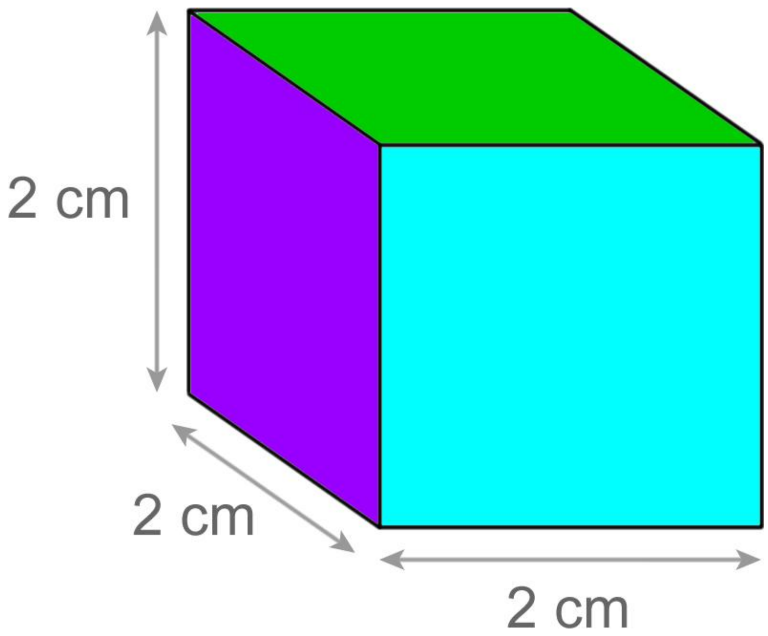
- Increasing the *frequency of effective collisions* between the particles increases the rate of the reaction.

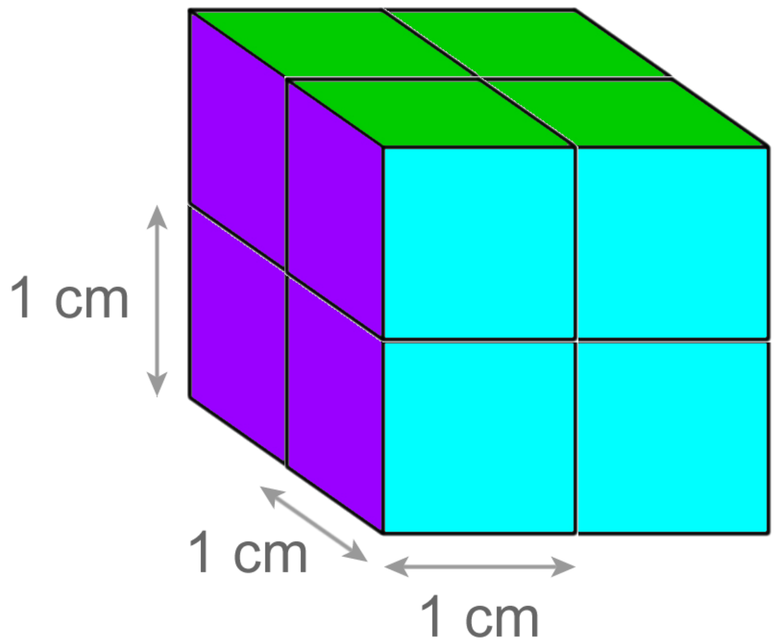
- **NOTE**: A change in surface area does **NOT** affect the *energy* of the collisions between the particles.

Rate of Reaction



Rate of Reaction



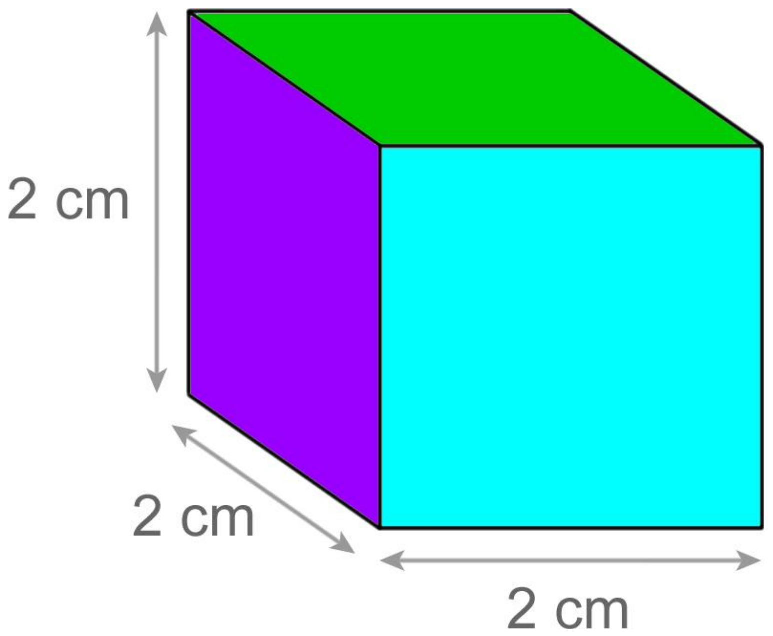


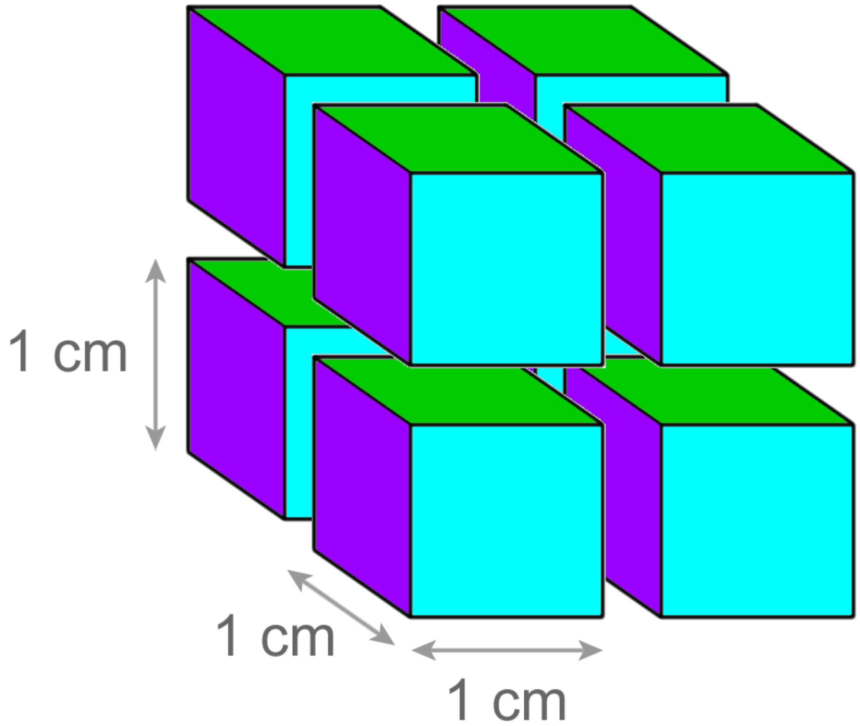
- Surface area of cube

  $= 6 \times (2 \times 2)$

  $= 24\ cm^2$

Rate of Reaction



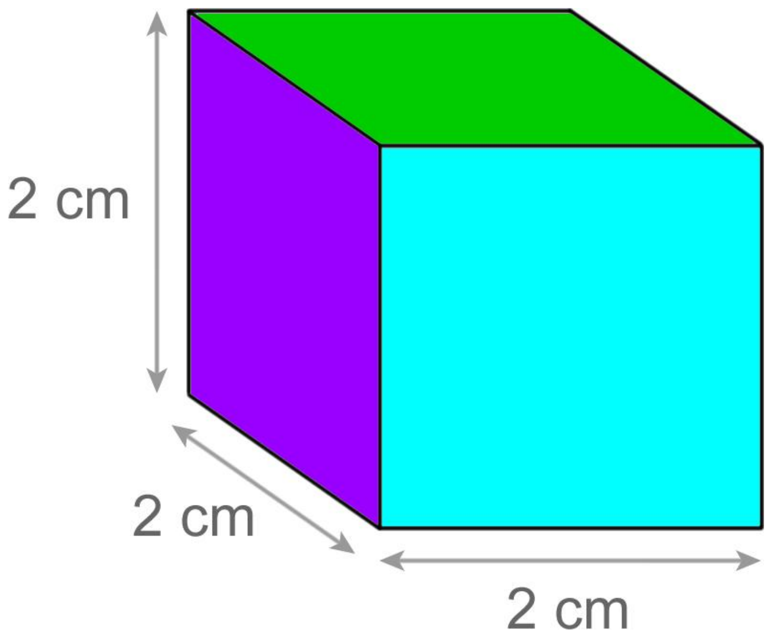


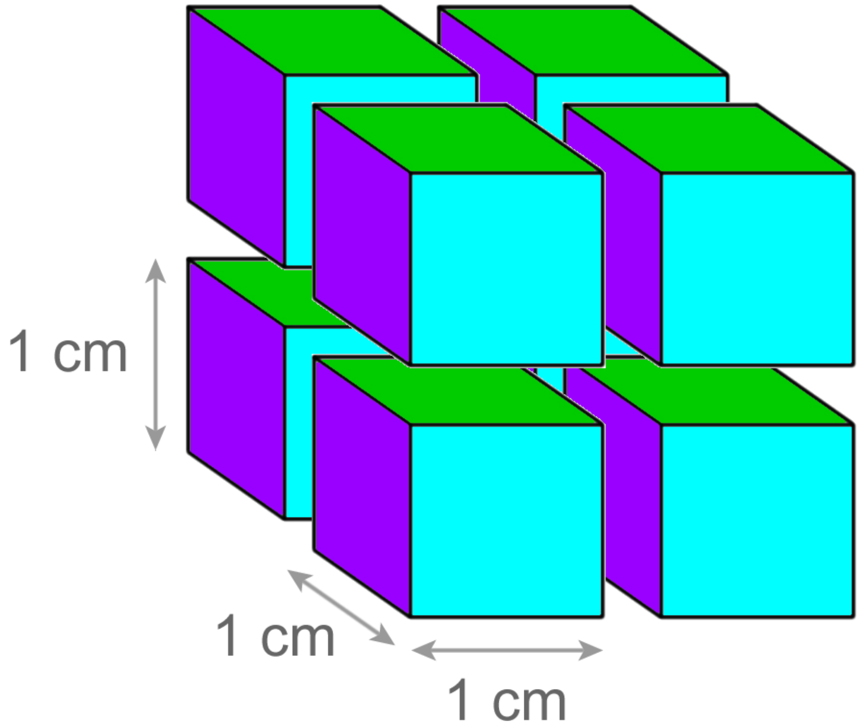
- Surface area of cube

  $= 6 \times (2 \times 2)$

  $= 24 \text{ cm}^2$

Rate of Reaction





- Surface area of cube

  $= 6 \times (2 \times 2)$

  $= 24\ cm^2$

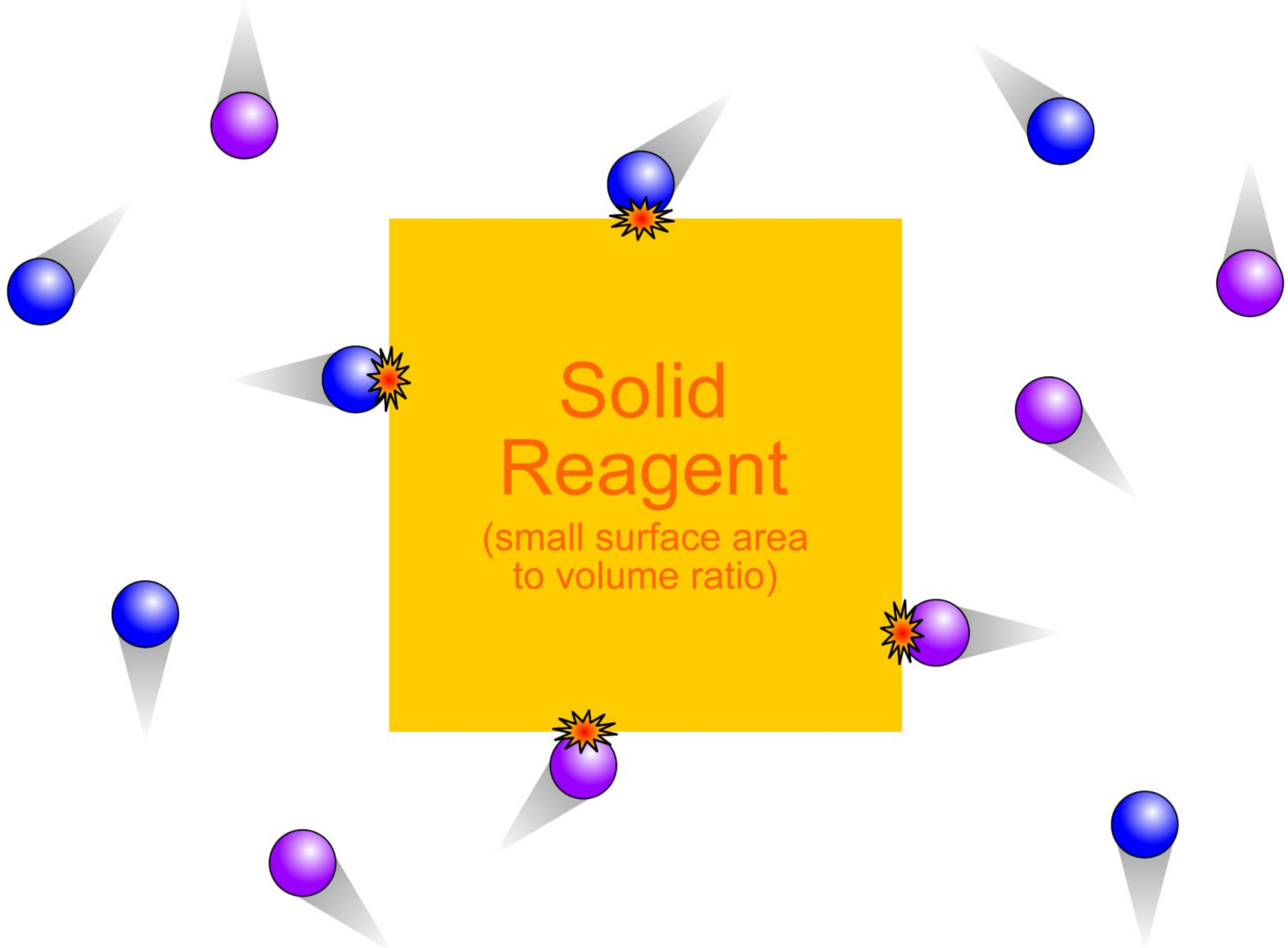
- Surface area of cubes

  $= 8 \times (6 \times (1 \times 1))$

  $= 48\ cm^2$

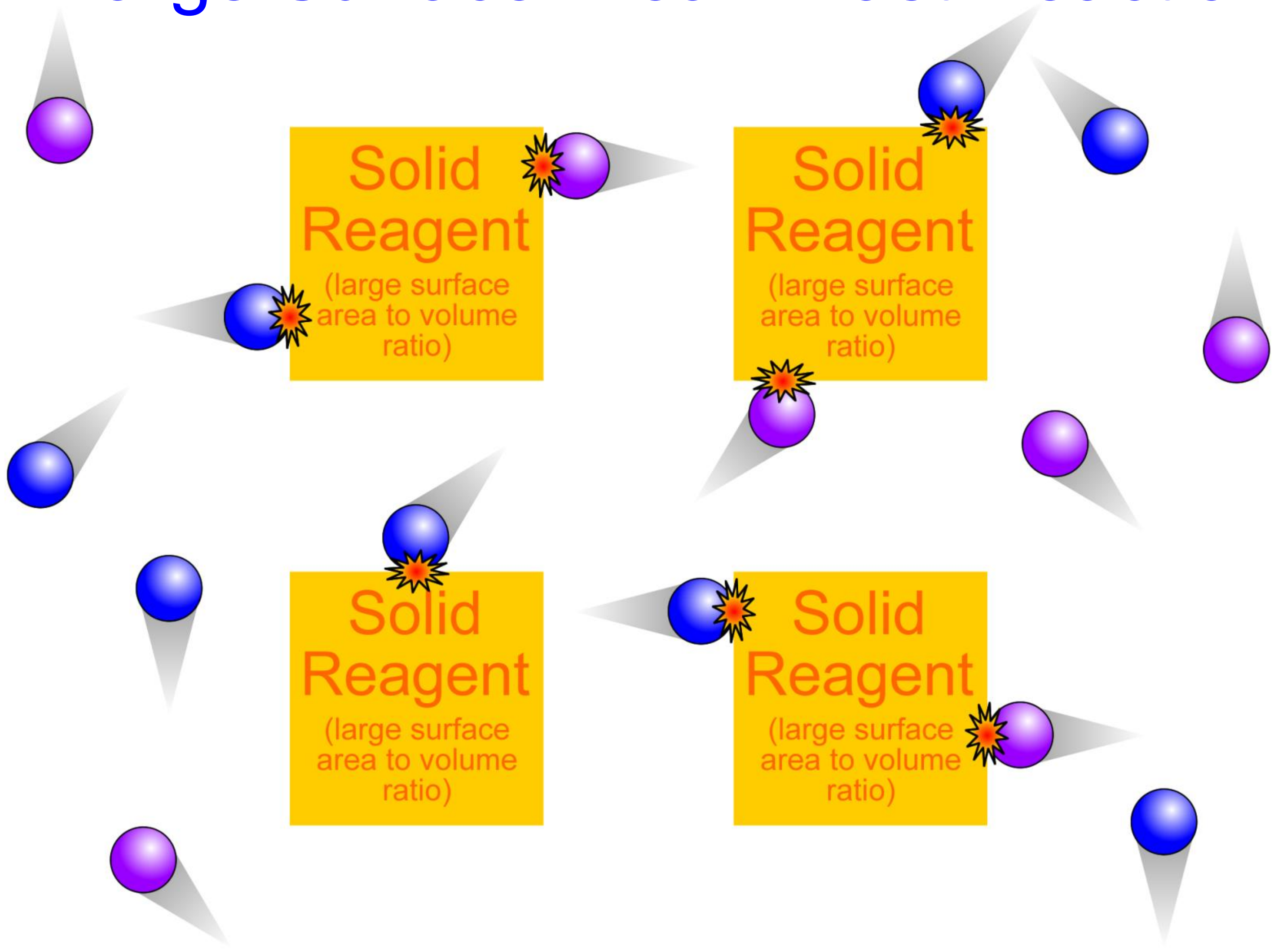
Rate of Reaction

- Small Surface Area – Slow Reaction



Rate of Reaction

- Large Surface Area – Fast Reaction



Rate of Reaction

- Imagine a long queue at the school canteen.

Stall
#1

Rate of Reaction

- Imagine a long queue at the school canteen.
- Opening more stalls will *increase the rate* at which students are served their food.

Stall #1

Stall #2

Rate of Reaction

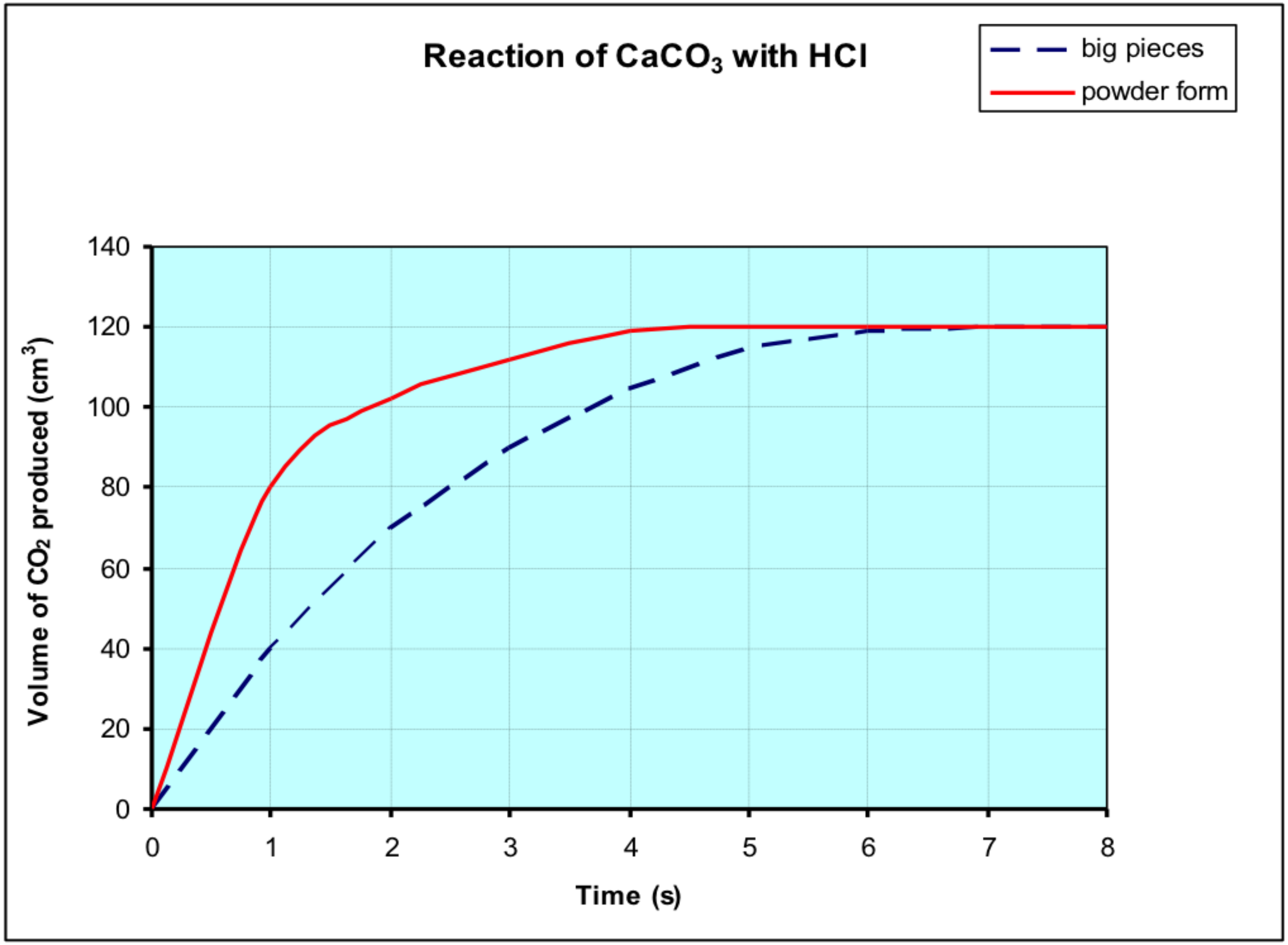
- Imagine a long queue at the school canteen.
- Opening more stalls will *increase the rate* at which students are served their food.

Stall #1

Stall #2

Stall #3

Rate of Reaction



Rate of Reaction

5) Catalyst (or enzyme).

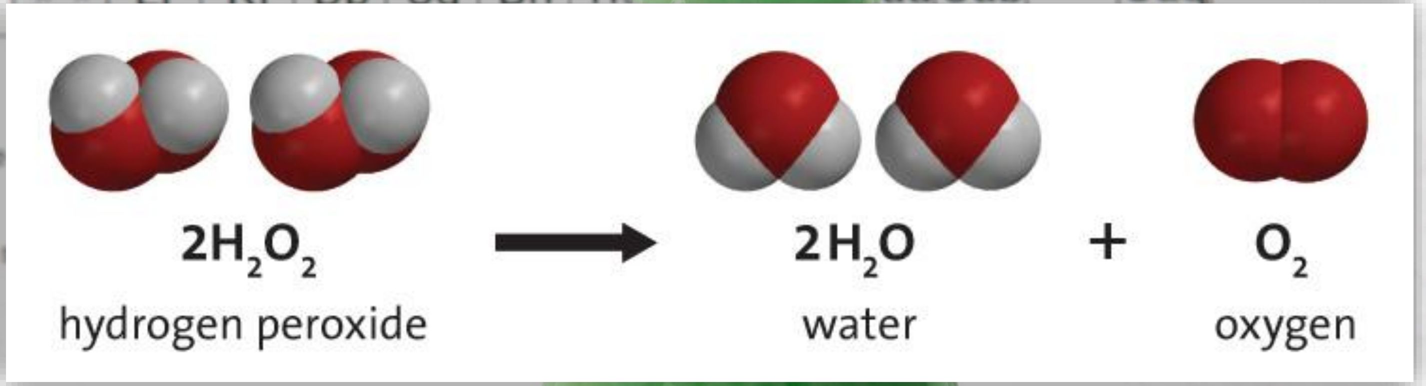
Addition of a catalyst will increase the rate of reaction.

- Catalysts (or enzymes) increase the speed of a reaction by providing an *alternative pathway* for the reaction to proceed. The alternative pathway of the catalysed reaction has a *lower activation energy* than the reaction without the catalyst.

- By lowering the activation energy, more collisions between particles will have an energy equal to or greater than the reaction's activation energy.

- Increasing the *frequency of effective collisions* between the particles increases the rate of the reaction.

Rate of Reaction



- Iodide ions (I^-) catalyse the decomposition of hydrogen peroxide into water and oxygen.

Rate of Reaction

"Elephant Toothpaste"

Catalytic Decomposition of Hydrogen Peroxide by Iodide Ions

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

- Duration: 16 sec.

Rate of Reaction

"Elephant Toothpaste"

Catalytic Decomposition of Hydrogen Peroxide by Iodide Ions

- Iodide ions from aqueous potassium iodide catalyse the decomposition of hydrogen peroxide in two steps:

Step 1: $H_2O_2(aq) + I^-(aq) \rightarrow H_2O(l) + IO^-(aq)$

Step 2: $H_2O_2(aq) + IO^-(aq) \rightarrow H_2O(l) + O_2(g) + I^-(aq)$

- Note that the iodide ions participate in the reaction (iodide ions are one of the reactants in Step 1) but are regenerated at the end of the reaction (iodide ions are one of the products in Step 2) allowing them to catalyse the reaction over again without having to be continuously added.

Rate of Reaction

What is a *catalyst*? How does it affect the rate of a chemical reaction?

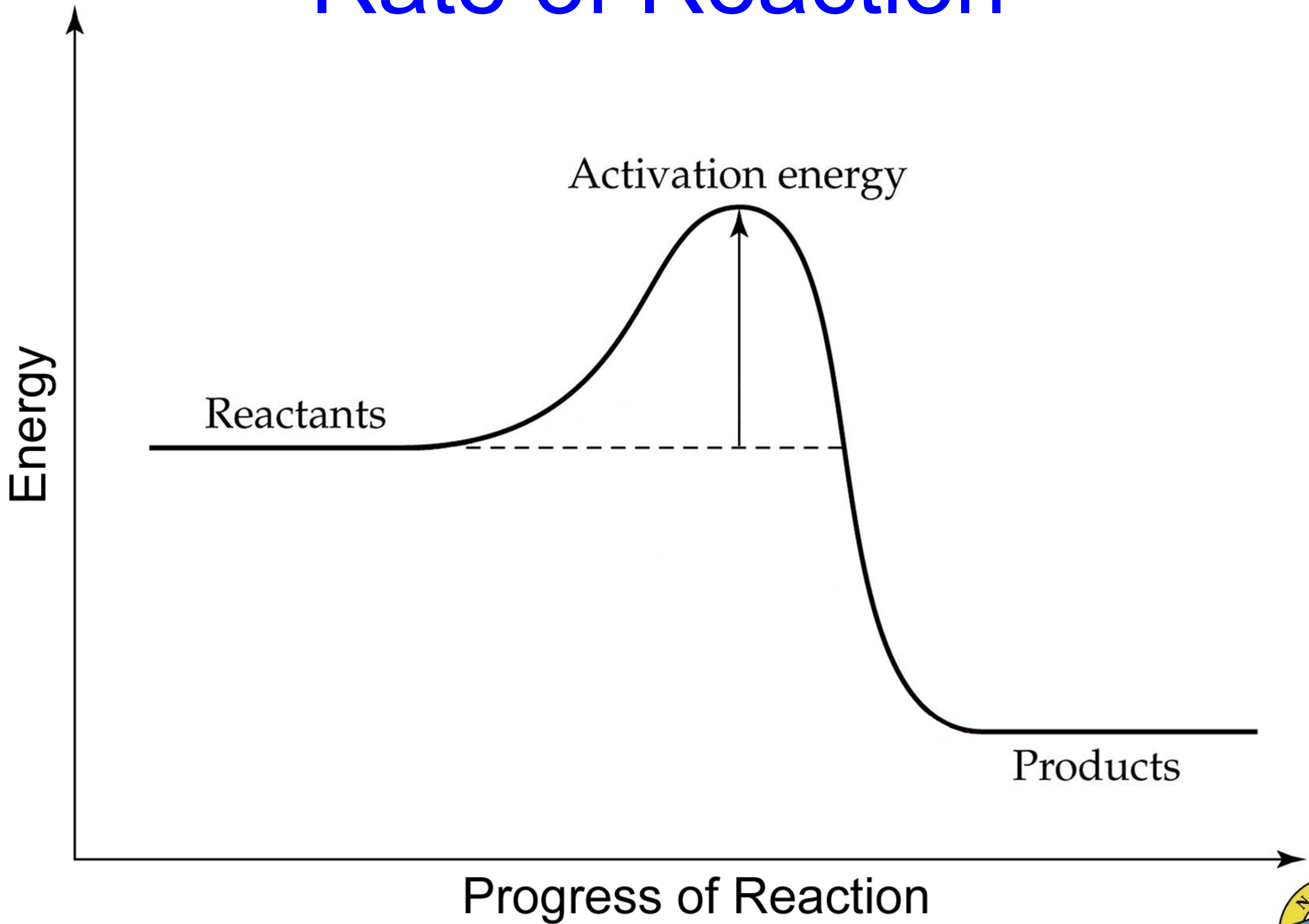
Rate of Reaction

Catalyst

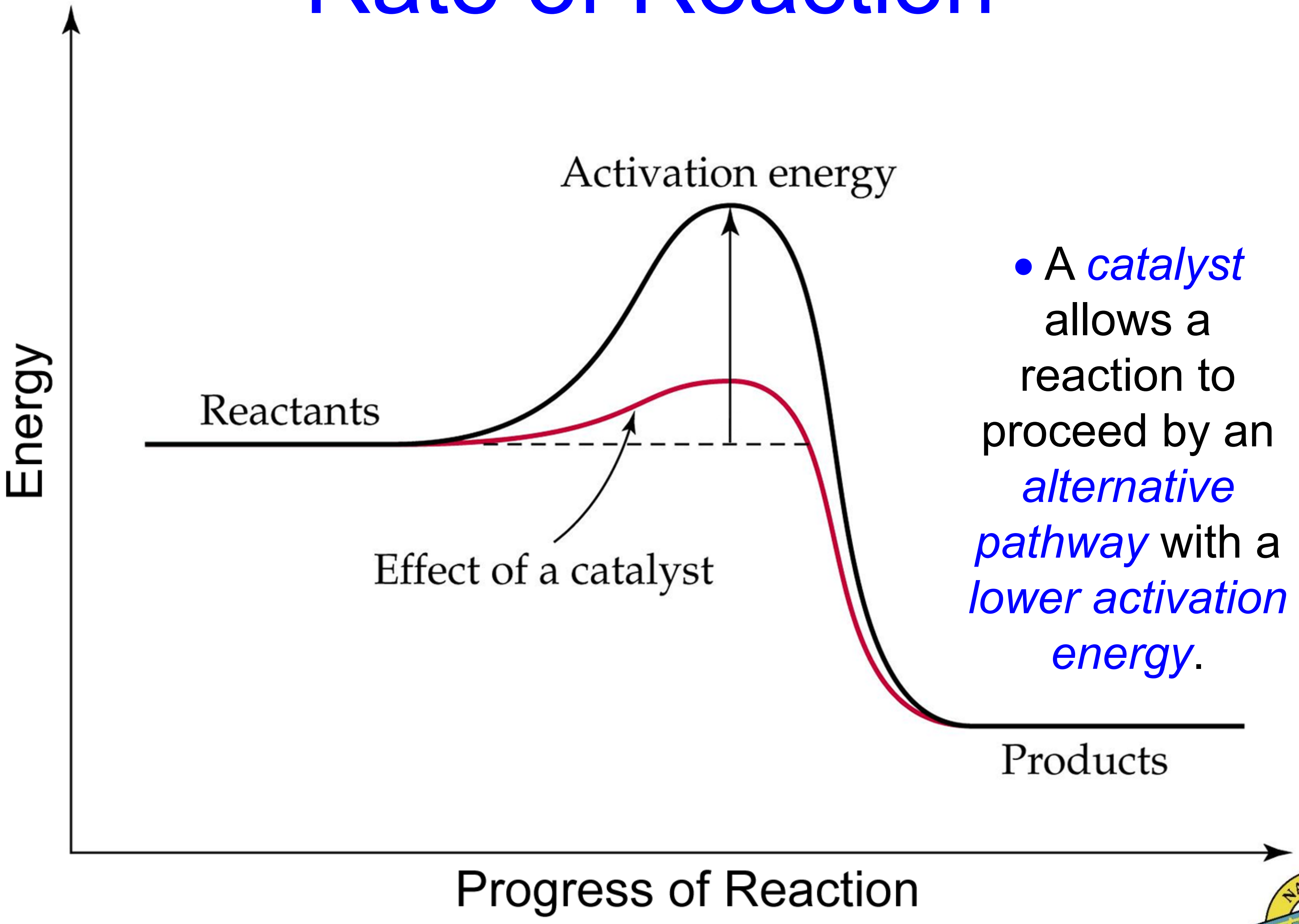
Rate of Reaction

- A catalyst is a chemical that *increases the rate* of a chemical reaction.
- Catalysts increase the speed of a reaction by providing an *alternative pathway* for the reaction to proceed. The alternative pathway of the catalysed reaction has a *lower activation energy* than the reaction without the catalyst.
- *More particles* will have energy *equal to or greater than the reaction's activation energy* and will therefore *react* when they *collide*. This *increases the frequency of effective collisions* between the particles.
- The catalyst remains *chemically unchanged* at the end of the chemical reaction.
- Only a *small amount* of catalyst is required and it can be recovered at the end of the experiment.

Rate of Reaction

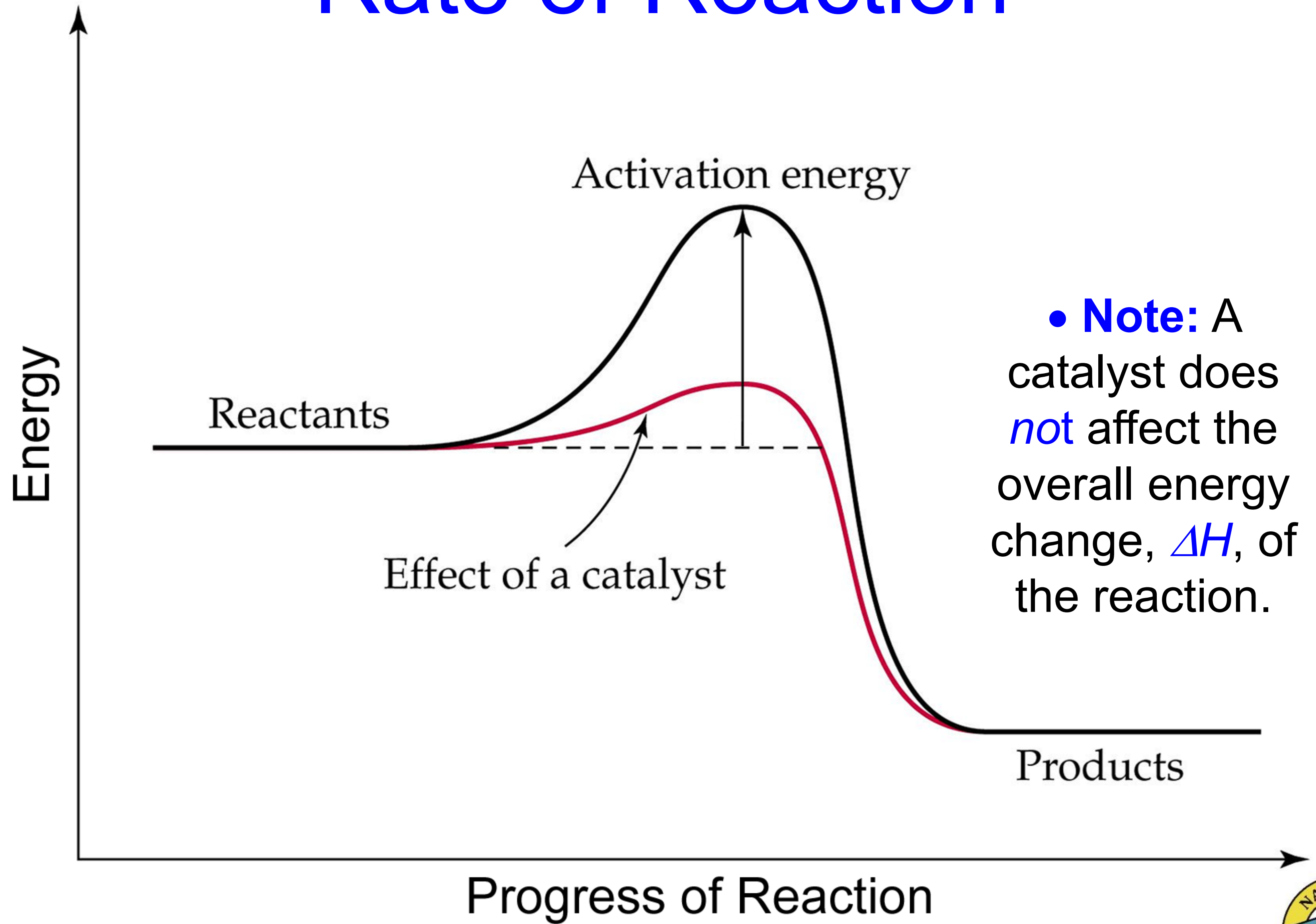


Rate of Reaction



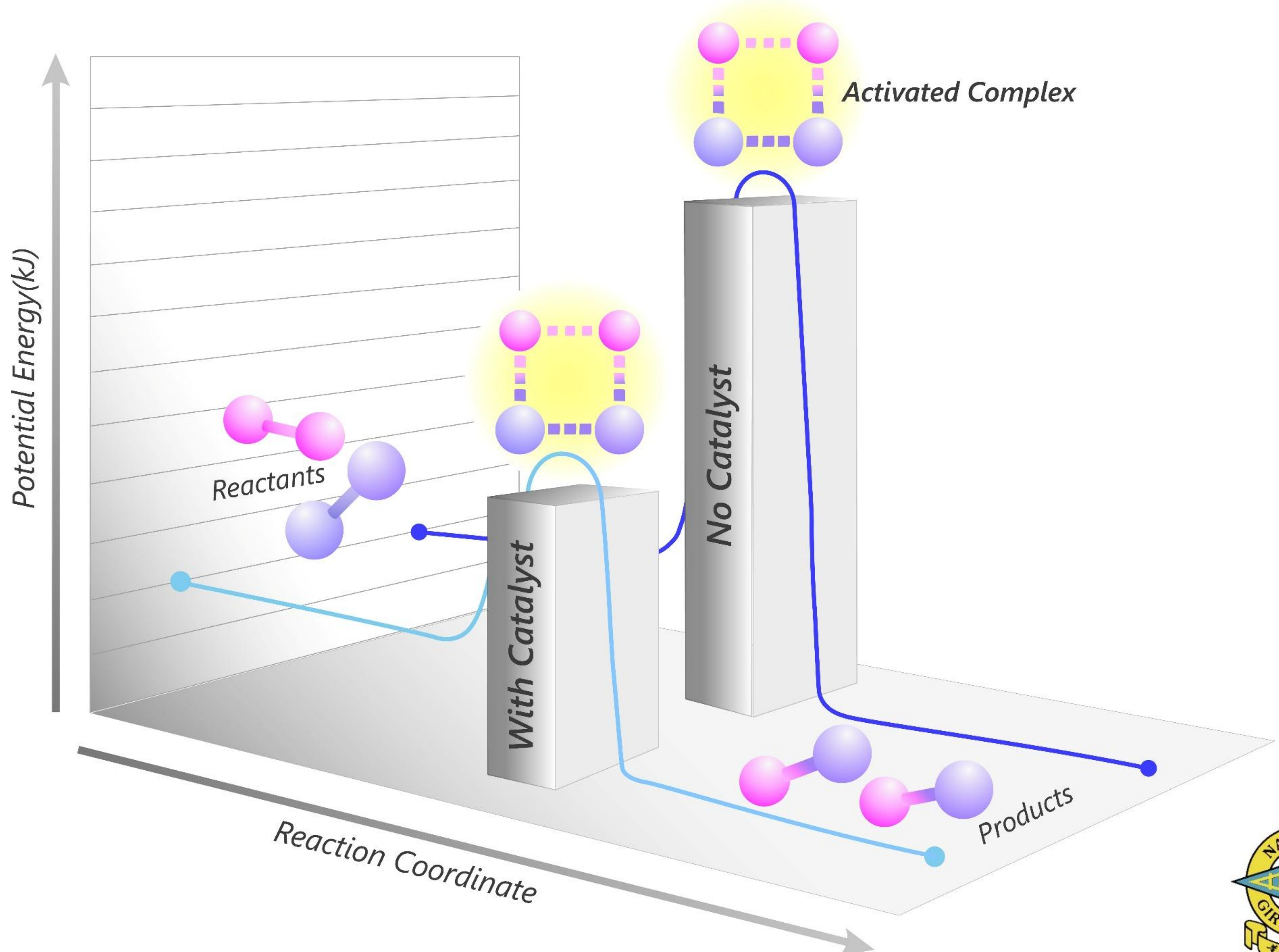
- A *catalyst* allows a reaction to proceed by an *alternative pathway* with a *lower activation energy*.

Rate of Reaction

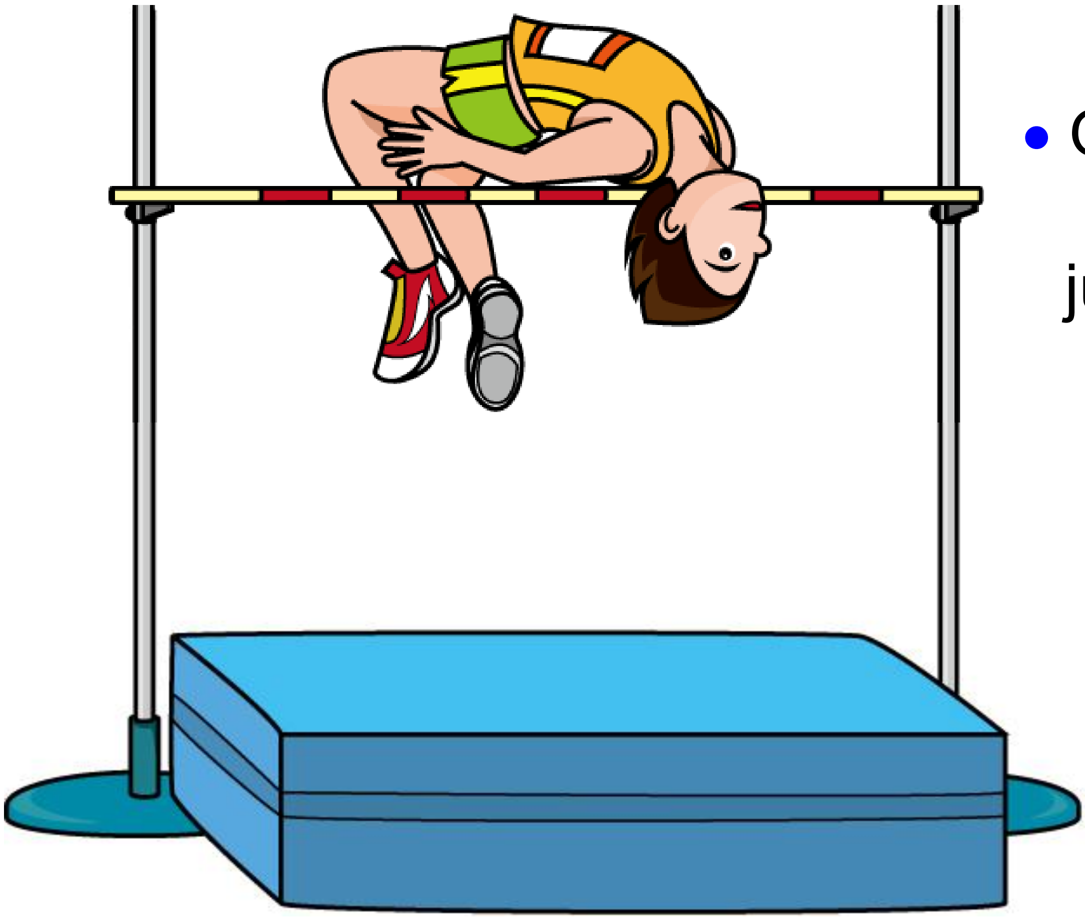


- **Note:** A catalyst does *not* affect the overall energy change, *ΔH*, of the reaction.

Rate of Reaction

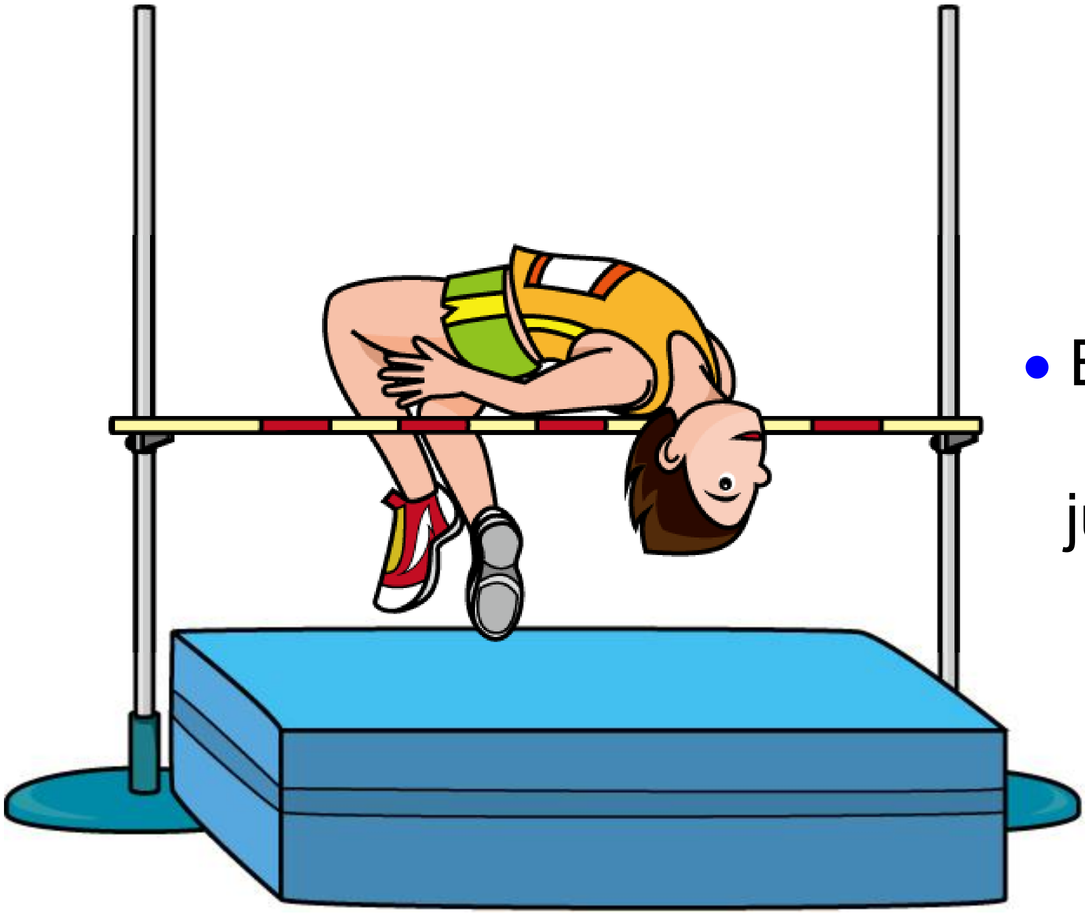


Rate of Reaction



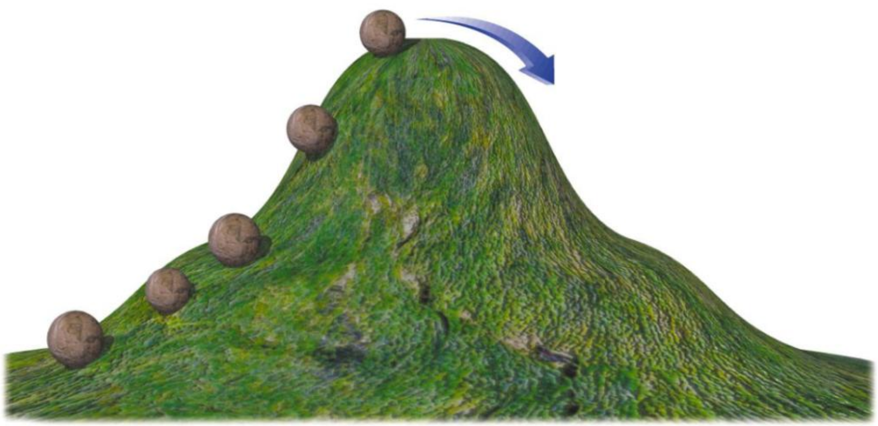
- Only 3 out of 30 students can jump this high.

Rate of Reaction



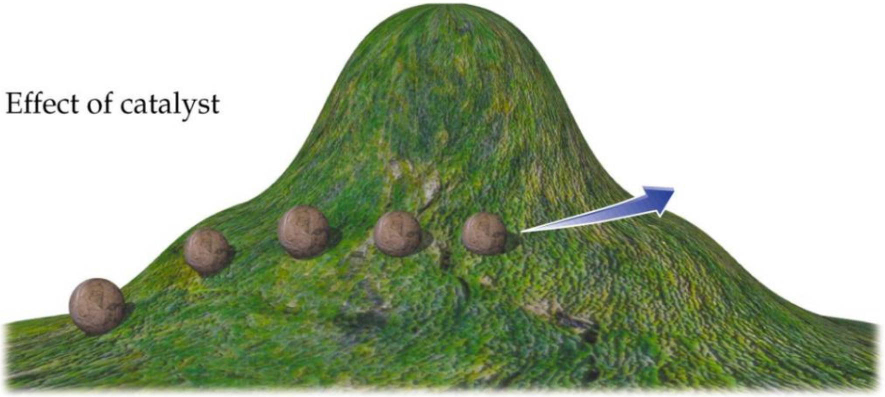
- But 24 out of 30 students can jump this high.

Rate of Reaction



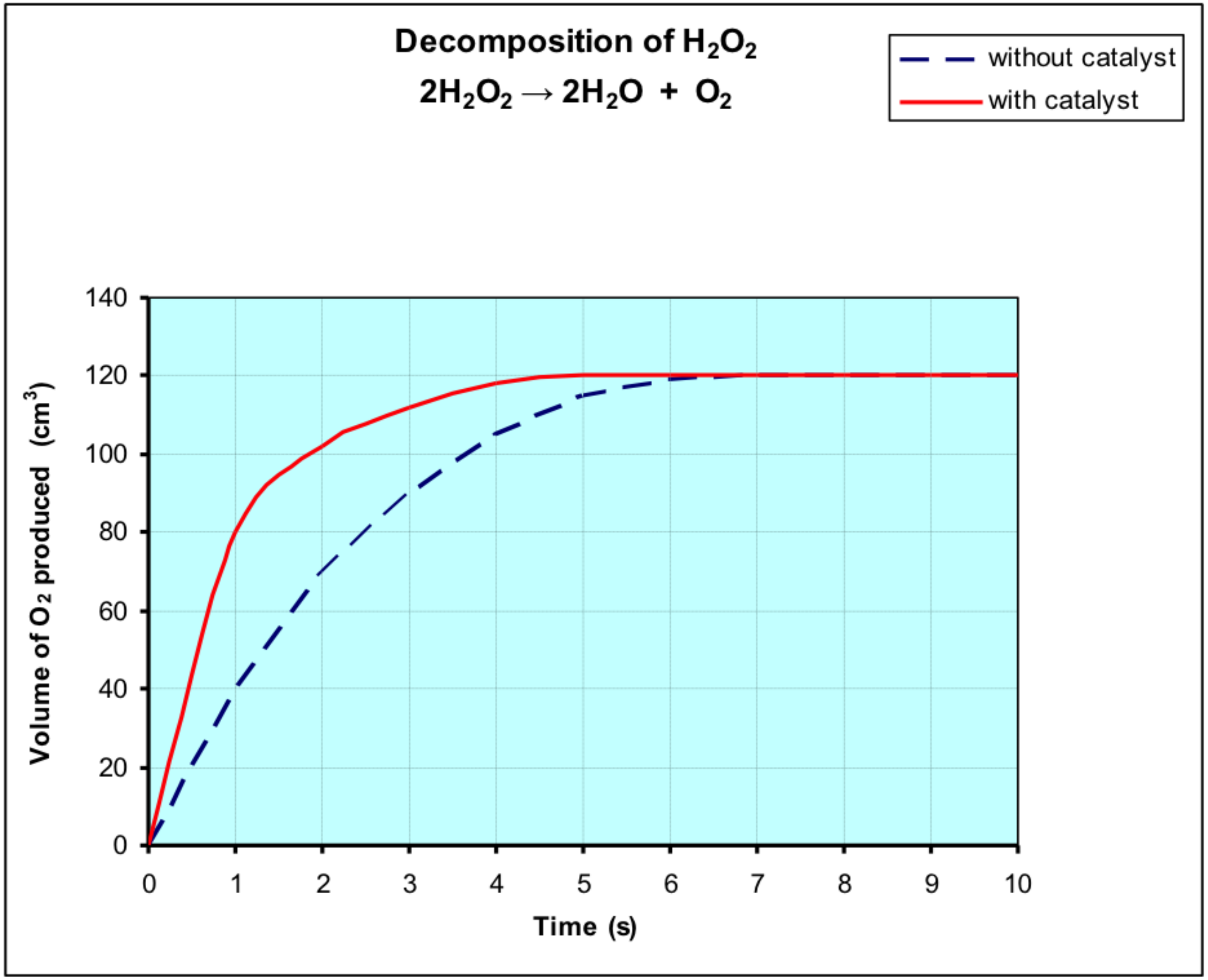
- Analogy for activation energy. *Less energy* is required to push the boulders *around* the side of the hill compared to the energy that is required to push the boulders *over* the hill.

Effect of catalyst

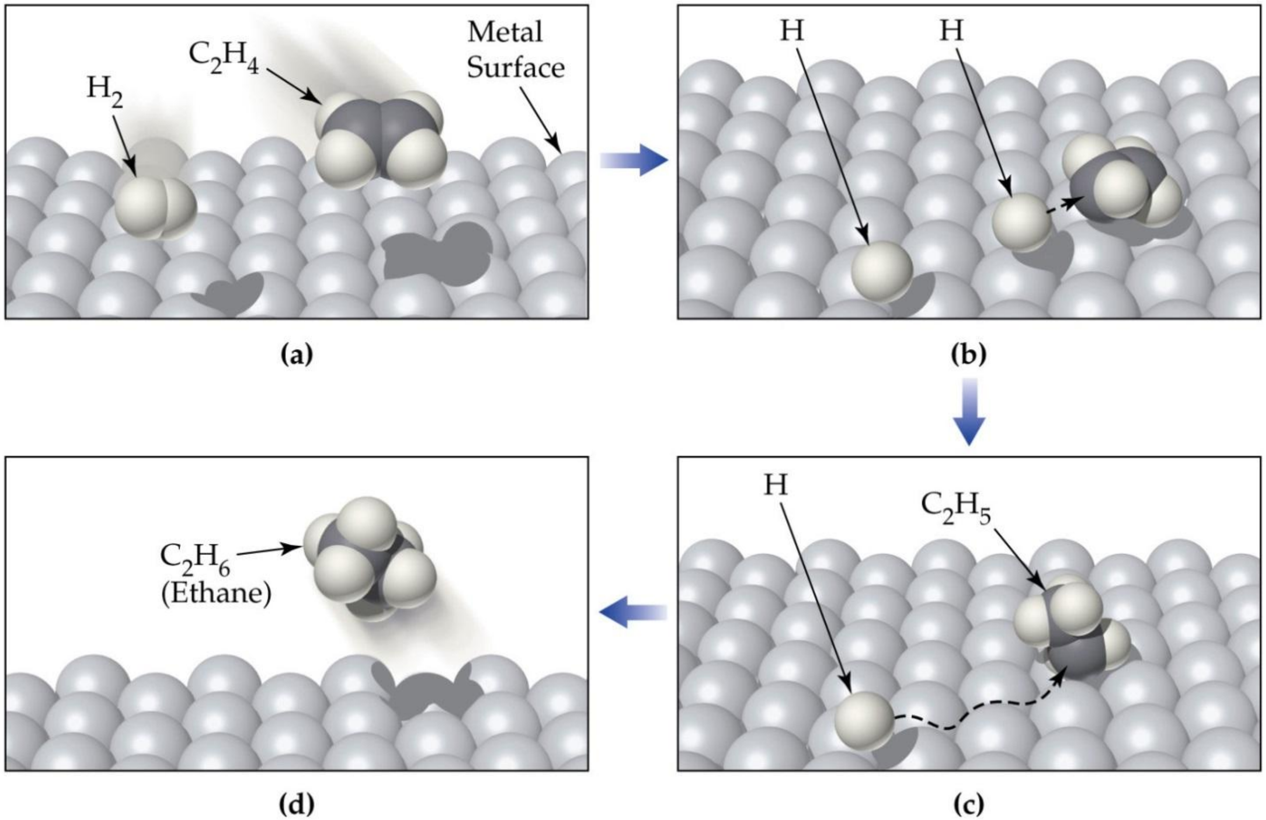


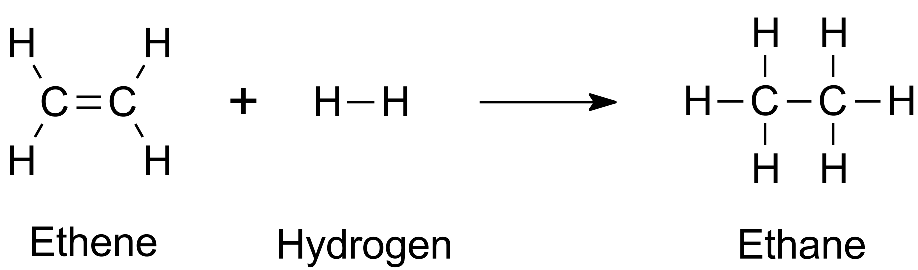
- Using the *same amount of energy*, more boulders can be pushed around the side of the hill compared to the number that can be pushed over the hill.

Rate of Reaction



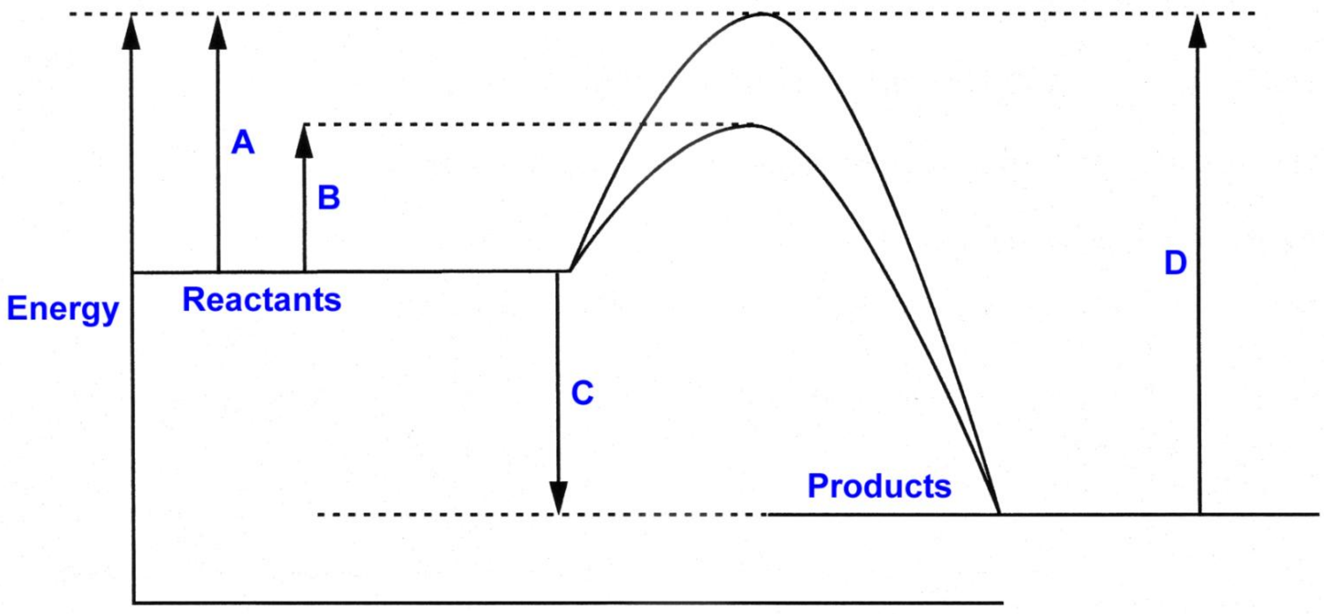
Rate of Reaction





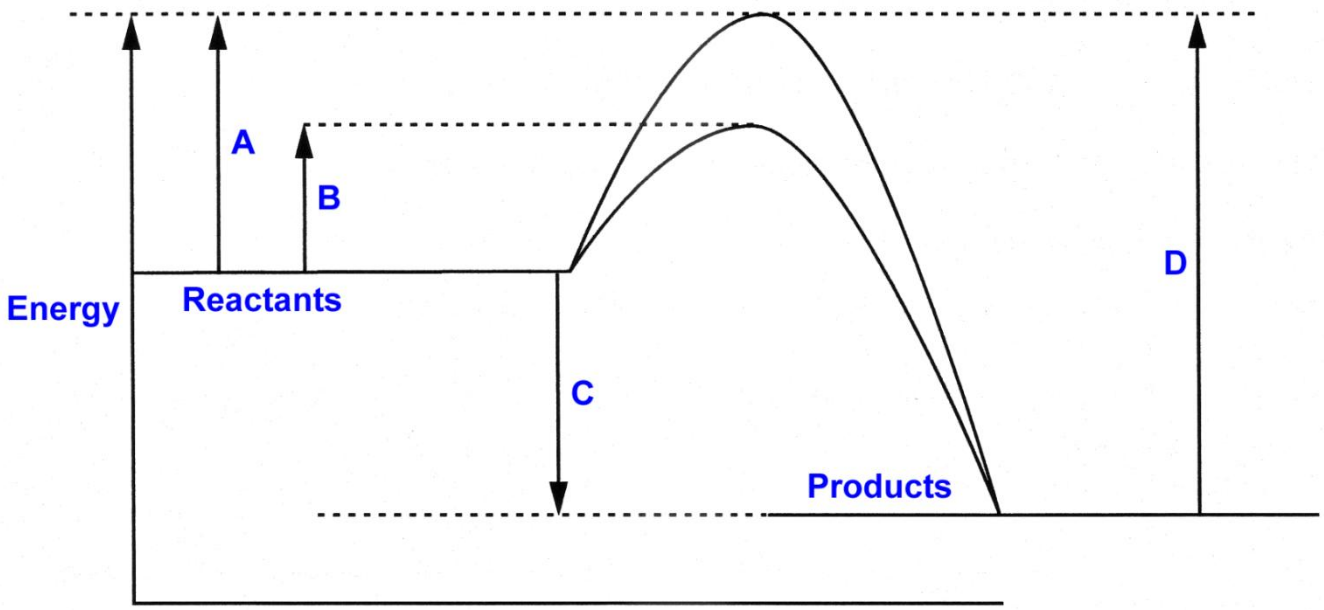
Rate of Reaction

- Study the energy change diagram shown below.
- Identify energy changes **A**, **B**, **C** and **D**.



Rate of Reaction

- Study the energy change diagram shown below.
- Identify energy changes **A**, **B**, **C** and **D**.



A Activation energy of the forward reaction.

B Activation energy of the forward reaction with a catalyst.

C Energy change (ΔH) of the forward reaction.

D Activation energy of the backward reaction.

Rate of Reaction

- Examples of catalysts include:

- **Iron – Fe:** Catalyst used in the industrial manufacture of ammonia (Haber process).

$$N_2(g) + 3H_2(g) \overset{Fe}{\rightleftharpoons} 2NH_3(g)$$

- **Nickel – Ni:** Catalyst used in the catalytic hydrogenation of plant oils during the manufacture of margarine.

$$C_2H_4(g) + H_2(g) \xrightarrow{Ni} C_2H_6(g)$$

- **Vanadium(V) Oxide – V_2O_5:** Catalyst used in the industrial manufacture of sulfuric acid (Contact process).

$$2SO_2(g) + O_2(g) \xrightarrow{V_2O_5} 2SO_3(g)$$

Rate of Reaction

- *Enzymes* are biological catalysts.

- Enzymes increase the rate of the reactions that take place inside living organisms.

- One example of an enzyme is the catalyst in *yeast* which converts glucose into ethanol and carbon dioxide during the process of fermentation.

$$C_6H_{12}O_6(aq) \rightarrow 2CH_3CH_2OH(aq) + 2CO_2(g)$$

- Unlike inorganic catalysts, enzymes are very sensitive to variables such as *temperature* and *pH*. The efficiency with which an enzyme catalyses a reaction falls rapidly at temperatures above 40 °C (the enzyme has been *denatured*). In addition, each enzyme only functions over a narrow range of pH, unique to that particular enzyme.

Rate of Reaction

- *Enzymes* are biological catalysts.

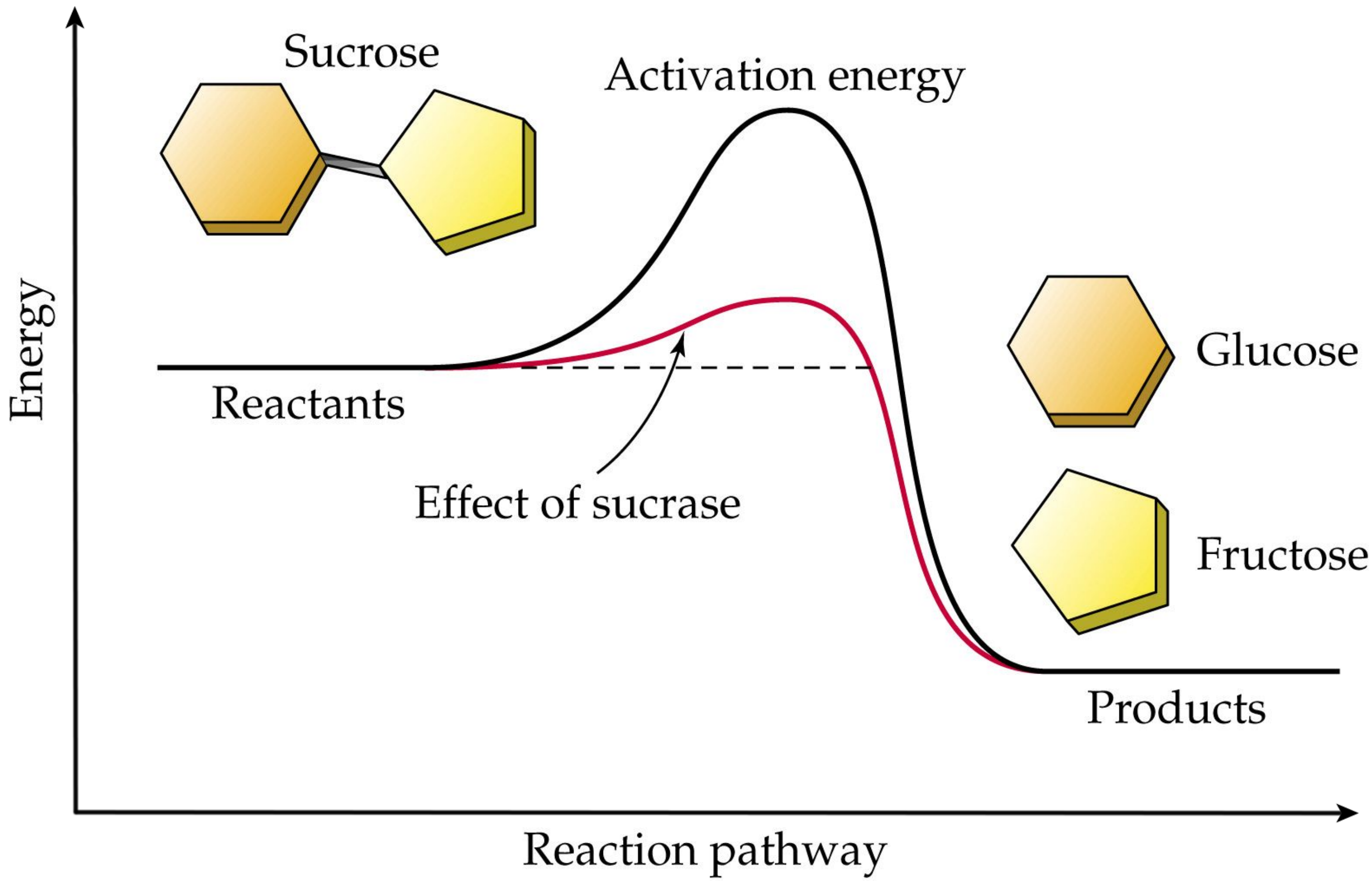
- Another example of an enzyme is *sucrase* which breaks down sucrose (sugar) into glucose and fructose.

$$\text{sucrose} + \text{water} \xrightarrow{\text{sucrase}} \text{glucose} + \text{fructose}$$

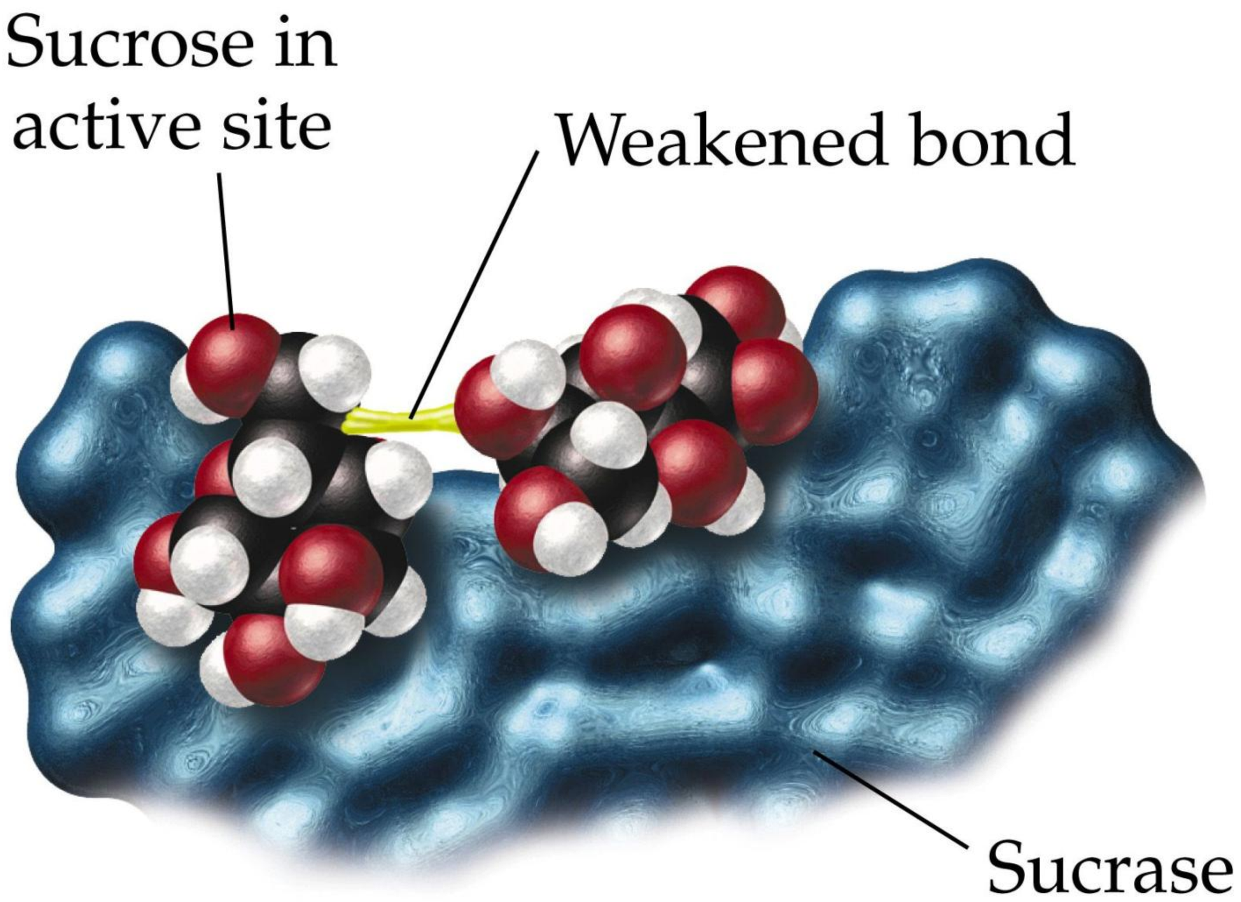
$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \xrightarrow{\text{sucrase}} C_6H_{12}O_6(aq) + C_6H_{12}O_6(g)$$

Rate of Reaction

Sucrose ⟶ Glucose and fructose



Rate of Reaction



Rate of Reaction

How does the *type of acid* used affect the rate of a chemical reaction?

Rate of Reaction

Chemical Properties of Reactants

- In addition to variables such as catalyst, concentration, pressure, surface area and temperature, the *chemical properties of the reactants* can also affect the rate of a chemical reaction.

Rate of Reaction

Chemical Properties of Reactants

- Assuming that all other variables remain constant, *i.e.* concentration of acid, surface area of magnesium and temperature, which one of the following reactions will be the *slowest*, and which one will be the *fastest* ?

$$Mg(s) + 2CH_3COOH(aq) \rightarrow Mg(CH_3COO)_2(aq) + H_2(g)$$

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

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

Rate of Reaction

Chemical Properties of Reactants

- All three reactions share the same ionic equation:

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

- The *concentration of $H^+(aq)$* will affect the rate of the reaction.

- The greater the concentration of $H^+(aq)$, the greater the number of $H^+(aq)$ per unit volume.

- This will *increase the frequency of effective collisions* between $H^+(aq)$ and $Mg(s)$ therefore increasing the rate of the reaction.

Rate of Reaction

Chemical Properties of Reactants

- Even though the concentrations of the three acids are the same, because the three acids vary slightly in their *chemical properties*, the concentration of $H^+(aq)$ in each acid is actually *different*.

Rate of Reaction

Chemical Properties of Reactants

- Ethanoic acid is a *weak monobasic* acid.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

- Ethanoic acid will only *partially ionize* to produce $H^+(aq)$ when it is dissolved in water.

- One molecule of ethanoic acid can produce a maximum number of *one* $H^+(aq)$.

- Out of the three acids, ethanoic acid has the *lowest* $H^+(aq)$ concentration and therefore has the *slowest rate of reaction* with Mg(s).

Rate of Reaction

Chemical Properties of Reactants

- Sulfuric acid is a *strong dibasic* acid.

$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

- Sulfuric acid will *fully ionize* to produce $H^+(aq)$ when it is dissolved in water.

- One molecule of sulfuric acid can produce a maximum number of *two $H^+(aq)$*.

- Out of the three acids, sulfuric acid has the *highest $H^+(aq)$* concentration and therefore has the *fastest rate of reaction* with Mg(s).

Rate of Reaction

Chemical Properties of Reactants

- Hydrochloric acid is a *strong monobasic* acid.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

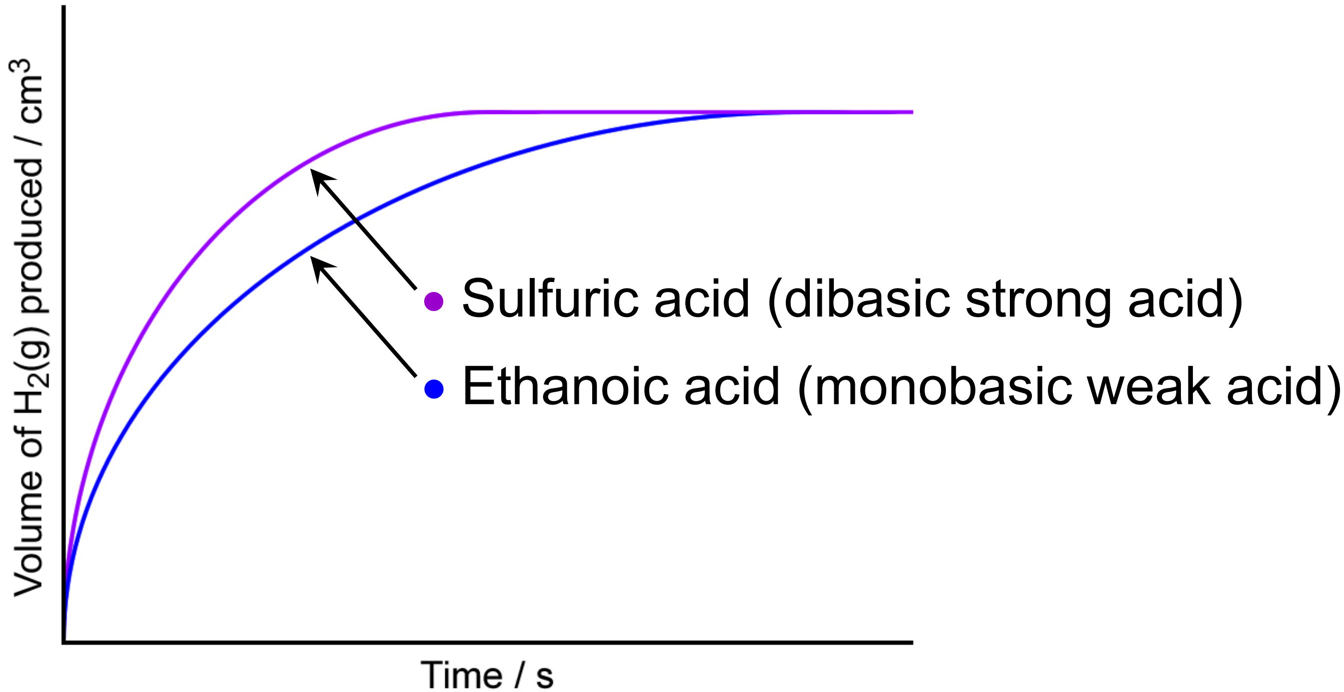
- Hydrochloric acid will *fully ionize* to produce $H^+(aq)$ when it is dissolved in water.

- One molecule of hydrochloric acid can produce a maximum number of *one $H^+(aq)$*.

- Out of the three acids, hydrochloric acid has a *faster* rate of reaction compared to ethanoic acid, but a *slower* rate of reaction compared to sulfuric acid.

Rate of Reaction

Chemical Properties of Reactants



- For the same volume and concentration of each acid reacting with the same mass of magnesium, the reaction between sulfuric acid and magnesium will be *faster* (producing a larger volume of hydrogen gas per unit time) than the reaction between ethanoic acid and magnesium.

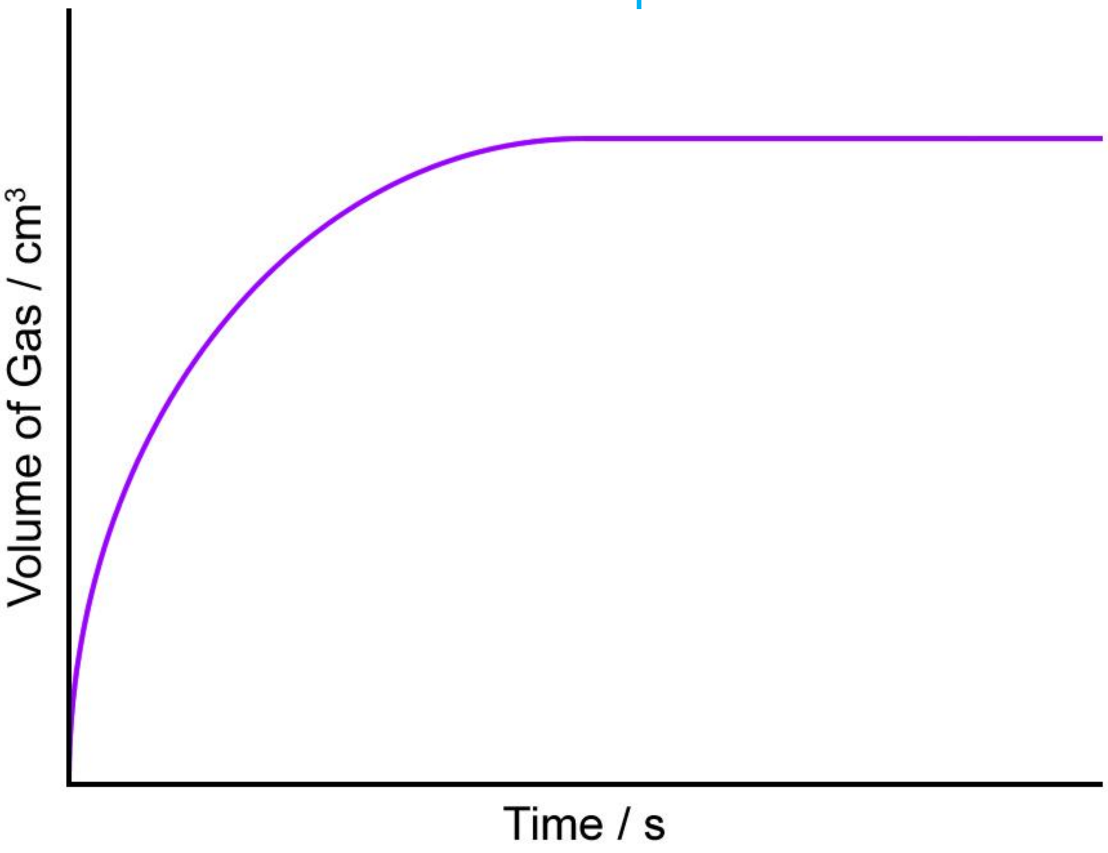
Rate of Reaction

What information can I extract from a *rate of reaction graph*?

Rate of Reaction

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

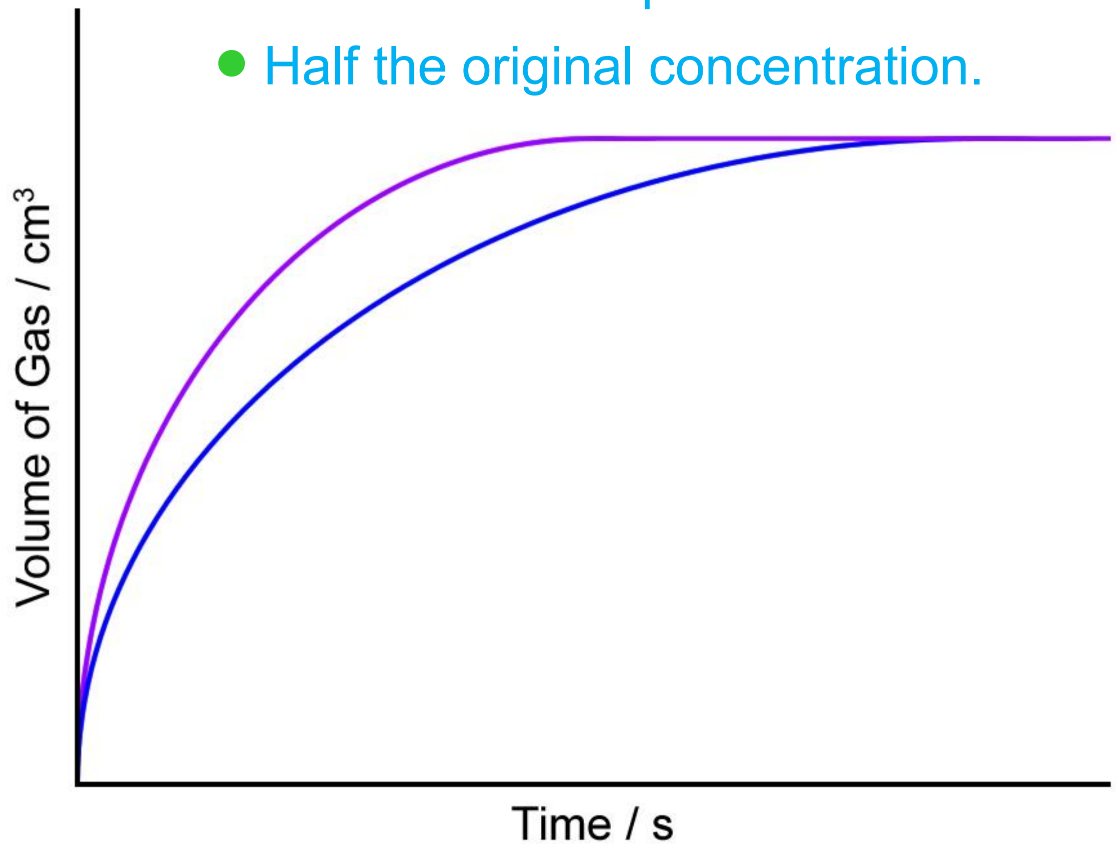
- Original graph.
- Lower temperature.



Rate of Reaction

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

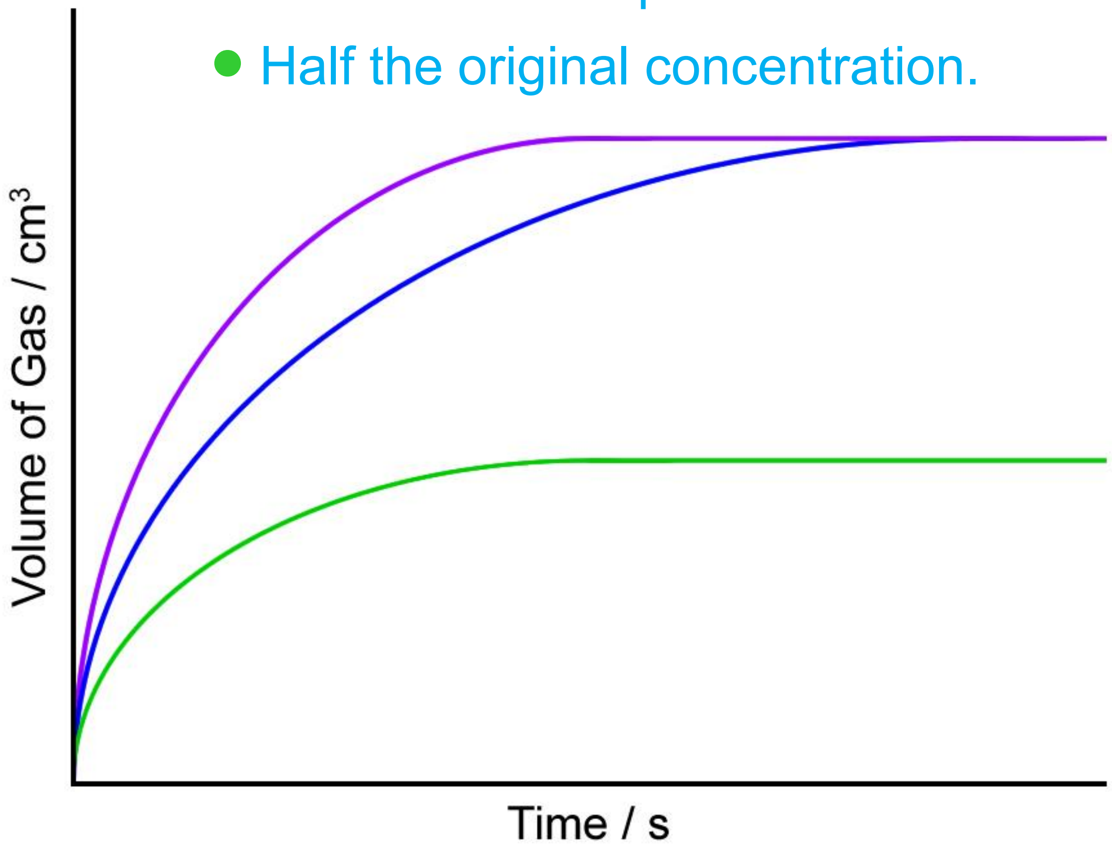
- Original graph.
- Lower temperature.
- Half the original concentration.



Rate of Reaction

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

- Original graph.
- Lower temperature.
- Half the original concentration.



Rate of Reaction

- When sketching a rate of reaction graph, consider the following points. These should help guide you towards drawing a graph with the correct shape.

a) What is the Initial Gradient?

If a variable* has been changed that will *increase* the rate of reaction (*e.g.* increase in temperature) then the initial gradient should be *steeper*, *i.e.* the line for the new graph should be draw to the *left* of the original graph.

If a variable* has been changed that will *decrease* the rate of reaction (*e.g.* decrease in concentration) then the initial gradient should be *more shallow*, *i.e.* the line for the new graph should be draw to the *right* of the original graph.

*Assume that only one variable is changed at-a-time.

Rate of Reaction

- When sketching a rate of reaction graph, consider the following points. These should help guide you towards drawing a graph with the correct shape.

b) When does the Reaction Stop?

The point at which a reaction is complete is indicated by the point at which the variable being measured against time (*independent variable*) becomes *constant*.

If a variable* has been changed that will *increase* the rate of reaction (*e.g.* addition of a catalyst) then the reaction will take a *shorter* period of time to complete / finish.

If a variable* has been changed that will *decrease* the rate of reaction (*e.g.* decrease in temperature) then the reaction will take a *longer* period of time to complete / finish.

*Assume that only one variable is changed at-a-time.

Rate of Reaction

- When sketching a rate of reaction graph, consider the following points. These should help guide you towards drawing a graph with the correct shape.

c) How Much Product is Formed?

For reactions in which a *solution* is the *limiting reagent* – increasing the *concentration** of the solution will not only increase the rate of reaction, it will also *increase the amount of product* formed (moles = **c** × v × 10^{-3}).

For reactions in which a *solution* is the *limiting reagent* – increasing the *volume** of the solution will *not* affect the rate of reaction, but it *will increase the amount of product* formed (moles = c × **v** × 10^{-3}).

*Assume that only one variable is changed at-a-time.

Rate of Reaction

- Changing the *temperature* (increase or decrease) of a chemical reaction *will* change the *rate* of the chemical reaction (correspondingly, increase or decrease), but *will not* affect the total *amount* of product formed.

- Adding a *catalyst* to a chemical reaction *will* increase the *rate* of the chemical reaction, but *will not* affect the total *amount* of product formed.

- Assuming that the mass of the solid reagent remains *constant* – increasing the *surface area* of a solid reagent *will* increase the *rate* of the chemical reaction, but *will not* affect the total *amount* of product formed.

Rate of Reaction

- Assuming that it is the *limiting reagent*, and assuming that its volume remains *constant* – changing the *concentration* of a *solution* (increase or decrease) *will* change the *rate* of the chemical reaction (correspondingly, increase or decrease) and *will also* change the total *amount of product* formed (correspondingly, increase or decrease).

  moles of chemical in solution = $c \times v \times 10^{-3}$

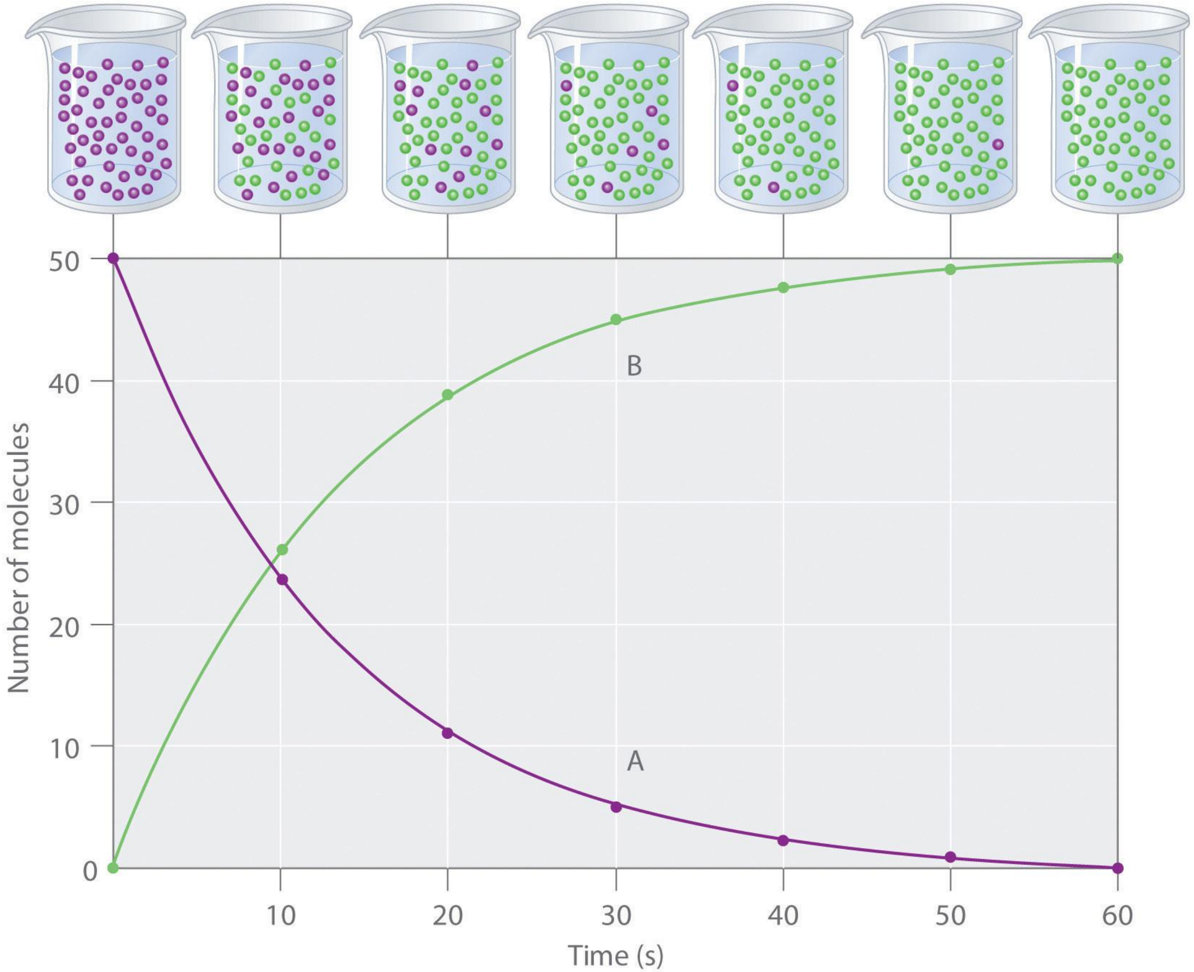
- Assuming that it is the *limiting reagent*, and assuming that its concentration remains *constant* – changing the *volume* of a *solution* (increase or decrease) *will not* change the *rate* of a chemical reaction but *will* change the total *amount of product* formed (correspondingly, increase or decrease).

  moles of chemical in solution = $c \times v \times 10^{-3}$

Rate of Reaction

Why do reactions *slow down* and eventually *stop*?

Rate of Reaction



- The *steeper* the gradient, the *faster* the reaction. The *more shallow* the gradient, the *slower* the reaction.

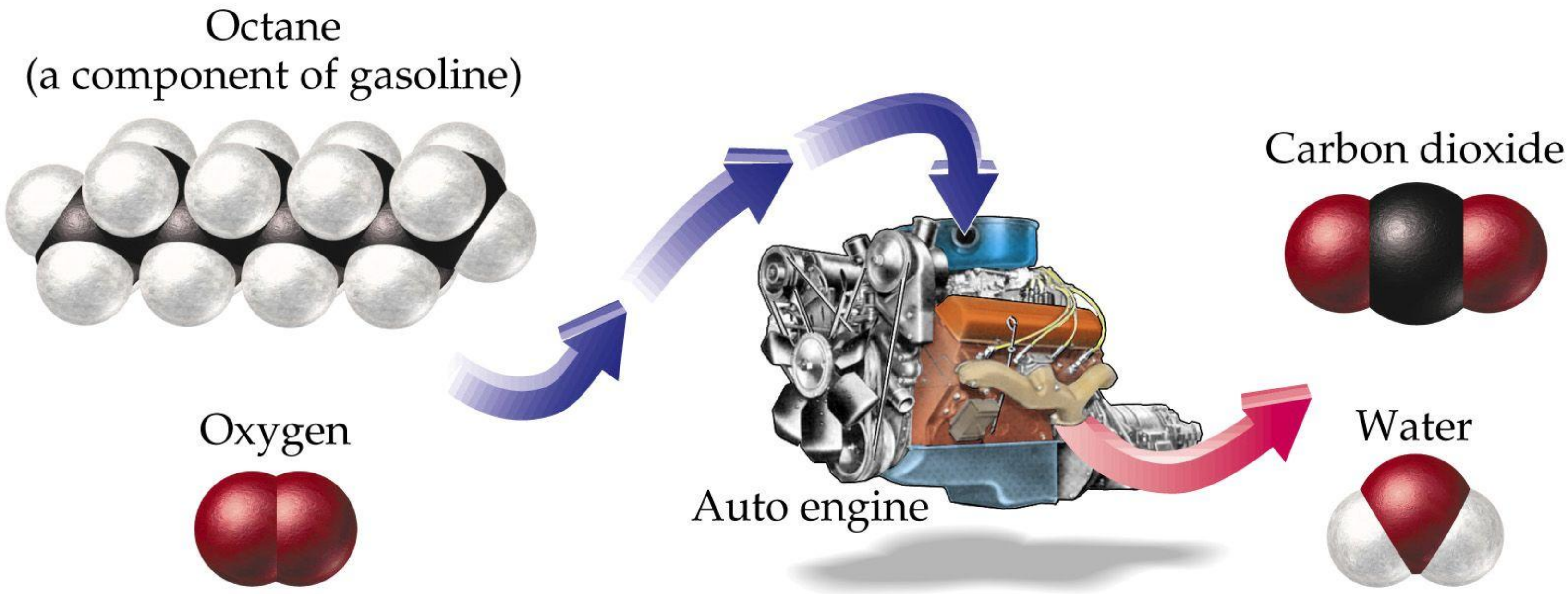
Rate of Reaction

- A reaction will slow down, and eventually stop, as the *amount* (*moles*) of the *limiting reagent* decreases to *zero*.
- This will decrease the frequency of effective collisions between the particles, decreasing the rate of the chemical reaction.

Rate of Reaction

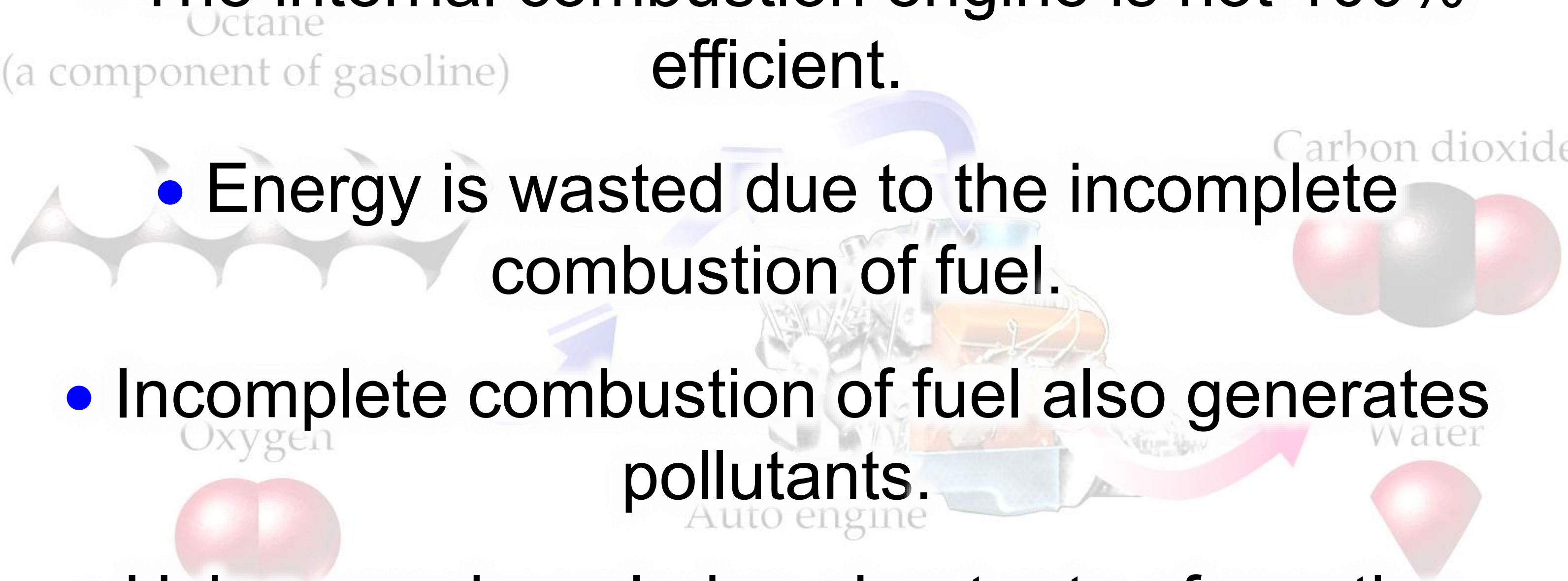
How can I *apply* my knowledge?

Rate of Reaction



Rate of Reaction

- $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
- The internal combustion engine is not 100% efficient.
- Energy is wasted due to the incomplete combustion of fuel.
- Incomplete combustion of fuel also generates pollutants.
- Using your knowledge about rate of reaction, propose ways of improving the combustion of fuel in the engine, thus producing more energy and fewer pollutants.



Rate of Reaction

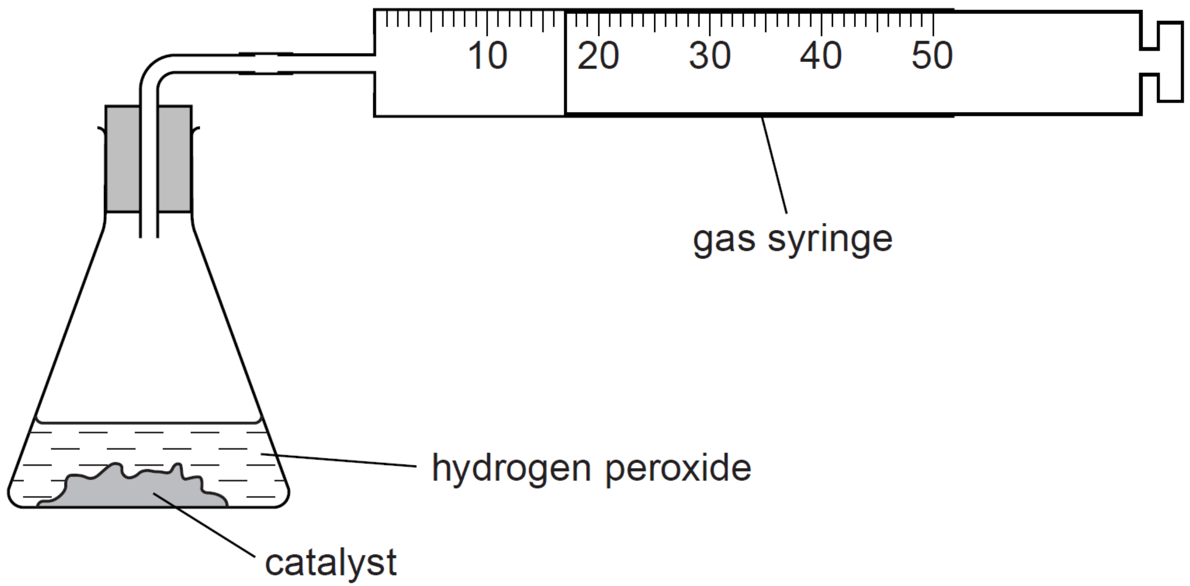
Could I please try a *question* on rate of reaction?

Rate of Reaction

Hydrogen peroxide, H_2O_2, decomposes slowly at room temperature to form water and oxygen:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

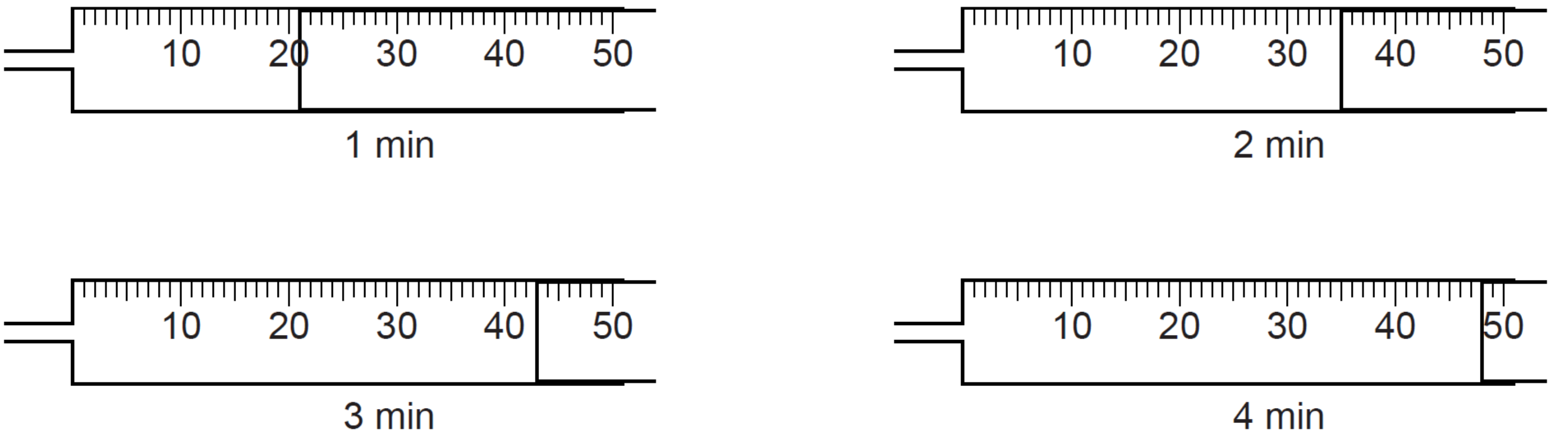
A student investigated how the rate of decomposition changed by using two catalysts, manganese(IV) oxide and copper. The volume of oxygen produced was measured at intervals using the apparatus shown below:



Rate of Reaction

The student carried out two experiments using the same volume of hydrogen peroxide but with the same mass of a different catalyst in each experiment.

a) The results for experiment **1** and some of the results for experiment **2** are shown. Use the diagrams to complete the results for experiment **2**.

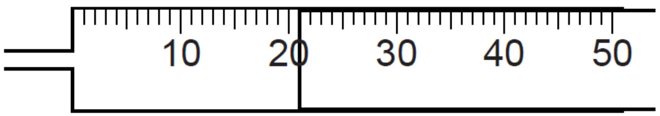


Time / min	1	2	3	4	5	6
Volume of oxygen collected in experiment **1** / cm^3	9	17	24	29	32	35
Volume of oxygen collected in experiment **2** / cm^3					50	50

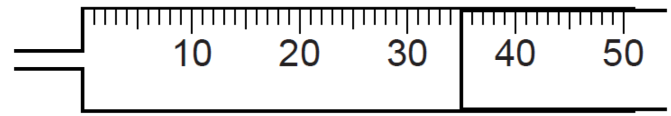
Rate of Reaction

The student carried out two experiments using the same volume of hydrogen peroxide but with the same mass of a different catalyst in each experiment.

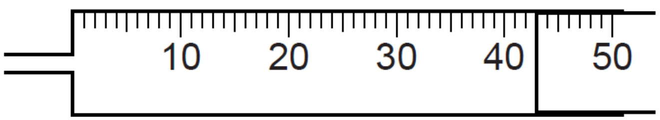
a) The results for experiment **1** and some of the results for experiment **2** are shown. Use the diagrams to complete the results for experiment **2**.



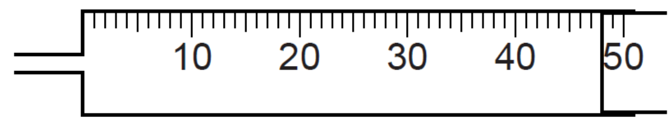
1 min



2 min



3 min

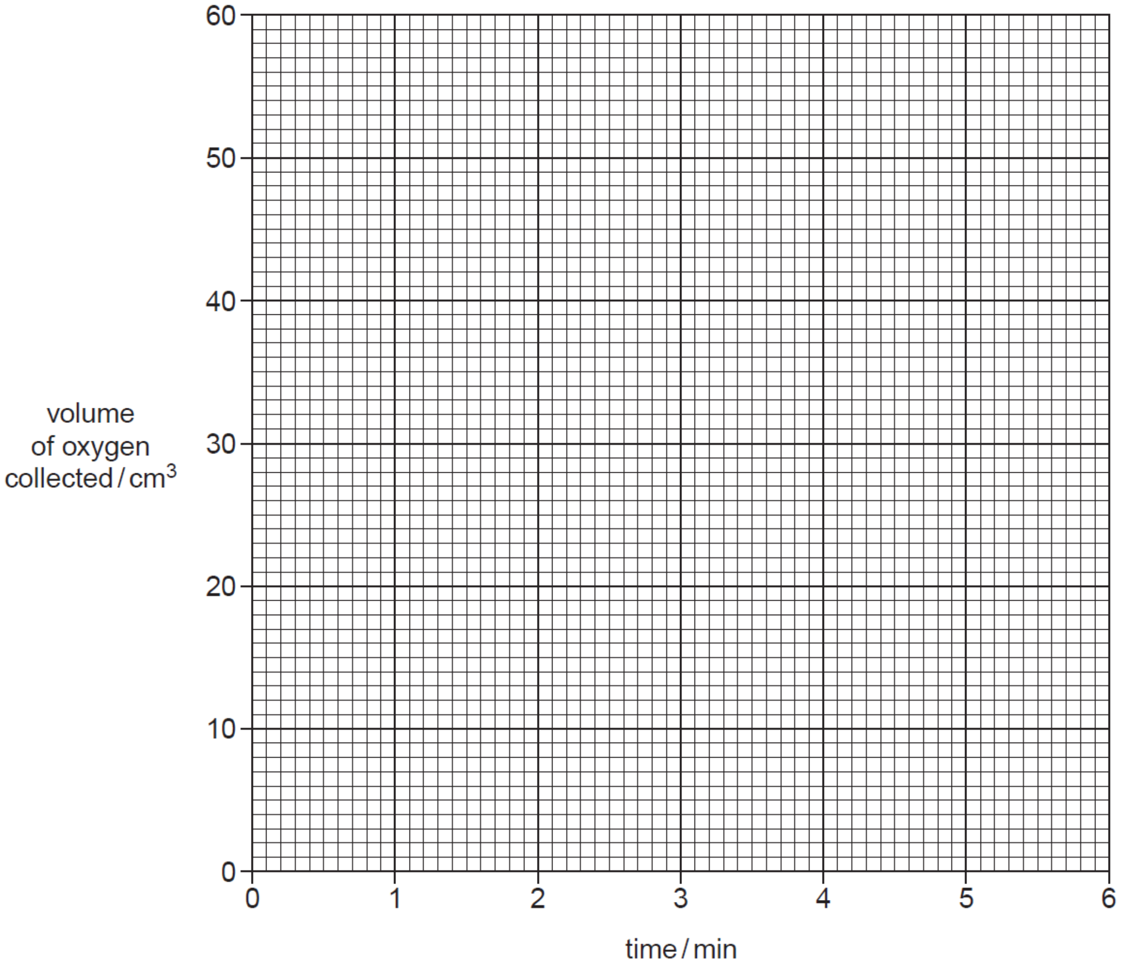


4 min

Time / min	1	2	3	4	5	6
Volume of oxygen collected in experiment **1** / cm^3	9	17	24	29	32	35
Volume of oxygen collected in experiment **2** / cm^3	21	35	43	48	50	50

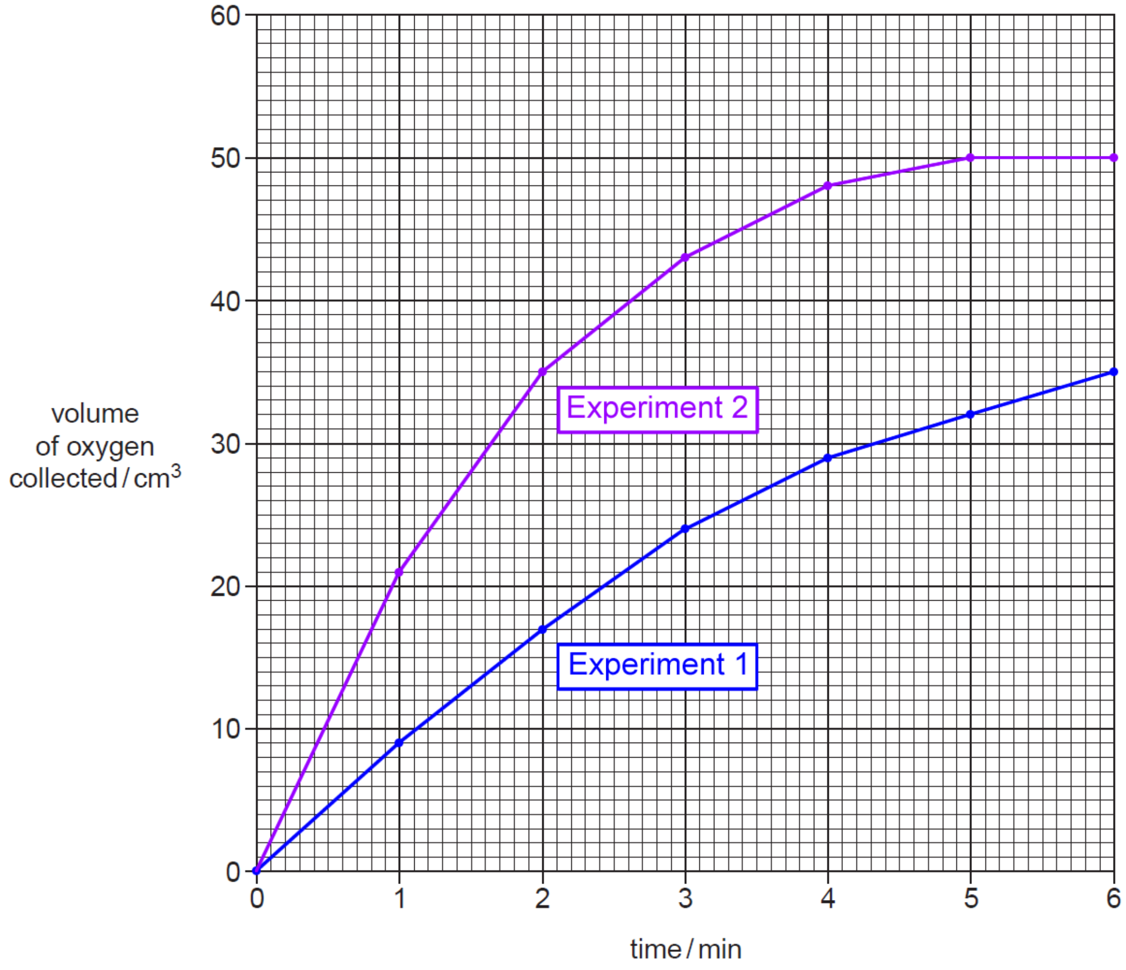
Rate of Reaction

b) Plot the results for experiment **1** and **2** on the grid below and draw a smooth curve through each set of points. Label the curves **1** and **2**.



Rate of Reaction

b) Plot the results for experiment **1** and **2** on the grid below and draw a smooth curve through each set of points. Label the curves **1** and **2**.



Note: You should draw a *smooth curve* through the points.

Rate of Reaction

c) Which of the experiments first reached completion? Explain your answer.

……………………………………………………………………………………

……………………………………………………………………………………

……………………………………………………………………………………

d) Use your graph to estimate the time taken in experiment **1** (using manganese(IV) oxide) to double the volume of oxygen produced from 15 cm^3 to 30 cm^3. Record your answers in the table below.

	Experiment **1**
Time taken to produce 30 cm^3 / min	
Time taken to produce 15 cm^3 / min	
Time taken to double the volume from 15 cm^3 to 30 cm^3 / min	

Rate of Reaction

c) Which of the experiments first reached completion? Explain your answer.

Experiment **2**. The last two volume readings are the same, which means that the reaction has finished. For experiment **1**, the volume readings are still changing, which means that the reaction is still taking place.

d) Use your graph to estimate the time taken in experiment **1** (using manganese(IV) oxide) to double the volume of oxygen produced from 15 cm^3 to 30 cm^3. Record your answers in the table below.

	Experiment **1**
Time taken to produce 30 cm^3 / min	4.20 minutes
Time taken to produce 15 cm^3 / min	1.70 minutes
Time taken to double the volume from 15 cm^3 to 30 cm^3 / min	4.2 – 1.7 = 2.50 minutes

Rate of Reaction

e) The rate of a reaction can be calculated using the formula:

$$\text{rate of reaction} = \frac{\text{volume of gas produced / cm}^3}{\text{time taken / min}}$$

Using the two graphs and the above formula, calculate the rate of each reaction after the first 2.5 minutes.

i) Rate of reaction using manganese(IV) oxide (experiment **1**).

ii) Rate of reaction using copper (experiment **2**).

iii) Using your answers to **i)** and **ii)**, suggest which is the better catalyst, manganese(IV) oxide or copper. Explain your answer.

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Rate of Reaction

e) The rate of a reaction can be calculated using the formula:

$$\text{rate of reaction} = \frac{\text{volume of gas produced / cm}^3}{\text{time taken / min}}$$

Using the two graphs and the above formula, calculate the rate of each reaction after the first 2.5 minutes.

i) Rate of reaction using manganese(IV) oxide (experiment **1**).

$21 \div 2.5 = 8.40 \text{ cm}^3 \text{ / min}$

ii) Rate of reaction using copper (experiment **2**).

$40 \div 2.5 = 16.0 \text{ cm}^3 \text{ / min}$

iii) Using your answers to **i)** and **ii)**, suggest which is the better catalyst, manganese(IV) oxide or copper. Explain your answer.

Copper is the better catalyst. The reaction is faster / the rate of the reaction is greater / the speed of the reaction is greater.

Rate of Reaction

f) At the end of experiment **2** the copper was removed from the solution by filtration. It was dried and weighed. How does this mass of copper compare with the mass of copper used at the start of the experiment? Explain your answer.

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g) Suggest two ways by which the rate of decomposition in either experiment could be further increased.

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Rate of Reaction

f) At the end of experiment **2** the copper was removed from the solution by filtration. It was dried and weighed. How does this mass of copper compare with the mass of copper used at the start of the experiment? Explain your answer.

The mass of the copper will be the same at the end of the experiment as it was at the start of the experiment. The copper is a catalyst, and catalysts increase the rate of a reaction without being consumed by the reaction.

g) Suggest two ways by which the rate of decomposition in either experiment could be further increased.

By increasing the temperature of the hydrogen peroxide.

By increasing the concentration of the hydrogen peroxide.

By increasing the surface area of the catalyst.

Rate of Reaction

Rate of Reaction

Presentation on
Rate of Reaction
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