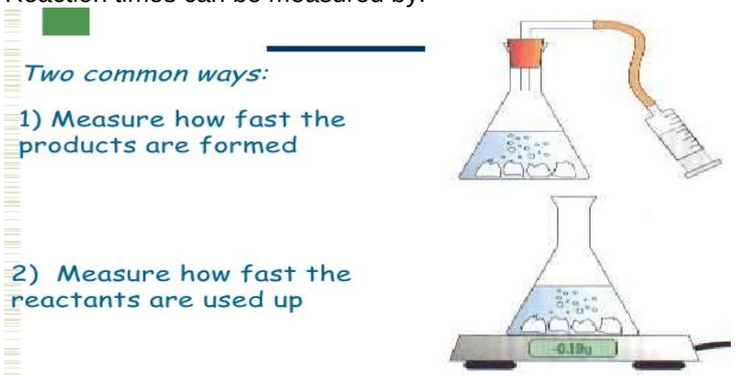
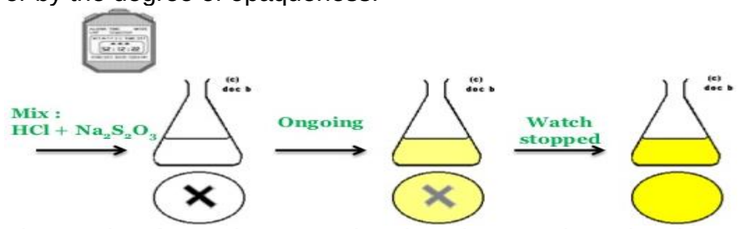
Chemistry knowledge organiser – unit 6 The rate and extent of a chemical reaction

Ways to measure reaction time

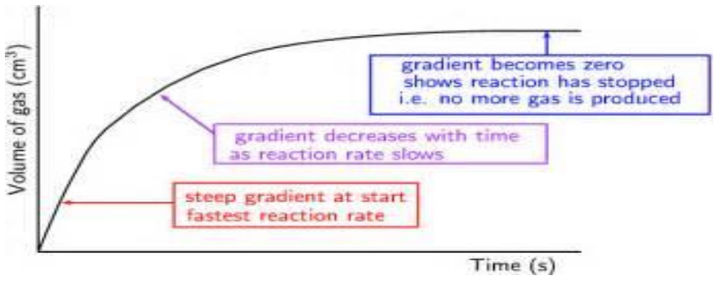
Reaction times can be measured by:



or by the degree of opaqueness:



The results of experiment can be plotted on graphs. The slope of the line shows how fast the reaction is. **The steeper the slope the faster the reaction.**



Limiting reactants are used up during the reaction (it stops when they run out)

The excess reactant is left over when the reaction stops.

If the **limiting reactant** (factor) is doubled the amount of product is also doubled as they are **directly proportional** in quantity.

Calculating rates or reaction

The reaction time is how long it takes for the reaction to finish.

The shorter the reaction time = the faster the reaction.

Rate of reaction can be calculated by:

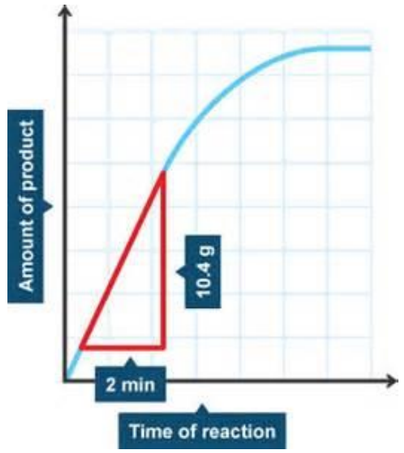
$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

OR

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

Reactions are usually faster at the start and begin to slow down once the limiting reactant starts to get used up.

The rate of reaction can be worked out using the gradient of the graph. **Construction lines** are added using the part of the graph where there is a **straight line**.



1. measure the value of y and x using the scale
2. calculate gradient = y/x

In this case 10.4/2 = 5.2 g/min

Collision Theory

In order for any reaction to take place particles must collide when the particles have the minimum energy required – this is called **activation energy.**

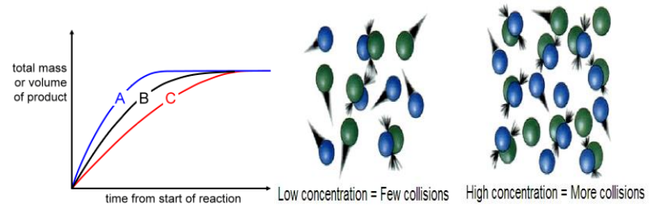
Factors affecting rates of reactions

There are four factors affecting the rate at which a reaction takes place:

- **concentration/pressure**
- **temperature**
- **surface area** of solid reactants
- **catalysts**

Increasing concentration/pressure: **more** particles in a **given** volume.

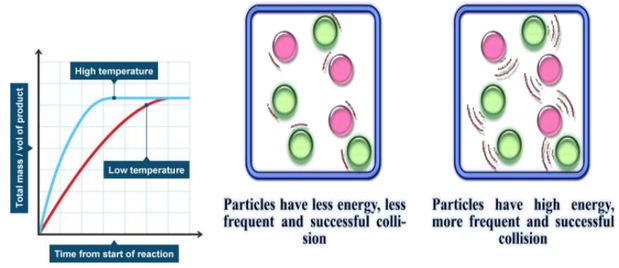
- Increases the number of particles with activation energy
- Increases the number of **successful** collisions and
- Increases the **rate** of the reaction.



A is a more concentrated compared to B or C as the reaction at the start is much faster. All 3 concentrations will reach the same total mass or volume produced due to the **limiting reactant**.

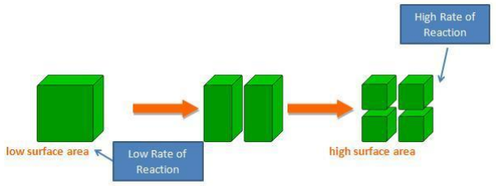
Increasing temperature: all of the particles gain kinetic energy.

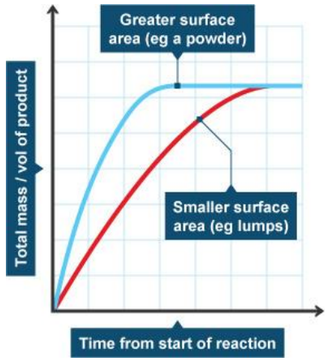
- More of the particles have the minimum activation energy
- Successful collisions will occur more often
- Increasing the rate of the reaction.



Surface area

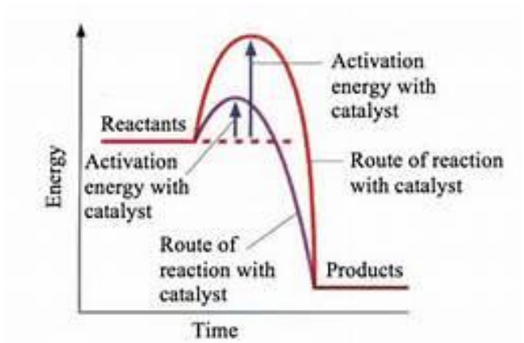
Cutting/crushing a solid into smaller pieces increases the number of surfaces and the reaction rate as there are more particles able to successfully collide.





Catalyst

- A **catalyst** changes the rate of the reaction by **lowering the activation energy**, but is not used up in the reaction. A lower activation energy allows **more** of the collisions to be successful.

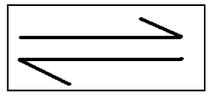


When planning a rates experiment consider:

-mass/volume of limiting reactant
-concentration of the limiting/excess reactant
-Temperature of the reactants
-Surface area of reactant
-How often you record mass change/gas collected (eg every X seconds)

Reversible reactions and energy changes

This symbol is used for reversible reactions:

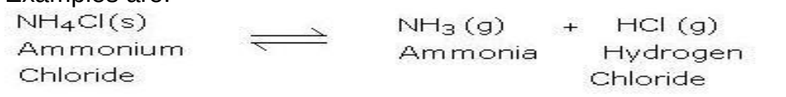


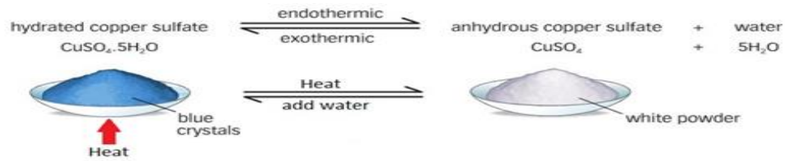
In a chemical reaction usually:

reactants → products

However, in a reversible reaction there is a **forward** and a **backwards** reaction taking place at the same time.

Examples are:





In **dynamic equilibrium** the forward and backwards reactions continue at **equal** rates so the concentrations of reactants and products do not change. The system needs to be **closed** – isolated from the outside world.

Le Chatelier's principle (HT)

-If you remove product as it is made then equilibrium will move to the right to counteract the change by making more.
-If you add more reactant then equilibrium will move to the right to counteract the change by removing the extra.

HTIER – Changing concentration and equilibrium

If the concentration of a reactant/product is changed. The concentration of **all** of the substances will change until the equilibrium is restored.

- Increase of reactant → more products are formed
- Decrease of product → more reactants are used.

HTier – Changing the temperature and equilibrium

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

-In this rection the forward reaction is **exothermic**.
-This means to increase the yield of ammonia the temperature must be **decreased.**

HTier – Changing the pressure and equilibrium

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

4 molecules → 2 molecules

An increase in pressure the equilibrium moves to the side with the **smallest** number of molecules (product).

1. What measurements do we need to find the rate of a fizzing reaction?
2. What piece of equipment is used to measure gas volumes?
3. How does the rate of reaction change over time?
4. What is reaction time? .
5. How can you measure the rate of a reaction that goes cloudy?
6. Explain what limiting reactant means.
7. How can you tell if a reactant is in excess?
8. What does directly proportional mean?
9. Limiting reactant or excess reactant. Which has the most moles?
10. Why does the rate of reaction slow down as the reaction goes on?
11. How is mean rate calculated using volumes of gas?
12. How does a graph show you when the reaction is fastest?
13. How is a gradient calculated from a graph?
14. Why is the gradient important?
15. What units are used for the rate of a mass loss reaction?
16. How does increasing temperature cause an increase in the frequency of successful collisions?
17. How does increasing concentration cause an increase in the frequency of successful collisions?
18. How does making a solid into a powder cause an increase in the frequency of successful collisions?
19. What is the minimum energy required for a successful collision called?
20. How can you increase the rate of a reaction that involves gases? .
21. How does a catalyst work?
22. Why is only a small amount of a catalyst required?
23. What does specificity mean?
24. Why are catalysts popular in industry?
25. How is the use of a catalyst shown on an energy profile?
26. What does Le Chatelier's principle mean?
27. What happens when a reaction is at equilibrium?
28. How does increasing the pressure of reacting gases shift the position of equilibrium?
29. How does increasing the temperature of a reaction shift the position of equilibrium?
30. How does increasing the concentration of a reactant change the position of equilibrium?