With 2020 NCEA Exam included

2021 Version

Chemistry AS 91166

C2.6 Chemical Reactivity

GZ Science Resources

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Achievement Criteria

AS 91166
C2.6

This achievement standard involves demonstrating understanding of chemical reactivity.

Rates of Reaction typically involves:

- ❑ factors affecting rates of reaction – restricted to changes in concentration, temperature, surface area, and the presence of a catalyst
- ❑ using particle theory to explain the factors (includes activation energy).

Equilibrium principles is limited to:

- ❑ the dynamic nature of equilibrium
- ❑ the effect of changes in temperature, concentration, pressure, or addition of a catalyst on equilibrium systems
- ❑ the significance of the equilibrium constant (Kc) for homogeneous systems. This may involve calculations.
- ❑ the nature of acids and bases in terms of proton transfer
- ❑ properties of aqueous solutions of strong and weak acids and bases including ionic species such as NH_4^+. The properties are restricted to conductivity, rate of reaction, and pH
- ❑ calculations involving Kw and pH (restricted to strong acids and bases).

Background Knowledge

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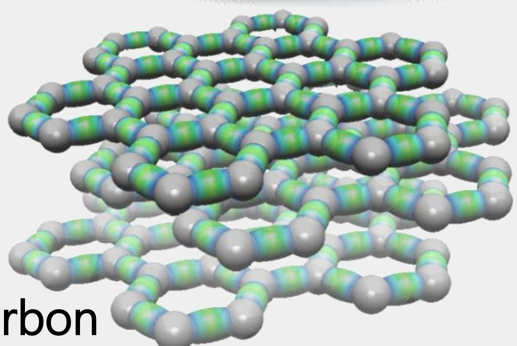
The Particle Theory of Matter

1. All matter is made up of **very small particles** (atoms, ions or molecules)
2. Each substance has **unique particles** that are different from particles of other substances
3. There are **spaces between the particles** of matter that are very large compared to the particles themselves
4. There are **forces** holding particles together
5. The **further apart** the particles, the **weaker** the forces holding them together
6. Particles are in **constant motion**
7. At **higher temperatures** particles on average **move faster** than at lower temperatures.



Carbon in the form of graphite

Background Knowledge

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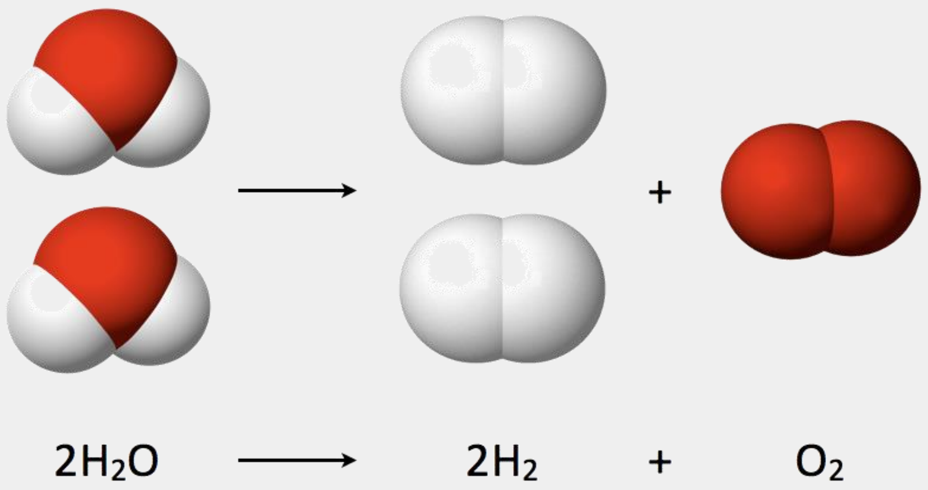
Chemical Reactions - reactants & products

Reactants → Products



A **chemical reaction** is a process that produces a chemical change to one or more substances.

A chemical reaction will produce a **new substance.**

Other observations may include a temperature change, a colour change or production of gas.

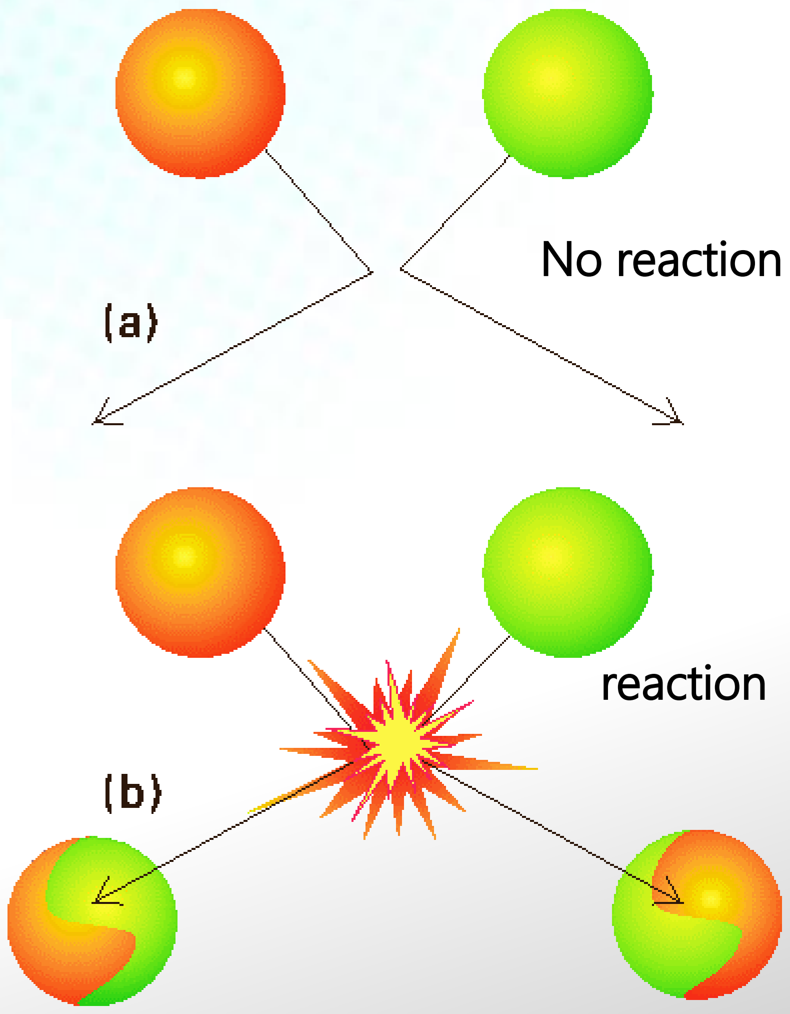
Chemicals that are used in a chemical reaction are known as **reactants**. Those that are formed are known as **products**.

Chemical reactions between particles involve breaking bonds and forming new bonds.

Collision Theory

Chemical reactions between particles of substances only occur when the following conditions have been met:

- ❑ **Particles must collide.**
- ❑ **With enough energy** (called activation energy E_a)
- ❑ **And with the correct orientation**

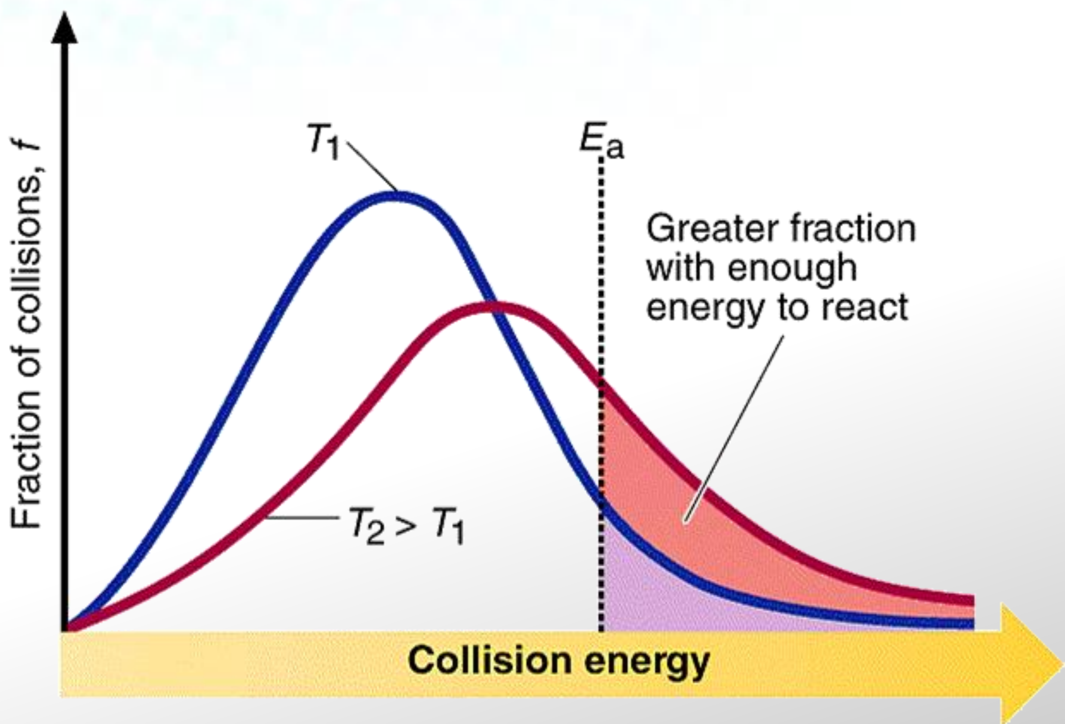


If these conditions are met the collision will be considered **successful**. (or effective)

Activation Energy (E_A)

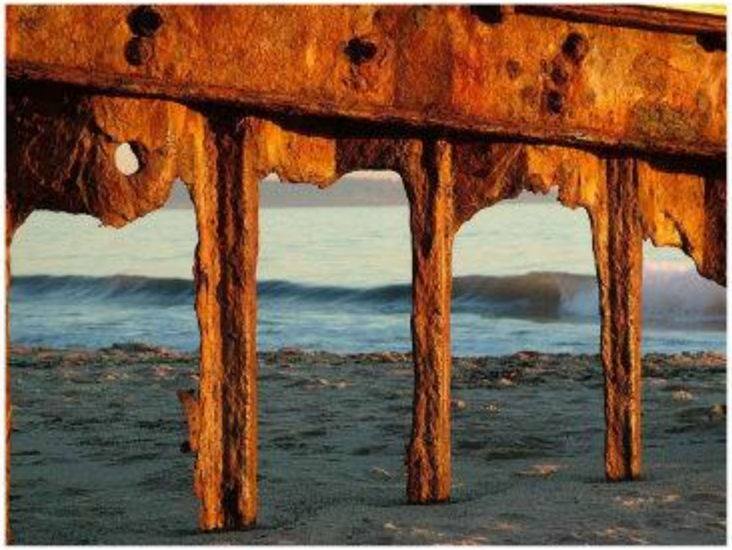
Activation energy is the initial energy required for a reaction to occur. It could be provided in the form of heat or kinetic energy.

At lower temperatures (T1) most particles will have the same collision energy when they collide. As the temperature is increased (T2) the range of collision energy is more spread out and a greater proportion of particles will have enough energy to cross the activation energy threshold and therefore react.



Reaction Rate

The reaction rate is the speed at which a chemical reaction occurs. This is measured by how quickly the reactants change into products or how quickly one of the reactants disappears.
Reactions can vary in their reaction rate.

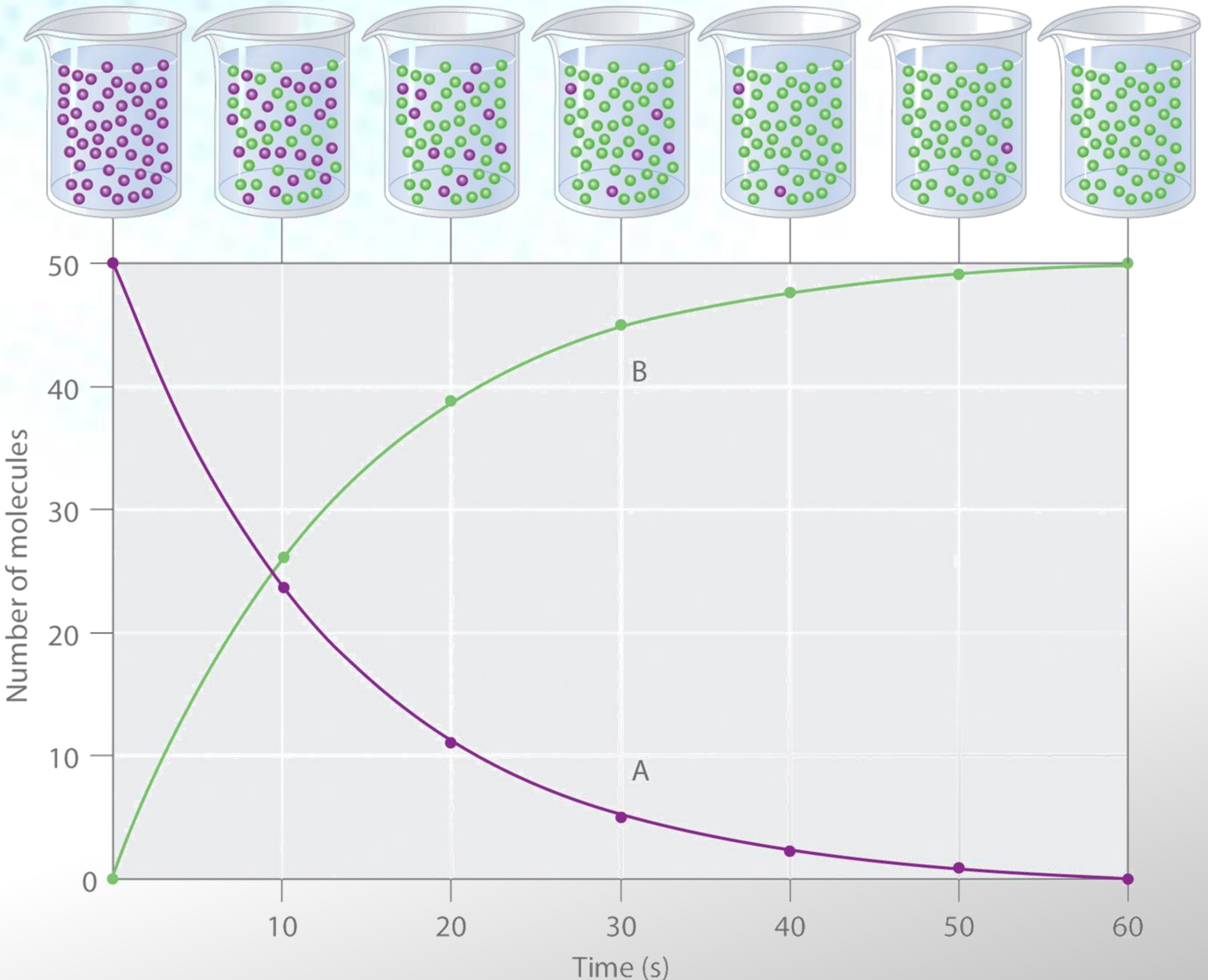


Iron oxidising

oxygen and hydrogen combusting

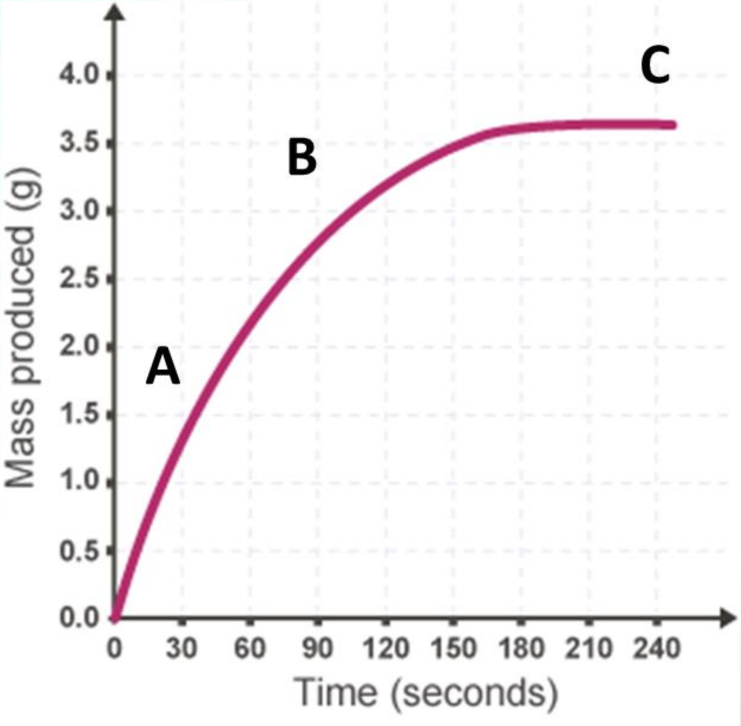
Reaction Rate

Reactions take place over time. As the amount of reactants decrease the amount of products increase. The reaction rate is shown as a curve because the amount of reactants at the start is greater and the reaction rate slows as they decrease.



Reaction rate is the speed at which a chemical reaction occurs

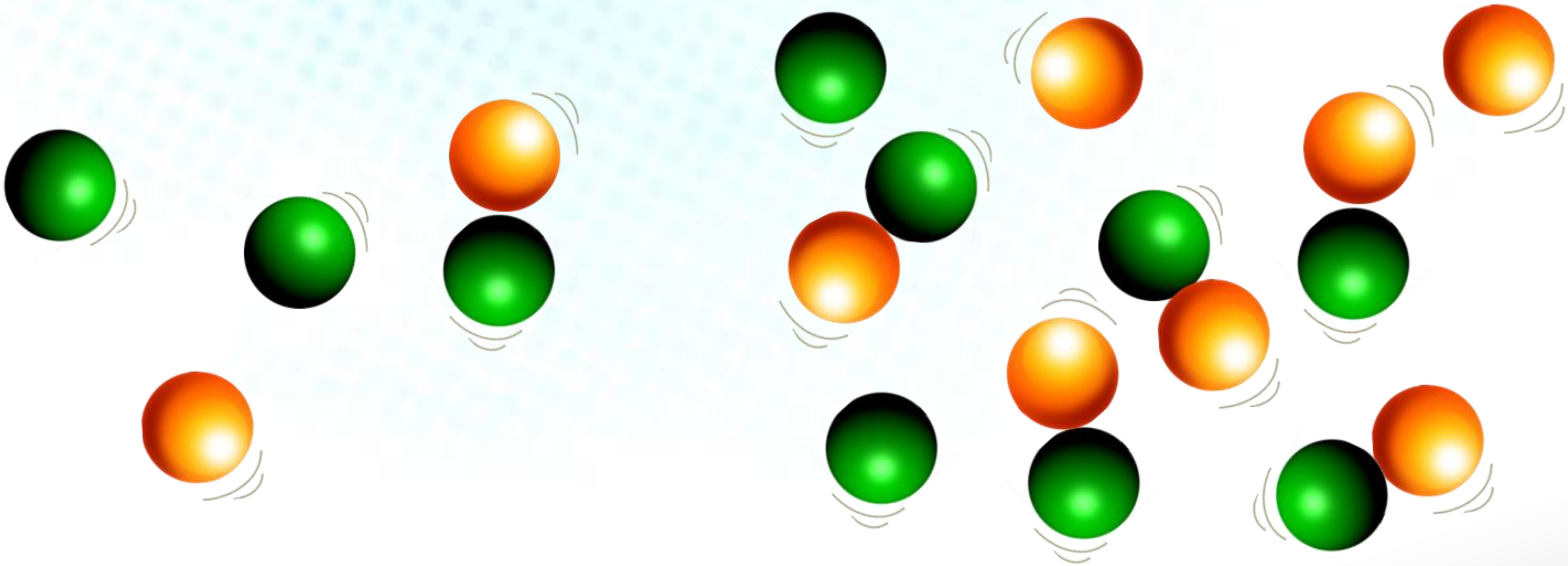
A. Reactions start out relatively fast because there is a much higher concentration of reactant particles available to collide and therefore the frequency of collisions will be high, resulting in more successful collisions; resulting in an increase in the rate of reaction. The gradient of the line on the graph for products formed will be high.



B. As the reaction proceeds there will be less reactant particles available to collide as many have already reacted to form products. The gradient of the line will be lower, and therefore the frequency of collisions will be less, resulting in less successful collisions; resulting in the slowing down in the rate of reaction.

C. When the reaction has come to completion, when all of the reactants has reacted to form particles, then there will be no further collisions and the gradient of the line will be zero. (note: for reactions the reach equilibrium the rate that products are made will match the rate reactants are reformed and therefore the gradient of the line will still zero)

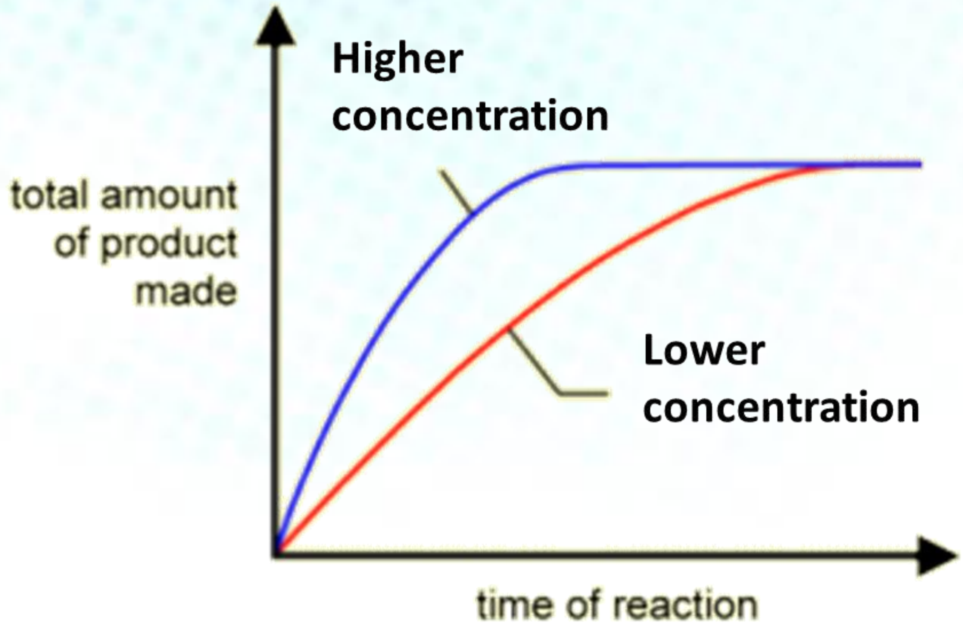
Reaction rate can be increased by increasing the concentration



Low concentration = few collisions **High concentration = more collisions**

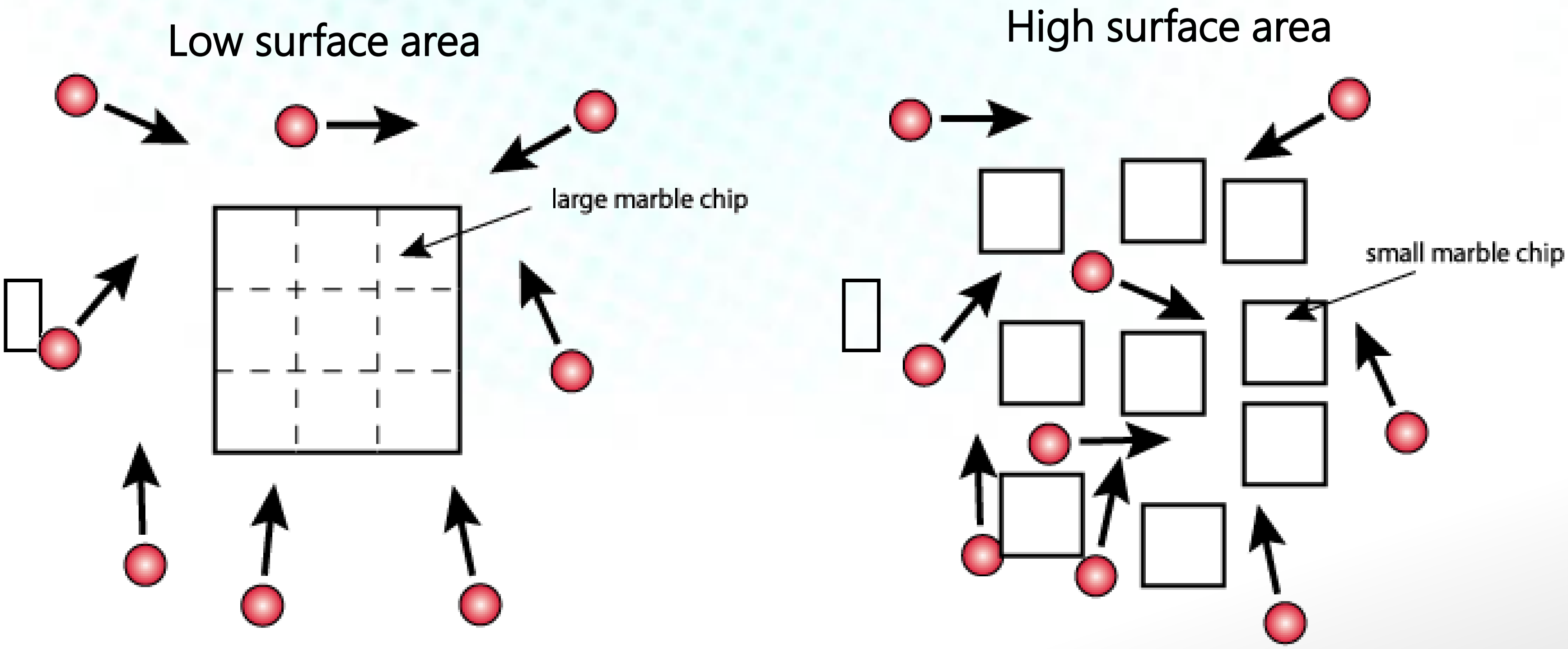
If there is a higher concentration of a substance in a system, there is a greater chance that molecules will collide, as there is less space between particles. The higher frequency of collisions means there are more effective collisions per unit of time and this will increase the rate of the reaction. If there is a lower concentration, there will be fewer collisions and the reaction rate will decrease.

Reaction rate can be increased by increasing the concentration



It is important to note that the total amount of product made depends upon the total amount of reactants at the start. A 1 $molL^{-1}$ solution will contain only half the particles of a 2 $molL^{-1}$ so twice the volume will be required to produce the same quantity of product.
Also note that at the end point of both reactions the **total amount of successful collisions does not change** by increasing the concentration (as both have produced the same amount of products) only the frequency (amount of collisions per unit of time) of collisions is increased.

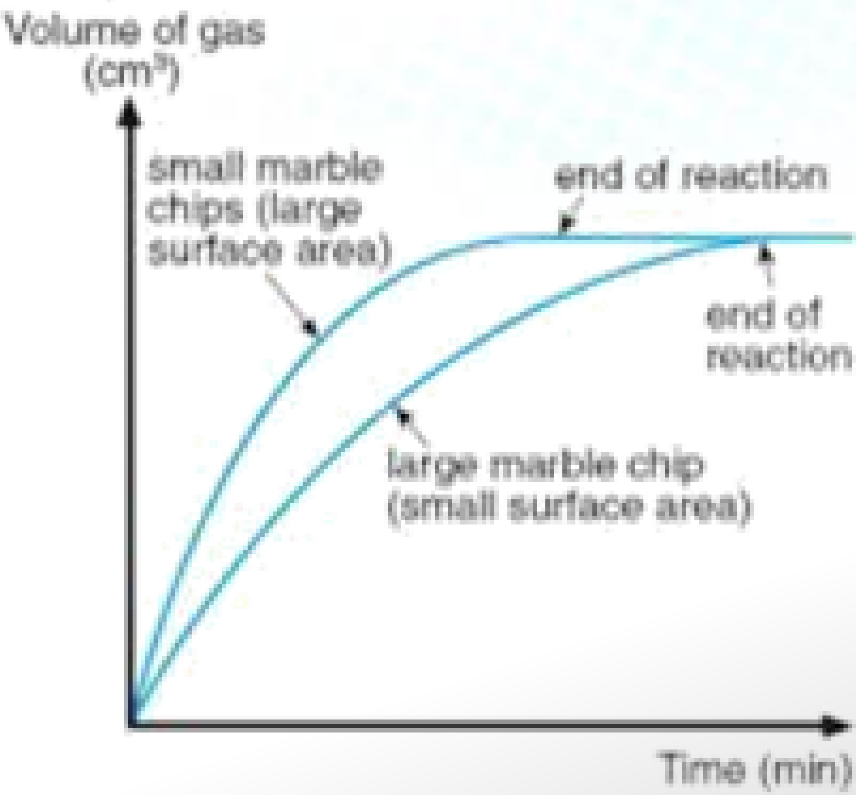
Reaction rate can be increased by increasing the Surface Area



Surface area can be increased by grinding and crushing large lumps into a finer powder. The smaller the pieces, the greater the surface area. The reactant(s) with a **greater surface area will have a faster rate of reaction** than the same amount of reactants with a smaller surface area.

Reaction rate can be increased by increasing the surface area

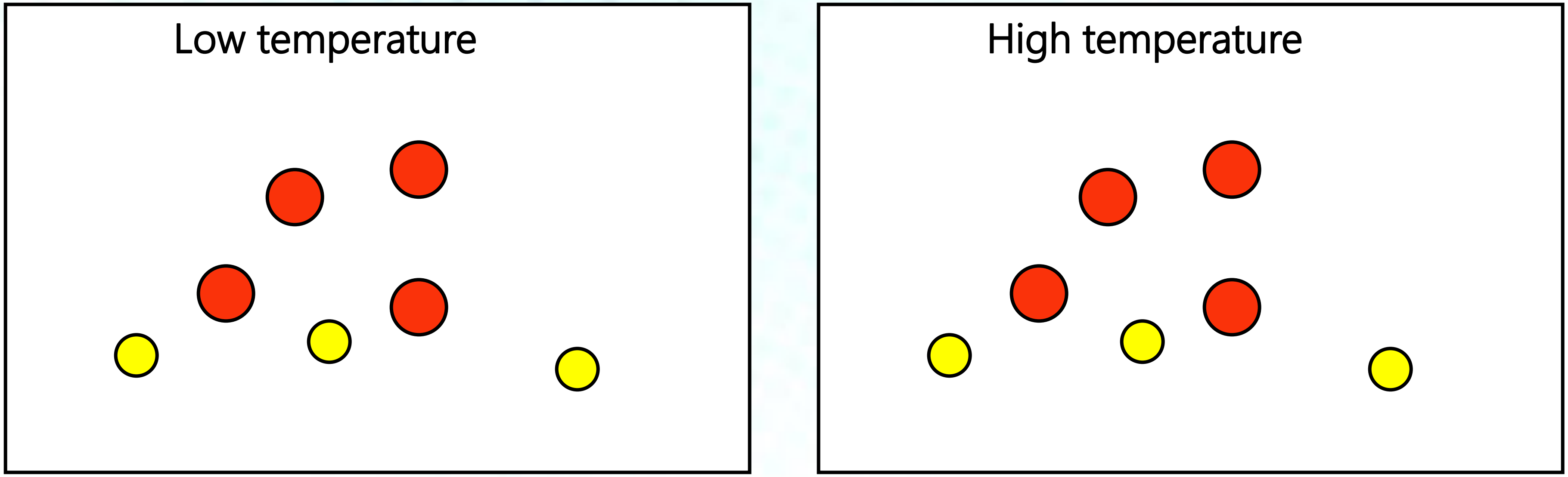
By increasing surface area a greater number of reactant particles are exposed and therefore able to collide. The frequency of collisions (number of collisions per unit of time) will increase and therefore the frequency of successful collisions so the reaction rate will also increase.



An example is comparing the reaction between marble (calcium carbonate) and hydrochloride acid to produce carbon dioxide gas.

Note: although the reaction rate is higher for the smaller marble chips the total amount of gas (CO_2) produced is the same for both reactions as they both started off with the same amount of reactants.

Reaction rate can be increased by increasing the Temperature

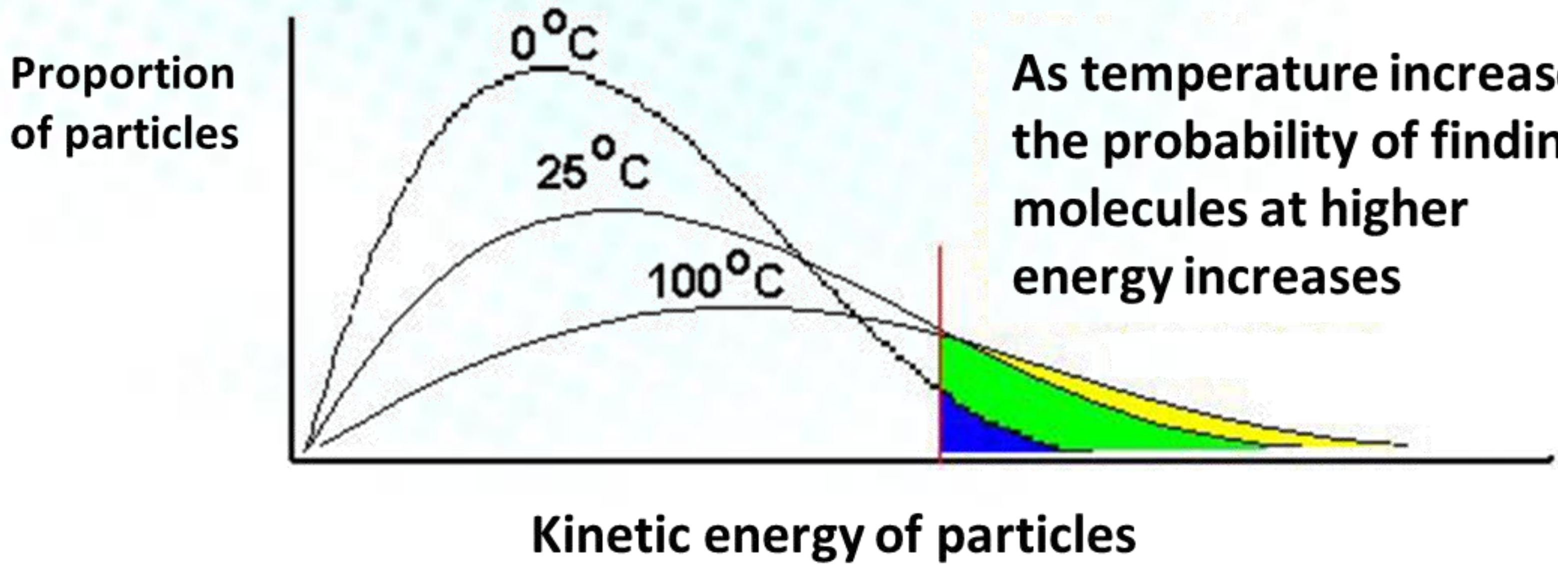


Increasing temperature effects the reaction rate in two ways.

Firstly, when you raise the temperature of a system, the particles move around a lot more (because they have more kinetic energy). When they move around more, they are **more likely to collide and the frequency of collisions increases**, therefore the reaction rate increases. When you lower the temperature, the molecules are slower and collide less frequently therefore the reaction rate decreases.

Secondly, at a higher temperature a **larger proportion of particles have sufficient energy to overcome the activation energy required during a collision for it to be successful** and therefore a reaction to occur. This increases the proportion of successful collisions and therefore the reaction rate.

Reaction rate can be increased by increasing the Temperature



As temperature increases the probability of finding molecules at higher energy increases

At any given temperature, there will be a range in the kinetic energy of particles. At a lower temperature a greater proportion of particles are likely to have insufficient kinetic energy during a collision in order for a successful collision, and therefore a reaction, to take place.

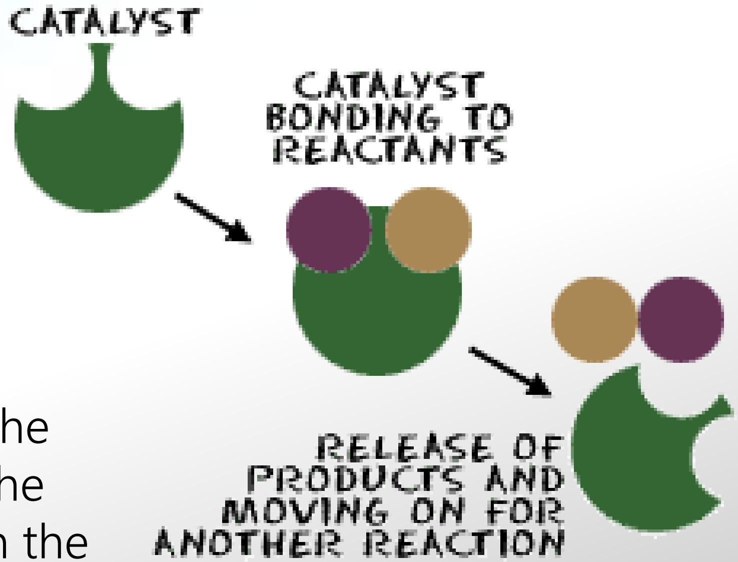
Increasing the temperature also increases the probability of a successful collision.

Reaction rate can be increased by using a catalyst.

A catalyst is a substance that increases the reaction rate without being used up or forming part of the products. Only some reactions have catalysts that are effective, but for many reactions there is no catalyst that works.

How does a catalyst work?

A catalyst lowers the activation energy pathway (the minimum amount of energy required for a reaction to take place). This means that the particles can successfully collide with less energy than they required before the catalyst was added. **A greater proportion of particles will successfully collide**, and therefore the reaction rate will be increased.



extra info

A catalyst provides a surface on which the reaction can take place. This increases the number of successful collisions between the particles of the substances as the particles come in **closer contact with each other and are more likely to be in the correct orientation**.

Reaction rate can be increased by using a catalyst.

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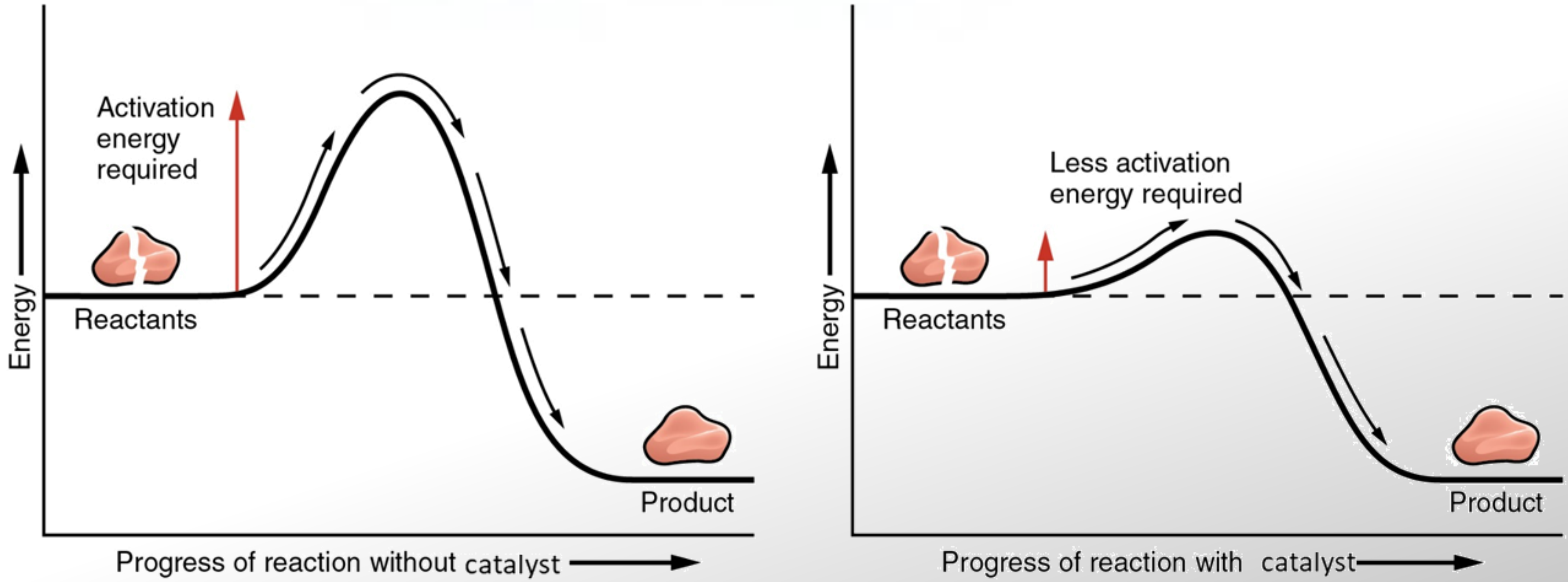
The reaction shown below is an exothermic reaction: the enthalpy level of the products is lower than the enthalpy of the reactions therefore energy is released during this reaction.

Activation energy is required before a collision is successful between particles.

A catalyst lowers the activation energy pathway and therefore a greater proportion of particles will have sufficient energy during collision for it to be successful.

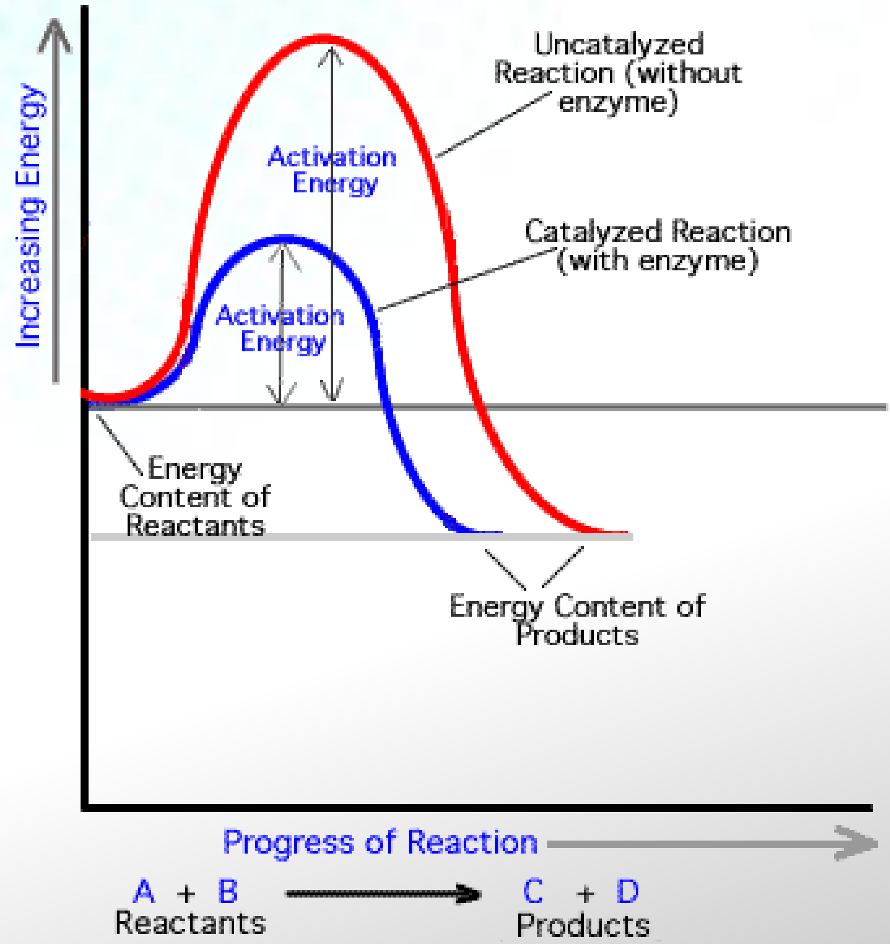


Biological catalysts

extra info

Cell metabolism consists of reactions between chemicals. Chemical reactions require an amount of energy in order for them to be successful. This energy is called activation energy.
Enzymes are biological catalysts, and are often produced by the cell in the form of proteins.

The human body temperature of around 37.5°C is often too cold for many metabolic reactions to take place at a sufficiently fast reaction rate so enzymes are needed to increase the reaction rate.



Summary of Factors affecting Reaction rate

<u>Increase the frequency of collisions</u>

- By **increasing surface area**: smaller pieces of reactant expose more reactant particles to collisions. Stirring will also increase the reaction rate
- By **increasing the concentrations**: more reactant particles exist in a given volume so more collisions occur

<u>Increase the energy of collisions</u>

>by **increasing temperature**: particles move faster so have more kinetic energy. More collisions will be effective.

Note: increasing temp also increases frequency of collisions

<u>Make it easier for reaction to occur</u>

>by **using a catalyst**: allows reaction to occur along a different pathway that requires less activation energy

Reaction rate can also be slowed down by decreasing these

Writing Reaction rate Answers

1. Particles need to **collide** with **sufficient kinetic energy** and in the **correct orientation** in order for an effective/successful collision to occur.
2. Increasing surface area, temperature and concentration of reactants increases the **number of collisions per unit of time** (frequency)
3. Increasing temperature increases **both** the number of collisions per unit of time **and** the average amount of kinetic energy the particles have, so more particles have sufficient energy to obtain the activation energy requirements. Discuss both effects.
4. Always **identify the factor involved**, ideally at the beginning of the answer: surface area, temperature, concentration or catalyst. If you are unsure look at the remaining questions as the same factor is rarely used twice.
5. **Link** the **increase in effective/successful collisions** to an **increase in reaction rate**.

Summary Revision - Reaction Rate

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What needs to occur for a collision to be successful?

Factor	Does it increase frequency of collisions	Does it increase the % of successful collisions?
1.		
2.		
3.		
4.		

Link answers to increase / decrease in reaction rate

Summary Revision - Reaction Rate

GZ Science Resources

What needs to occur for a collision to be successful?

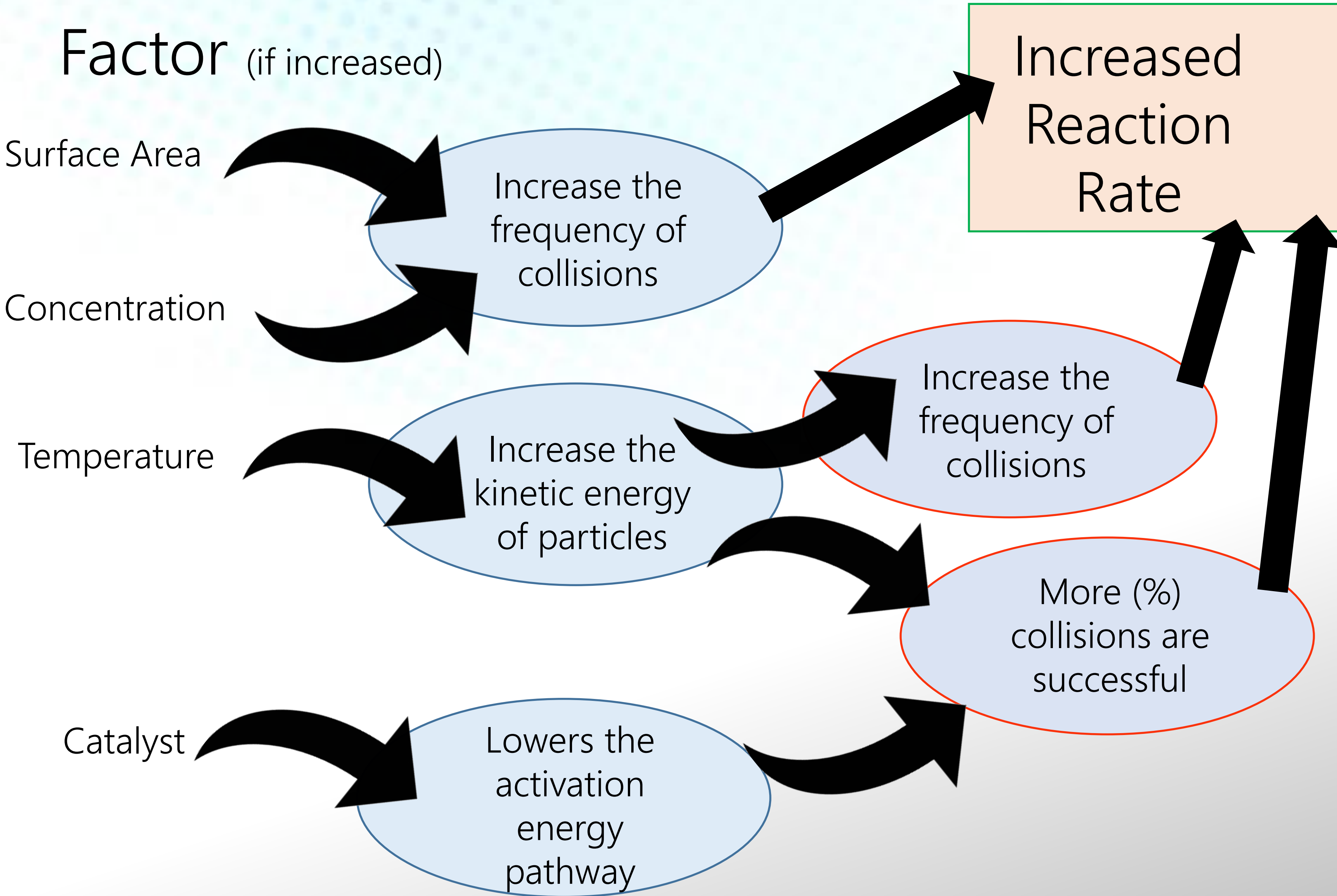
- ❑ Particles must collide with enough energy – to overcome activation energy requirements
- ❑ Collide in the correct orientation

Factor (increased)	Does it increase frequency of collisions	Does it increase the % of successful collisions?
1. Concentration	Yes More particles in a given area, therefore more chance of colliding	No But more successful collisions (per unit time) – as more frequent collisions
2. Surface Area	Yes More particles in a given area, therefore more chance of colliding	No But more successful collisions(per unit time) – as more frequent collisions
3. Temperature	Yes Particles have more kinetic energy – move faster therefore more chance of colliding	Yes More particles have required energy to overcome activation energy therefore result in successful collision
4. Catalyst	No	Yes A lower activation energy pathway available – "lowers the bar' and a greater proportion of collisions become successful. So catalysts also assist orientation

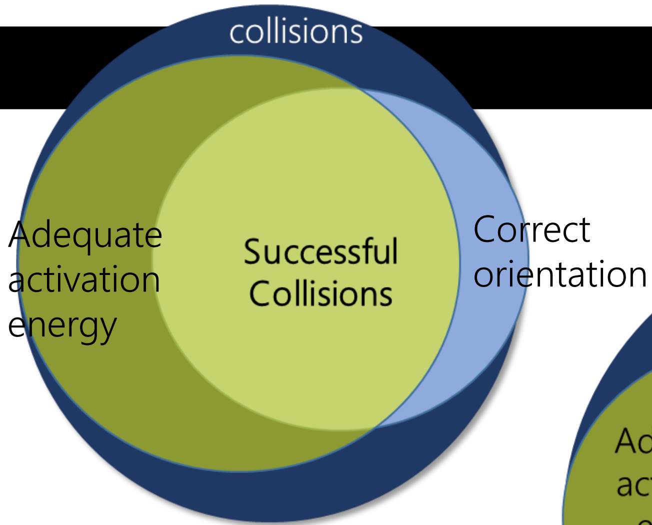
Link answers to increase / decrease in reaction rate

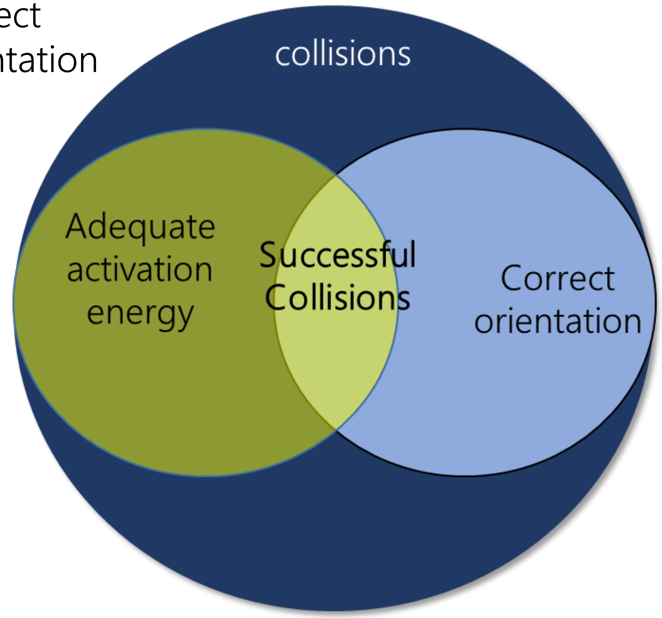
Reaction rates

GZ Science Resources



Reaction rates





Temperature

- ❑ Frequency of collisions increased

ALSO

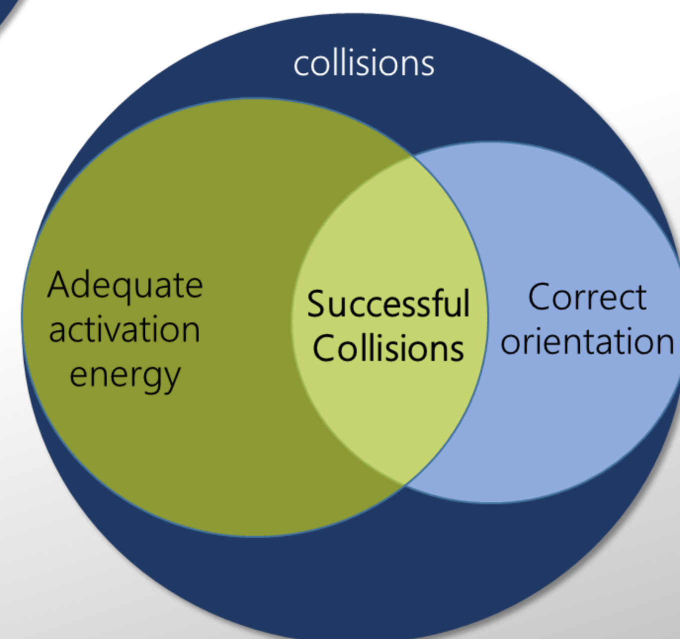
- ❑ Increase in % of successful collisions

Catalyst

- ❑ Frequency of collisions not changed
- ❑ Increase in % of successful collisions
- ❑ Due to lowering activation energy pathway
- ❑ Some catalysts also ‘align’ reactants for collision

Concentration
Surface Area

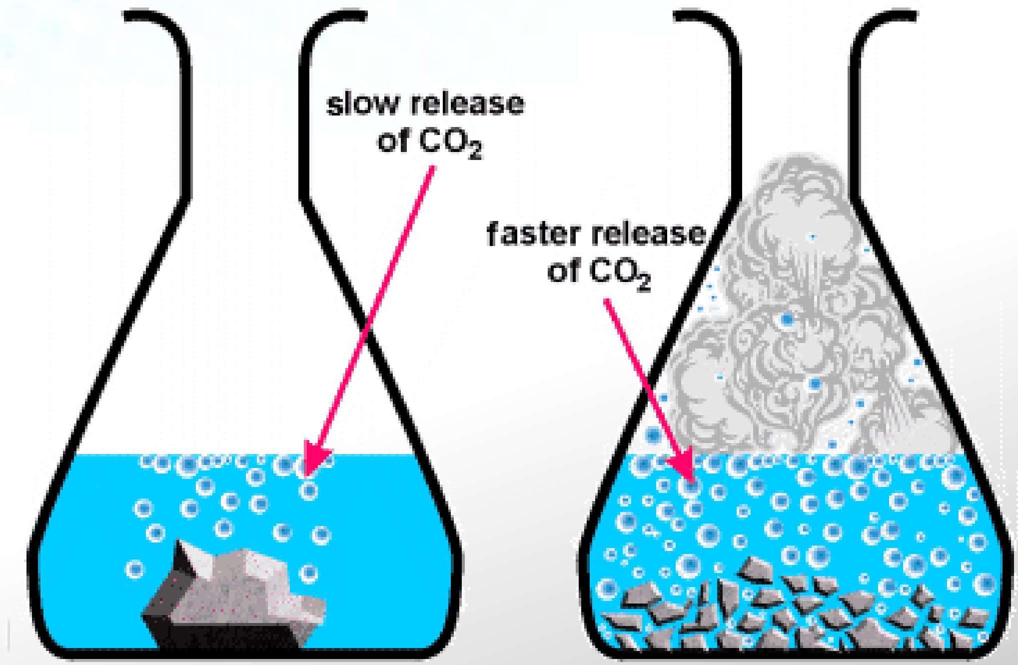
- ❑ Frequency of collisions increased
- ❑ Overall % of successful collisions not changed BUT RATE of successful collisions is



NCEA 2013 Reaction rate

Question 1a: Hydrochloric acid was reacted with calcium carbonate in the form of marble chips (lumps) and powder (crushed marble chips) in an experiment to investigate factors affecting the rate of a chemical reaction. (i) Identify the factor being investigated.
(ii) Explain why the hydrochloric acid would react faster with the powder.

When the marble chips are crushed there is a **greater surface area**. This means there are now **more particles for collisions** to occur between the acid and the calcium carbonate. Because more collisions can now occur **more frequently** the **reaction rate is faster**.



NCEA 2013 Reaction rate (PART ONE)

Question 1b: A particular reaction is complete when the solution turns cloudy and the paper cross under the flask can no longer be seen. The following experiments were carried out, and the times taken for the cross to disappear recorded.

experiment		**Temperature /ºC**	**Time for cross to disappear**
1	No Cu^{2+} present	25	42
2	No Cu^{2+} present	50	23
3	Cu^{2+} present	25	5

Elaborate on why the reactions in **Experiment 2** and **Experiment 3** occur faster than the reaction in **Experiment 1**.

In Experiment 2, the only change is an increase in temperature. An increase in temperature means a faster rate of reaction. For a **chemical reaction to occur, the reactants must collide effectively**. This means they must collide with enough energy to **overcome the activation energy** of the reaction. The activation energy is the energy that is required to start a reaction. When the temperature is higher, the particles have **more kinetic energy**; the particles are **moving faster**. Because the particles are moving faster, there will be **more frequent collisions**. Also because the particles are moving with more kinetic energy, it will be more likely that when collisions occur they are **more likely to be effective**, Therefore the **rate of reaction is faster**, as <u>more effective collisions are occurring more frequently</u>.

NCEA 2013 Reaction rate - (PART TWO)

Question 1b: A particular reaction is complete when the solution turns cloudy and the paper cross under the flask can no longer be seen. The following experiments were carried out, and the times taken for the cross to disappear recorded.

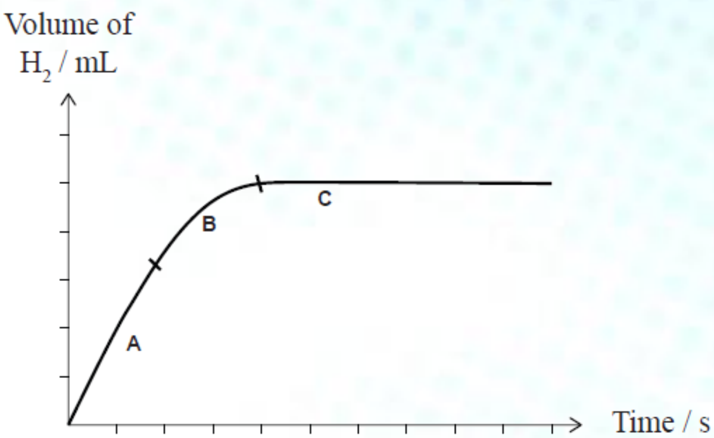
experiment		Temperature /ºC	Time for cross to disappear
1	No Cu^{2+} present	25	42
2	No Cu^{2+} present	50	23
3	Cu^{2+} present	25	5

Elaborate on why the reactions in **Experiment 2** and **Experiment 3** occur faster than the reaction in **Experiment 1**.

In Experiment Three, a catalyst is used (the copper ions). **Use of a catalyst speeds up the rate of chemical reaction**. For a chemical reaction to occur, the reactants must collide effectively. This means they must collide with enough energy to overcome the activation energy of the reaction. The activation energy is the energy that is required to start a reaction. When a catalyst is used, the **activation energy is lowered**. This is because the catalyst provides an **alternative pathway** for the reaction to occur in which the activation energy is lowered. Now that the activation energy has been lowered, **more reactant particles will collide** with **sufficient energy** to overcome this lowered activation energy. Therefore, the **rate of reaction is faster** as <u>more effective collisions are occurring more frequently</u>.

NCEA 2014 Reaction Rates - (PART ONE)

Merit Question



Question 3a: The equation for the reaction between zinc granules (lumps), $Zn_{(s)}$, and sulfuric acid, $H_2SO_{4(aq)}$, is represented by:

$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

The graph below shows how the volume of hydrogen gas produced changes with time, when zinc is reacted with excess sulfuric acid at 20°C.

Explain the changes in the reaction rate during the periods **A**, **B** and **C**.

In your answer you should refer to collision theory.

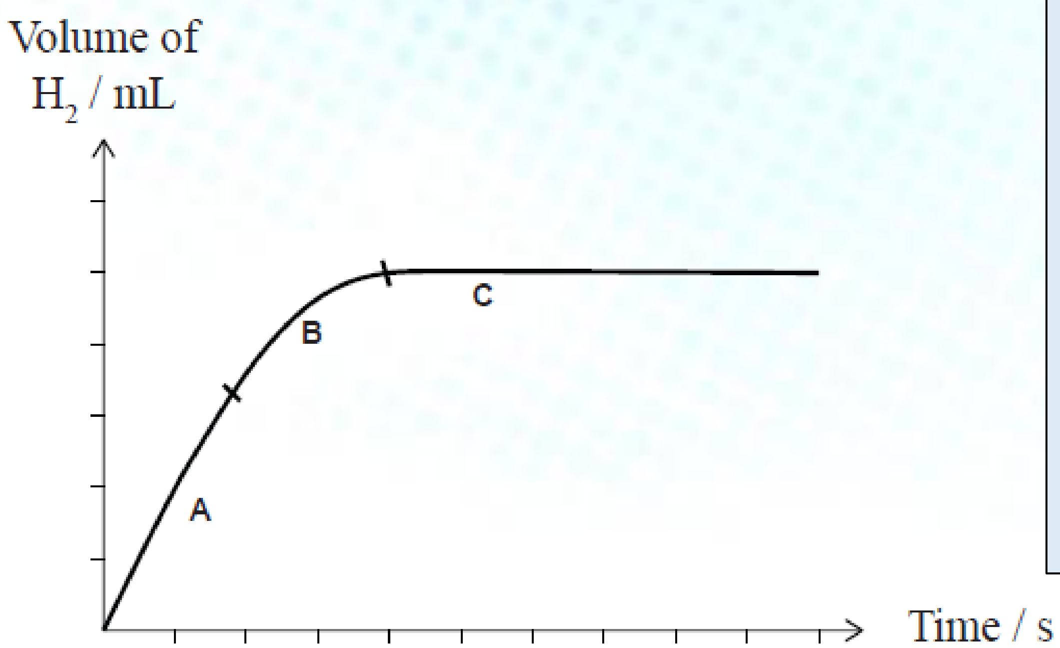
A: Reaction rate starts off high (steep slope) since there is a high concentration of reactants or many Zn particles at the start, so the frequency of collisions is highest, therefore the reaction rate is highest.

B: As zinc and sulfuric acid react to form zinc sulfate & hydrogen gas, the number of Zn particles decrease as it is used up. As the concentration of this reactant drops, the frequency of collisions will be less since there are fewer particles per unit volume to collide with the zinc. Thus the rate of reaction starts to slow down and the gradient of the graph becomes shallower.

C: Once all of the zinc reactant has been used up (sulfuric acid is in excess), then the reaction will stop since there is no more zinc to collide with the sulfuric acid and produce zinc sulfate & hydrogen gas. Thus there will be no more hydrogen gas formed; the maximum volume of hydrogen gas has been obtained and so the line remains horizontal.

NCEA 2014 Reaction Rates - (PART TWO)

Excellence Question



Question 3b: The rate of the reaction between zinc and sulfuric acid can be changed by the addition of small pieces of copper, Cu(*s*), as a catalyst.
Explain the role of the copper catalyst in the reaction between zinc and sulfuric acid.
In your answer you should refer to collision theory.

Answer: Increases the rate of reaction.
The catalyst, copper, provides an alternative pathway for the reaction between zinc and sulfuric acid, which involves lower activation energy. Therefore more particles will collide with sufficient energy to overcome the activation energy and result in successful collisions, so the rate of reaction will increase.

NCEA 2015 Reaction Rates

Achieved Question

Question: 1a: The 'elephant toothpaste' demonstration shows the decomposition of hydrogen peroxide, H_2O_2, into water and oxygen gas.

$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

This reaction can be observed by adding detergent to the hydrogen peroxide solution. As oxygen gas is produced, the detergent foams up, as seen in the photograph on the right. The time taken for the foam to reach the top of the measuring cylinder can be used to measure the rate of the reaction.

Three experiments were carried out to investigate factors that change the rate of the reaction.

(a) The decomposition reaction of hydrogen peroxide, H_2O_2, is very slow. By adding a small amount of powdered manganese dioxide, MnO_2, the rate of the reaction can be increased.

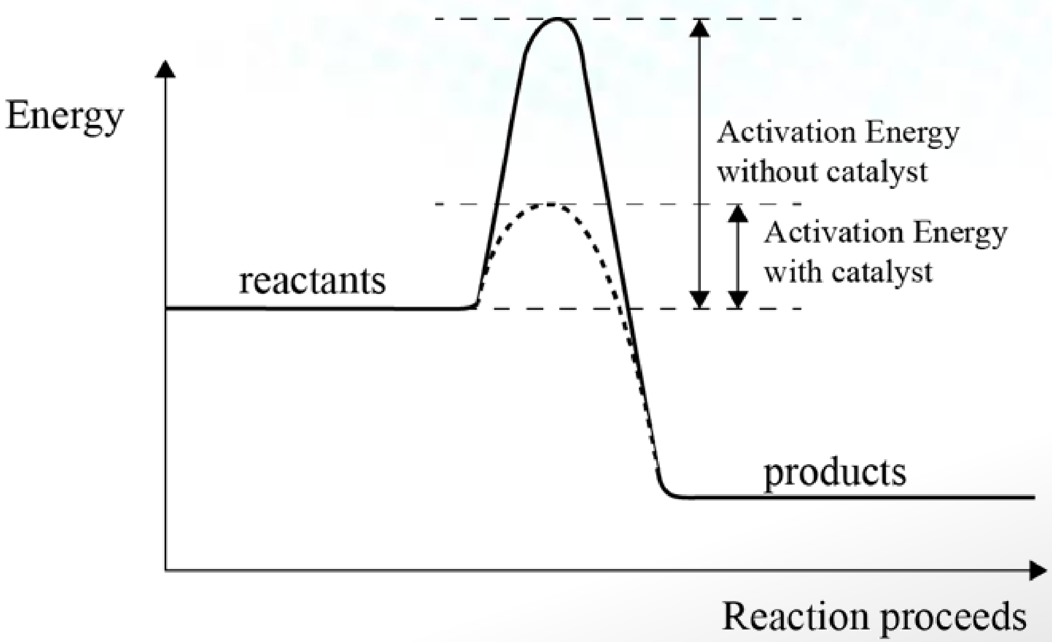
(i) Explain why only a small amount of manganese dioxide is needed to increase the rate of the reaction.

Answer: 1a: The added MnO_2 acts as a catalyst and is added in small amounts because it is not used up in the reaction, so can be reused over and over again in the chemical reaction.

NCEA 2015 Reaction Rates

Excellence Question

(ii) The diagram below shows the energy diagram for the decomposition reaction without manganese dioxide.
Label this diagram and use it to help you explain how the addition of manganese dioxide speeds up the rate of the reaction.



Answer: The MnO_2 speeds up the rate of reaction by lowering the activation energy required. It does this by providing an alternative pathway for the reaction to occur. Once the activation energy barrier is lowered, more reactants will have sufficient energy to overcome the activation energy, resulting in an increase in the rate of reaction

NCEA 2015 Reaction Rates

Excellence Question

Experiment	Concentration of H_2O_2	Temperature °C	Presence of small amount of MnO_2
1	20%	20	yes
2	20%	30	yes
3	30%	20	yes

Question 1(b): Compare Experiment 2 with Experiment 1.
In your answer, you should:
- identify the factor being changed, and the effect this will have on the rate of reaction
- explain the effect on the rate of reaction by referring to the collision of particles and activation energy, where appropriate.

Answer: In Experiment 2, the only change is an increase in temperature. An increase in temperature means an increase in the rate of reaction.
Increased temperature increases the speed of movement of the particles, and thus increases the frequency of collisions.
Increased temperature also increases the kinetic energy of the particles, so the collisions that occur are more likely to be successful (more likely to have sufficient activation energy). So the rate of reaction is increased.

NCEA 2015 Reaction Rates

Excellence Question

Experiment	Concentration of H_2O_2	Temperature °C	Presence of small amount of MnO_2
1	20%	20	yes
2	20%	30	yes
3	30%	20	yes

Question 1 (c): Compare Experiment 3 with Experiment 1.
In your answer, you should:
- identify the factor being changed, and the effect this will have on the rate of reaction
- explain the effect on the rate of reaction by referring to the collision of particles and activation energy, where appropriate.

Answer: In Experiment 3, the concentration of hydrogen peroxide has been increased. This will increase the rate of reaction because there are more hydrogen peroxide molecules per unit volume. This means there will be more frequent collisions in a given time due to having more reactant particles available to collide. This will increase the rate of decomposition of the hydrogen peroxide.

NCEA 2016 Reaction Rates

Merit Question

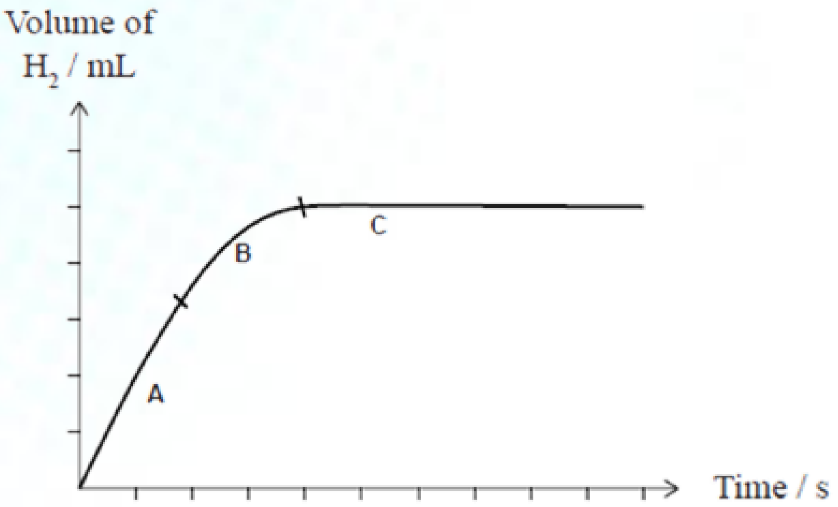
Question 1 (a): Cleaned magnesium ribbon, Mg(s), reacts with a solution of hydrochloric acid, HCl(aq). The reaction is represented by the equation:

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

The reaction is monitored by measuring the volume of hydrogen gas produced over a given period of time. This is shown in the graph below.

Explain the changes in the rate of reaction between magnesium, Mg(s), and hydrochloric acid, HCl(aq), in terms of collision theory.

Refer to parts A, B, and C of the graph in your answer.



Answer: Initially the rate of the reaction is fast, as shown by the steepness of Part A in the graph. This is because the concentrations of the reactants, Mg and HCl, are at their highest, resulting in a high frequency of effective collisions, so the rate of reaction is highest there.

As time proceeds, the reactants are forming the products $MgCl_2$ and H_2 gas, so the concentration of the reactants decreases, resulting in fewer reactant particles available to react. This results in a decrease in the frequency of effective collisions between reacting particles and a decrease in the production of H_2 gas, as shown in Part B of the graph where the gradient of the graph becomes less steep. Consequently, the rate of reaction decreases.

Once all of one reactant (or both) has been used up, the reaction will stop, so no H_2 gas is produced, as shown on the graph in Part C as a horizontal line.

NCEA 2016 Reaction Rates

Excellence Question

Question 1 (b): Compare and contrast the reactions of 0.5 g of **magnesium ribbon**, Mg(s), with 50.0 mL of 0.100 mol L–1 hydrochloric acid, HCl(aq), and 0.5 g of **magnesium powder**, Mg(s), with 50.0 mL of 0.100 mol L–1 hydrochloric acid, HCl(aq). Refer to collision theory and rates of reaction in your answer.

Answer: In the reaction of hydrochloric acid with Mg ribbon and Mg powder, both form the same products, magnesium chloride and hydrogen gas.

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

However, since Mg powder has a larger surface area than Mg ribbon, the powder will have more Mg particles immediately available to collide, there will be more effective collisions per second and more H2 gas will be produced initially, resulting in a faster rate of reaction.

Mg ribbon will take longer to react because fewer particles are immediately available to collide, so will have a slower rate of reaction.

Both reactions will eventually produce the same volume of hydrogen gas as the same amounts of each reactant are used.

NCEA 2016 Reaction Rates

Excellence Question

Question 1(c): The decomposition reaction of hydrogen peroxide solution, $H_2O_{2(aq)}$, is a slow reaction. This reaction is represented by the equation:

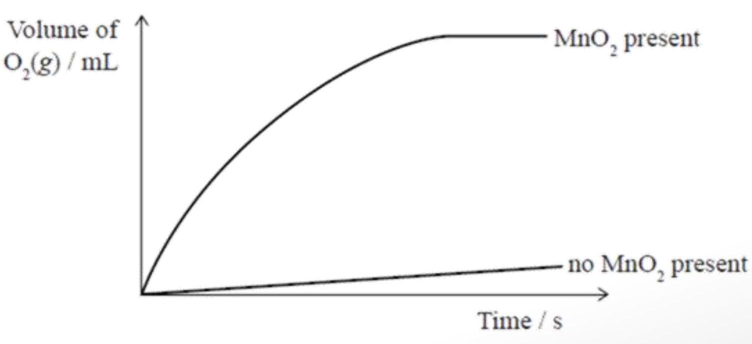
$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

The rate of the decomposition reaction can be changed by adding a small amount of manganese dioxide, $MnO_{2(s)}$. The graph below shows the volume of oxygen gas formed in the reaction with and without manganese dioxide, $MnO_{2(s)}$.

(i) State the role of manganese dioxide, $MnO_{2(s)}$, in this reaction.

(ii) Elaborate on how manganese dioxide, $MnO_{2(s)}$, changes the rate of the decomposition reaction of the hydrogen peroxide, $H_2O_{2(aq)}$.

In your answer you should refer to the activation energy and collision theory.



Answer: the MnO_2 is a catalyst.

MnO_2 provides an alternative pathway with lower activation energy for the decomposition of H_2O_2.

Therefore, (*) more reacting particles will collide with sufficient energy, resulting in a higher frequency of successful collisions; resulting in an increase in the rate of reaction.

(Only a small amount of MnO_2 is required because catalysts are not used up in this reaction.)

NCEA 2017 Reaction Rates

Excellence Question

Question 2a: The addition of a small amount of iron to a mixture of nitrogen and hydrogen gases helps to speed up the production of ammonia gas.

$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

(a) Identify and explain the role of iron in this reaction.

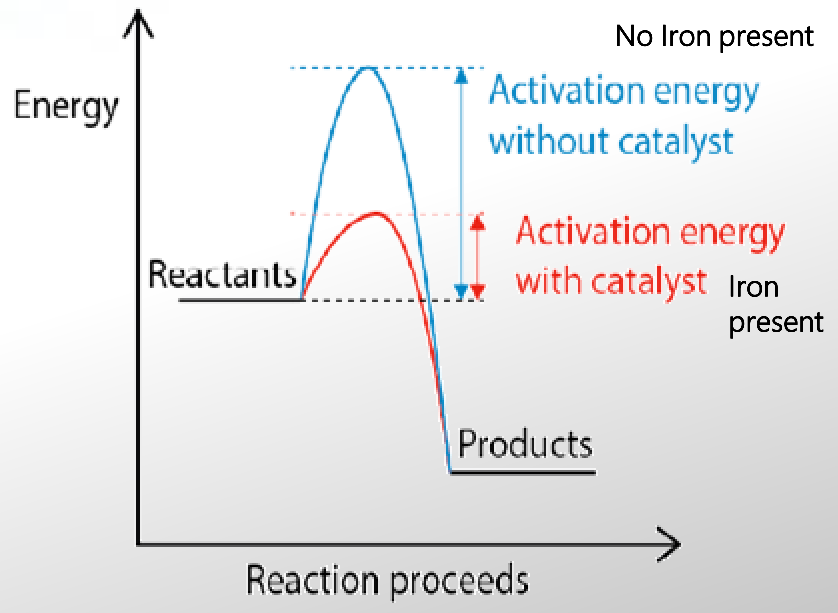
In your answer, you should refer to activation energy and collision theory.

You may include a diagram or diagrams in your answer.

Answer: The iron is a catalyst.

Iron provides an alternative pathway with lower activation energy for the production of ammonia gas. Therefore, (*) more reacting particles will collide with sufficient energy, resulting in a higher frequency of successful collisions; resulting in an increase in the rate of reaction.

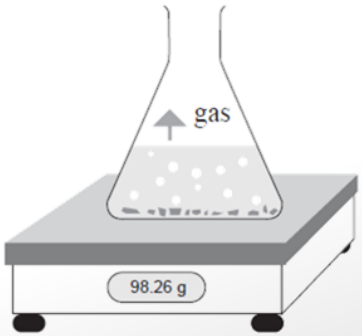
(Only a small amount of Iron is required because catalysts are not used up in this reaction.)

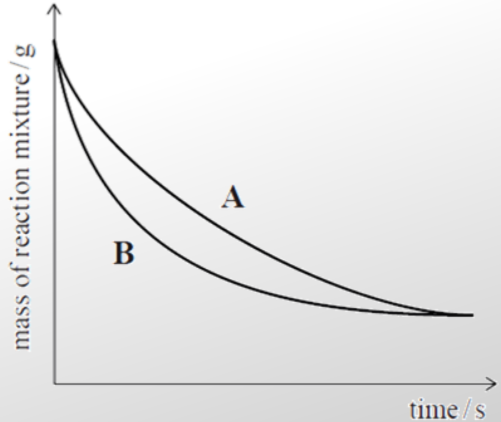


NCEA 2017 Reaction Rates

Excellence Question

Question 3a: Consider the reaction between calcium carbonate powder, $CaCO_{3(s)}$, and a solution of hydrochloric acid, $HCl_{(aq)}$.
As the reaction proceeds, the mass of the reaction mixture decreases as carbon dioxide gas, $CO_{2(g)}$, escapes.
This is represented on the graph below.
Line A represents the reaction occurring at 20°C and line B represents the reaction occurring at 40°C.
Compare and contrast the reaction between calcium carbonate powder, $CaCO_{3(s)}$, and a solution of hydrochloric acid, $HCl_{(aq)}$ at two temperatures: 20°C and 40°C, assuming all other conditions are kept the same.
Your answer should refer to collision theory and rates of reaction.

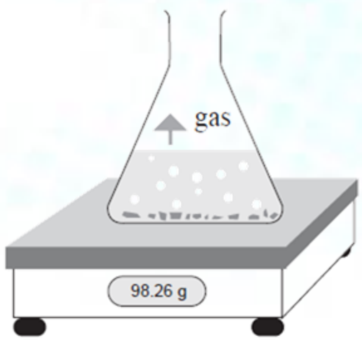


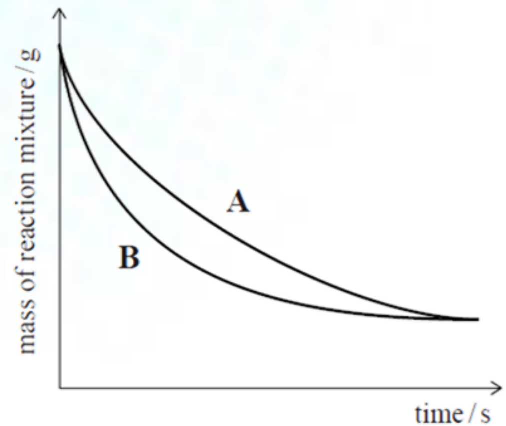


NCEA 2017 Reaction Rates

Excellence Question

Question 3a: Compare and contrast the reaction between calcium carbonate powder, $CaCO_{3(s)}$, and a solution of hydrochloric acid, $HCl_{(aq)}$ at two temperatures: 20°C and 40°C, assuming all other conditions are kept the same.





The **increased temperature** means an increase in the rate of the reaction because the **kinetic energy of the particles has increased**. This means the **particles move faster**, **increasing the frequency of collisions / more collisions per second**, resulting in the CO_2 gas being lost in a shorter period of time at 40°C than 20°C. In addition, the collisions are more likely to be **successful / effective** because the **average kinetic energy of the particles has increased**, so a **greater proportion of particles have enough / sufficient energy to overcome the activation energy**. This causes the rate of reaction to increase. Overall, the same mass is lost in both reactions.

NCEA 2018 Reaction Rates

Excellence Question

Question 1a: In the iodine clock reaction, a solution of hydrogen peroxide is mixed with a solution containing potassium iodide, starch, and sodium thiosulfate.
After some time, the colourless mixture suddenly turns dark blue.
The table shows the time taken for the reaction performed at different temperatures. The concentration of all reactants was kept constant.
Explain the effect of changing the temperature on the rate of reaction.
Refer to collision theory and activation energy in your answer.

Temperature / °C	Time for dark blue colour to appear / s
20	15
30	9
40	4

Increased temperature **increases the kinetic energy** of the particles, causing them to **move faster** and **collide more frequently**. The collisions that occur are more likely to have **sufficient energy to overcome the activation energy barrier** / more particles have sufficient energy to overcome the activation energy barrier. So the rate of reaction is increased as there will be **greater frequency / more successful collisions per second / unit time**.

NCEA 2018 Reaction Rates

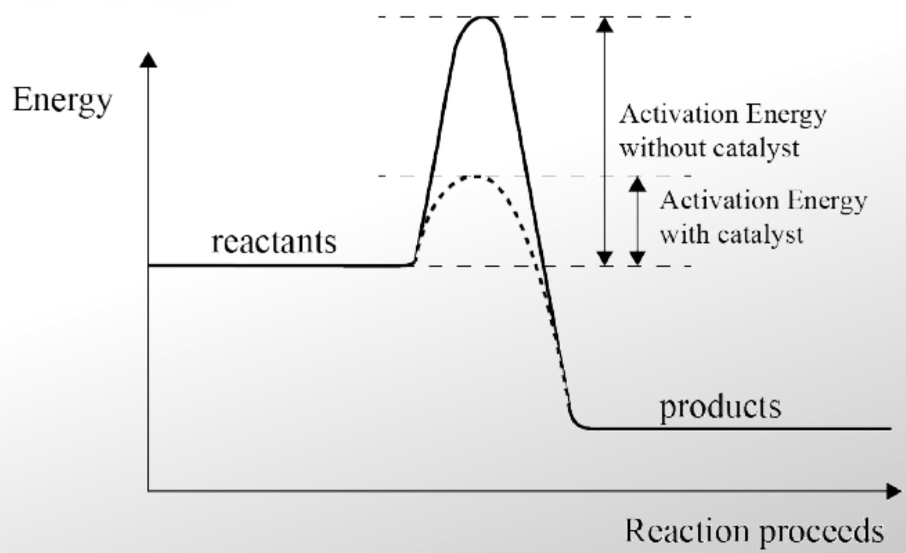
Excellence Question

Question 1b: Consider the following observations in another experiment using hydrogen peroxide:

- When hydrogen peroxide is mixed with solution **X**, which contains universal indicator, the colour changes from blue to green to yellow to orange-red **over a time of one hour**.
- If a crystal of **ammonium molybdate** is added to solution **X** before the hydrogen peroxide is added, the same colour changes will be seen in **three to four minutes**.

(i) Identify and explain the role of ammonium molybdate.
Use a diagram and refer to activation energy in your answer.

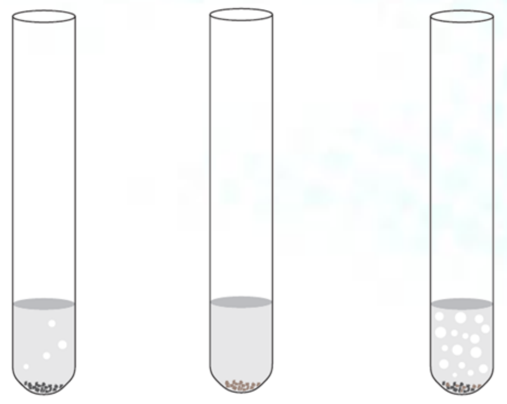
The ammonium molybdate is a **catalyst**, which speeds up the rate of reaction by **providing an alternative pathway** for the reaction to occur with a **lower activation energy** barrier to be overcome.
Now **more reactants will have sufficient energy to overcome the activation energy**, resulting in an increase in the rate of reaction.



NCEA 2019 Reaction Rates

Excellence Question

Question 1a: The same volume and concentration of hydrochloric acid, $HCl_{(aq)}$, was added to each of three test tubes. Metal samples were added, according to the table and diagram below.



Test tube	Contents	Observations
1	20 mL hydrochloric acid, $HCl_{(aq)}$, and 1 g zinc granules, $Zn_{(s)}$	Slow rate of bubbles
2	20 mL hydrochloric acid, $HCl_{(aq)}$, and 1 g copper granules, $Cu_{(s)}$	No observable reaction
3	20 mL hydrochloric acid, $HCl_{(aq)}$, 1 g zinc granules, $Zn_{(s)}$, and 1 g copper granules, $Cu_{(s)}$	Fast rate of bubbles

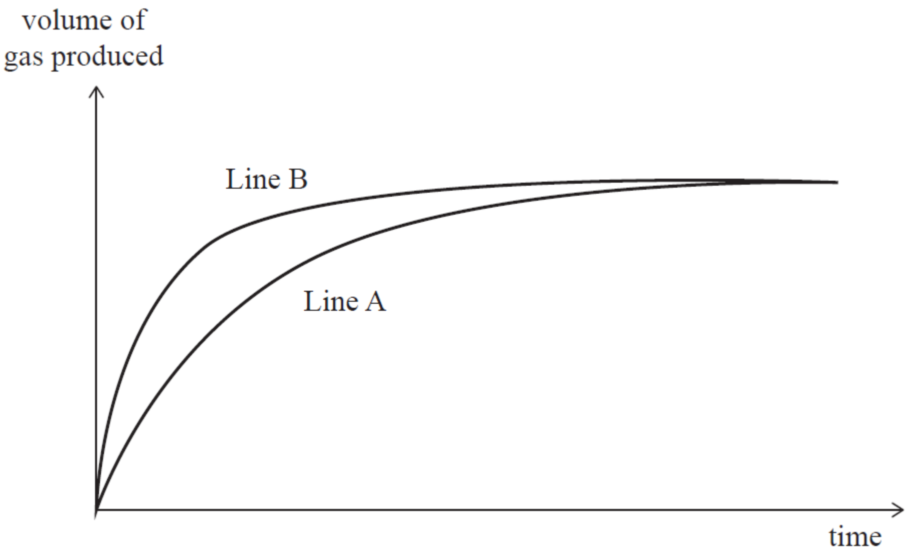
(i) Identify the role of the copper granules, $Cu_{(s)}$, in test tube 3.
(ii) Explain the role of copper, $Cu_{(s)}$, in this reaction.
You should refer to activation energy and collision theory in your answer.

Cu is a catalyst.
Cu provides an **alternative pathway** with **lower activation energy** for the reaction. Therefore, **more reacting particles will collide with sufficient (kinetic) energy above activation energy**, (resulting in a higher frequency of successful collision) resulting in an increase in the rate of reaction.

NCEA 2019 Reaction Rates

Merit Question

Question 1b: In a second investigation, two 20 mL samples of 0.2 mol L^{-1} sulfuric acid, $H_2SO_{4(aq)}$, were placed in separate conical flasks. One of the flasks was placed in a water bath at 40°C and the other was placed in a water bath at 20°C. To each conical flask, 5.0 g of zinc granules, $Zn_{(s)}$, were added. The gas produced was collected and measured over time and the following graph was produced.



(i) Identify which line on the graph represents the reaction at 40°C, and explain why the two lines still finish in the same position.

Line B.

The lines finish in the same position as there are **(1)** the same amounts of reactants, thereby producing **(2)** the same amount of products. **(3)** The only difference is the rate at which the products are produced, i.e. reaction at higher temperature produces products at a faster rate.

NCEA 2019 Reaction Rates

Excellence Question

Question 1b: In a second investigation, two 20 mL samples of 0.2 mol L^{-1} sulfuric acid, $H_2SO_{4(aq)}$, were placed in separate conical flasks. One of the flasks was placed in a water bath at 40°C and the other was placed in a water bath at 20°C. To each conical flask, 5.0 g of zinc granules, $Zn_{(s)}$, were added. The gas produced was collected and measured over time
(ii) Elaborate on the effect of increasing temperature on the rate of reaction.
Refer to collision theory and activation energy in your answer.

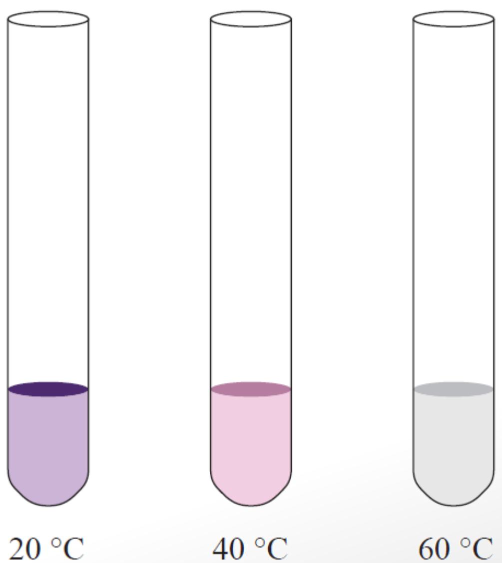
An increased temperature means an increase in the rate of the reaction because the **kinetic energy (E_K) of the particles has increased**. This means the particles **move faster**, **increasing the frequency of (successful / effective) collisions / more collisions per second**.
In addition, a **greater percentage / proportion of collisions are likely to be successful because more particles have enough kinetic energy to overcome the activation energy**. This causes the rate of reaction to increase / more gas to be produced.

NCEA 2020 Reaction Rates

Excellence Question

Question 2a: When oxalic acid solution, $H_2C_2O_{4(aq)}$, reacts with purple acidified potassium permanganate solution, H^+ / $MnO_4^-{}_{(aq)}$, the purple colour fades and the reaction is complete when the mixture turns colourless.
The picture shows the colour changes after 45 seconds for three different temperatures.
(a) Explain how the rate of reaction for this experiment is affected by the temperature at which the reaction occurs. In your answer refer to the information in the picture, collision theory, and activation energy.

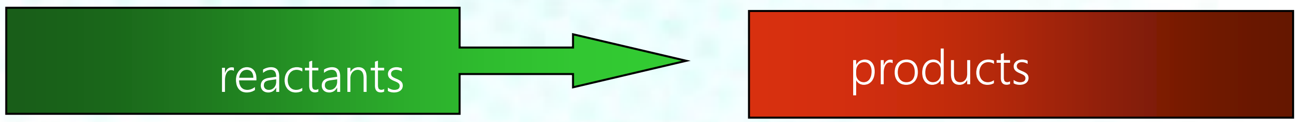
Colour change after 45 seconds



The warmer the solution, the more complete the reaction / the greater the colour change after 45 seconds.
An increased temperature means an increase in the rate of the reaction because the kinetic energy of the particles has increased. The particles are moving faster, increasing the frequency of collisions.
More of these collisions will have enough energy to be successful collisions, which means their energy is greater than the activation energy. Overall, this leads to a greater frequency of effective collisions, so the rate of reaction is increased, and the solution turns colourless more quickly.

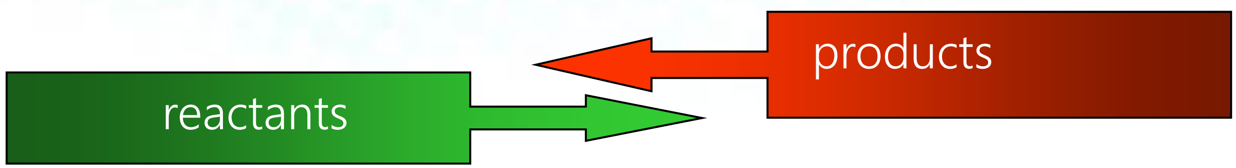
Equilibrium

Some reactions go to completion



Reaction stops when one of the reactants is used up

Other reactions are reversible



Products are also forming reactants. Reaction continues

Equilibrium is a state of dynamic balance

where the rates of formation of product = equals the rate of formation of reactants

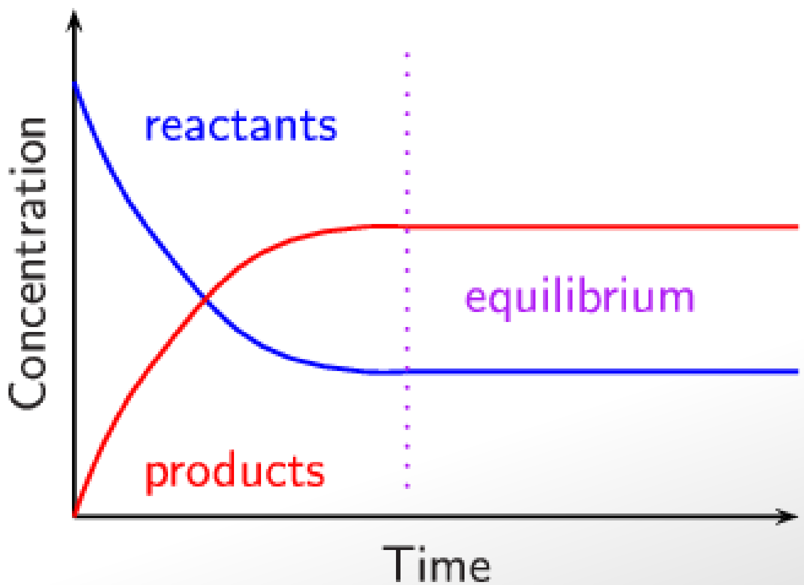
At equilibrium the concentrations of reactants and products are constant.

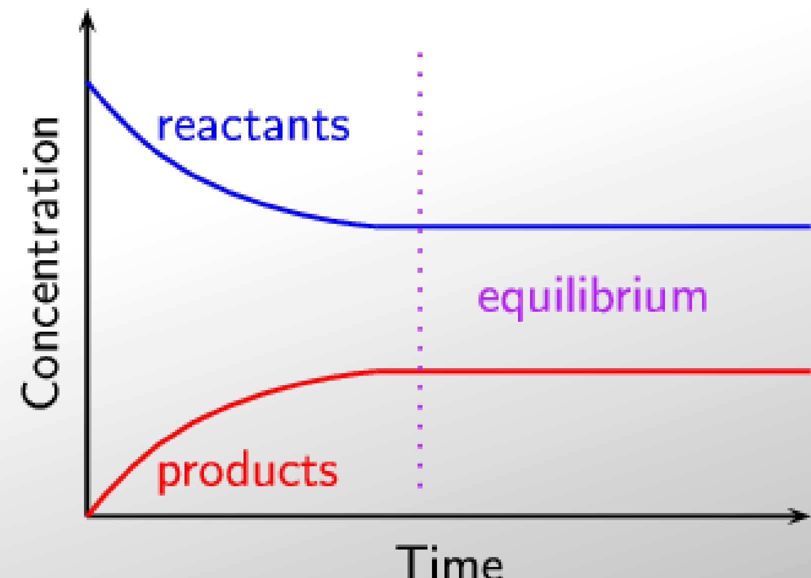
However, both the forward and reverse reactions are continuing

Equilibrium

When a reaction has reached equilibrium then the **proportion of reactants is fixed in relation to the proportion of products**. Reactants particles are still colliding to form products but the same number of products are colliding (or breaking apart) to form reactants.
The proportion of reactants to products depends upon the reaction and the environmental conditions of a reaction such as temperature, pressure and concentration.

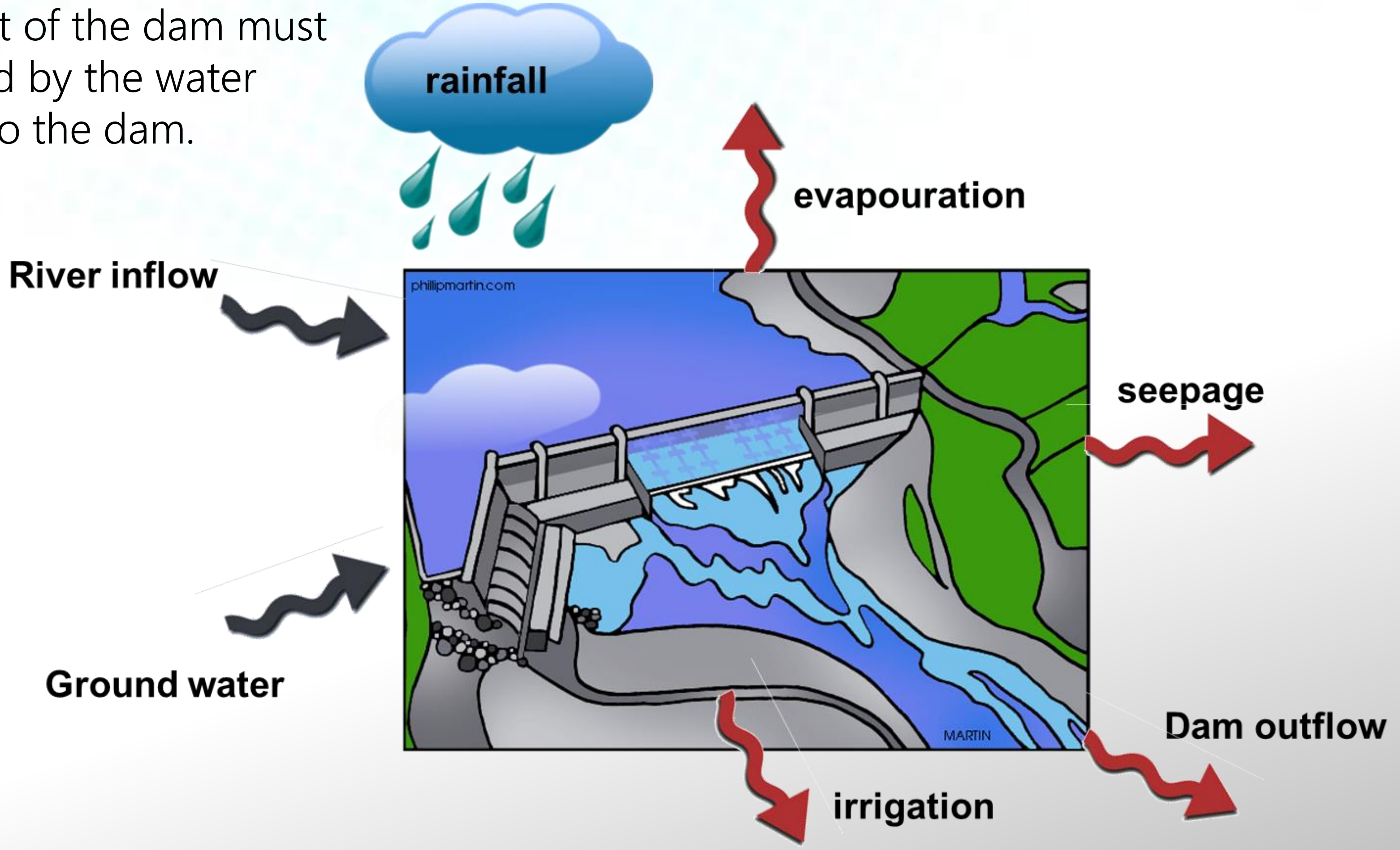
On the left hand side, the proportion of products will be higher than the reactants and on the right hand side, the proportion of reactants will be higher than products.





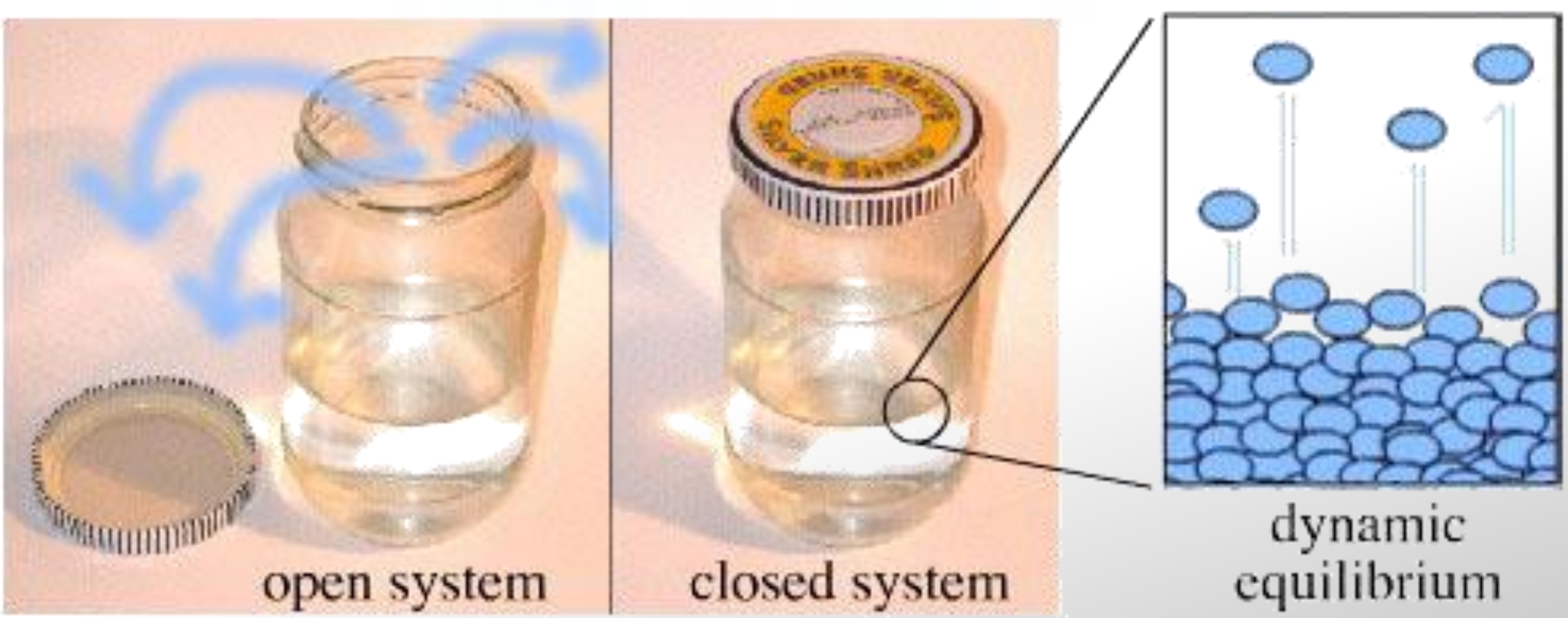
Equilibrium – Water Dam case study

In order for the level of the water to stay constant in the dam there must be a state of equilibrium. The total water moving out of the dam must be replaced by the water moving into the dam.



Equilibrium – Evaporation case study

A dynamic equilibrium must occur in a closed system where all reactants and products are retained in an area where particles can collide with each other.
The example below shows a system where liquid water is evaporating into a gas.
In an **open system,** the gas will escape and gradually the water level will decrease.
In a **closed system**, where the lid prevents the gas escaping, the proportion of liquid to gas will become fixed at a **dynamic equilibrium**. Liquid will evaporate into gas at the same rate that gas condenses into a liquid.



Equilibrium Constant: K_C

An equilibrium equation can be written as an expression (K_c) in which concentrations of products and reactants can be placed in to give us a value. The value will indication the proportion of reactants to products in any given reaction.

Note: only reactants and products in gas state or aqueous can be placed into an equilibrium expression. **Do not place solids or liquids** into the expression.

Given $aA + bB \rightleftharpoons cC + dD$

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

$$K_c = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

[ ] = concentration in $molL^{-1}$ at equilibrium

e.g. $$K_c = \frac{[NH_3]^2}{[N_2] \times [H_2]^3}$$

Products are divided by reactants and the number of moles in the equation is written to the power of each reactant and product.

Calculating the Equilibrium Constant: K_C

To calculate the equilibrium expression (K_C) place the given concentrations in mol L^{-1} into the correct position in the expression. (remember do not put solids in as they have a fixed density and their concentration does not change)

e.g. $N_{2(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$ at room temperature

If the values given for concentration are:

$[NH_3] = 0.325 mol L^{-1}$, $[N_2] = 0.625 mol L^{-1}$ and $[H_2] = 2.45 mol L^{-1}$

$$K_c = \frac{[0.325]^2}{[0.625] \times [2.45]^3}$$

Step 1. place the values in the expression – make sure products are on top and reactants below

Step 2. expand out equation

$$K_c = \frac{0.106}{0.625 \times 14.7}$$

Step 3. divide equation

$$K_c = \frac{0.106}{9.19} \qquad K_c = 0.0115$$

The small K_c value indicates that there is a small amount of products compared to reactants.

Analysing the K_c value

If the K_c is less than 1 then there will be more (higher concentration) reactants than products at equilibrium. If the K_c is greater than 1 then there will be more products than reactants at equilibrium.

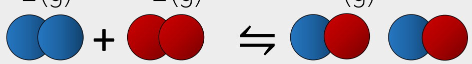
A K_c value of 1 means reactants = products at equilibrium.

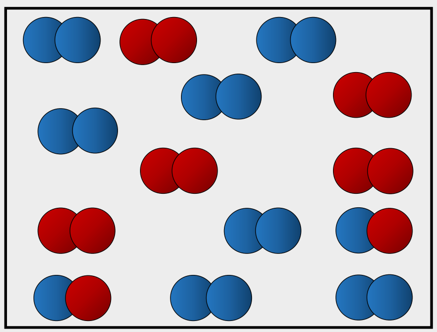
1

More reactants

K_c is less than 1

$N_{2(g)} + O_{2(g)} \leftrightharpoons 2NO_{(g)}$





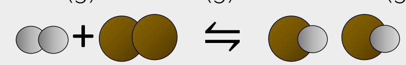
$$Kc = \frac{[NO]^2}{[N_2]\ [O_2]}$$

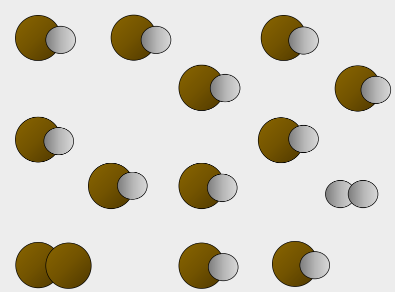
= small number

More products

K_c is greater than 1

$H_{2(g)} + Br_{2(g)} \leftrightharpoons 2HBr_{(g)}$





$$Kc = \frac{[HBr]^2}{[H_2]\ [Br_2]}$$

= large number

Some questions will ask you to calculate a value (Q) using the equilibrium constant and provided concentrations [ ]

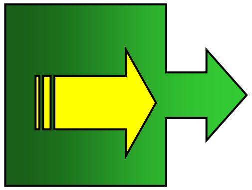
You then need to compare this value to K_c (at a particular temperature). Use the scale above to compare positions of Q to K_c to see if that value indicates the reaction is at equilibrium (they are the same) or more reactants/products

Changes in Equilibrium

A system stays in equilibrium unless a change is made

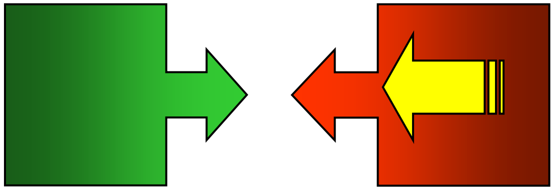
A change made to a system in equilibrium will either

Increase the rate of the **forward reaction**



OR

Increase the rate of the **reverse reaction**



Eventually equilibrium is re-established, and the rate of forward reaction again equals rate of reverse reaction

A temperature change will permanently shift the equilibrium

Le Chatelier's Principle

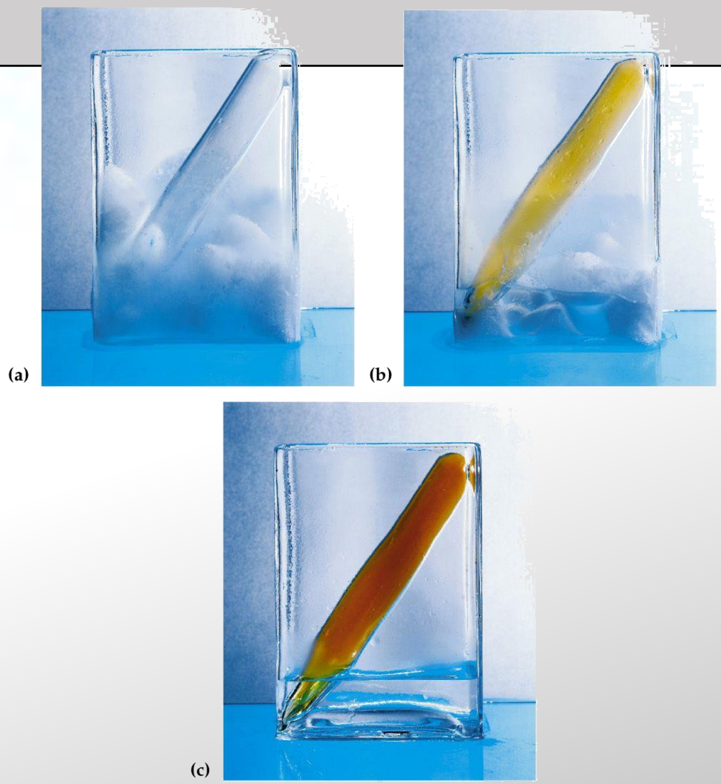
When a change is applied to a system at equilibrium, the system responds so that the effects of the change are minimised

Change in conditions	Direction of change in equilibrium position
Concentration - increase products	In the reverse direction
- decrease products	In the forward direction
- increase reactants	In the forward direction
- decrease reactants	In the reverse direction
Pressure Increase	In the direction with the least no. of moles of gas
Decrease	In the direction with the greater no. of moles of gas
Temperature Increase	In the direction of the endothermic reaction
Decrease	In the direction of the exothermic reaction
Catalyst added	No change in equilibrium position or in K_c Equilibrium is reached more quickly (ie reaction rate changes)

Observable changes

At a **particular temperature the equilibrium constant always has the same value** for a set reaction. It is unaffected by a change in concentration, pressure or whether or not you are using a catalyst.
If changes are made to the pressure or concentration of reactants or products, which changes the equilibrium position, where the ratio of reactants to products is changed, then **the equilibrium responds until the equilibrium ratio is re-established at the K_c value.**

Questions asked about changes to a system ask for **observations immediately after a change has been made**.
Remember, that unless a temperature change is made, the eventual observation of the system will be that it is the same as prior to the change due to the equilibrium being re-established.



$N_2O_4(g)$ + heat $\rightleftharpoons$ 2 NO_2
Colorless Brown

Changes in Equilibrium – Increasing Concentration

GZ Science Resources

If the concentration of either the reactants or the products increases which **temporarily moves the position of the equilibrium** then the **rate of reaction will increase in one direction** only to partially undo this change. The original equilibrium position is then re-established eventually.

For example: if the concentration of the reactants is increased then the rate of reaction changing reactants in products will increase (due to collision theory) favouring the forward reaction. The effect of this will change more of the reactant into product – lowering the concentration of reactant and increasing the concentration of product.

Brown gas $\rightleftharpoons$ colourless gas

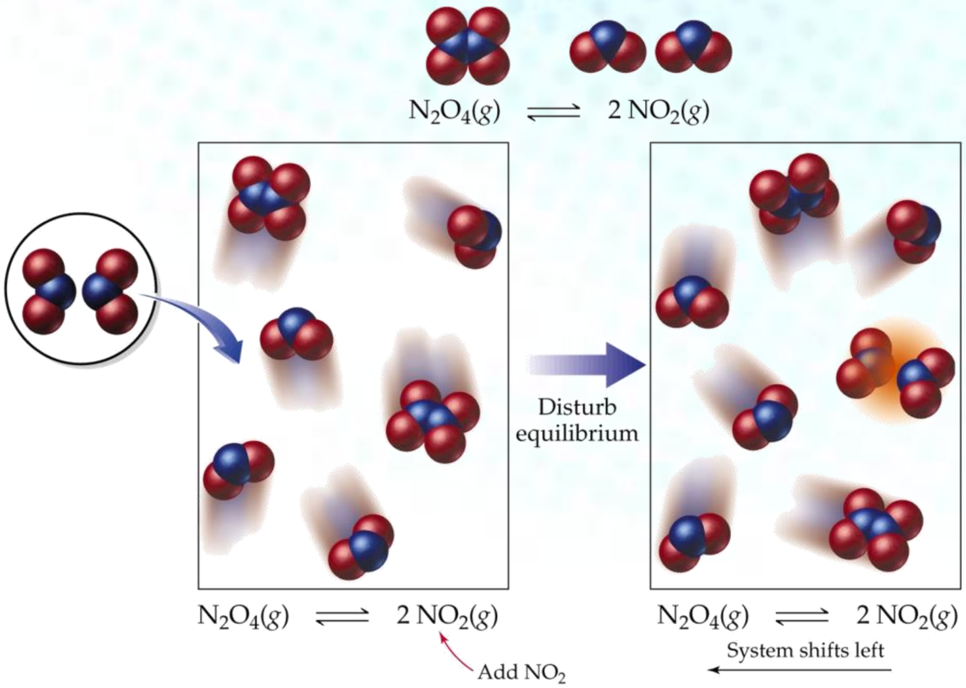
Add more NO_2 to increase concentration **colour gets darker**

Forward reaction increases which reduces concentration

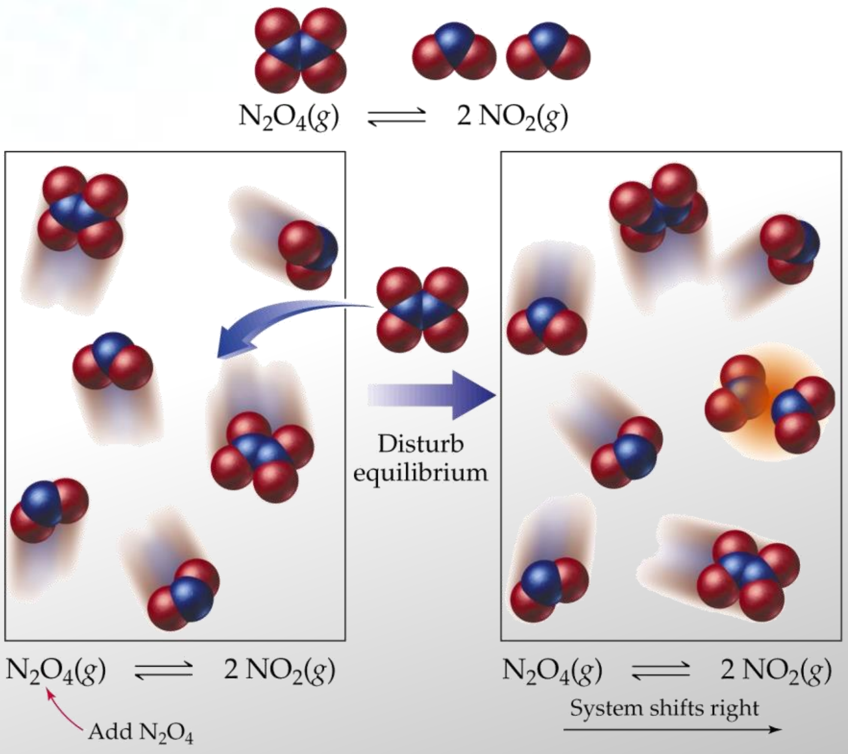
More N_2O_4 is produced **colour gets lighter again**

Note: observations are compared to the situation straight after a change was made **not** in comparison to the original system

Changes in Equilibrium – Increasing Concentration



Adding more reactants shifts the equilibrium position left (more reactants to products).



The system will then increase the forward reaction to make more products – therefore moving the equilibrium position back to it's original position

Changes in Equilibrium – Decreasing Concentration

If the concentration of either the reactants or the products decreases (by being taken away) that **temporarily moves the position of the equilibrium** then the **rate of reaction will increase in one direction** only to partially undo this change. The original equilibrium position is then re-established eventually.

<u>For example:</u> if the concentration of the product is decreased then the rate of reaction changing reactants in products (due to collision theory) will increase **favouring the forward reaction**. The effect of this will change more of the reactant into product – to replace the lost product, lowering the concentration of reactant and increasing the concentration of product.

Nitrogen		Hydrogen		Ammonia
$N_{2(g)}$	+	$3H_{2(g)}$	$\rightleftharpoons$	$2NH_{3(g)}$

Remove NH_3 to decrease concentration of product

Forward reaction increases to replace removed NH_3

More NH_3 is produced **N_2 and H_2 reduce in concentration**

GZ Science Resources

Changes in Equilibrium – Increasing Pressure

With gases only

Le Chatelier's Principle states that either a forward or reverse reaction will be favoured if it lessens the impact of a change to the system. Only when a reaction has a different number of moles on either the reactants or products side and the pressure is changed then the reaction will shift to lessen the impact.

For example: In the reaction below the reactants have 2 moles compared to the product with 1 mole. If the pressure is increased the side with the least moles, the product in this example, will be favoured. The immediate effect straight after the pressure increase will be the gas mixture becoming lighter.

Brown gas $2NO_{2(g)}$ $\rightleftharpoons$ colourless gas $N_2O_{4(g)}$

Note: if both sides have the same number of moles then there will be no change.

Increase pressure to the system → Forward reaction increases favour the least amount of moles → More N_2O_4 is produced **The brown colour becomes lighter**

Changes in Equilibrium – Decreasing Pressure

With gases only

When the pressure is decreased in a system that has unequal number of moles in the reactant and products then the reaction direction that produces more moles is favoured.

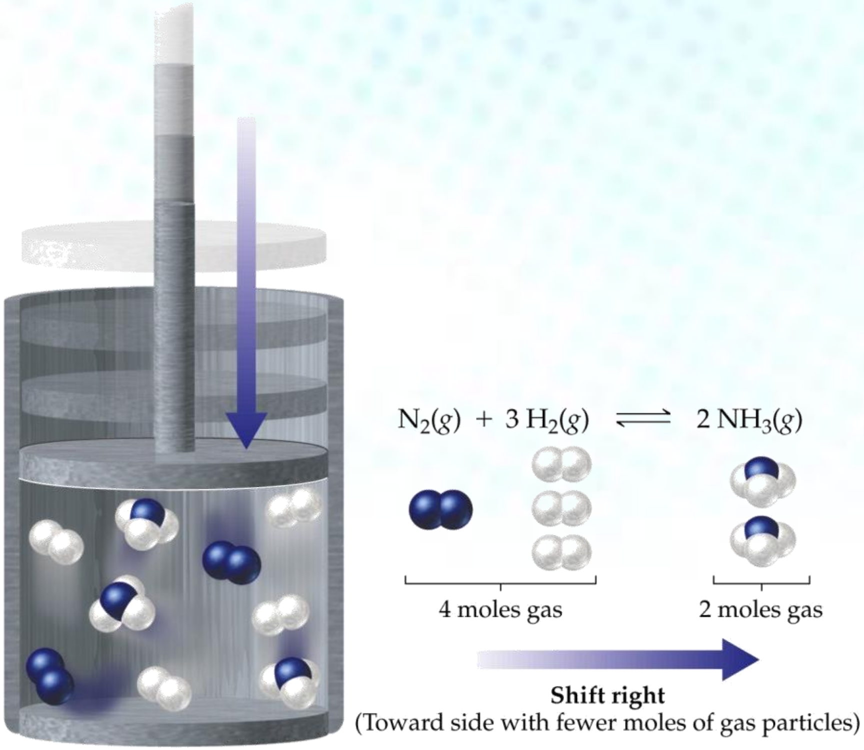
For example: In this reaction the NO_2 has 2 moles compared to 1 mole of the N_2O_4 so the reaction rate from the product to the reactant is increased.

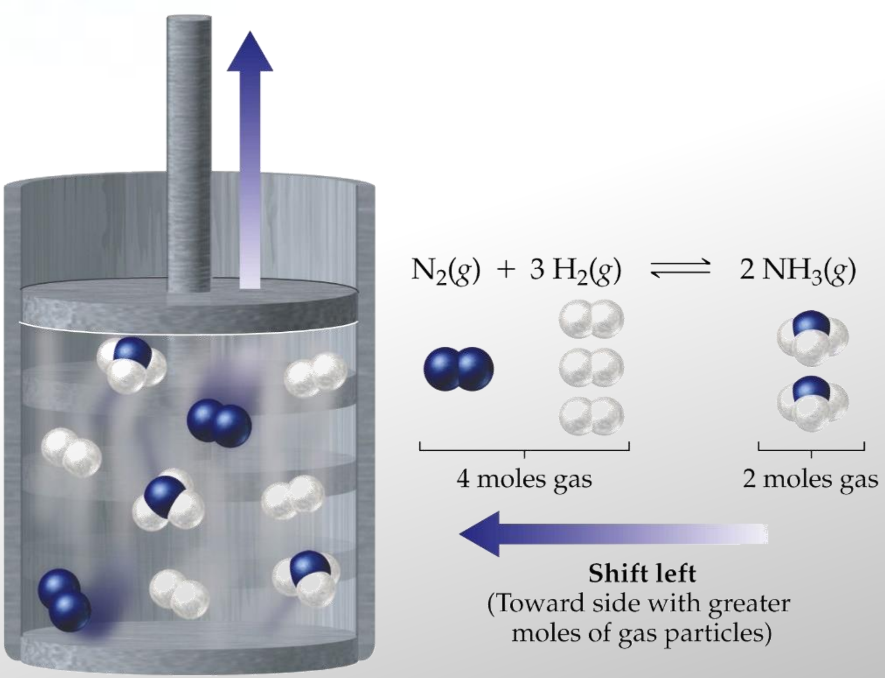
Brown gas $2NO_{2(g)}$ $\rightleftharpoons$ colourless gas $N_2O_{4(g)}$

Decrease pressure to the system → Reverse reaction increases favour the most amount of moles → More NO_2 is produced **The brown colour becomes darker**

Note: a change in pressure is caused by changing volume. The effective concentration is changed and creates the same equilibrium response.

Changes in Equilibrium – Summary pressure changes



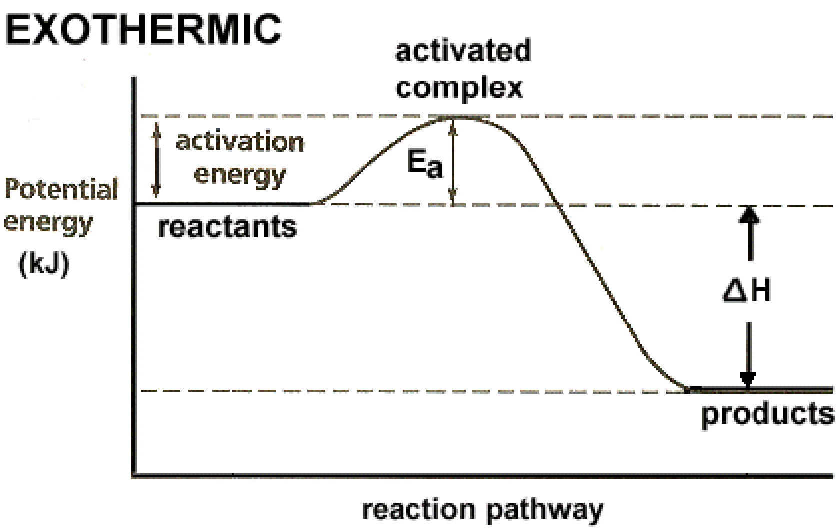


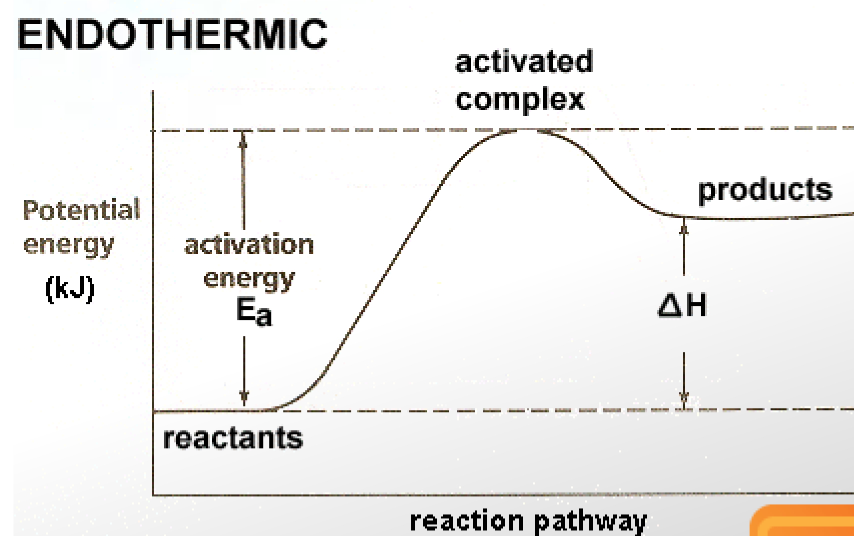
Endothermic and Exothermic reactions

A reaction that is exothermic in one direction will be endothermic in the opposite direction due to the Law of conservation of energy

An endothermic reaction will absorb heat energy from the surrounding area as the products contain more enthalpy than the reactants. The temperature of the closed system will decrease.

An exothermic reaction will release heat energy into the surrounding area as the products contain less enthalpy than the reactants. The temperature of the closed system will increase.





There is a difference in activation energy between the endothermic and exothermic direction, There will uneven proportions of energy involved in activation energy and substance creation if the equilibrium is shifted. This changes the K_c value permanently.

extra info

Changes in Equilibrium – Increasing Temperature

A change in temperature will permanently change the equilibrium , and K_c, of the system

When a reaction is exothermic, it releases heat energy in the forward reaction when products are formed. An increase in temperature will favour the reverse reaction which is endothermic and more reactants will be made. (An endothermic reaction will favour the forward reaction if temperature is increased)

sulfur		oxygen		Sulfur dioxide	
$S_{(g)}$	+	$O_{2(g)}$	$\rightleftharpoons$	$O_{2(g)}$	-297kJmol^{-1}

Note: negative means exothermic

Increase temperature to the system

Reverse reaction increases to favour the side that absorbs heat energy

More S and O_2 is produced
The temperature of the system reduces

Changes in Equilibrium – Decreasing Temperature

A change in temperature will permanently change the equilibrium , and K_c, of the system

When a reaction is exothermic it releases heat energy in the forward reaction when products are formed. An decrease in temperature will favour the forward reaction so heat energy is released to increase the temperature of the system and more products will be made.

sulfur		oxygen		Sulfur dioxide	
$S_{(g)}$	+	$O_{2(g)}$	$\rightleftharpoons$	$SO_{2(g)}$	$-297 kJmol^{-1}$

Decrease temperature to the system

forward reaction increases to favour the side that releases heat energy

More SO_2 is produced
The temperature of the system increases

Changes in Equilibrium Constant K_C – Changing Temperature

A decrease in the value of K_C means the concentration of the product is reduced. If the K_C value reduces when the temperature increases then the reverse reaction to form reactants must be endothermic, hence the forward reaction is exothermic.

K_C **decreases** when temperature **increases**

↓

Exothermic reaction

↑

K_C **increases** when temperature **decreases**

K_C **increases** when temperature **increases**

↓

Endothermic reaction

↑

K_C **decreases** when temperature **decreases**

Changes in Equilibrium – Adding catalyst

A catalyst will increase the reaction rate of both the forward and reverse reaction at the same rate. The **equilibrium position does not change when a catalyst is added.** A catalyst makes a system reach equilibrium faster.

NCEA 2013 Equilibrium

Excellence Question

Question 2c: The two reactions shown in the following table are both at equilibrium.

Reaction	**Equation**	**Affected by increased pressure**
One	$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$	no
Two	$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$	yes

Compare and contrast the effect of increasing the pressure on both reactions, with reference to the equilibrium positions.

When a change is made to a system that is at equilibrium, the system responds to **reduce the effect** of that change. **If there is an increase in pressure, the system responds by decreasing the pressure**. This occurs by **favouring the reaction that produces fewer gas particles**. Because there are now fewer particles hitting the sides of the container, there is less pressure.
In Reaction One there are two moles of gas particles on each side of the equation. Because there are the **same numbers of gas particles on both sides** of the reaction, then a change in pressure will have **no effect** as neither reaction will be favoured. In Reaction Two however, there are four moles of gas particles on the reactant side of the equation and two moles of gas particles on the product side of the equation. Therefore, when there is an increase in pressure, the system would shift and **favour the forward reaction** meaning there are **now fewer gas particles overall** and hence fewer gas particles hitting the sides of the container and **therefore less pressure overall**.

NCEA 2014 Equilibrium Expression

Excellence Question

Question 2a (i) : Hydrogen can be produced industrially by reacting methane with water.
An equation for this reaction can be represented by: $CH_{4(g)} + H_2O_{(g)} \leftrightharpoons CO_{(g)} + 3H_{2(g)}$
K_c = 4.7 at 1127°C
(a) (i) Complete the equilibrium constant expression for this reaction:

Answer:

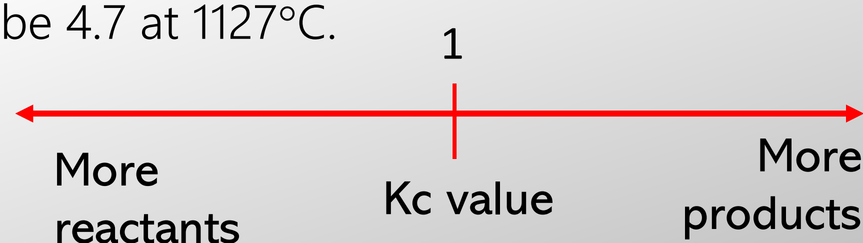
$$K_c = \frac{[H_2]^3[CO]}{[CH_4][H_2O]}$$

Place these values into expression

Gas	CH_4	H_2O	CO	H_2
Concentration/mol L^{-1}	0.0300	0.0500	0.200	0.300

Question 2a (ii): The concentrations of the four gases in a reaction mixture at 1127°C are found to be: (see above)
Use these values to carry out a calculation to determine if the reaction is at equilibrium.
Kc = 4.7 at 1127°C

No the system is not at equilibrium.
Calculated K_c = 3.6 which is < (less than) the actual K_c of 4.7; therefore the reaction is not at equilibrium since if the reaction was at equilibrium, the K_c would be 4.7 at 1127°C.



NCEA 2014 Equilibrium

Merit Question

Question 2b: The reaction shown in the equation below is at equilibrium.

$CO_{(g)} + 2H_{2(g)} \leftrightharpoons CH_3OH_{(g)}$

Describe the effect of each of the following changes on the equilibrium concentration of methanol (increase, decrease, stay the same).

Justify your answers using equilibrium principles.

(i) A copper oxide, CuO, catalyst is added. Amount of $CH_3OH_{(g)}$ would:

Answer: CuO catalyst:

Amount of methanol **stays the same**.

A catalyst, such as CuO, will speed up the rate of both the forward and the reverse reactions. The proportions of all of the reactants and products remain the same.

(ii) $H_{2(g)}$ is removed. Amount of $CH_3OH_{(g)}$ would:

Answer: H_2 removed:

Amount of methanol will **decrease**.

As hydrogen gas is removed, the system will oppose the change and the position of the equilibrium will shift in the reverse direction as more hydrogen is formed. This means the amount of methanol will decrease.

NCEA 2014 Equilibrium - (PART ONE)

Merit Question

Question 2c: In a reaction, the brown gas nitrogen dioxide, $NO_{2(g)}$, exists in equilibrium with the colourless gas dinitrogen tetroxide, $N_2O_{4(g)}$.
The equation for this reaction is represented by:

$2NO_{2(g)} \leftrightharpoons N_2O_{4(g)}$

brown gas ⠀⠀⠀⠀ colourless gas

The table below shows the observations when changes were made to the system.
Analyse these experimental observations.
In your answer you should:

- link all of the observations to equilibrium principles
- justify whether the formation of dinitrogen tetroxide from nitrogen dioxide is endothermic or exothermic.

Change		**Observations**
Pressure	increased (by decreasing the volume of the container)	Colour faded
	decreased (by increasing the volume of the container)	Colour darkened
Temperature	container with reaction mixture put into hot water	Colour darkened
	container with reaction mixture put into ice water	Colour faded

NCEA 2014 Equilibrium (PART TWO)

Excellence Question

Answer 2c:

Increasing pressure resulted in the colour fading. This was due to the position of the equilibrium shifting in the forward direction to counteract this change. The system shifts in the **direction of the least number of moles of gas** since this will decrease the pressure. This forms more N_2O_4 (*g*), which is colourless, so the colour fades.

When the pressure is decreased again the system adjusts to increase the pressure, hence shifts in the direction of a greater number of moles, i.e. in the reverse direction, forming more NO_2, resulting in a darker colour due to the brown colour of this gas.

When the reaction container is placed in **hot water**, the system will adjust so that some of the heat is used up; therefore it will shift in the **endothermic direction**. In this case, the colour darkened, indicating that this favoured the reverse reaction which must be the endothermic direction. When the container was **cooled**, the colour faded indicating that the forward reaction, forming colourless N_2O_4, **must be exothermic**.

Change		**Observations**
Pressure	increased (by decreasing the volume of the container)	Colour faded
	decreased (by increasing the volume of the container)	Colour darkened
Temperature	container with reaction mixture put into hot water	Colour darkened
	container with reaction mixture put into ice water	Colour faded

NCEA 2015 Equilibrium

Question 3(a): The equilibrium constant for a reaction involving compounds A, B, C, and D is shown as:

$$K_c = \frac{[C]^3[D]}{[A]\,[B]^2}$$

Write the chemical equation for this reaction.

Answer: A + 2B $\rightleftharpoons$ 3C + D

Question 3(b)(i) : The reaction between ethanoic acid and ethanol is reversible. Ethyl ethanoate and water are the products formed. In a closed system, a dynamic equilibrium is set up.

ethanoic acid + ethanol $\rightleftharpoons$ ethyl ethanoate + water

$CH_3COOH_{(aq)} + C_2H_5OH_{(aq)} \rightleftharpoons CH_3COOC_2H_{5(aq)} + H_2O_{(l)}$

(i) Explain, using equilibrium principles, the effect of adding more ethanol to the reaction mixture.

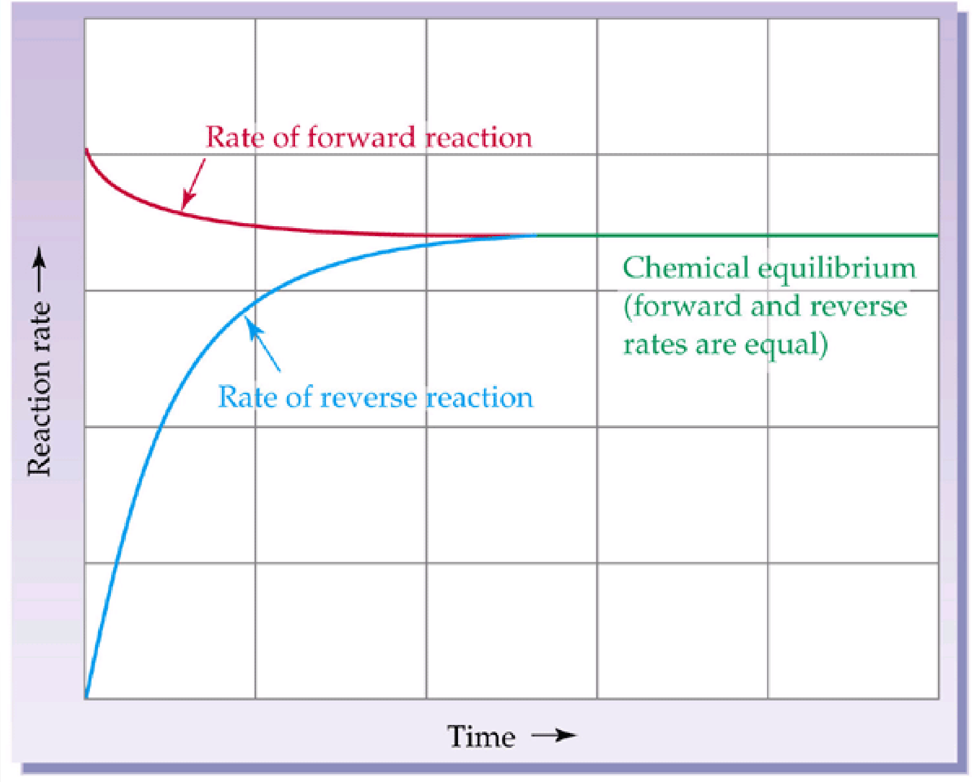
Merit Question

Answer: Adding more ethanol causes the equilibrium to move in the forward direction in order to use the extra added ethanol. This is because the equilibrium has to re-establish itself with the added reactant in order to maintain K_c.

NCEA 2015 Equilibrium

Merit Question

Question 3(b)(ii): The reaction between ethanoic acid and ethanol is reversible. Ethyl ethanoate and water are the products formed. In a closed system, a dynamic equilibrium is set up.
The reaction is quite slow, so a small amount of concentrated sulfuric acid is added as a catalyst.
Explain, using equilibrium principles, the effect of adding this catalyst to the equilibrium mixture.



Answer: A catalyst speeds up the rate of the reaction so both forward and backward reaction will speed up but no particular reaction is favoured.

NCEA 2015 Equilibrium Expression

Excellence Question

Question 3(c)(i): The following chemical equation represents a reaction that is part of the Contact Process which produces sulfuric acid.

$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$ ΔH = −200 kJ mol^{-1}, Kc= 4.32 at 600°C

(i) Write an equilibrium constant expression for this reaction.

Answer:

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

1

More reactants

Kc value

More products

Question 3(c)(ii): A reaction mixture has the following concentration of gases at 600°C:

$[SO_{2(g)}]$ = 0.300 mol L^{-1}

$[O_{2(g)}]$ = 0.100 mol L^{-1}

$[3_{(g)}]$ = 0.250 mol L^{-1}

Justify why this reaction mixture is not at equilibrium.

In your answer you should use the equilibrium expression from part (c)(i) and the data provided above to show that the reaction mixture is not at equilibrium.

Answer:

$$Q = \frac{0.250^2}{0.300^2 \times 0.100} = 6.94$$

Since K_c = 4.32, $Q \neq K_c$, so this reaction mixture is not at equilibrium. This number is greater than he K_c value, 4.32, which indicates that the reaction lies to the products side as the larger the K_c or Q value, the greater the numerator (products).

NCEA 2015 Equilibrium

Excellence Question

Question 3(c)(iii): The reaction on the previous page was repeated at 450°C. Explain, using equilibrium principles, how the change in temperature will affect:

- the value of *Kc*
- the position of equilibrium.

Answer:

At 450°C, the temperature has decreased. This reaction is exothermic, as shown by the negative enthalpy. This means that if the temperature is decreased, the reaction will move in the direction that produces more heat. Because this is an exothermic reaction, the exothermic direction is forwards. This will lead to more products and an increase in K_c.

K_c **decreases** when temperature **increases**

Exothermic reaction

K_c **increases** when temperature **decreases**

K_c **increases** when temperature **increases**

Endothermic reaction

K_c **decreases** when temperature **decreases**

NCEA 2016 Equilibrium Expression

Achieved Question

Question 3(a): The equilibrium constant expression for a reaction is:

$$K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$$

Write the equation for this reaction.

$CO + 2H_2 \rightleftharpoons CH_3OH$

Merit Question

Question 3(b): The ionisation of water is represented by the equation:
$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
Give an account of the extent of ionisation of water, given $Kw = 1 \times 10^{-14}$.

The value of K_w at 1×10^{-14} is very small, which means that very little water has ionised / dissociated because there is very little product (i.e. very few ions).

NCEA 2016 Equilibrium

Excellence Question

Question 3(c): When acid is added to a yellow solution of chromate ions, $CrO_4^{2-}(aq)$, the following equilibrium is established.

$2CrO_4^{2-}(aq) + 2H^+(aq) \leftrightharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$

yellow **orange**

Analyse this equilibrium using equilibrium principles to explain the effect on the colour of the solution when:

(i) more dilute acid is added:

Adding dilute acid increases the concentration of the acid, so the reaction moves in the forward direction / favours the products to use up the added acid, so the colour of the solution will become more orange.

(ii) dilute base is added:

Adding base means that acid that reacts with the base is removed from the equilibrium / concentration of the acid decreases. This will drive the equilibrium in the backwards direction / favours the reactants to replace the H^+ used up, causing the solution to become more yellow in colour.
May include the equation $H^+ + OH^- \rightarrow H_2O$ in your answer.

Merit Question

NCEA 2016 Equilibrium Expression (PART ONE)

Question 3(d):
When hydrogen gas, $H_2(g)$, and iodine gas, $I_2(g)$ are mixed, they react to form $HI(g)$, and an equilibrium is established.

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
$Kc = 64$ at 445°C.

(i) Calculate the concentration of HI in an equilibrium mixture at 445°C when the concentrations of $H_2(g)$ and $I_2(g)$ are both 0.312 mol L^{-1}.

$$64 = \frac{[HI]^2}{0.312 \times 0.312} \qquad [HI] = \sqrt{64 \times 0.312^2} = 2.50 \text{ mol L}^{-1}$$

NCEA 2016 Equilibrium (PART TWO)

Merit Question

Question 3(d): When hydrogen gas, $H_2(g)$, and iodine gas, $I_2(g)$ are mixed, they react to form $HI(g)$, and an equilibrium is established.

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

$Kc = 64$ at 445°C.

(ii) Explain the effect on the position of equilibrium if the overall pressure of the equilibrium system is increased.

There would be no effect on the equilibrium if pressure was increased because there are equal numbers of moles of gas on either side of the equilibrium / in the reactants and products.

NCEA 2016 Equilibrium (PART THREE)

Excellence Question

Question 3(d): When hydrogen gas, $H_{2(g)}$, and iodine gas, $I_{2(g)}$ are mixed, they react to form $HI_{(g)}$, and an equilibrium is established.

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

$K_c = 64$ at 445°C.

(iii) When the temperature of the equilibrium system is increased to 510°C, the K_c value decreases to 46.

Justify, using equilibrium principles, whether the forward reaction is exothermic or endothermic.

Answer: (Temperature is the only factor that can change the *K* value of an equilibrium).

When the temperature increases, the reaction moves in the endothermic direction to absorb the added heat. In this reaction, the value of *K* decreased, indicating the ratio of products to reactants (numerator to denominator) decreased. Since there will be fewer products and more reactants, adding heat is favouring the backwards reaction.

Therefore, the forward reaction is exothermic.

NCEA 2017 Equilibrium Expression (PART ONE)

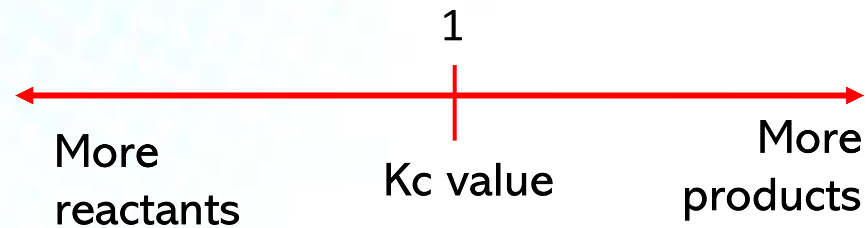
Merit Question

Question 2b: The reaction described below is an equilibrium reaction, as represented by the following equation:

$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$

(i) Write the equilibrium constant expression for this reaction.

$$K_c = \frac{[NH_3(aq)]^2}{[N_2(aq)][H_2(aq)]^3}$$



(ii) The value of the equilibrium constant, K_c, is 640 at 25°C.
Show, by calculation, using the concentrations of the gases given in the table below, whether or not the reaction is at equilibrium.
Explain your answer.

Gas	N_2	H_2	NH_3
Concentration (mol L^{-1})	0.0821	0.0583	0.105

$$K_c = \frac{0.105^2}{0.0821 \times 0.0583^3}$$
$$= 677$$

No, the reaction is not at equilibrium because 677 > 640 (values must be equal for a reaction to be at equilibrium). *Accept answers between 676 – 678.*

NCEA 2017 Equilibrium (PART TWO)

Excellence Question

Question 2c: As the temperature increases, the value of the equilibrium constant, K_c, decreases from 640 at 25°C to 0.440 at 200°C.
Justify whether the formation of ammonia, $NH_{3(g)}$, is an endothermic or exothermic reaction.

$$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$$

Answer: (Temperature is the only factor that can change the *K* value of an equilibrium).
When the temperature increases, the reaction moves in the endothermic direction to absorb the added heat. In this reaction, the value of *K* decreased, indicating the ratio of products to reactants (numerator to denominator) decreased. Since there will be fewer products and more reactants, adding heat is favouring the backwards reaction.

Therefore, the forward reaction is exothermic.

NCEA 2017 Equilibrium

Excellence Question

Question 3b: Two different cobalt(II) complex ions, $[Co(H_2O)_6]^{2+}$ and $[CoCl_4]^{2-}$, exist together in a solution in equilibrium with chloride ions, $Cl^-_{(aq)}$.
The forward reaction is endothermic; ΔH is positive. The equation for this equilibrium is shown below.

$$\underset{\text{Pink}}{[Co(H_2O)_6]^{2+}_{(aq)}} + 4Cl^-_{(aq)} \rightleftharpoons \underset{\text{blue}}{[CoCl_4]^{2-}_{(aq)}} + 6H_2O_{(l)}$$

Explain using equilibrium principles, the effect on the colour of the solution if:
(i) more water is added to the reaction mixture

Adding water to this equilibrium means there has been an **increase in (**concentration of) **a product**. The system will react to reduce this change, so the backward reaction will be favoured to **use up** *(idea required for Excellence)* some of the extra product.
This results in an increased concentration of the pink $[Co(H_2O)_6]^{2+}$ ion, so the solution will turn pink or the pink colour will intensify.

Question 3b: (ii) a test tube containing the reaction mixture is placed in a beaker of ice-cold water.

The ice-cold water will cause the reaction to move in the exothermic direction to compensate for the loss of heat energy / **release heat energy** *(for Excellence)* into surroundings. Because this reaction is endothermic (positive ΔH value), the exothermic direction will be backwards, so the colour of the solution will become pink or the pink colour will intensify.

NCEA 2017 Equilibrium

Question

Question 3c: Brown nitrogen dioxide gas, $NO_{2(g)}$, exists in equilibrium with the colourless gas, dinitrogen tetroxide, $N_2O_{4(g)}$.

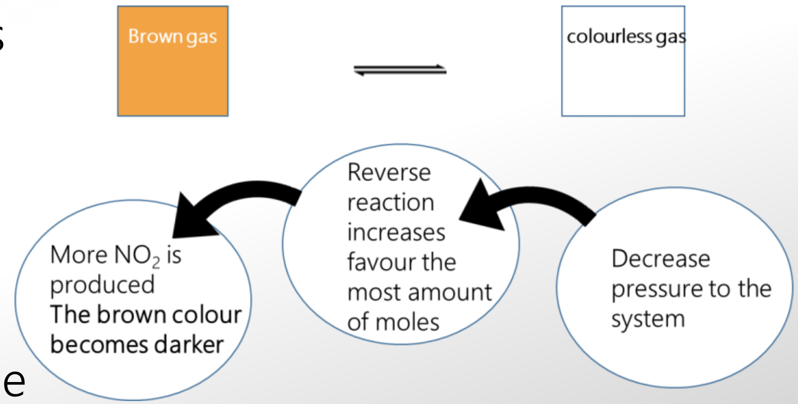
$2NO_{2(g)} \leftrightharpoons N_2O_{4(g)}$

brown colourless

Explain using equilibrium principles, the effect of decreasing the volume of the container (therefore increasing the pressure) on the observations of this equilibrium mixture.

Increasing pressure resulted in the colour fading. This was due to the position of the equilibrium shifting in the forward direction to counteract this change. The system shifts in the **direction of the least number of moles of gas** since this will decrease the pressure. This forms more N_2O_4 (*g*), which is colourless, so the colour fades.

When the pressure is decreased again the system adjusts to increase the pressure, hence shifts in the direction of a greater number of moles, i.e. in the reverse direction, forming more NO_2, resulting in a darker colour due to the brown colour of this gas.



NCEA 2018 Equilibrium Expression (part ONE)

Achieved Question

Question 2a: The Contact Process is used industrially in the manufacture of sulfuric acid. One step in this process is the oxidation of sulfur dioxide, $SO_{2(g)}$, to sulfur trioxide, $SO_{3(g)}$.

$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$

(a) Write the equilibrium constant expression for this reaction.

$$K_c = \frac{[SO_3]^2}{[SO_2]^2\ [O_2]}$$

Question 2b: (i) Calculate the equilibrium constant (*Kc*) for this reaction at 600°C using the following concentrations:

$[SO_2] = 0.100 \text{ mol L}^{-1}$

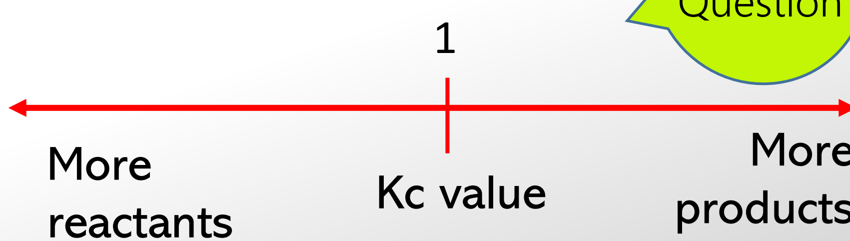
$[O_2] = 0.200 \text{ mol L}^{-1}$

$[SO_3] = 0.0930 \text{ mol L}^{-1}$

$$K_c = \frac{0.093^2}{0.1^2 \times 0.2}$$

$$K_c = 4.32$$

Merit Question



Question 2b: (ii) Explain what the size of the K_c value indicates about the extent of the reaction at equilibrium.

The K_c value is larger than 1 so there are more products than reactants at equilibrium.

NCEA 2018 Equilibrium (part TWO)

Excellence Question

Question 2c: Explain, using equilibrium principles, why it is important for an industrial plant to continue to remove the sulfur trioxide gas, $SO_{3(g)}$, as it is produced.
$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$

If the sulfur trioxide is removed as it is produced, **$[SO_3]$ will decrease**, so the equilibrium will move to minimise the change (stress placed on the system). This means the reaction will move forward to **replace the lost sulfur trioxide**. This will **increase the yield of the desired product.**

Question 2d: Predict, using equilibrium principles, the effect on the concentration of sulfur trioxide gas, $SO_{3(g)}$, of carrying out the reaction in a **larger** reaction vessel.

Increasing the size of the reaction vessel **decreases the pressure** of the system. In order to minimise this change / stress, the reaction moves to **increase the number of gaseous particles**. For this reaction, the greatest number of gaseous particles is the reactants side so the reaction will **move backwards towards the reactants.** This has the effect of **decreasing the amount of sulfur trioxide.**

NCEA 2018 Equilibrium (part THREE)

Excellence Question

Question 2e: When the reaction is carried out at 450ºC, the K_c value is higher than the value at 600°C.
Justify whether the oxidation of sulfur dioxide gas, $SO_{2(g)}$, to sulfur trioxide gas, $SO_{3(g)}$, is exothermic or endothermic.
$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$

When the **temperature decreases** to 450°C, the reaction moves in the **exothermic direction to produce more heat**. Since the K_c **value increased**, **more products** and less reactants are present. This means the reaction produces more products when the temperature drops. This means the oxidation of sulfur dioxide to sulfur trioxide is an exothermic reaction.

K_c **decreases** when temperature **increases**

Exothermic reaction

K_c **increases** when temperature **decreases**

K_c **increases** when temperature **increases**

Endothermic reaction

K_c **decreases** when temperature **decreases**

NCEA 2019 Equilibrium (part ONE)

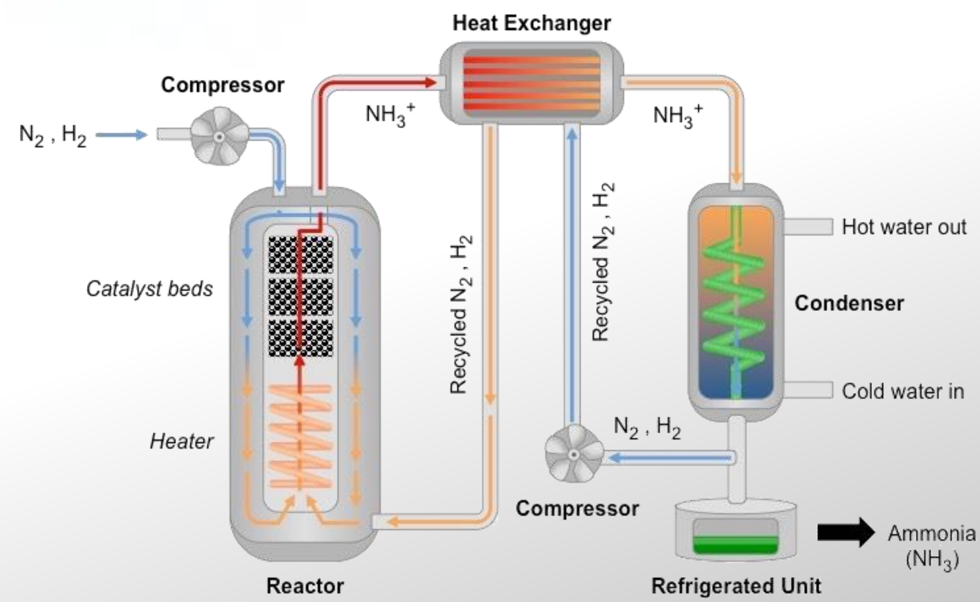
Excellence Question with (iii)

Question 2a: The Haber process combines nitrogen, $N_{2(g)}$, from the air with hydrogen, $H_{2(g)}$, to form ammonia, $NH_{3(g)}$, which is then used in the manufacture of fertiliser.
The equation for this process is $N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$
(i) Write the equilibrium constant expression for this reaction.

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

(ii) Using equilibrium principles, explain why carrying out the Haber process at high pressure is an advantage to the manufacturer.

There are four moles of **gas** particles on the reactant side of the equation, and two moles of **gas** particles on the product side of the equation.
Therefore, when there is an increase in pressure, the system would shift towards the products (to minimise the stress on the equilibrium) since there are fewer gas molecules on the product side. This would increase the yield of ammonia, so would be an advantage for the manufacturer.



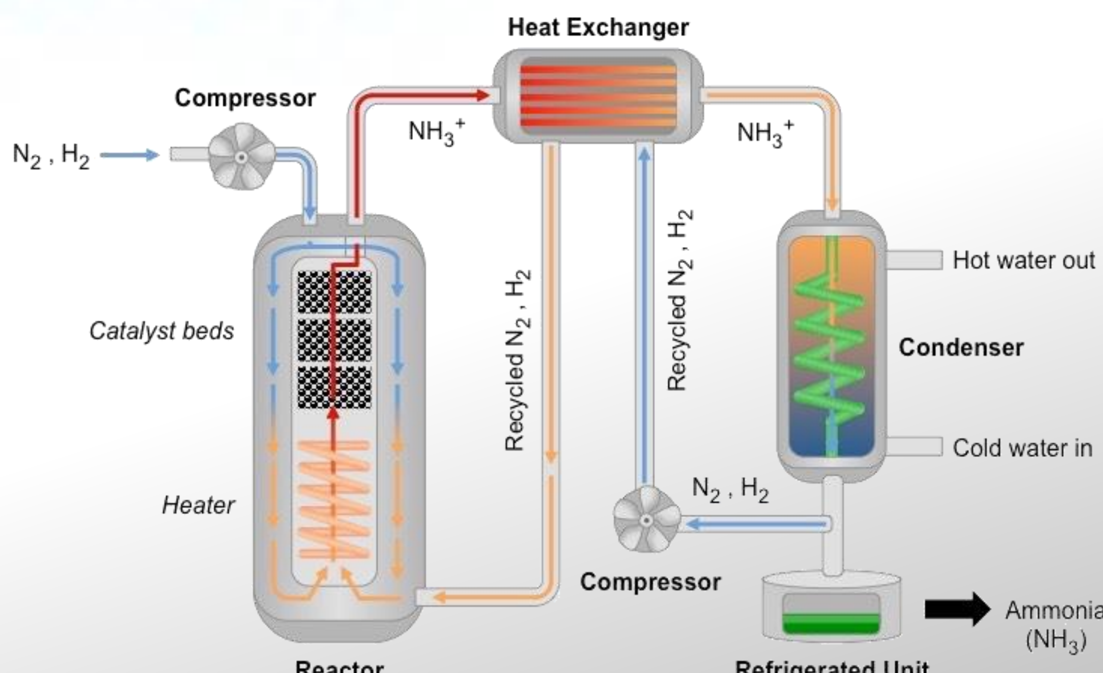
NCEA 2019 Equilibrium (part TWO)

Excellence Question with (ii)

Question 2a: The Haber process combines nitrogen, $N_{2(g)}$, from the air with hydrogen, $H_{2(g)}$, to form ammonia, $NH_{3(g)}$, which is then used in the manufacture of fertiliser.
The equation for this process is $N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$
(iii) In another part of the process, the ammonia, $NH_{3(g)}$, is removed as it is produced. Justify this step using equilibrium principles to explain why this would be an advantage to a manufacturer.

As ammonia gas is removed, the concentration of the products decreases. The system will oppose the change by shifting in the forward direction to form more ammonia/replace ammonia. In industry, this is an advantage as it maximises the amount of ammonia produced.

Haber process



Excellence Question

NCEA 2019 Equilibrium (part THREE)

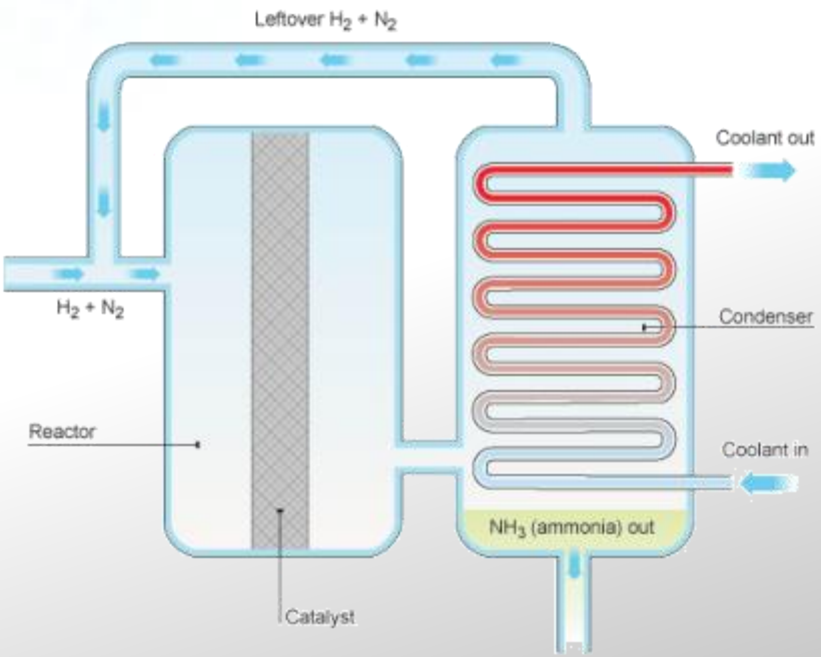
Question 2b: The Haber process

$N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)} \quad \Delta_r H = -92 \text{ kJ mol}^{-1}$

Explain, using equilibrium principles, whether the value of K_c would increase or decrease if the temperature of the reaction is increased.

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

As the temperature increases, the system will act to reduce the temperature by favouring the endothermic direction to absorb some of the extra heat energy. Since the reaction has a negative $\Delta_r H$, this means that the forward reaction is exothermic and produces heat energy. So, an increase in temperature will cause the equilibrium to shift towards the reactants and therefore the concentration of reactants will increase. A higher concentration of reactants (compared to products) will cause the K_c value to decrease.



NCEA 2019 Equilibrium Expression

Excellence Question with (ii)

Question 2c: (i) Nitrogen, $N_{2(g)}$, can also be reacted with oxygen, $O_{2(g)}$, to give nitrogen dioxide, $NO_{2(g)}$, and the following K_c expression would apply.

$$K_c = \frac{[NO_2]^2}{[N_2][O_2]^2}$$

The K_c for the reaction at 25°C is 8.30×10^{-10}.
Calculate the concentration of nitrogen dioxide, NO_2, if the concentration of oxygen, O_2, is 0.230 mol L^{-1} and the concentration of nitrogen, N_2, is 0.110 mol L^{-1}.
Give your answer to appropriate significant figures.

$$[NO_2] = \sqrt{K_c[N_2][O_2]^2}$$

$$= \sqrt{(8.3\times10^{-10})\times0.110\times0.230^2}$$

$$= \sqrt{4.830\times10^{-12}}$$

$$= 2.20\times10^{-6}\text{ mol L}^{-1}\quad 3\text{sf}$$

NCEA 2019 Equilibrium Expression

Excellence Question with (i1)

Question 2c: (ii) Explain the effect on K_c if the concentration of nitrogen, $N_{2(g)}$, is increased to 0.200 mol L^{-1} at 25°C (no calculations are necessary).

$N_{2(g)} + 2O_{2(g)} \leftrightharpoons 2NO_{2(g)}$

$$K_c = \frac{[NO_2]^2}{[N_2][O_2]^2}$$

When $N_{2(g)}$ is added, the system will oppose the change (increase in concentration of $N_{2(g)}$) and therefore the position of the equilibrium will shift in the forward direction to use up some of the added $N_{2(g)}$. This means more $NO_{2(g)}$ will be produced.

However, the ratio of the concentrations of the reactants and products will remain the same, consequently the value of K_c remains unchanged. Only a change in temperature will affect the value of K_c.

NCEA 2020 Equilibrium Expression

Merit Question

Question 3a: (i) Write the equilibrium constant expression, K_c , for the conversion of gaseous carbonyl fluoride, $COF_{2(g)}$, to the gas carbon tetrafluoride, $CF_{4(g)}$ and carbon dioxide, $CO_{2(g)}$.
$2COF_{2(g)} \leftrightharpoons CF_{4(g)} + CO_{2(g)}$

$$K_c = \frac{[CF_4][CO_2]}{[COF_2]^2}$$

Question 3a: (ii) At equilibrium, carbonyl fluoride, COF_2, has a concentration of 0.040 mol L^{-1}.
The concentration of both carbon tetrafluoride, CF_4, and carbon dioxide, CO_2, is 0.80 mol L^{-1}. Calculate the K_c for this equilibrium.

$$K_c = \frac{0.8 \times 0.8}{0.40^2} = 400$$

Excellence Question

NCEA 2020 Equilibrium Expression

Question 3a: $2COF_{2(g)} \leftrightharpoons CF_{4(g)} + CO_{2(g)}$
(iii) At a **different** temperature, the Kc value is 50.
Explain what the value of the Kc indicates about the extent of this reaction.

The value of K is significantly over 1, so there are far more products than reactants, which means the equilibrium favours the products.

Question 3a: (iv) The enthalpy change, Δ_rH, for the decomposition of carbonyl fluoride is –24 kJ mol^{-1}.
Explain what happens to the value of K_c when the temperature is decreased.

The reaction is exothermic. If the temperature is decreased, the reaction moves in the exothermic direction to produce more heat energy. For this reaction, it will favour the forward reaction. This leads to fewer reactants / increased products so the value of K_c will increase.

NCEA 2020 Equilibrium (part ONE)

Excellence Question (i) , (ii), and (iii)

Question 3b: The following equilibrium was established in the laboratory by mixing iron(III) nitrate solution, $Fe(NO_3)_{3(aq)}$, with potassium thiocyanate solution, $KSCN_{(aq)}$.

$Fe_3^{+}{}_{(aq)}$	$+ \ SCN^{-}{}_{(aq)}$	$\leftrightharpoons$	$[FeSCN]^{2+}{}_{(aq)}$
Orange	colourless		dark red

The forward reaction produces heat.
Explain, using equilibrium principles, the effect on the colour of the solution if:
(i) More potassium thiocyanate solution, $KSCN_{(aq)}$, is added to the reaction mixture.

Adding thiocyanate ions to the equilibrium means there is an increase in the concentration of a reactant. The system will react to reduce this change so the forward reaction will be favoured to use up the added SCN^{-} ions producing more red $[FeSCN]^{2+}$. This means the dark red colour will intensify.

NCEA 2020 Equilibrium (part TWO)

Excellence Question (i) , (ii), and (iii)

Question 3b: The following equilibrium was established in the laboratory by mixing iron(III) nitrate solution, $Fe(NO_3)_{3(aq)}$, with potassium thiocyanate solution, $KSCN_{(aq)}$.

$Fe_3{}^{+}{}_{(aq)}$	$+ \ SCN^{-}{}_{(aq)}$	$\leftrightharpoons$	$[FeSCN]^{2+}{}_{(aq)}$
Orange	colourless		dark red

The forward reaction produces heat.
Explain, using equilibrium principles, the effect on the colour of the solution if:
(ii) Solid sodium fluoride is added to the mixture. The added F^{-} ions react with the Fe^{3+} ions.

The added fluoride ions reacts with the Fe^{3+} ions, and this decreases the concentration of the Fe^{3+} in this equilibrium. The system will react by favouring the backward reaction to replace the lost orange Fe^{3+} ions, while using up red $[FeSCN]^{2+}$. This means the dark red colour will lighten and it will become more orange.

(iii) A test tube containing the reaction mixture is placed in a beaker of recently boiled water.

The forward reaction produces heat so when the mixture is put into hot water, the reaction moves in the endothermic direction to absorb the added heat energy. This will favour the backward reaction using up red $[FeSCN]^{2+}$ and producing orange Fe^{3+}, which means the dark red colour will lighten and the mixture will be more orange.

Background Knowledge

Acids – their characteristics

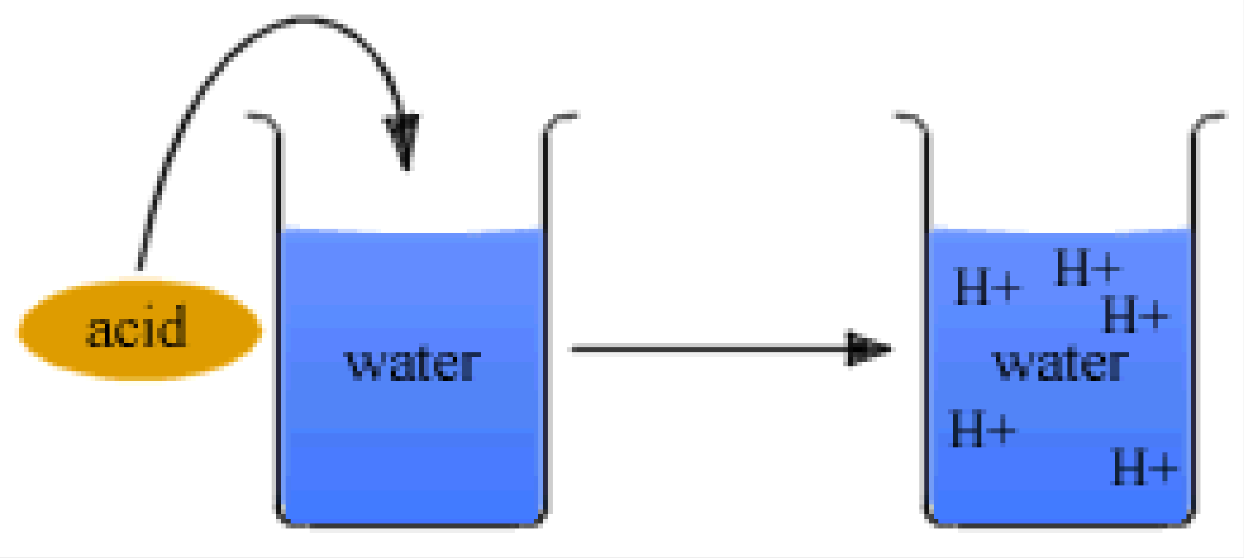
Acids are a family of substances which all show **acidic characteristics** or properties. These properties relate to how the acids react with other chemicals.

They have a **sour taste** and react with metals. Acids can be found in nature and called **organic acids** or manufactured in the laboratory and called **mineral acids**.

Background Knowledge

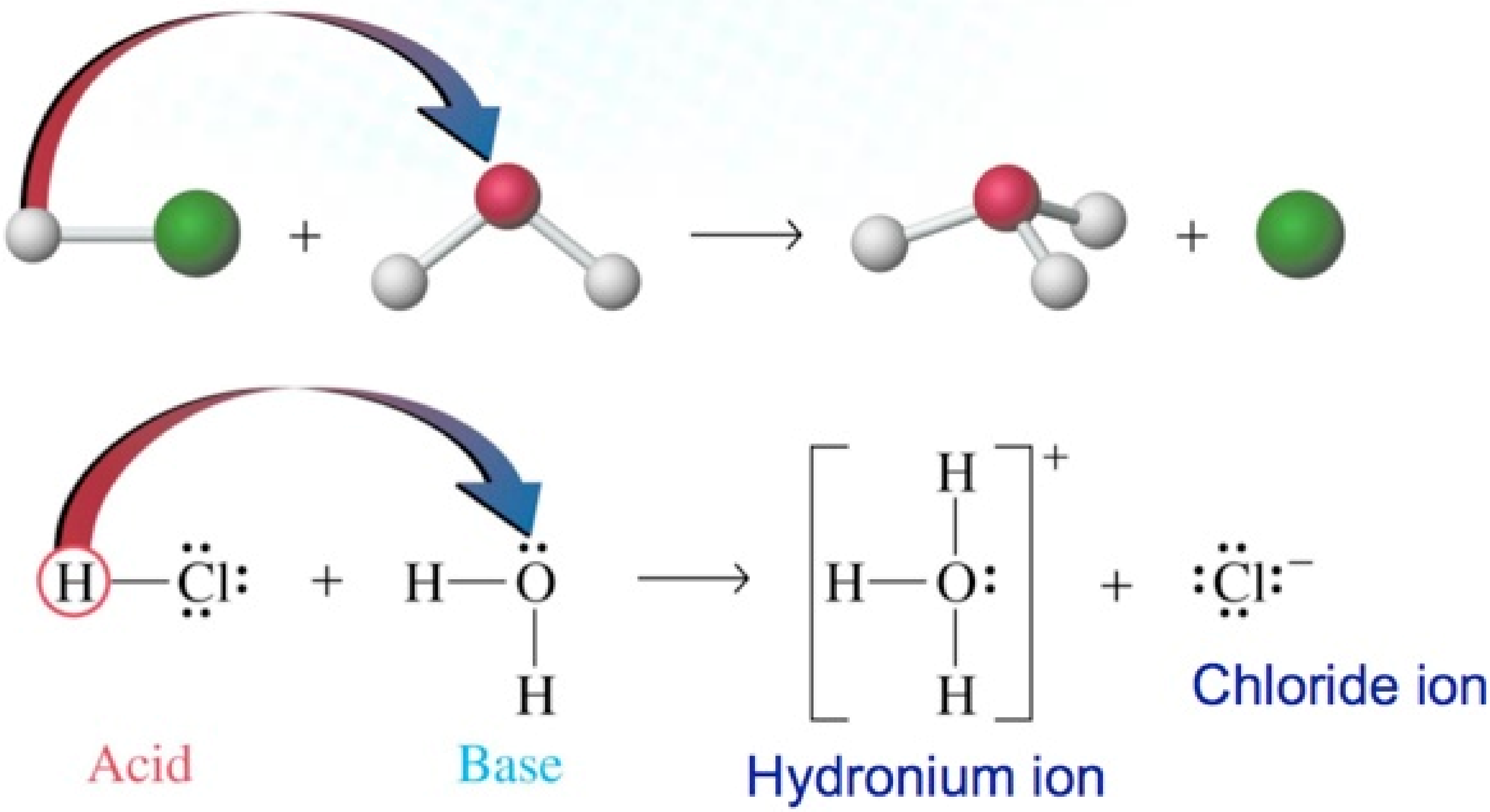
Acids – their characteristics

An **Acid** donates its **Hydrogen ion** (H+) , which is really just a proton - the electron remains behind.
Common acids include the strong acids HNO_3 - nitric acid, HCl - hydrochloric acid, and H_2SO_4 – sulfuric acid.



Acids – their characteristics

An **Acid** donates its **Hydrogen ion** (H+) , which is really just a proton - the electron remains behind. H_3O^+ ions are produced
Common acids include the strong acids HNO_3 - nitric acid, HCl - hydrochloric acid, H_2SO_4 – sulfuric acid, and the weak acid CH_3COOH – ethanoic acid.



Background Knowledge

Bases – their characteristics

GZ Science Resources

Bases are a family of chemicals that can remove acid particles (H^+) from a solution. They have opposite properties from acids.

Bases have a slippery feel to them and common house hold bases include floor clearers and antacid tablets to fix indigestion. **Bases that dissolve into water are called an alkali, and produce OH^- ions.**

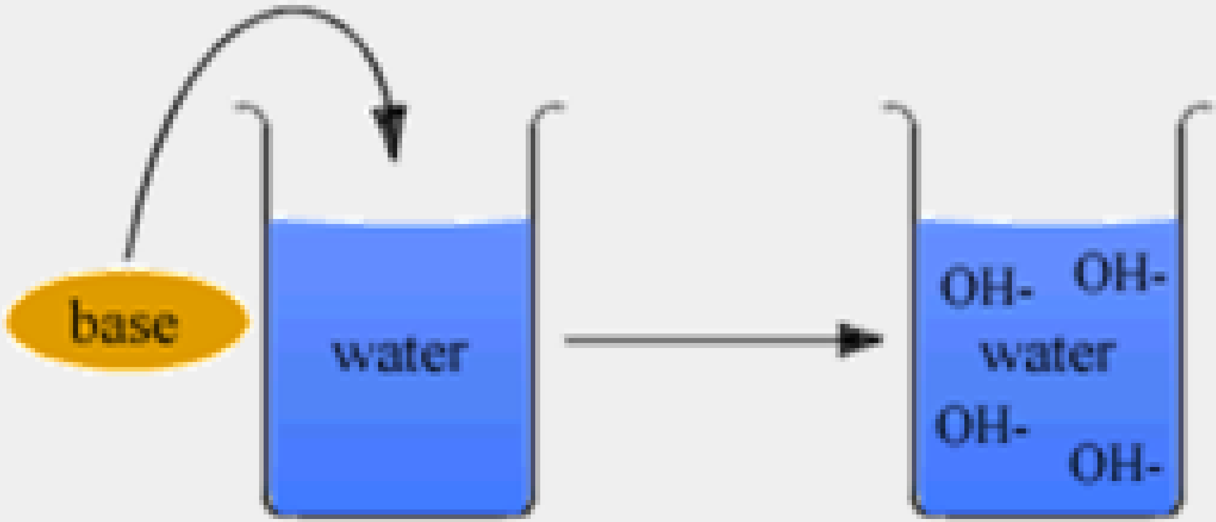
Background Knowledge

GZ Science Resources

Bases – their characteristics

A **Base** accepts a **Hydrogen ion** that have been donated from an Acid. They release hydroxide ions into solution. Common bases include the strong bases NaOH – sodium hydroxide, and other metal oxides, hydroxides, carbonates and hydrogen carbonates

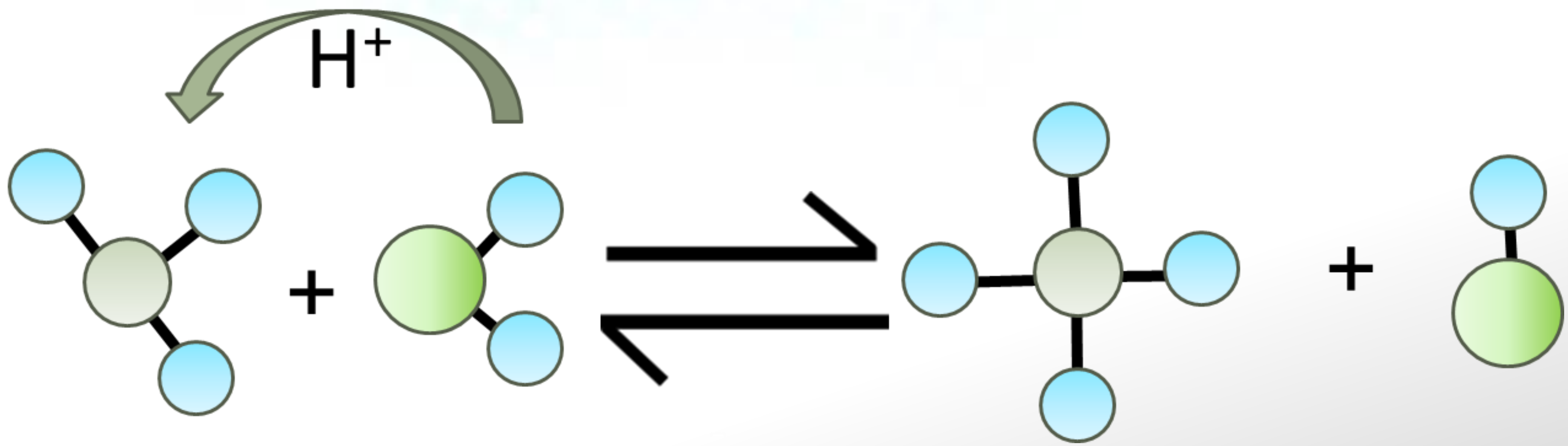
Some substances such as water are **amphiprotic** and can act as both an acid or a base depending on what other substance the water is with.



Bases – their characteristics

A **Base** accepts a **Hydrogen ion** that have been donated from an Acid.
Common bases include the strong bases NaOH – sodium hydroxide, KOH – potassium hydroxide and the weak base NH_3 – ammonia.
An **alkali** is a base that is soluble in water and contains hydroxide ions.

Some substances such as water are **amphiprotic** and can act as both an acid or a base depending on what other substance the water is with.

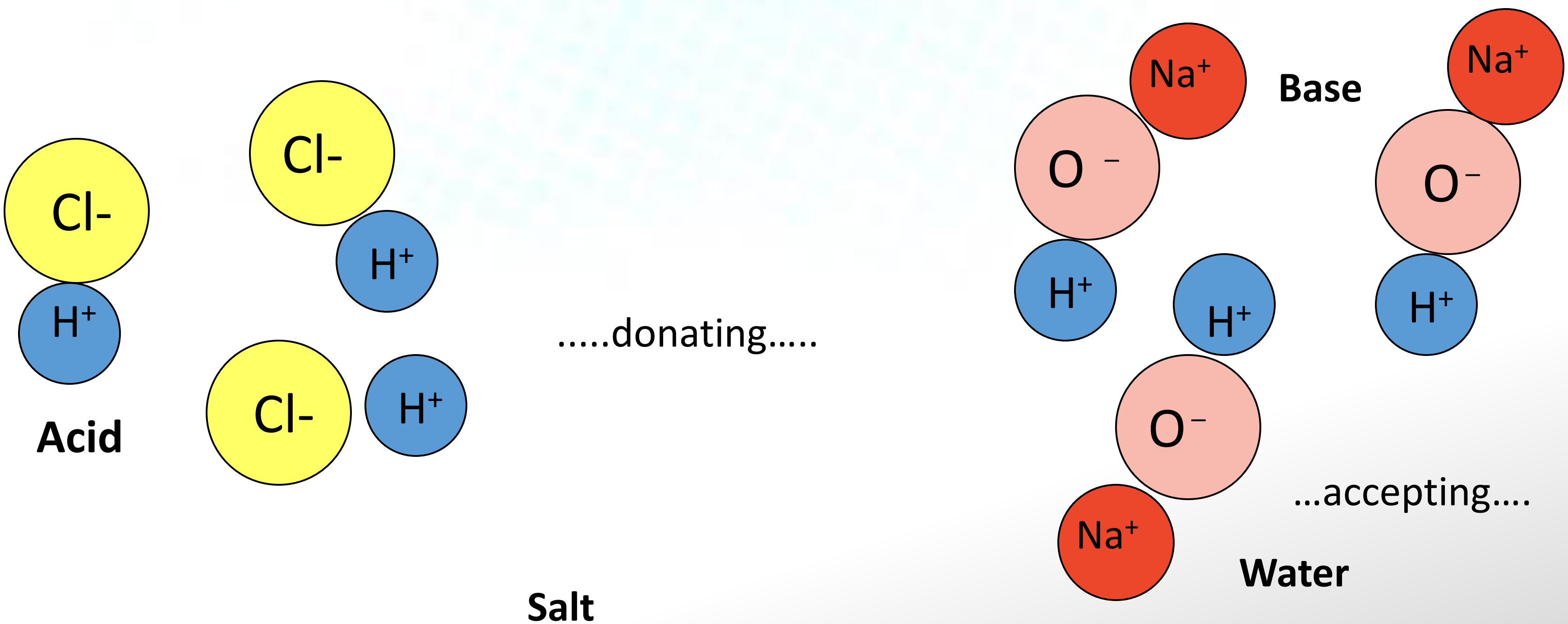


$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

Definition of Acid and Bases

Acids are substances that release (donate) hydrogen ions, H^+, in water.

Bases are substances that accept hydrogen ions.



Bases neutralise acids

a salt and water are formed

Brønsted–Lowry theory of Acid and Base reactions

Acid-Base reactions involve the **transfer of Hydrogen ions**, H^+

A hydrogen ion, H^+ is simply a lone proton (an H with the electron removed)

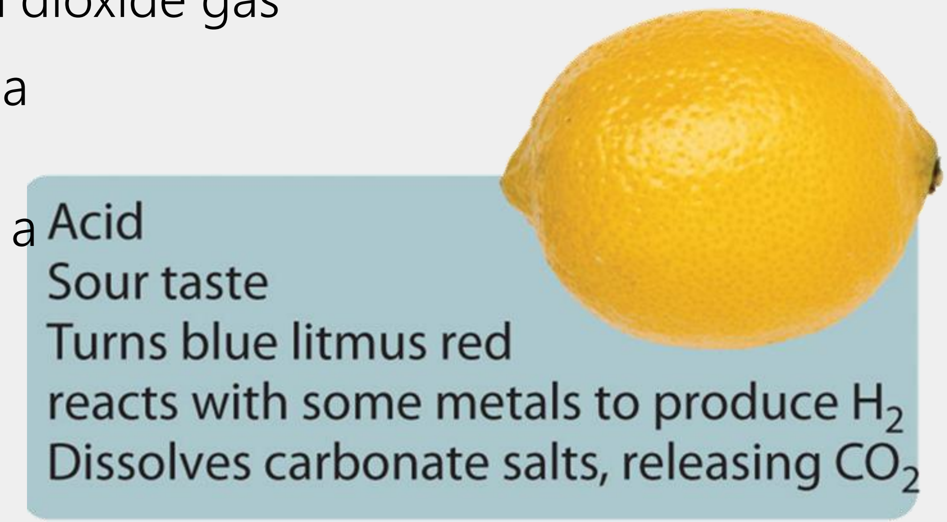
In water (or aqueous solutions) H^+ ions exist as an H_3O^+ ion, called **hydronium**.

Acids	Bases
Acids are substances that donate protons (H^+) in solution	Bases are substances that accept protons (H^+) in solution
$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$	$NH_{3(g)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$
HCl gas dissolved in water	NH_3 gas dissolved in water
HCl has donated a H^+ so is acting as an acid	NH_3 has accepted a H^+ so it is acting as a base
H_2O has accepted a H^+ so it is acting as a base	H_2O has donated a H^+ so is acting as an acid
Solution becomes acidic since H_3O^+ ions form	Solution becomes basic since OH^- ions form.

Background Knowledge

Physical and Chemical Properties of Acids

- pH < 7
- Turn litmus paper **red**
- Turn phenolphthalein **colourless**
- Turn methyl orange **red**
- Neutralised by bases
- React with carbonates to form a metal salt, water and carbon dioxide gas
- React with most metals to form a metal salt and hydrogen gas
- React with metal oxides to form a metal salt and water
- React with metal hydroxides to form a metal salt and water
- Have a sour taste
- Conduct electricity



Background Knowledge

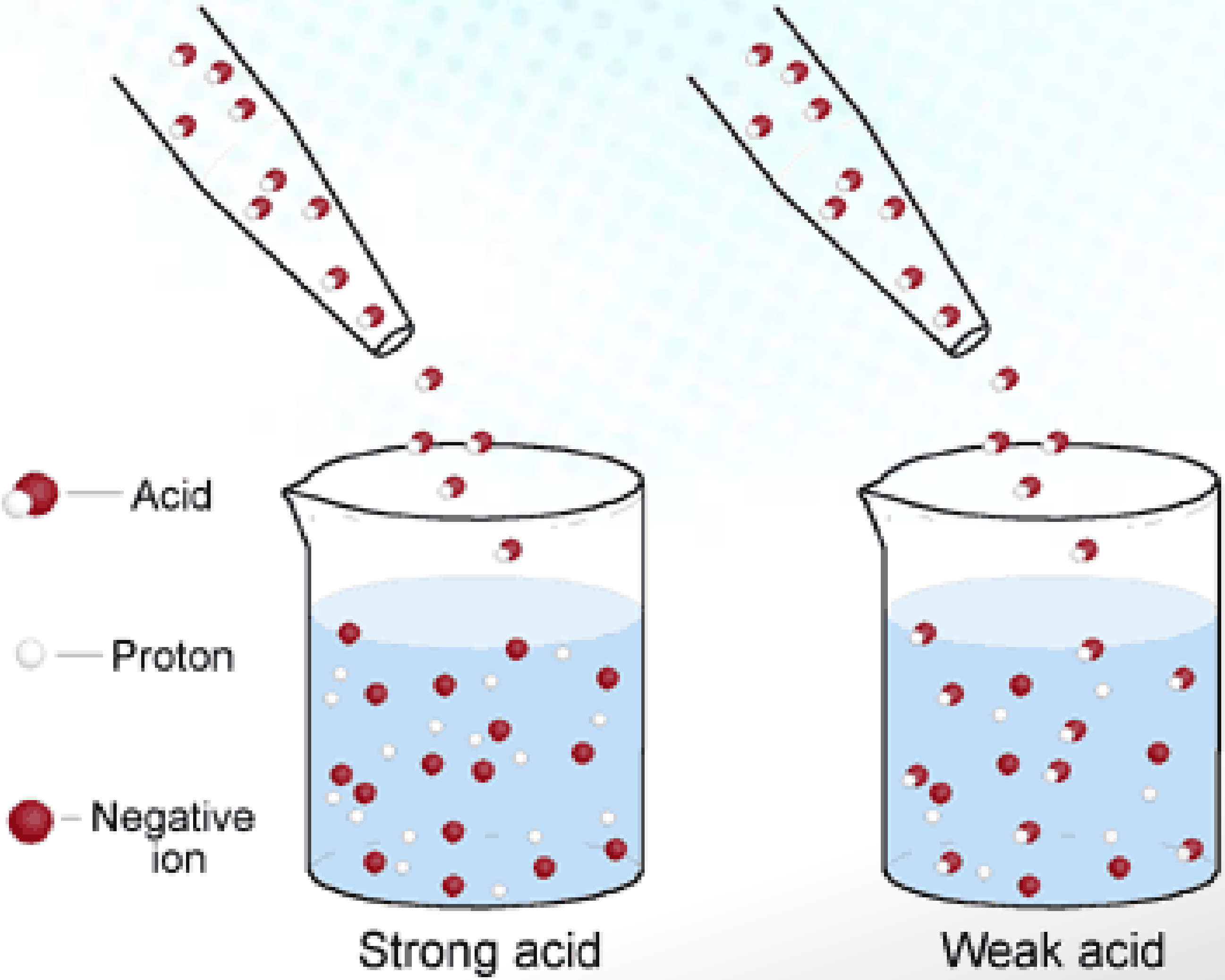
GZ Science Resources

Chemical and Physical Properties of Bases

- pH > 7
- Turn litmus paper **blue**
- Turn phenolphthalein **pink**
- Turn methyl orange **yellow**
- Neutralise acids
- Feel soapy
- Conduct electricity
- Bases that dissolve in water are called alkalis, they form OH^- ions.

Base
Bitter taste
Turns red litmus blue
Slippery to the touch

Understand the difference between Strong and Weak acids and bases



You can define acids and bases as being "strong" or "weak". **Strong acids** are compounds that **completely dissociate** (break up) in water. All of the H^+ ions (protons) break away from the original acid molecule in water. A **weak acid** only **partially dissociates** and loses just some of its H^+ ions (protons) in water.

GZ Science Resources

Strong and Weak Acids

The strength of an acid is determined by how readily it will donate its H^+ ions. Strong acids will have a low pH (0-3) and include HCl, H_2SO_4 and HNO_3. Weak acids will have a higher pH (4-6). They are mostly organic acids and include CH_3COOH.

Strong acids

Donate protons (H^+) in aqueous solution to become completely dissociated.

$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

HCl gas dissolved in water

HCl has donated a H^+ so is acting as an acid

H_2O has accepted a H^+ so it is acting as a base

Solution contains **virtually no** intact HCl molecules after reaction.

Weak acids

Donate protons (H^+) in aqueous solution to become partially dissociated.

$CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

CH_3COOH dissolved in water

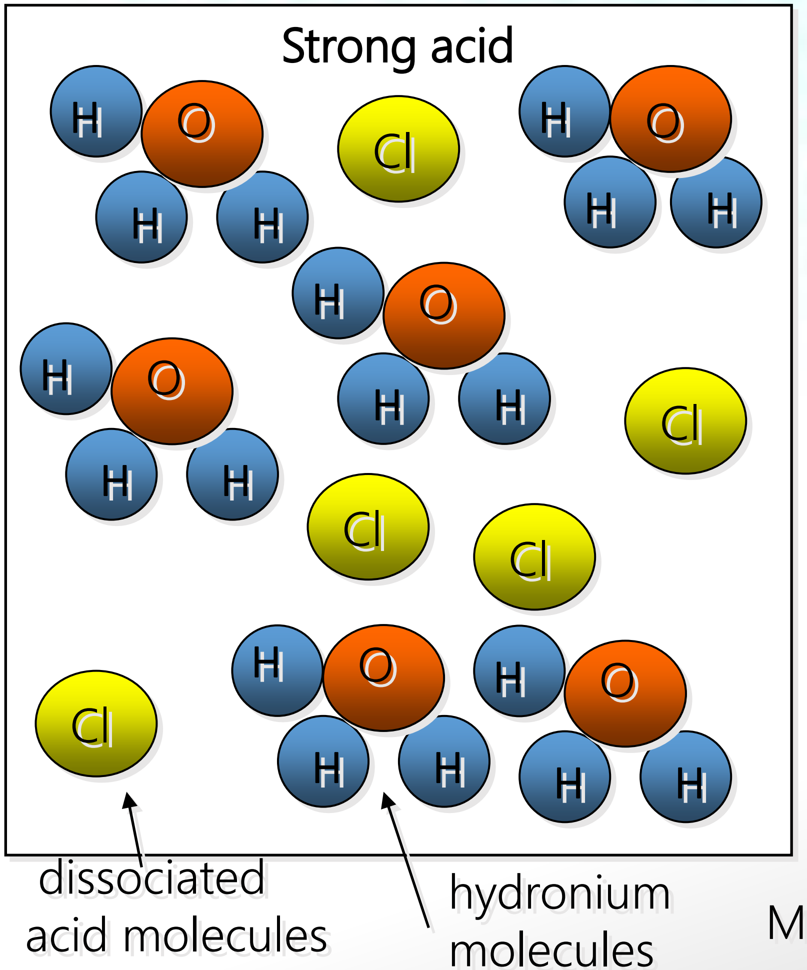
Only some of the acetic acid molecules dissociate into acetate ions (CH_3COO^-)

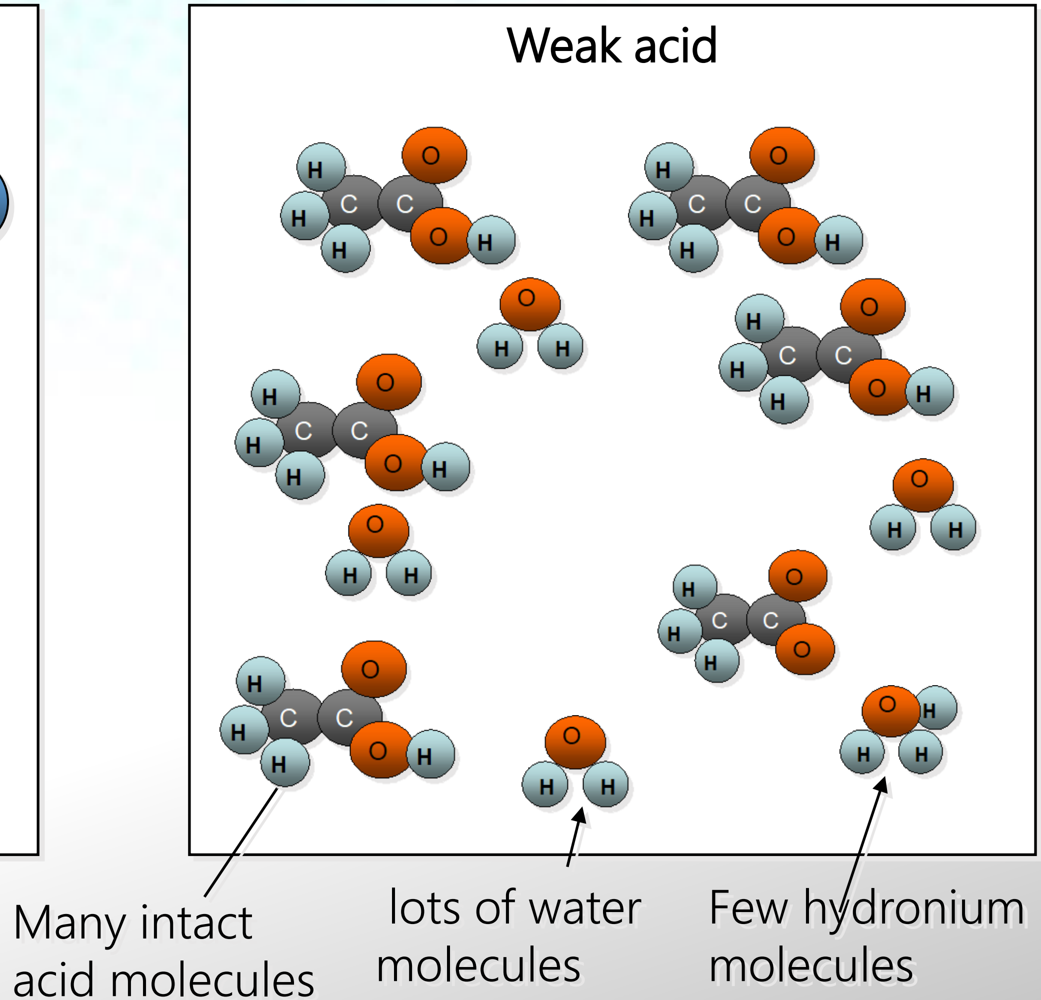
Because the acetate ion is a strong base (conjugate pairs) it will readily accept H^+ (from H_3O^+) and become acetic acid.

Solution contains **mostly** intact CH_3COOH molecules.

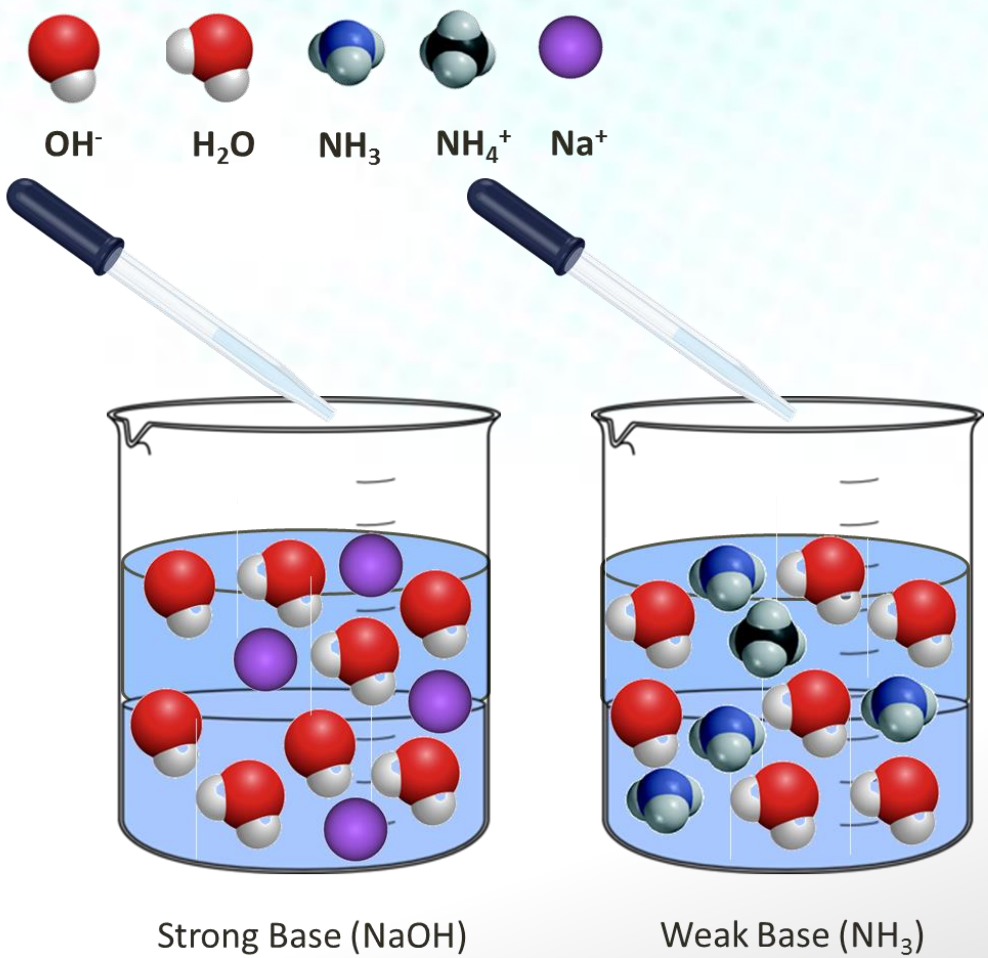
Strong and Weak Acids

In reality the strong acid molecules would be almost completely dissociated in an aqueous solution. The Cl^- would remain in solution and free H^+ ions would join with available water to form hydronium ions.





Strong and Weak bases



You can define bases as being "strong" or "weak". **Strong bases** are compounds where each molecule will accept an H^+ ion. A **weak base** is a compound where only some of the molecules will accept a H^+ ion. Most weak base molecules remain unreacted.

Note: For strong alkalis, all of the OH^- ions break away from the molecule in water.

Strong and Weak Bases

The strength of an base is determined by how readily it will accept H^+ ions. Strong bases will have a high pH (12-14) and include NaOH and KOH. Weak acids will have a lower pH (8-11). They include NH_3.

Strong Bases

Accept protons (H^+) in aqueous solution to become completely dissociated.

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

NaOH completely dissociates

The OH- ions will readily accept H^+ ions.

Solution contains very **few** intact NaOH molecules after reaction.

Weak Bases

Accept protons (H^+) in aqueous solution to become partly dissociated.

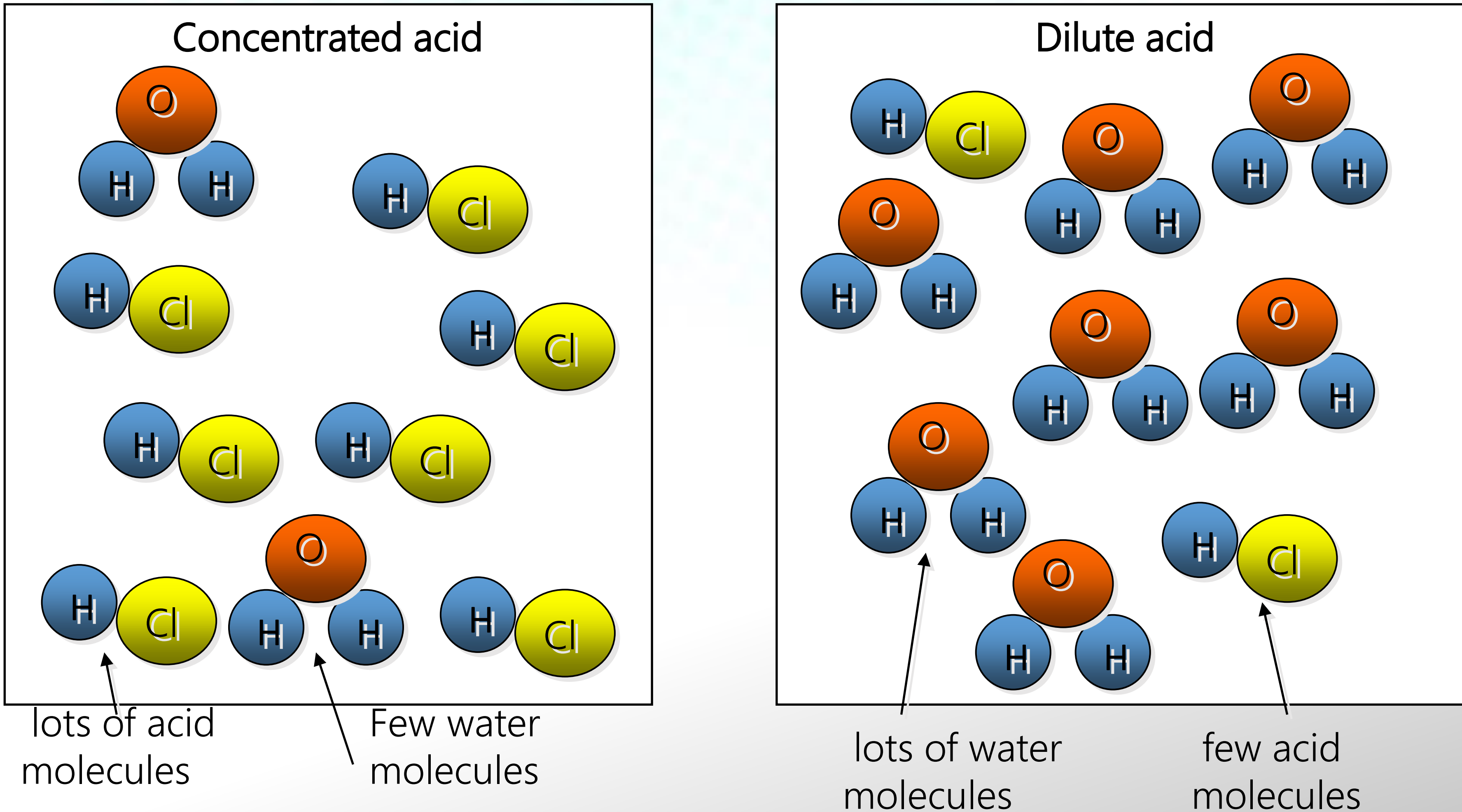
$$NH_{3(g)} + H_2O \leftrightharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Only some of the ammonia molecules dissociate into ammonium ions (NH_4^+)

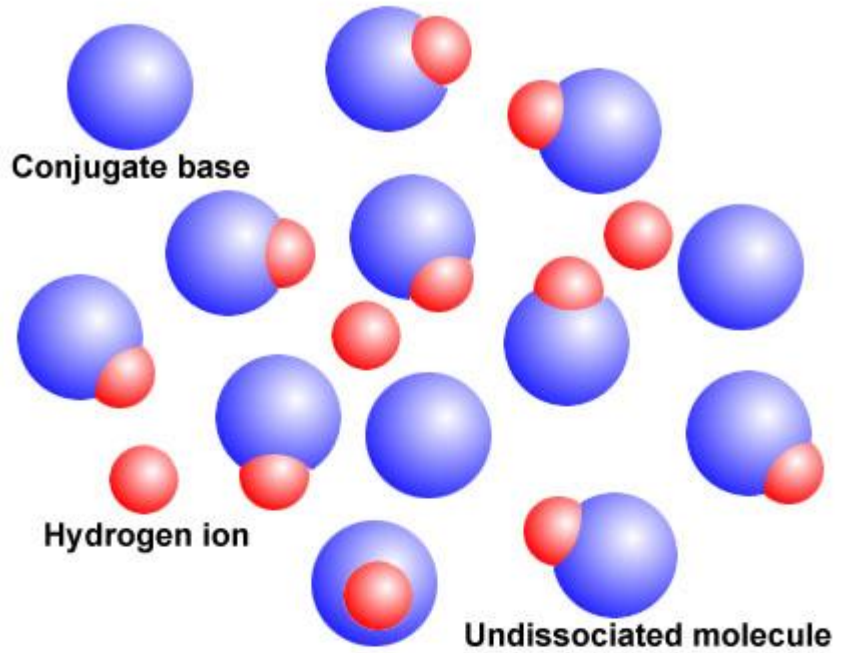
Because ammonium is a reasonably strong acid (conjugate pairs) it will readily donate H^+ and become ammonia. Solution contains **mostly** intact NH_3 molecules.

Concentrated and Dilute Acids

Concentration is determined by the number of molecules present per volume of solution. In <u>aqueous</u> solutions the solvent is water. $c = n/v$



Transfer of hydrogen ions in conjugate pairs



Weak acids dissociate only slightly in aqueous solution. The majority of molecules remain undissociated.

When a base accepts a proton, it becomes an acid because it now has a proton that it can donate. And when an acid donates a proton it becomes a base, because it now has room to accept a proton.

These are what we call **conjugate pairs** of acids and bases.

When an acid gives up its proton, what remains is called the **conjugate base** of that acid. When a base accepts a proton, the resulting chemical is called the conjugate acid of that original base.

ACID - Conjugate Acid and Base pairs

If 2 species differ by just 1 proton they are classed as a conjugate acid-base pair.
Examples of acid-base pairs are H_2SO_4/HSO_4^-, and NH_4^+/NH_3. The acid is always the species with the additional proton. It can also be said that NH_3 is the conjugate base of NH_4^+.

Acid		Conjugate Base
HCl	hydrochloric acid	Cl^-
H_2SO_4	sulfuric acid	HSO_4^-
HNO_3	nitric acid	NO_3^-
CH_3COOH	acetic acid	CH_3COO^-
NH_4^+	ammonium ion	NH_3
H_3PO_4	phosphoric acid	$H_2PO_4^-$

Conjugate Acid and Base pairs

←Relative acid strength increasing

Relative base strength increasing→

	ACID	BASE	
negligible	OH^-	O^{2-}	strong
negligible	HS^-	S^{2-}	strong
weak	H_2O	OH^-	weak
weak	HPO_4^{2-}	PO_4^{3-}	weak
weak	HCO_3^-	CO_3^{2-}	weak
weak	NH_4^+	NH_3	weak
weak	HCN	CN^-	weak
weak	$H_2PO_4^-$	HPO_4^{2-}	weak
weak	HSO_3^-	SO_3^{2-}	weak
weak	H_2S	HS^-	weak
weak	H_2CO_3	HCO_3^-	weak
weak	$C_5H_5NH^+$	C_5H_5N	weak
weak	CH_3CO_2H	$CH_3CO_2^-$	weak
weak	HF	F^-	weak
weak	H_3PO_4	$H_2PO_4^-$	weak
weak	H_2SO_3	HSO_3^-	weak
weak	HSO_4^-	SO_4^{2-}	weak
strong	H_3O^+	H_2O	weak
strong	HNO_3	NO_3^-	negligible
strong	H_2SO_4	HSO_4^-	negligible
strong	HCl	Cl^-	negligible
strong	HBr	Br^-	negligible

The stronger an acid, the weaker its conjugate base, and, conversely, the stronger a base, the weaker its conjugate acid.

A strong acid like HCl donates its proton so readily that there is essentially no tendency for the conjugate base Cl^- to reaccept a proton. Consequently, Cl^- is a very weak base. A strong base like the H^- ion accepts a proton and holds it so firmly that there is no tendency for the conjugate acid H_2 to donate a proton. Hence, H_2 is a very weak acid.

GZ Science Resources

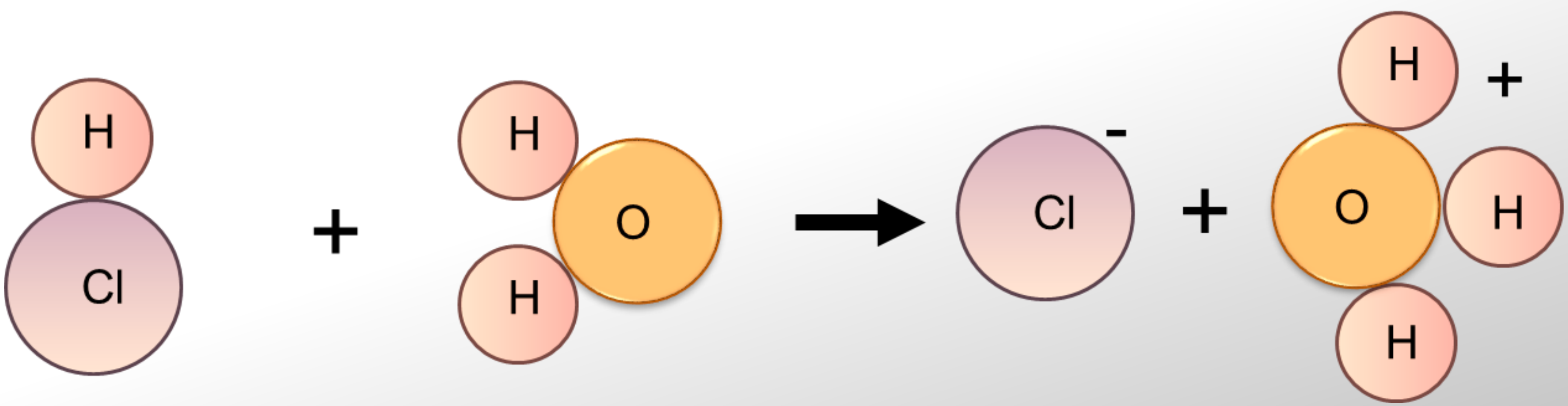
Conjugate Acid and Base pairs (Strong Acid)

HX is a symbol used for a strong acid. A conjugate acid can be seen as the chemical substance that releases a proton in the **backward** chemical reaction.
The base produced, X–, is called the **conjugate base** and it absorbs a proton in the **backward** chemical reaction.

paired　　　　paired

Strong acid		base		conjugate base		conjugate acid
HX	+	H_2O	$\longrightarrow$	X^-	+	H_3O^+

transfer of H^+



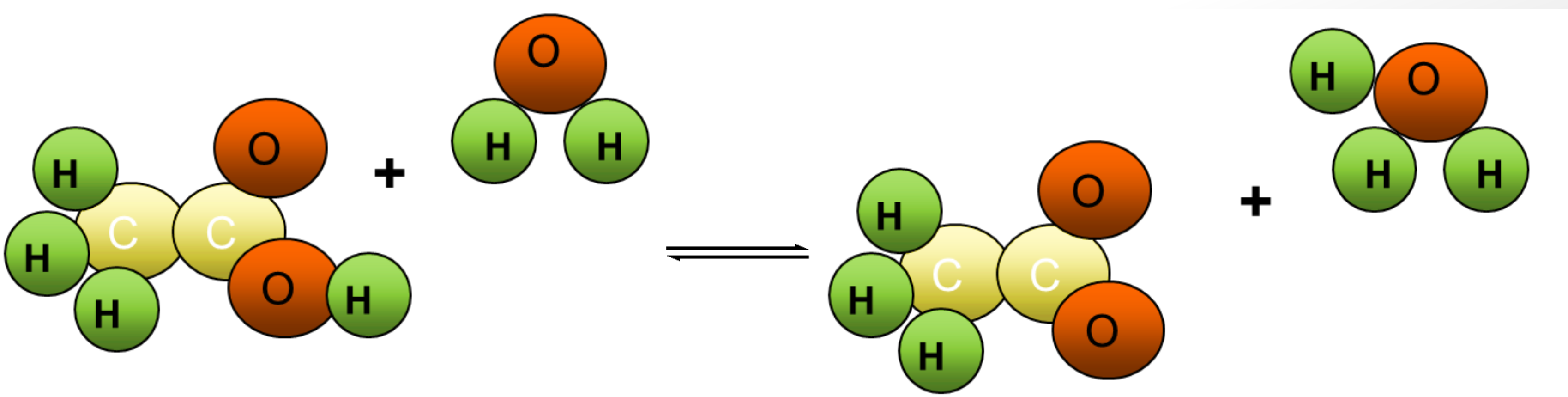
Conjugate Acid and Base pairs (weak acid)

HA is a symbol used for weak acid. Note the use of the double arrow. Because the weak acid only partially dissociates an equilibrium reaction occurs with a fixed amount of an acid and its conjugate remain in solution.

paired paired

Weak acid		base		conjugate base		conjugate acid
HA	$+$	H_2O	$\rightleftharpoons$	A^-	$+$	H_3O^+

transfer of H^+



BASE - Conjugate Acid and Base pairs

If 2 species differ by just 1 proton they are classed as a conjugate acid-base pair. Examples of acid-base pairs are H_2SO_4/HSO^{4-}, and NH_4^+/NH_3. The acid is always the species with the additional proton. It can also be said that NH_3 is the conjugate base of NH_4^+.

Base		Conjugate Acid
H_2O	water	H_3O^+
SO_4^{2-}	sulfate ion	HSO_4^-
NH_3	ammonia	NH_4^+
OH^-	hydroxide ion	H_2O
HCO_3^-	hydrogen carbonate ion	H_2CO_3
CO_3^{2-}	carbonate ion	HCO_3^-

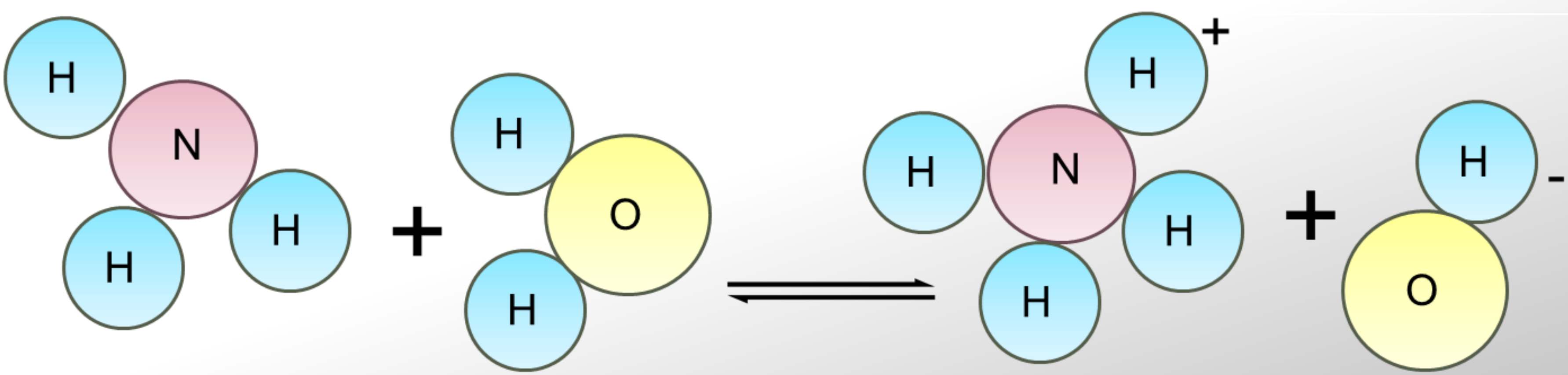
Conjugate Acid and Base pairs (Base)

B is a symbol used for a base. The base now accepts the hydrogen ion from the water. The hydroxide ion, OH^-, is the paired conjugate of the water once the H^+ has been removed. Strong bases use a single direction arrow and weak bases use a double arrow.

paired paired

base		acid		conjugate acid		conjugate base
B	$+$	H_2O	$\rightleftharpoons$	BH^+	$+$	OH^-

transfer of H^+



GZ Science Resources

Acid and Base Dissociation

Strong Acid

paired

paired

Strong acid		base		conjugate base		conjugate acid
HX	$+$	H_2O	$\longrightarrow$	X^-	$+$	H_3O^+

transfer of H^+

Weak Acid

paired

paired

Weak acid		base		conjugate base		conjugate acid
HA	$+$	H_2O	$\rightleftharpoons$	A^-	$+$	H_3O^+

transfer of H^+

Weak Base

paired

paired

base		acid		conjugate acid		conjugate base
B	$+$	H_2O	$\rightleftharpoons$	BH^+	$+$	OH^-

transfer of H^+

NCEA 2014 Acids and Bases

Merit Question

Question 1a: Ammonia, NH_3, is dissolved in water and the resulting solution has a pH of 11.3.

(i) Complete the equation by writing the formulae of the two products.

$NH_{3(aq)} + H_2O_{(l)} \rightarrow \qquad\qquad +$

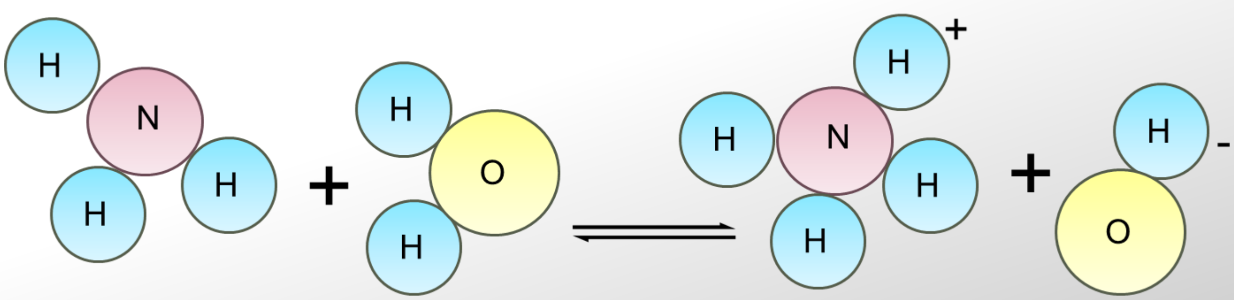
(ii) Explain what is occurring during this reaction.

In your answer you should: identify the acid and its conjugate base, identify the base and its conjugate acid, describe the proton transfer that occurs.

Answer:

$NH_3 + H_2O \rightarrow NH_4^+ \; + \; OH^-$

Protons are being transferred from the acid, H_2O to the base, NH_3. When the NH_3 gains a proton, it forms its conjugate acid, NH_4^+. When H_2O loses its proton, it forms its conjugate base, OH^-.



NCEA 2015 Acids and Bases

Excellence Question

Question 2(a): Ammonia solution, $NH_{3(aq)}$, is a common chemical in the school laboratory.
(i) Explain, using an equation, whether ammonia solution is acidic or basic.

Answer: Ammonia is basic, $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$.
When ammonia (partially) ionises in water, it produces hydroxide ions. (When hydroxide ions are in greater concentration than H_3O^+, the pH will be above 7 and therefore basic.)

Question 2(a): (ii) Bottles of ammonia solution are often labelled ammonium hydroxide, $NH_4OH_{(aq)}$.
Explain why both names, ammonia and ammonium hydroxide, are appropriate.

Answer: Ammonia is a weak base, and so doesn't ionise fully. This means that there are still many ammonia molecules in the reaction mixture with just some ammonium and hydroxide ions. Because all 3 species are present, the labels of ammonia or ammonium hydroxide are equally valid.

NCEA 2015 Acids and Bases

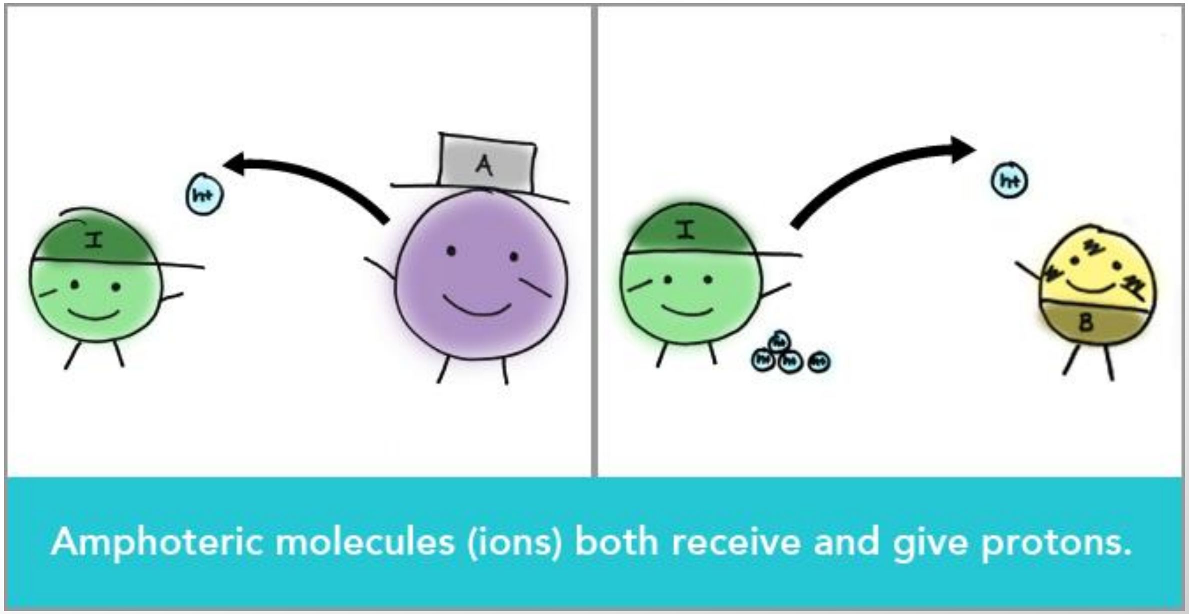
Achieved Question

Question 2(b): The hydrogen carbonate ion, HCO_3^-, is an amphiprotic species because it can donate or accept a proton, therefore acting as an acid or base. Write equations for the reactions of HCO_3^- with water: one where it acts as an acid, and one where it acts as a base.

Answer:

$HCO_3^- + H_2O \rightarrow CO_3^{2-} + H_3O^+$

$HCO_3^- + H_2O \rightarrow H_2CO_3 + OH^-$



NCEA 2016 Acids and Bases

Achieved Question

Question 2(a): Water is an amphiprotic substance because it can accept or donate a proton, therefore acting as an acid or a base.
Complete the equations for the reactions of water, H_2O, with ammonia, NH_3, and the ammonium ion, NH_4^+, in the box below.

H_2O acting as	Equation
an acid	$H_2O(\ell) + NH_3(aq) \rightleftharpoons$ $H_2O(\ell) + NH_3(aq) \rightleftharpoons OH^-(aq) + NH_4^+(aq)$
a base	$H_2O(\ell) + NH_4^+(aq) \rightleftharpoons$ $H_2O(\ell) + NH_4^+(aq) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$

NCEA 2016 Acids and Bases

Merit Question

Question 2(b): Sodium carbonate, $Na_2CO_3(s)$, is a salt. When dissolved in water, it dissociates into ions.
Explain whether a solution of sodium carbonate would be acidic or basic.
In your answer you should include TWO relevant equations.

Answer:

When sodium carbonate dissolves in water, it dissociates into ions.

EQUATION ONE:

$Na_2CO_3(s) \rightarrow 2Na^+(aq) + CO_3^{2-}(aq)$

The carbonate ions then react in water.

EQUATION TWO:

$CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

The OH^- ions produced make this **a basic (alkaline) solution**.

NCEA 2017 Acids and Bases

Merit Question

Question 1a: Propanoic acid, C_2H_5COOH, is dissolved in water and the resulting solution has a pH of 4.2.

(i) Complete the equation by writing the formulae of the two products.

$C_2H_5COOH_{(aq)} + H_2O_{(l)} \leftrightharpoons C_2H_5COO^-_{(aq)} + H_3O^+_{(aq)}$

(ii) Explain the proton, H^+, transfer in this reaction, and identify the two conjugate acid-base pairs.

Propanoic acid transfers a proton / H^+ to the water so it is the acid and it forms a conjugate base, $C_2H_5COO^-$. The water, H_2O, accepts the proton and therefore acts as a base, forming H_3O^+, hydronium ion, which is the conjugate acid.

NCEA 2017 Acids and Bases

Merit Question

Question 1b: Sodium ethanoate, $CH_3COONa_{(s)}$, is a salt. When dissolved in water, it dissociates into ions.
Explain, including TWO relevant equations, whether a solution of sodium ethanoate is acidic or basic.

EQUATION ONE: $CH_3COONa_{(s)} \rightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)}$
The salt dissociates / dissolves into ions in water.
The ethanoate ions then react with water to form $CH_3COOH_{(aq)}$ and $OH^-_{(aq)}$.

EQUATION TWO: $CH_3COO^-_{(aq)} + H_2O_{(\ell)} \leftrightharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$
The OH^- ions produced make this a basic solution.

Excellence Question

Question 1b (iv) : Another chemical in solution **X** is a salt, sodium ethanoate, CH_3COONa. When solid sodium ethanoate is dissolved in water, it separates into ions. Use TWO relevant equations to explain whether **the solution** is acidic or basic.

When sodium ethanoate dissolves in water, the ions separate.
EQUATION ONE: $CH_3COONa_{(s)} \rightarrow CH_3COO^-_{(aq)} + Na^+_{(aq)}$
The ethanoate ions then react with water.

EQUATION TWO:
$CH_3COO^-_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$

The OH^- ions produced make this a **basic solution.**

NCEA 2018 Acids and Bases

Achieved Question

Question 3a: The hydrogensulfate ion, HSO_4^-, is an amphiprotic species because it can both accept or donate a proton, thus acting as an acid or base. Complete the equations for the reactions of the hydrogensulfate ion, HSO_4^-, with water in the box below.

HSO_4^- acting as	Equation
an acid	$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons$ $SO_4^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$
a base	$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons$ $H_2SO_{4(aq)} + OH^-{}_{(aq)}$

NCEA 2019 Acids and Bases

Merit Question

Question 3a: (a) Nitric acid, $HNO_{3(aq)}$, and ethanoic acid, $CH_3COOH_{(aq)}$, are both acids.
(i) Write equations to show their reactions with water, $H_2O(l)$.

$HNO_3(aq) + H_2O(\ell)$

$CH_3COOH(aq) + H_2O(\ell)$

(ii) Use these equations to explain why they are classified as acids.

When these substances dissociate / ionise in water, they produce hydronium ions / donate protons.
OR
When hydronium ions are in greater concentration than hydroxide ions (OH^-), the pH will be below 7 and therefore acidic.

Achieved Question

Question 1a: (i) Sodium hydrogen carbonate, $NaHCO_3$, is a salt and will dissociate into ions when dissolved in water.
Write an equation for this process.

$$NaHCO_3 \rightarrow Na^+ + HCO_3^-$$

(ii) One of the ions formed from the dissociation is amphiprotic because it can either accept or donate a proton.
Write equations for each of these reactions..

Acting as:	**Equation**
an acid	$HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$
a base	$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$

Weak and Strong – Acids and Bases

Strong Acid	Weak Acid
Single arrow $\rightarrow$ produces H_3O^+	double arrow $\rightleftharpoons$ Produces H_3O^+
$HCl + H_2O \rightarrow Cl^- + H_3O^+$	$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$
$H_2SO_4 + H_2O \rightarrow HSO_4^- + H_3O^+$	$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$
$HNO_3 + H_2O \rightarrow NO_3^- + H_3O^+$	$HF + H_2O \rightleftharpoons F^- + H_3O^+$

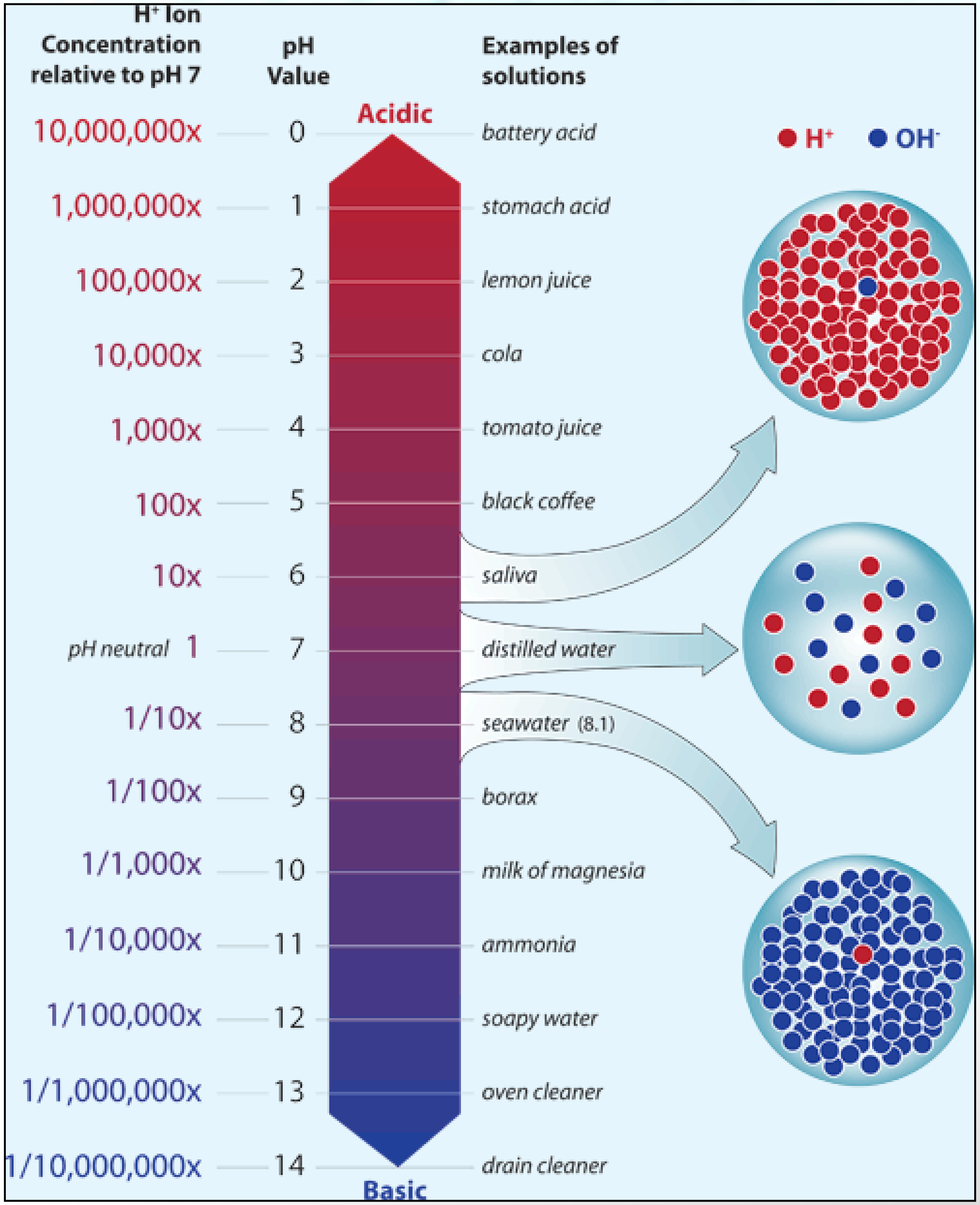
Strong Base	Weak Base
Single arrow $\rightarrow$ Produces OH^-	double arrow $\rightleftharpoons$ Produces OH^-
$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$	$SO_4^{2-} + H_2O \rightleftharpoons HSO_4^- + OH^-$
	$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
	$CH_3COO^-{}_{(aq)} + H_2O_{(\ell)} \leftrightharpoons CH_3COOH_{(aq)} + OH^-{}_{(aq)}$

Weak and Strong – Acids and Bases

H_2SO_4 SO_4^{2-} HCl NH_4^+ HCO_3^- OH^- CO_3^{2-} NH_3
CH_3COOH HNO_3

Strong Acid	Weak Acid	Strong Base	Weak Base
Single arrow $\rightarrow$ Produces H_3O^+	double arrow $\rightleftharpoons$ Produces H_3O^+	Single arrow $\rightarrow$ Produces OH^-	double arrow $\rightleftharpoons$ Produces OH^-

The pH scale measures level of acidity and alkalinity



The **pH scale** measures how acidic or alkaline a substance is. Substances with a pH of 7 are **neutral**, substances with a pH greater than 7 are **alkaline** (or 'basic') and substances with a pH lower than 7 are **acidic**. Alkalis are 'bases' that are soluble in water. (All alkalis are bases but not all bases are alkalis.) The pH of a substance is determined by the concentration of hydrogen ions. The higher the concentration of hydrogen ions the lower the pH.

Measuring the Acidity or Alkalinity of solutions

Strong Acids	**Weak Acids**	**Neutral solution**	**Weak Bases**	**Strong Bases**
Readily donate all their protons when dissolved	donate only a small proportion of protons		Accept only a small proportion of protons	Readily accept protons

Concentration of H_3O^+ ions is **greater** than that of OH^- ions		Concentration of H_3O^+ ions is the **same** as that of OH^- ions		Concentration of H_3O^+ ions is **less** than that of OH^- ions

← Increasing acidity | Increasing alkalinity →

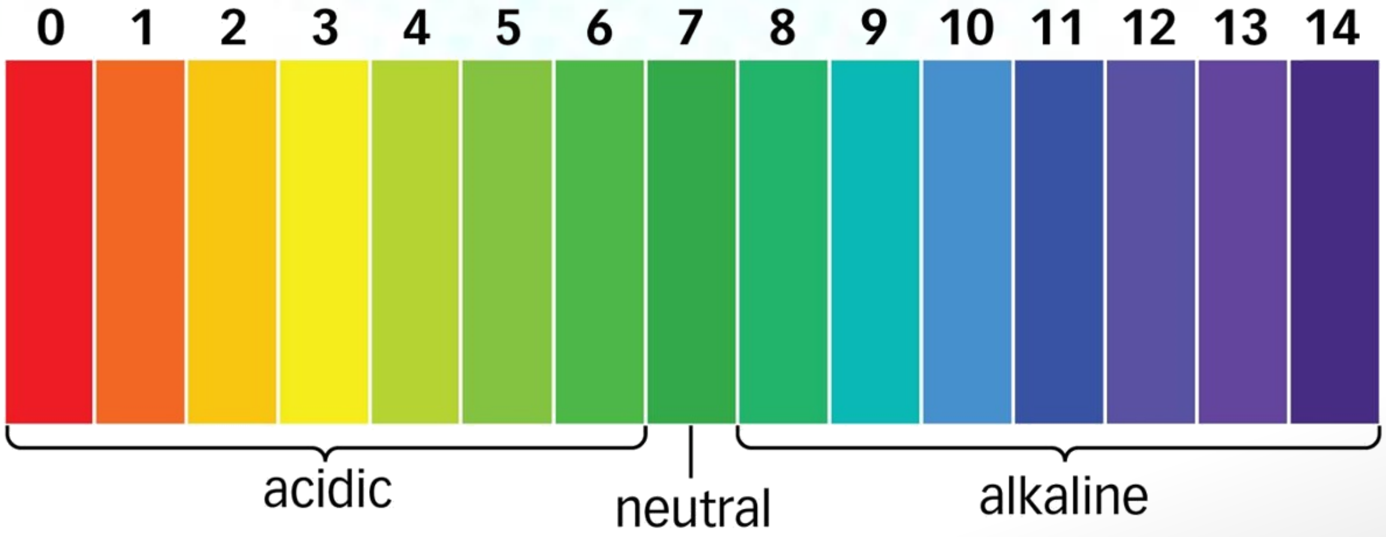
Amphiprotic Substances can act as acid or base e.g. H_2O

Describe solutions as acidic, alkaline or neutral in terms of the pH scale.

Acids have a pH less than 7

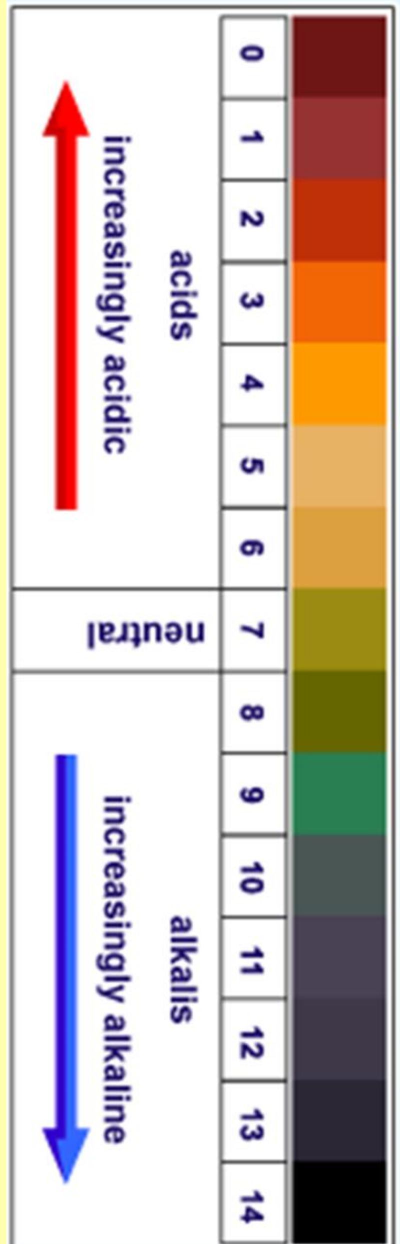
Neutral substances have a pH of 7

Alkalis have pH values greater than 7



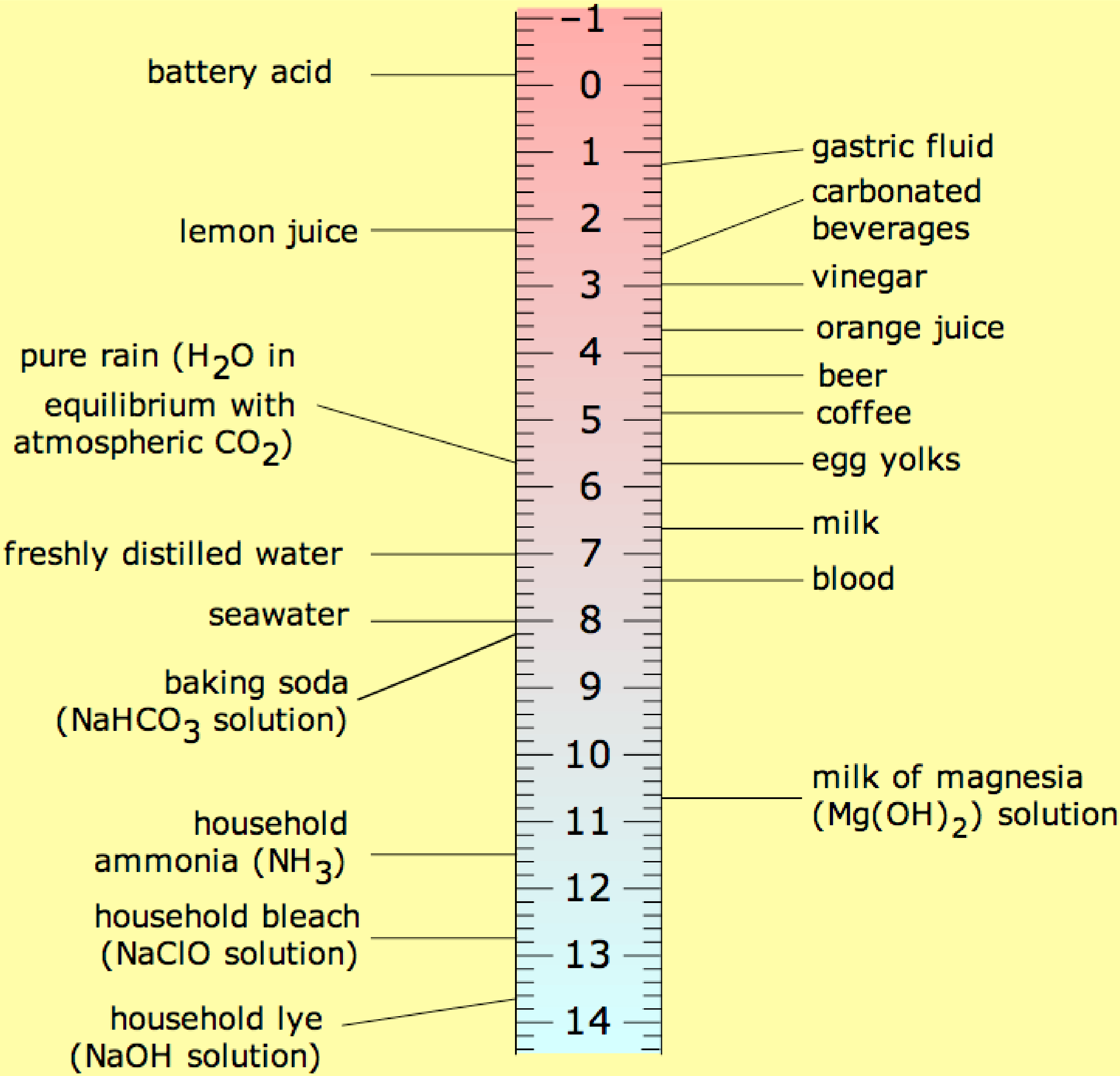
The pH scale is **logarithmic** and as a result, each whole pH value below 7 is ten times more acidic than the next higher value. For example, pH 4 is ten times more acidic than pH 5 and 100 times more acidic than pH 6. Universal indicator can be used to give the colours above, and therefore show the pH of a solution.

pH values of common substances



Substance group	Example
Strong Mineral Acids	HNO_3, HCl
Weak Organic Acids	CH_3COOH
Water, Chlorides	$NaCl$, H_2O
Carbonates	$NaHCO_3$
Oxides	MgO
Most Hydroxides	$Mg(OH)_2$
Ammonia	NH_3
Sodium Hydroxide	$NaOH$

Describe solutions as acidic, alkaline or neutral in terms of the pH scale.



Pure water is **neutral**. But when chemicals are mixed with water, the mixture can become either acidic or basic. Examples of acidic substances are vinegar and lemon juice. Lye, milk of magnesia, and ammonia are examples of basic substances

GZ Science Resources

Putting it all together

Blue litmus					
Red litmus					
Universal indicator					
pH	**1-3**	**4-6**	**7**	**8-10**	**11-14**
description	**Strong Acids** Readily donate all their protons when dissolved	**Weak Acids** donate only a small proportion of protons	**Neutral solution**	**Weak Bases** Accept only a small proportion of protons	**Strong Bases** Readily accept protons
H_3O^+ / OH^- concentration	Concentration of H+ ions is **greater** than that of OH- ions		Concentration of H+ ions is the **same** as that of OH- ions		Concentration of H+ ions is **less** than that of OH- ions

Increasing acidity ⟵⟶ Increasing alkalinity

Acid and Base Ratio of $[OH^-]$ or $[H_3O^+]$

When a strong acid dissociates into its conjugate and H_3O^+ ions then 100% of its molecules will dissociate.

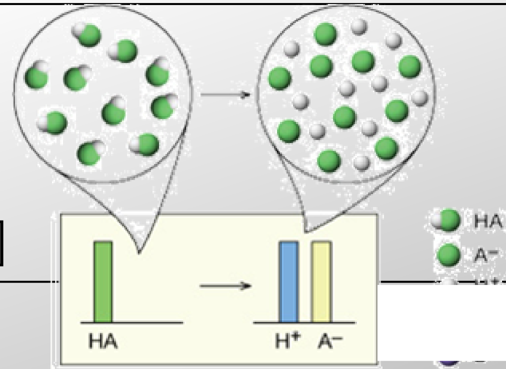
The concentration of H_3O^+ ions produced depends upon the Strong Acid involved.

HCl and HNO_3 acid concentration produces the same concentration of H_3O^+ ions. i.e. $0.123 molL^{-1}$ of HCl will produce $0.123 molL^{-1}$ $[H_3O^+]$ 1:1 ratio as those acids have one H^+ ion per molecule to donate

Note: 1 mol of H_2SO_4 acid produces 1 mol of H^+ ions and 1 mol of HSO_4^- ions. The HSO_4^- acts as a weak acid and further dissociation only produces a small amount of H^+ ions.

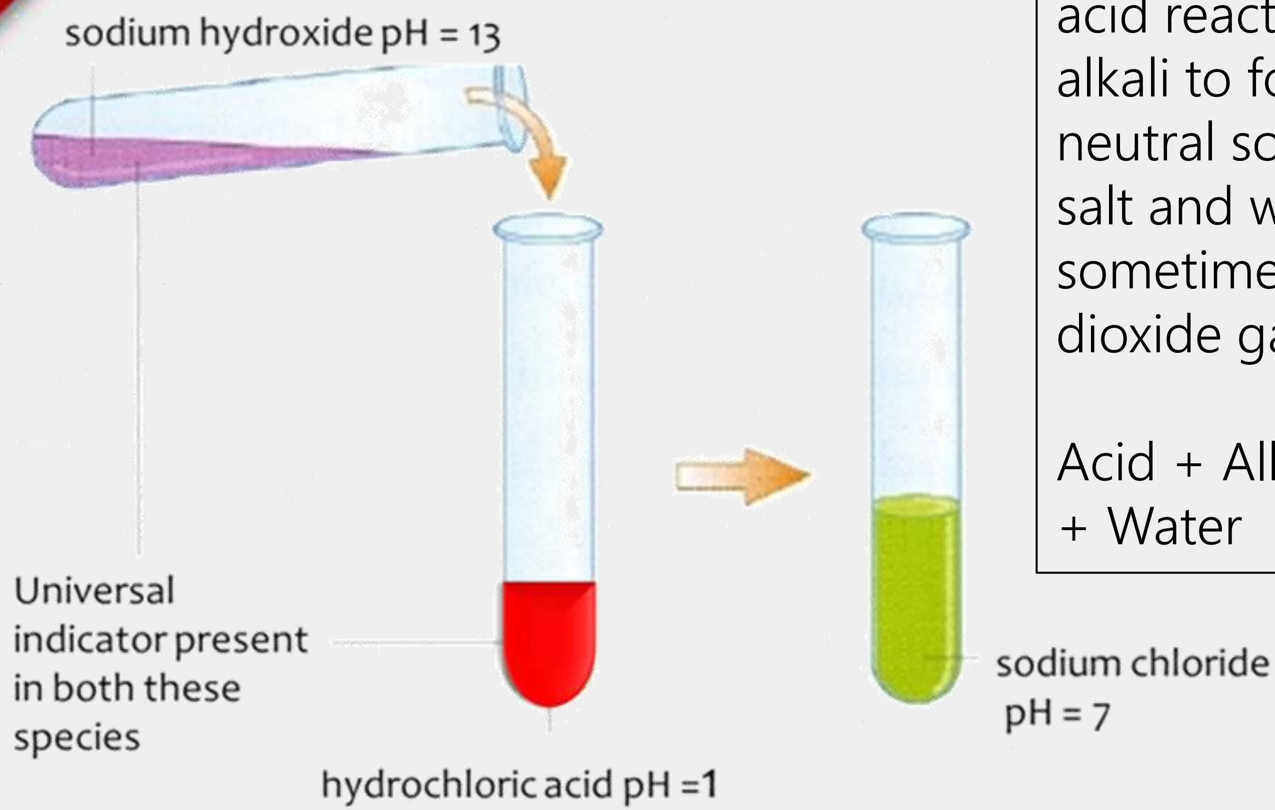
However, at the completion of the reaction with NaOH the ratio will be 2:1, acid:base

Strong Alkalis produce OH^- ions.
NaOH Alkali (Base) concentration produces the same concentration of OH^- ions
i.e. $0.267 molL^{-1}$ of NaOH will produce $0.267 molL^{-1}$ $[OH^-]$



Background Knowledge

During neutralisation reactions, hydrogen ions combine with hydroxide ions to form water molecules.

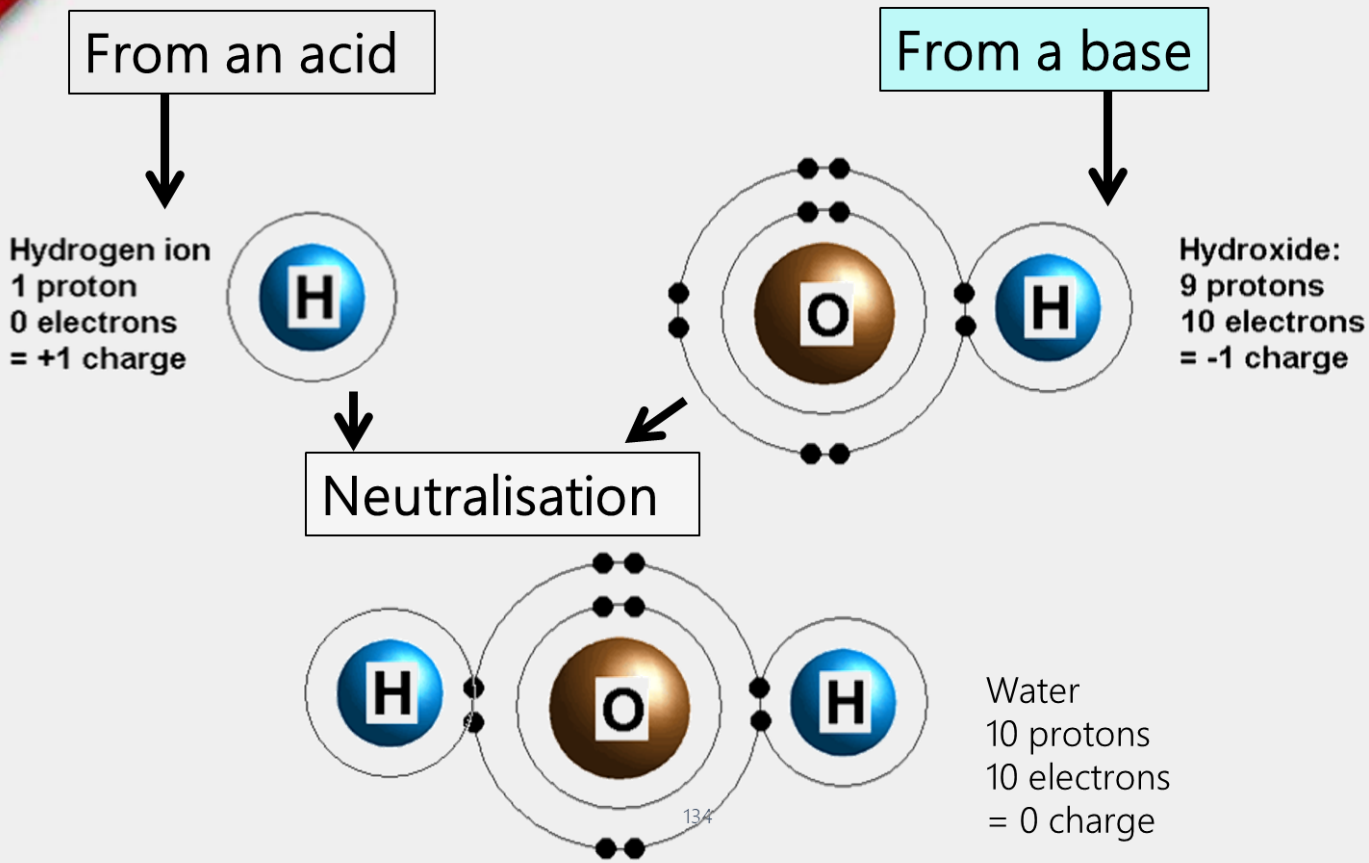


Neutralisation is a reaction where an acid reacts with an alkali to form a neutral solution of a salt and water. (and sometimes carbon dioxide gas)

Acid + Alkali → Salt + Water

Background Knowledge

During neutralisation reactions hydrogen ions combine with hydroxide ions to form water molecules



Background Knowledge

GZ Science Resources

Names of salts

When salts are formed the name depends upon the **acid reacted** and the **metal** that forms part of the base compound.

Name of acid	Name of salt formed	Formula of ion
hydrochloric acid	chloride	Cl^-
sulfuric acid	sulfate	SO_4^{2-}
nitric acid	nitrate	NO_3^-

One mole of acid is required to neutralize one mole of base

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Hydrochloric Acid | Sodium Hydroxide | Sodium Chloride | Water

Species in solution

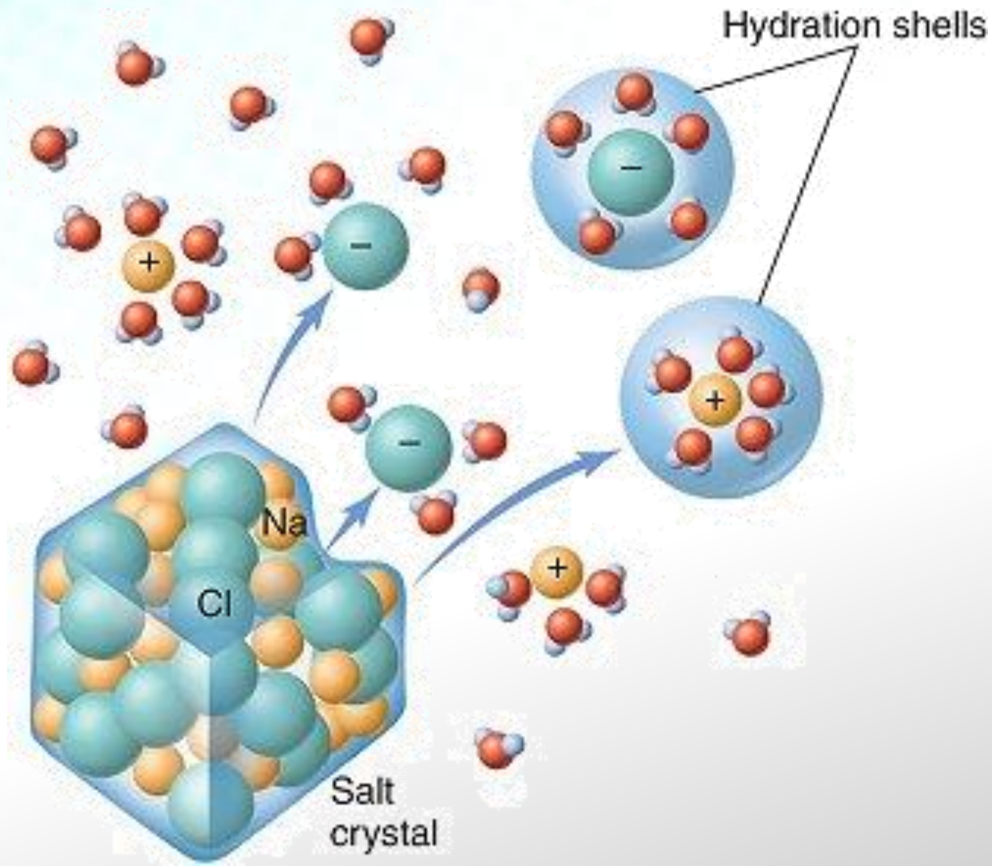
A solution is formed by mixing a **solute** (a dissolved substance) into a **solvent** (the solution that dissolves the solute.

The solvent is water, and the solute can be an acid, base or ionic salt.

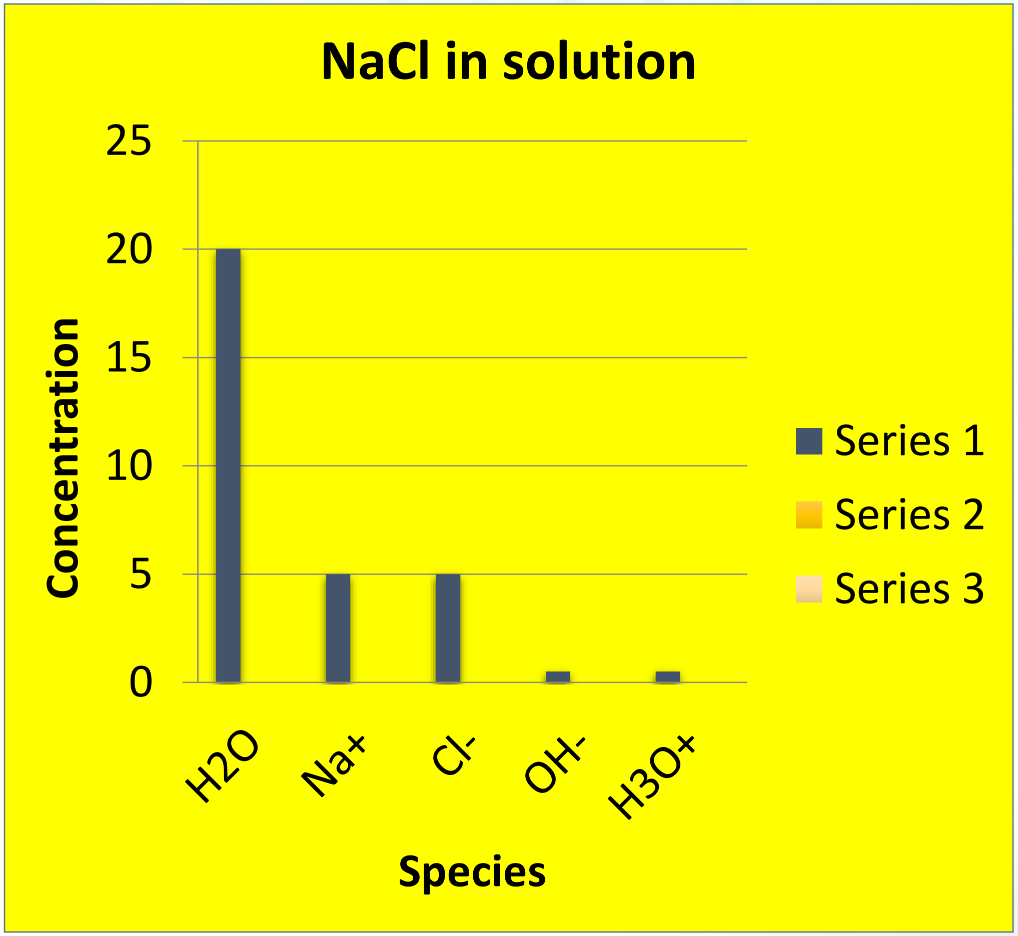
A solute dissolves by bonds being broken between solute particles (endothermic) and new bonds being formed between solute and solvent (exothermic).

A small amount of H_3O^+ and OH^- will always be present in water due to $Kw = [OH^-] [H_3O^+] = 1 \times 10^{-14}$

Water will always be present in large concentrations.



Concentration of Species in solution



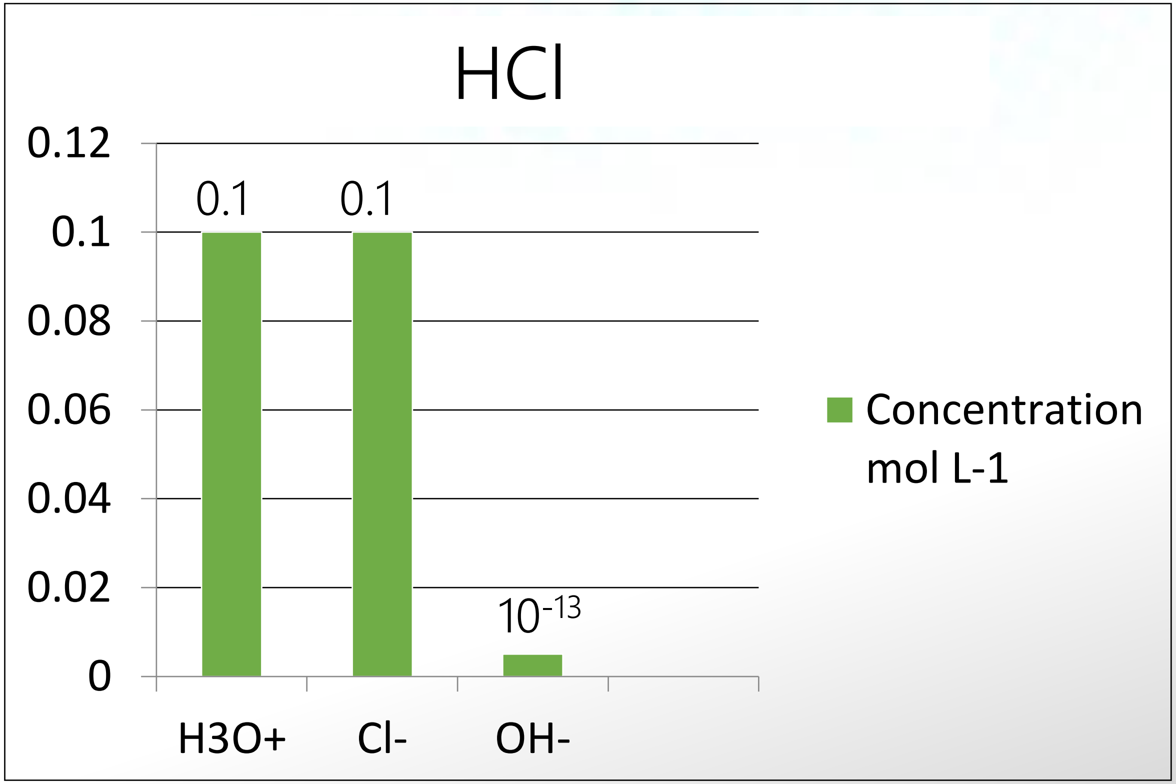
The relative concentration of the species in solution at equilibrium will depend upon the type of substances dissolved into water initially.

- In aqueous solutions water will almost always be present in the highest concentration.
- Small quantities of H_3O^+ and OH^- will also be present, according to the K_w = $[H_3O^+]$ $[OH^-]$ = 1×10^{-14}
- Information on relative concentration can often be presented in a bar graph.

Concentration of ions in solution – Strong Acid

Strong Acid i.e. HCl reacting with water
$HCl + H_2O \rightarrow Cl^- + H_3O^+$

Strong acids will provide **good conductivity** and **pH 1-2** due to the high presence of H_3O^+ ions



No strong acid will be left in the final mixture.

H_3O^+ and Cl^- are produced in equal concentrations – in the same concentration as the original strong acid.

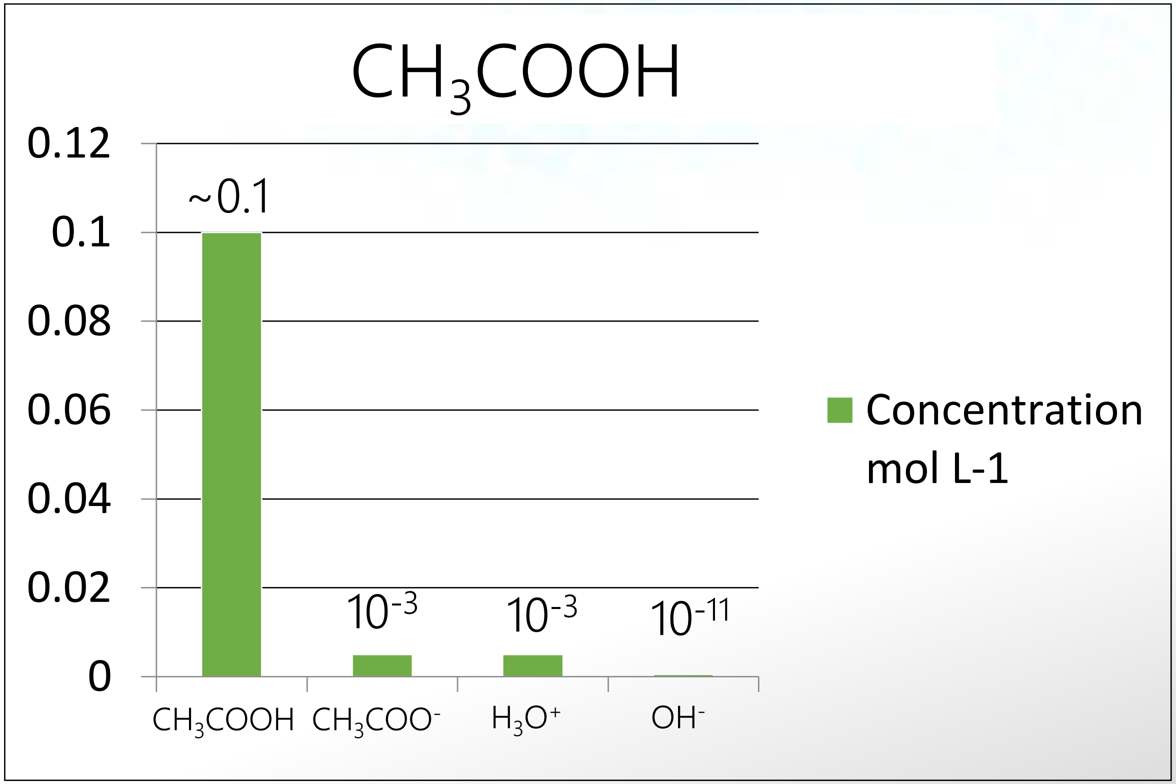
A small amount of OH^- is present as water dissociates into H_3O^+ and OH^-

Concentration of ions in solution – Weak Acid

Weak Acid i.e. CH_3COOH reacting with water

$$CH_3COOH + H_2O \leftrightharpoons CH_3COO^- + H_3O^+$$

Weak acids will provide **poor conductivity** and **pH 3-6** due to the low presence of H_3O^+ ions



Most weak acid will be left in the final mixture.

Small amounts of H_3O^+ and CH_3COO^- are produced in equal concentrations as only a small amount of the weak acid had dissociated.

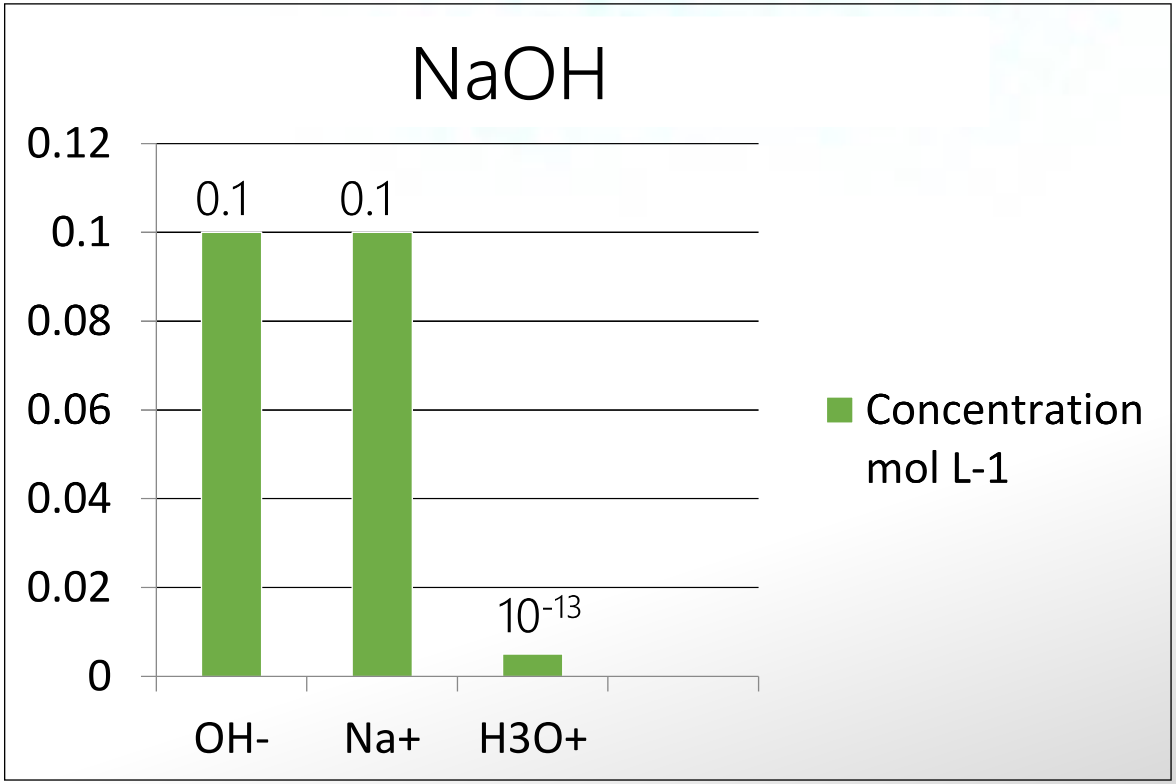
A small amount of OH^- is present as water dissociates into H_3O^+ and OH^-

Concentration of ions in solution – Strong Base

Strong Base i.e. NaOH reacting with water

$NaOH + H_2O \rightarrow Na^+ + OH^-$

Strong bases will provide **good conductivity** and **pH 12 – 14** due to the high presence of OH^- ions and Na^+ ions



No strong base will be left in the final mixture.

OH^- and Na^+ are produced in equal concentrations – in the same concentration as the original strong base.

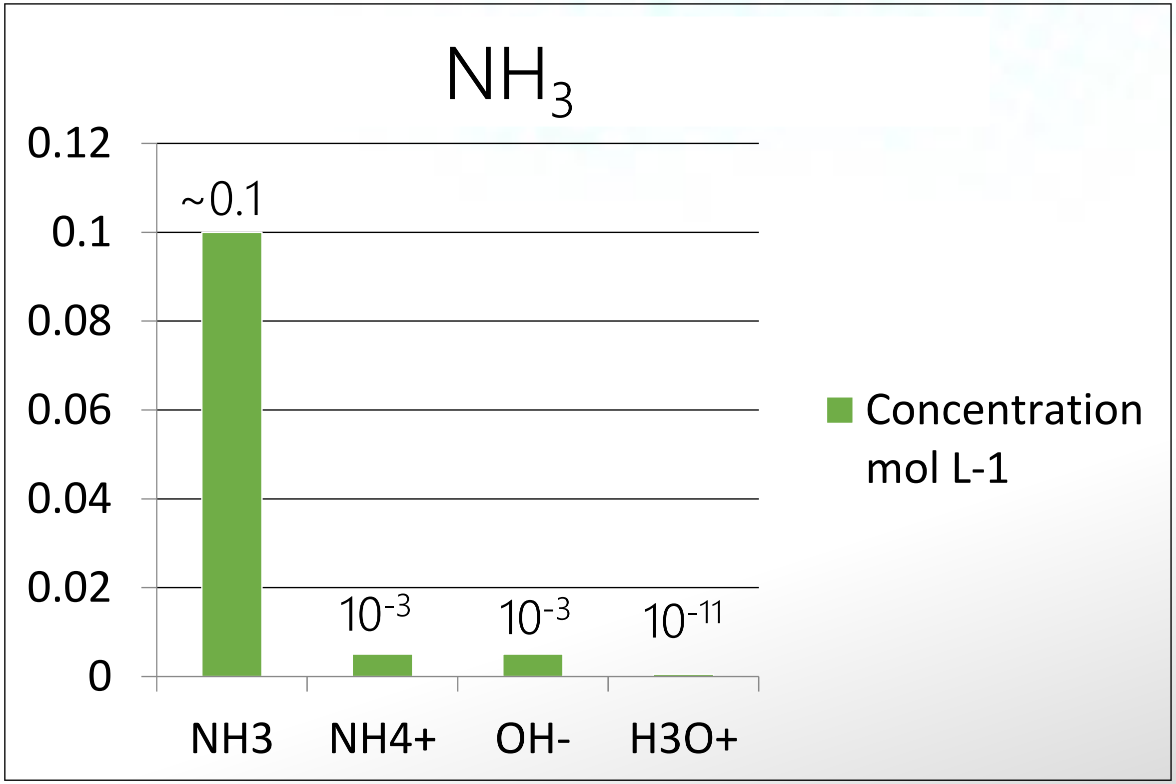
A small amount of H_3O^+ is present as water dissociates into H_3O^+ and OH^-

Concentration of ions in solution – Weak Base

Weak Base in water i.e. NH_3

$NH_3 + H_2O \leftrightharpoons NH_4^+ + OH^-$

Weak bases will provide **poor conductivity** and **pH 8 – 11** due to the low presence of OH^- ions, and the weak base remaining is a neutral substance.



Most weak base will be left in the final mixture.

OH^- and NH_4^+ are produced in equal concentrations – a small amount of the weak base had dissociated.

A small amount of H_3O^+ is present as water dissociates into H_3O^+ and OH^-

pH of salt solutions

All cations that are the conjugate acids of weak bases act as weak acids and lower the pH of the solution. This means that a salt solution containing this cation could be acidic. For example, a solution of ammonium chloride, NH_4Cl, contains the cation NH_4^+ and the anion Cl^-. The Cl^- ion acts as a neutral species and does not affect the pH (as it is the conjugate base of a strong acid and is so weakly basic that it effectively has no reaction with water). The NH_4^+ ion is the conjugate acid of the weak base NH_3 and so itself is a weak acid.

The ionic salt will first dissolve into its two ions. This equation needs to be shown. There will then be a further equation as the ion acting as a weak acid or base undergoes a acid/base reaction with water.
The non-reacting ion is left off as the spectator.

$$NH_4Cl_{(s)} \leftrightharpoons NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \leftrightharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$$

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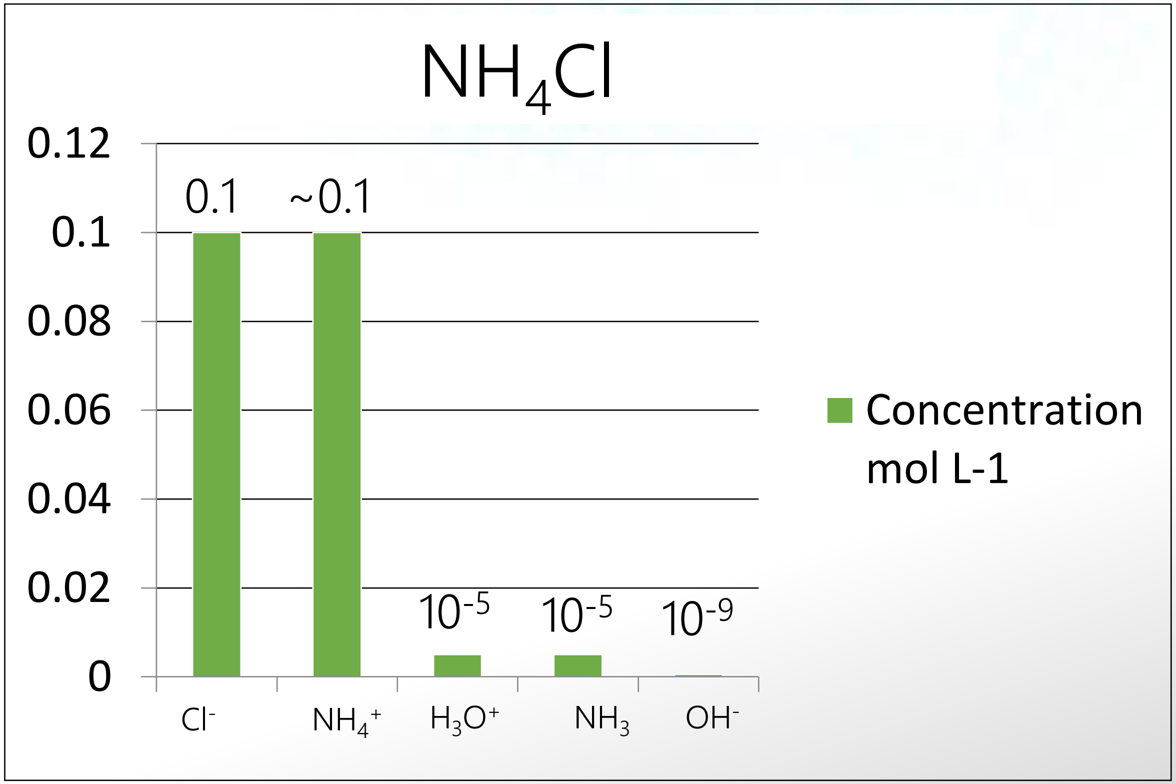
Concentration of ions in solution – acid salt

Acid Salt i.e. NH_4Cl

$NH_4Cl \rightarrow Cl^- + NH_4^+$

$NH_4^+ + H_2O \leftrightharpoons NH_3 + H_3O^+$

Acid salts will provide **good conductivity** and **pH < 7** due to the high presence of ions from dissolving and to a lesser extent H_3O^+ ions



The spectator ion will be left in the highest concentration followed by the weak acid.

H_3O^+ and NH_3 are produced in equal concentrations – a small amount of the weak acid had dissociated.

A small amount of OH^- is present as water dissociates into H_3O^+ and OH^-

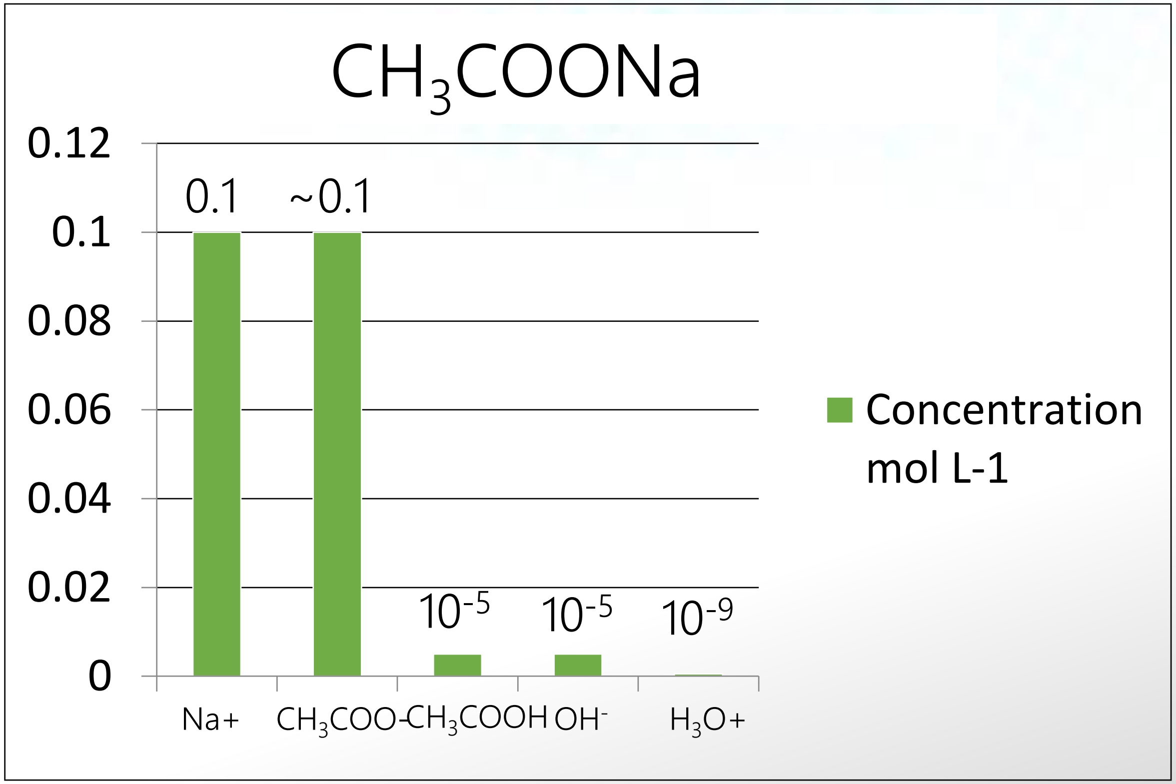
Concentration of ions in solution – base salt

Base Salt i.e. CH_3COONa

$CH_3COONa \rightarrow Na^+ + CH_3COO^-$

$CH_3COO^- + H_2O \leftrightharpoons CH_3COOH + OH^-$

Base salts will provide **good conductivity** and **pH > 7** due to the high presence of ions from dissolving and to a lesser extent OH^- ions



The spectator ion will be left in the highest concentration followed by the weak base.

OH^- and CH_3COOH are produced in equal concentrations – a small amount of the weak base had dissociated.

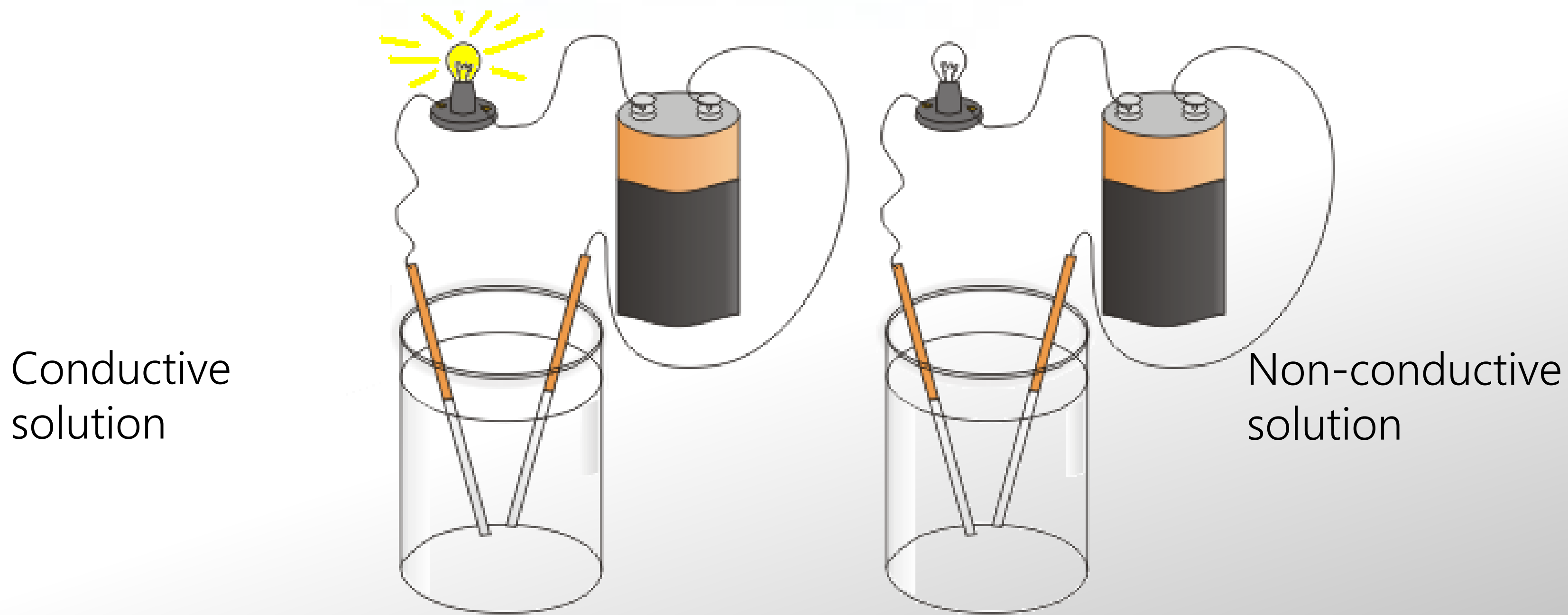
A small amount of H_3O^+ is present as water dissociates into H_3O^+ and OH^-

Conductivity of Acids and Bases

Conductivity is related to the availability of free moving charged particles. The presence of H_3O^+ or OH^- ions in solution and the concentration of them determine conductivity, along with any other ions.

<u>High conductivity – strong acids and Bases</u>

A strong electrolyte (solution containing ions) is created when a strong acid /strong base is added to water and fully dissociates. The H_3O^+ or OH^- ions carry the charge

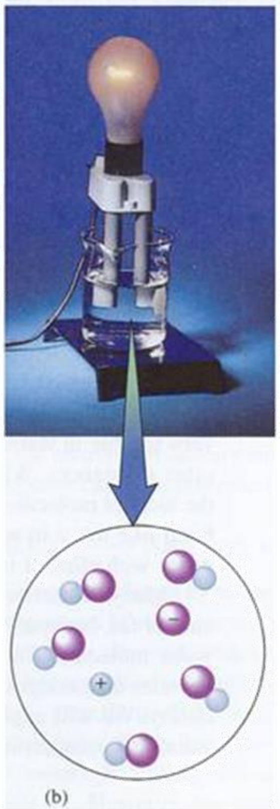


Conductivity of Acids and Bases

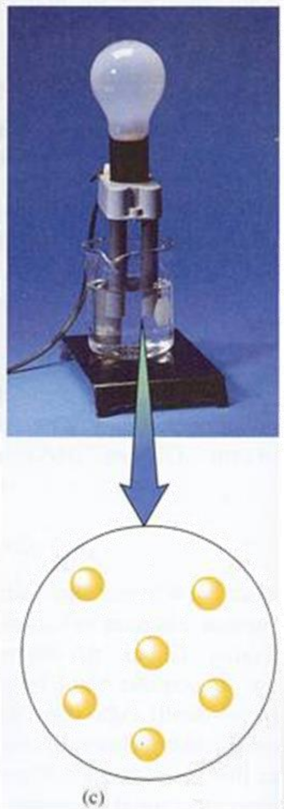
Low Conductivity – Weak acids and Bases

A weak electrolyte is formed from a weak acid or base that only partially dissociates. Only a small concentration of H_3O^+ or OH^- ions are created to carry charge. (such as ethanoic acid)

Conductive acids and bases	Poorly conductive acids and bases
HCl (hydrochloric acid) H_2SO_4 (sulfuric acid) HNO_3 (nitric acid) NaOH (sodium hydroxide) $NaCO_3$ (sodium carbonate) Na_2O (sodium oxide)	NH_3 (ammonia) CH_3COOH (ethanoic acid)



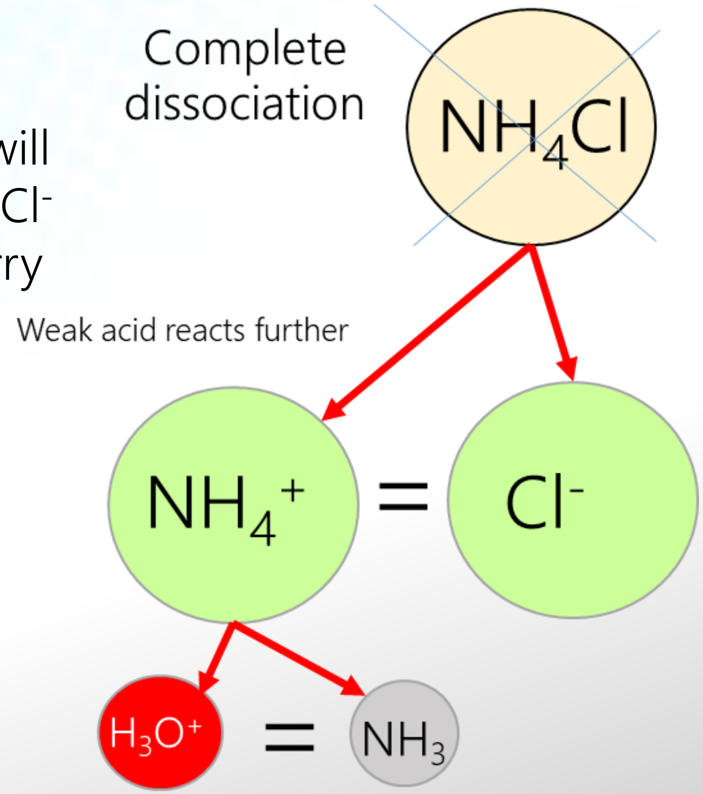
Acetic acid (weak acid)



Molecular substance

Conductivity of salts (weak acids – ethanoate salts)

If the acid used is a **soluble ionic solid**, for example ammonium chloride, then before it reacts with the base it dissolves in the water into its two ions; NH_4^+ and Cl^-. This means that even though a NH_4^+ is a weak acid only dissociating to release a small proportion of H_3O^+ ions (and its conjugate NH_3) it will still **be a good conductor** because of the NH_4^+ and Cl^- ions in solution. These are free moving ions that carry charge.
When drawing equations remember to show NH_4Cl dissolving into the two ions first before NH_4^+ ion further reacts as an acid. (2 equations required)
The Cl- ion remains as a spectator ion in the acid base reaction, and doesn't react further.

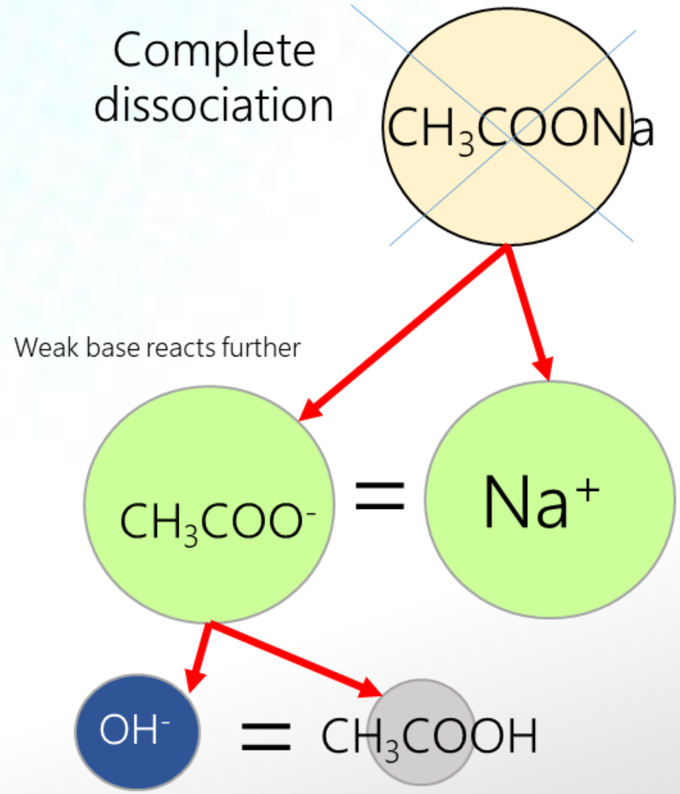


$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$
$$NH_4^+ + H_2O \leftrightharpoons NH_3 + H_3O^+$$

Note the single arrow for salt dissolving and double arrow for weak acid dissociation

Conductivity of salts (weak bases – ethanoate ion)

If the base used is a **soluble ionic solid**, for example sodium ethanoate, then before it reacts with the acid it dissolves in the water into its two ions; Na^+ and CH_3COO^-. This means that even though an ethanoate ion is a weak base only dissociating to release a small proportion of OH- ions / conjugate acid, it will still **be a good conductor** because of the Na^+ and CH_3COO^- ions in solution. These are free moving ions that carry charge.
When drawing equations remember to show CH_3COONa dissolving into the two ions first before CH_3COO^- ion further reacts as a base. (2 equations required)



$CH_3COONa \rightarrow Na^+ + CH_3COO^-$
$CH_3COO^- + H_2O \leftrightharpoons CH_3COOH + OH^-$

Note the single arrow for salt dissolving and double arrow for weak base dissociation

Reaction Rates of Acids (and Bases)

The reaction rate of acid and bases depends upon the concentration of available ions (either H_3O^+ with acids or OH^- with alkalis). If there is a higher concentration of these ions in a system, there is a greater chance that they will collide with other reactants, as there is less space between particles. The higher frequency of collisions means there are more effective collisions per unit of time and this will increase the rate of the reaction. If there is a lower concentration, there will be fewer collisions and the reaction rate will decrease.

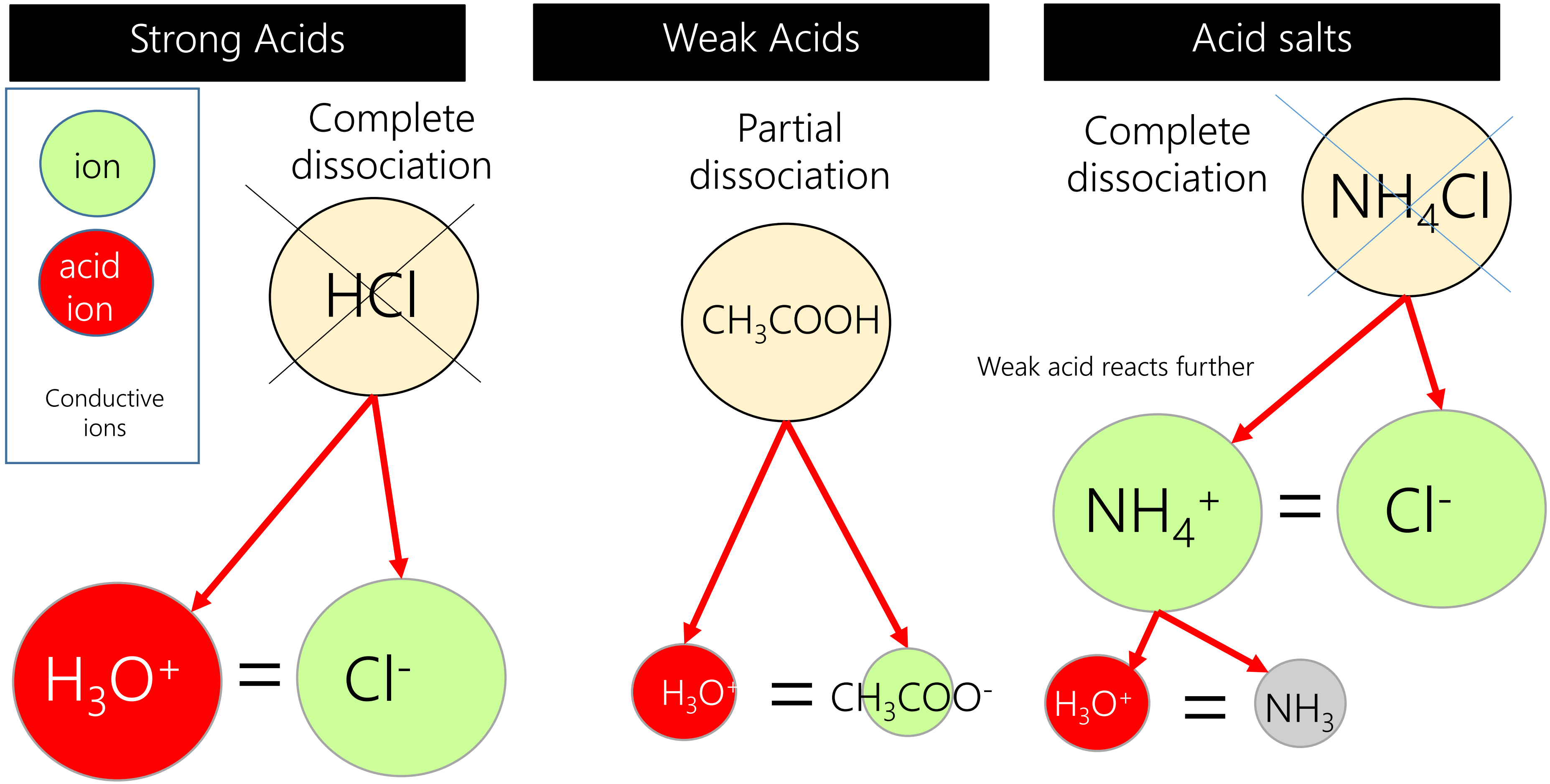
Strong Acids (HX)

Expect reaction to be more vigorous; rapidly produces gas / bubbles – since the concentration of hydrogen ions is high, there will be more frequent collisions resulting in a faster rate of reaction.

Weak Acids (HA)

Expect a slower reaction, taking longer to produce the same volume of gas – since the concentration of hydrogen ions is low, there will be less frequent collisions resulting in a slower rate of reaction.

Summary of Species/conductivity in Solution - Acid



No strong acid remains
High conductivity
Low pH (1-2)

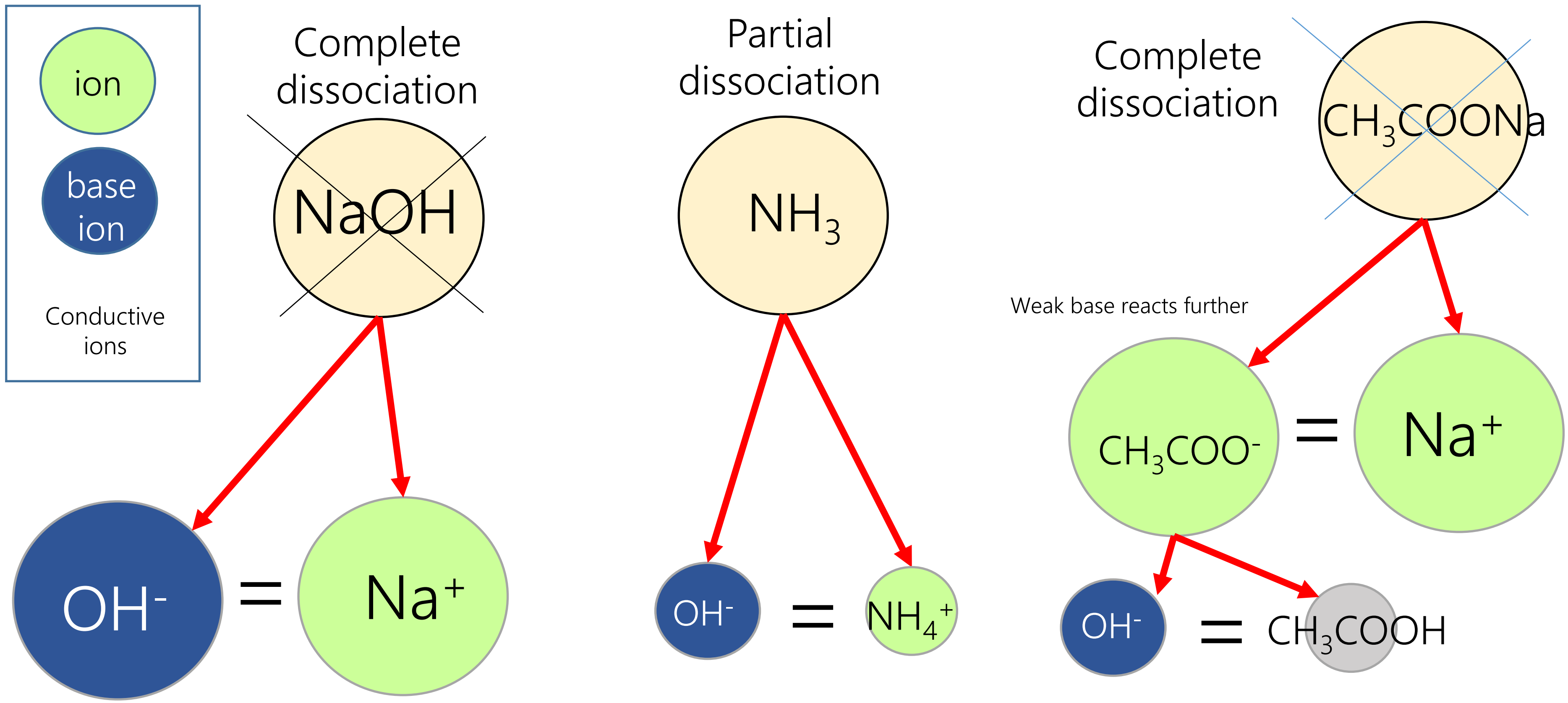
Most weak acid remains
Low conductivity
High pH (3-6)

No salt remains
High conductivity
High pH (3-6)

Water concentration is assumed to remain constant so is left out

Summary of Species / conductivity in Solution - Base

Strong Bases | Weak Bases | Base salts



No strong base remains
High conductivity
High pH (12-14)

Most weak base remains
Low conductivity
Low pH (8-11)

No salt remains
High conductivity
Low pH (8-11)

water concentration is assumed to remain constant so is left out

NCEA 2012 Conductivity and Ions (part one)

Question 3d: Some properties of three aqueous solutions A, B and C, of equal concentration are shown in the table below. It is known that the solutions are $NH_{3(aq)}$, $HCl_{(aq)}$ and $NH_4Cl_{(aq)}$

Solution	**A**	**B**	**C**
pH	5.15	11.6	1.05
Electrical conductivity	good	poor	good

Justify the identification of all three solutions.
- refer to both pH and electrical conductivity of the solutions
- link your answers to appropriate chemical equations.

$NH_4Cl_{(aq)}$ is solution **A**: good conductor of electricity – it fully dissociates in solution into ammonium and chloride ions, which conduct electricity.
$NH_4Cl \rightarrow NH_4^+ + Cl^-$
Its pH is that of a weak acid, as the ammonium ion is a weak acid and partially dissociates in water, producing hydronium ions:
$NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$

NCEA 2012 Conductivity and Ions (part two)

Question Some properties of three aqueous solutions A, B and C, of equal concentration are shown in a table. It is known that the solutions are $NH_{3(aq)}$, $HCl_{(aq)}$ and $NH_4Cl_{(aq)}$

$NH_{3(aq)}$ is solution **B**: its pH is that of a weak base as NH_3 is a weak base and it partially dissociates in water, producing hydroxide ions:

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

Poor conductor of electricity as it is only partially dissociated into ions in water. The remaining NH3 mlecules are neutral and do not conduct electricity.

$HCl_{(aq)}$ is solution **C**: low pH is that of a strong acid, HCl fully dissociates in water, producing hydronium ions:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Good conductor of electricity as it fully dissociates into ions in solution which conduct electricity.

In order to conduct electricity there needs to be the presence of **free moving charged particles**. The more charged particles there are available the better conductivity there will be. Ions in solution provide the charged particle. A strong acid and alkaline will allow good conductivity. A weak acid or base will provide poor conductivity. An ionic solid will dissolve in water into ions and provide good conductivity

NCEA 2013 Reaction rates of Acids

Question 3d: The following table shows the concentration and pH of three acids, and the relative rate of reaction with magnesium (Mg) metal.

acid	Concentration / $molL^{-1}$	pH	Relative rate of reaction with Mg
HA	0.100	3.4	slow
HB	0.0100	2	fast
HC	1.00×10^{-5}	5	Very slow

Compare and contrast the reactivity of the three acids with magnesium. In your answer:

- determine the concentration of hydronium ions, H_3O^+, in each acid
- compare the concentration of hydronium ions to the concentration of the acid
- explain the relative rate of reaction for each acid with magnesium

HA $[H_3O^+] = 10^{-3.4} = 3.98 \times 10^{-4}$ $molL^{-1}$ HB $[H_3O^+] = 10^{-2} = 0.0100$ $molL^{-1}$ HC $[H_3O^+] = 10^{-5} = 1.00 \times 10^{-5}$ $molL^{-1}$

The concentration of hydronium ions is the same as the concentration of the acid for HB and HC, which means they are strong acids, since they **both completely dissociate** when they react with water. The concentration of hydronium ions in HA is less than the concentration of the acid HA, therefore it is a weaker acid, since it only partially dissociates in water. HB will react faster with the Mg, as it has the greatest concentration of hydronium ions since it has the lowest pH. Because the concentration of hydronium ions is greater in HB, there are more hydronium ion particles in the same volume to collide with the Mg. Therefore, there will be more frequent collisions, and hence a faster rate of reaction. Even though HA has a higher concentration than HB, it is a weaker acid, so it only partially dissociates and not all of the HA particles donate hydrogen ions. Therefore the concentration of hydronium ions is low, so there are fewer hydronium ion particles

NCEA 2014 Conductivity and ions

Excellence Question

Question 1c: The table below shows the relative electrical conductivity of five solutions of the **same concentration**, and the colour of pieces of litmus paper which have been dipped into each solution. Identify a strong base and a neutral salt, using the information in the table above.
In your answer you should justify your choices by referring to the properties of the identified solutions.

Solution	**A**	**B**	**C**	**D**	**E**
Electrical conductivity	poor	good	good	poor	good
Red litmus paper	turns blue	stays red	stays red	stays red	turns blue
Blue litmus paper	stays blue	turns red	stays blue	turns red	stays blue

Answer: Strong base = **E**; red litmus paper turns blue and blue litmus remains blue indicates base; good conductivity indicates large number of charged particles to conduct current – strong base completely dissociates and produces high $[OH^-]$.

Neutral salt = **C**; no colour change in either litmus, indicating neutral solution; good conductivity indicates large number of charged particles to conduct current – salt will dissolve to give ions in solution or solution has ions.

NCEA 2014 Reaction rates of Acids (PART ONE)

Question 3c(i): The pH values of 0.100 mol L–1 solutions of two acids, HA and HB, are given in the table below.

Solution	pH
0.100 mol L^{-1} HA(*aq*)	1.0
0.100 mol L^{-1} HB(*aq*)	2.2

Compare the relative strengths of the two acids, HA(*aq*) and HB(*aq*), using the information given above.
Your answer should include equations and calculations.

$HA + H_2O \rightarrow A^- + H_3O^+$
$[H_3O^+] = 0.100 \text{ mol L}^{-1}$

$HB + H_2O \rightleftharpoons B^- + H_3O^+$
$[H_3O^+] = 0.00631 \text{ mol L}^{-1}$

HA is a strong acid since it fully dissociates, as shown by concentration of hydronium ions in HA solution – same as original concentration of HA (both 0.100 mol L^{-1}).
HB is a weak acid since it only partially dissociates; as shown by the concentration of hydronium ions in HB solution – concentration is only 0.00631 mol L^{-1}.

NCEA 2014 Reaction rates of Acids - (PART TWO)

Question 3c(i): (ii) Predict and compare, with reasons, what would be observed when two 5 g samples of calcium carbonate chips, $CaCO_{3(s)}$, are reacted, separately, with excess HA and HB.

Solution	**pH**
0.100 mol L^{-1} HA(*aq*)	1.0
0.100 mol L^{-1} HB(*aq*)	2.2

HA

Expect reaction to be more vigorous; rapidly produces gas / bubbles (CO_2) – since the concentration of hydrogen ions is high, there will be more frequent collisions resulting in a faster rate of reaction.

HB

Expect a slower reaction, taking longer to produce the same volume of gas – since the concentration of hydrogen ions is low, there will be less frequent collisions resulting in a slower rate of reaction.

NCEA 2015 Conductivity and ions

Excellence Question

Question 2(d): Ethanoic acid solution, $CH_3COOH_{(aq)}$, and ammonium chloride solution, $NH_4Cl_{(aq)}$, are both weakly acidic.
Identify and justify, using equations, which acid solution has greater electrical conductivity.

Answer: NH_4Cl is a better conductor of electricity because it completely dissolves / dissociates into ions:

$NH_4Cl_{(s)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$

(Then the NH_4^+ ions react in water to produce hydronium ions, which makes it a weak acid:

$NH_4^+{}_{(aq)} + H_2O_{(\ell)} \leftrightharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$

Ethanoic acid does not dissociate before reacting with water, so produces fewer ions than ammonium chloride, due to the formation of ions being dependent on the position of equilibrium, which for a weak acid like ethanoic acid, lies to the left, resulting in only a small number of ions being formed in this solution.

$CH_3COOH_{(aq)} + H_2O_{(\ell)} \leftrightharpoons CH_3COO^-{}_{(aq)} + H_3O^+{}_{(aq)}$

The ability to act as a conductor depends upon the concentration of ions. Ammonium chloride has a greater concentration of ions, so is a better conductor of electricity.

NCEA 2015 Reaction rates of Acids

Excellence Question

Question 2(e): The table shows the pH of two acidic solutions, methanoic acid, HCOOH, and hydrochloric acid, HCl, which both have a concentration of 0.1 mol L^{-1}.
Compare and contrast the pH of each solution, and their expected rate of reaction with a 2 cm strip of cleaned magnesium ribbon, Mg.

Solution	HCOOH(*aq*)	HCl(*aq*)
pH	2.4	1

Answer: HCl is a strong acid because it fully dissociates in water.
$HCl_{(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ $[H_3O^+] = 0.1\text{mol L}^{-1}$
Whereas HCOOH is a weak acid, it does not readily dissociate in water.
$HCOOH_{(aq)} + H_2O_{(\ell)} \leftrightharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$
$[H_3O^+] = 0.00398\text{mol L}^{-1}$
In the resulting solutions, HCl has a higher concentration of H_3O^+, and therefore a lower pH (1) than HCOOH, which has a lower concentration of H_3O^+, and therefore a higher pH (2.4).
Both acids will react with the cleaned Mg ribbon, but with a higher concentration of H_3O^+, HCl will have a faster rate of reaction with Mg than HCOOH, as there are more H_3O^+ ions available to react in a given volume.

NCEA 2016 Conductivity and ions

Excellence Question

Question 2(d): (ii) Explain why the solution of ammonium chloride, $NH_4Cl(aq)$, is a good conductor of electricity, while the solution of propanoic acid, $C_2H_5COOH(aq)$, is a poor conductor of electricity.

Answer: Charged particles that are free to move are needed to conduct electricity. Ammonium chloride is a good conductor of electricity because it is a salt that will completely dissolve / dissociate / ionise in water, releasing its ions NH_4^+ and Cl^- into solution, i.e. relatively high concentration of ions. Propanoic acid is a weak acid, so will only partially dissociate (ionise) in water to form a relatively small proportion of ions, resulting in it being a poor conductor of electricity

$$NH_4Cl_{(s)} \xrightleftharpoons{H_2O} NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$$

$$CH_3CH_2COOH_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3CH_2COO^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

NCEA 2016 pH and ions

Excellence Question

Question 2(d): The table shows the pH of three acidic solutions, ammonium chloride, NH_4Cl, propanoic acid, C_2H_5COOH, and hydrogen chloride, HCl.
(i) Explain why each of the three solutions in the table above has the same concentration, but a different pH.
Use equations to support your answer.

	$NH_4Cl(aq)$	$C_2H_5COOH(aq)$	$HCl(aq)$
Concentration / mol L^{-1}	0.1	0.1	0.1
pH	5.62	3.44	1.0

pH reflects the concentration of H_3O^+ ions. The higher the pH, the lower the $[H_3O^+]$. Both ammonium ions, NH_4^+, and propanoic acid are weak acids, so both only partially dissociate in water and produce fewer H_3O^+ ions in water than HCl, which is a strong acid. As a strong acid, HCl completely dissociates in water, resulting in a high concentration of H^+ (H_3O^+) ions and a low pH.

$NH_4^+(aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq$

$C_2H_5COOH(aq) + H_2O(\ell) \rightleftharpoons C_2H_5COO^-(aq) + H_3O^+(aq)$

$HCl(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + Cl^-(aq)$

NH_4^+ has a higher pH than propanoic acid, which means that it has a lower $[H_3O^+]$ in solution.

NCEA 2017 Conductivity and ions

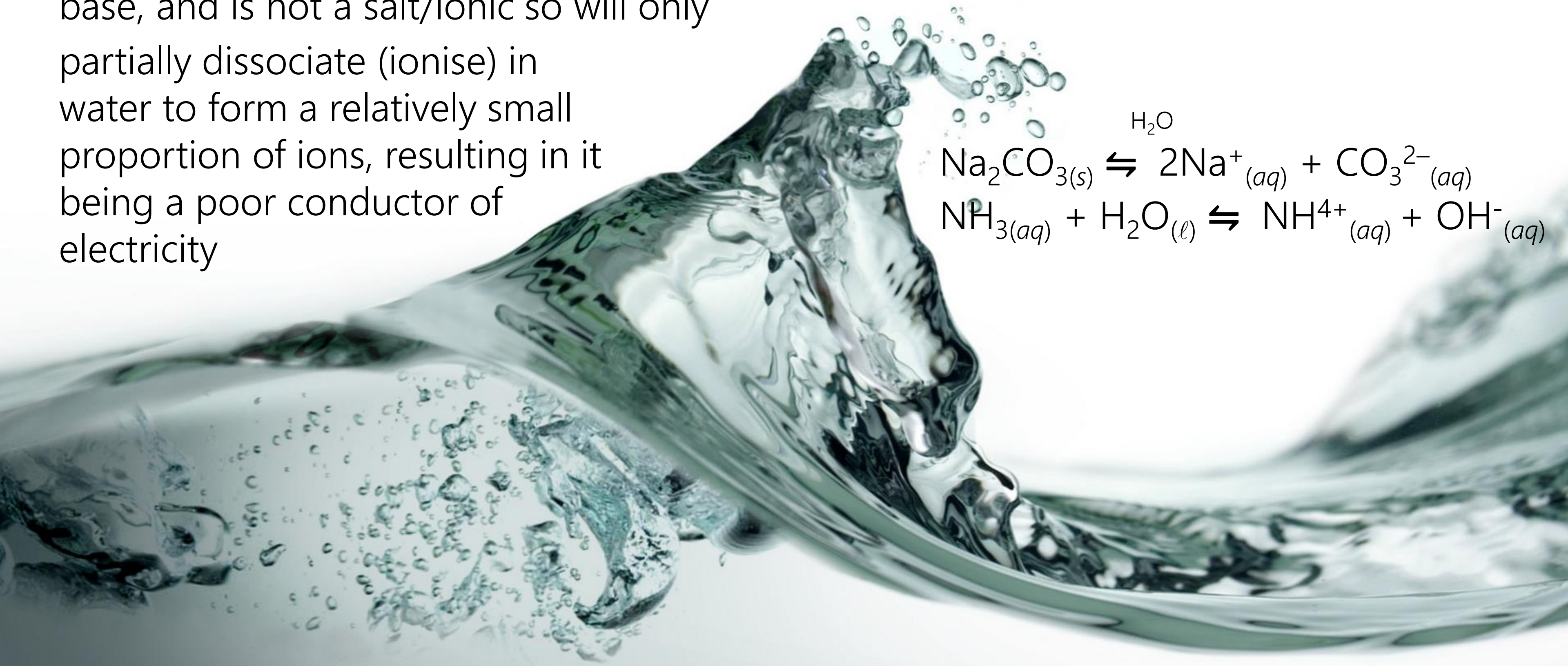
Excellence Question

Question 1d: Solutions of ammonia, $NH_{3(aq)}$, and sodium carbonate, $Na_2CO_{3(aq)}$, are both basic. Compare and contrast the electrical conductivity of these two solutions.

Answer: Charged particles that are free to move in solution are needed to conduct electricity. Sodium carbonate is a good conductor of electricity because it is a (base) salt that will completely dissolve / dissociate / ionise in water, releasing its ions CO_3^{-2} and Na^+ (x 2) into solution, i.e. relatively high concentration of ions. Ammonia is a weak base, and is not a salt/ionic so will only partially dissociate (ionise) in water to form a relatively small proportion of ions, resulting in it being a poor conductor of electricity

$$Na_2CO_{3(s)} \xrightleftharpoons{H_2O} 2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$$NH_{3(aq)} + H_2O_{(\ell)} \rightleftharpoons NH^{4+}{}_{(aq)} + OH^-{}_{(aq)}$$



NCEA 2018 Conductivity / pH and ions

Excellence Question

Question 3a: The pH and relative electrical conductivity of aqueous solutions of potassium hydroxide, $KOH_{(aq)}$, and ammonia, $NH_{3(aq)}$, are shown in the table below. Both have concentrations of 0.100 mol L^{-1}.
Explain the difference in pH and conductivity of these two solutions.
Use relevant equations in your answer.

Chemical	**pH**	**Conductivity**
$KOH(aq)$	13	good
$NH_3(aq)$	11.1	poor

KOH is a **strong base**, so it will ionize / **dissociate completely**, producing **many OH^- ions**.
$KOH_{(s)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$
NH_3 is a **weak base**, so it only **partially ionizes / dissociates** in water producing **fewer OH^- ions**.
$NH_{3(aq)} + H_2O_{(\ell)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$
pH depends on the $[H_3O^+]$ ions. **As $[OH^-]$ increases, $[H_3O^+]$ decreases and pH is higher.**
Since NH_3 produces fewer $[OH^-]$ ions, its pH is smaller than KOH.
Conductivity of a solution depends on presence of ions. KOH solution has **more ions (not just** OH^- **ions) than** NH_3 solution, so is a better conductor.

NCEA 2018 Reaction rates of Acids

Excellence Question

Question 3c (ii) : Use the concentrations below to predict the rate of reaction of each acid with a 2 cm strip of cleaned magnesium ribbon, Mg.
Refer to the collision theory in your answer.
$[H_3O^+]$ of $HNO_3 = 10^{-0.7} = 0.200$ mol L^{-1}
$[H_3O^+]$ of $CH_3COOH = 10^{-2.73} = 0.00186$ mol L^{-1}

Both acids will react with the cleaned Mg ribbon, but since HNO_3 completely dissociates, it has a **higher concentration of H_3O^+**, so HNO_3 will have a **faster rate** of reaction with Mg than CH_3COOH. There are **more H_3O^+ ions** available to react in a given volume so there will **be more effective collisions per second.**

NCEA 2019 pH and ions (Part ONE)

Achieved Question

Question 3c: The table below provides information about solutions **A** to **D**.

Solution	A	B	C	D
Concentration (mol L^{-1})	0.100	0.100	0.100	0.100
pH	5.62	1	7	13

The solutions are known to be hydrochloric acid, $HCl_{(aq)}$, ammonium chloride, $NH_4Cl_{(aq)}$, sodium hydroxide, $NaOH_{(aq)}$ and sodium chloride, $NaCl_{(aq)}$.

(i) Identify solutions **A** to **D**.

A: NH_4Cl **B**: HCl
C: NaCl **D**: NaOH

NCEA 2019 pH and ions (Part TWO)

Excellence Question

Question 3c: (ii) Justify your choices by comparing relative amounts of hydronium ion concentrations, $[H_3O+]$, in the solutions.
Include relevant equations in your answer.

Solution	A	B	C	D
Substance	NH_4Cl	HCl	NaCl	NaOH
Concentration (mol L^{-1})	0.100	0.100	0.100	0.100
pH	5.62	1	7	13

A: pH reflects the concentration of H_3O^+ ions. The higher the pH, the lower the $[H_3O^+]$.
Ammonium chloride fully dissociates into ions.
$NH_4Cl_{(s)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$
The NH_4^+ partially dissociates/ionises in water, producing fewer H_3O^+ ions, so is a weak acid where pH does not equal the – $\log[H_3O^+]$, so **A** is NH_4Cl.
(1) $NH_4^+ + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$

B: As a strong acid, HCl completely dissociates in water, resulting in a high $[H_3O^+]$ ions and a low pH. For a strong acid, pH = $-\log[H_3O+]$. Hence **B** is HCl.
(2) $HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$

C is NaCl as it is a neutral salt that completely dissociates in water. Neither Na or Cl ions react further when dissolved in water. Hence the $[H_3O^+]$ remains the same as $[OH^-]$, and so the pH is 7.
(3) $NaCl_{(s)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)}$

D is NaOH. NaOH is a strong base and so completely dissociates in water resulting in a high $[OH^-]$, low $[H_3O^+]$ and a high pH.
(4) $NaOH_{(s)} \rightarrow Na^+{}_{(aq)} + OH^-{}_{(aq)}$

NCEA 2019 Conductivity and ions (Part THREE)

Excellence Question

Question 3c: (iii) Elaborate on the electrical conductivity of the four solutions.

Solution	A	B	C	D
Substance	NH_4Cl	HCl	$NaCl$	$NaOH$
Concentration (mol L^{-1})	0.100	0.100	0.100	0.100
pH	5.62	1	7	13

Charged particles that are free to move are needed to conduct electricity. NH_4Cl, HCl, $NaOH$, and $NaCl$ are all good conductors of electricity because they completely dissociate in water, releasing ions into solution giving a relatively high concentration of ions to carry charge.
(Relates conductivity for ALL solutions to the degree of dissociation in water and the relative amounts of ALL ions (not just 'charged particles' or H_3O^+) in solution.

A: $NH_4Cl_{(s)} \rightarrow NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$ and $NH_4^+ + H_2O_{(l)} \leftrightharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$

B: $HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$

C: $NaCl_{(s)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)}$

D: $NaOH_{(s)} \rightarrow Na^+{}_{(aq)} + OH^-{}_{(aq)}$

NCEA 2020 pH and ions (Part ONE)

Excellence Question

Question 1c: The table below shows the concentration and pH of three basic solutions, sodium ethanoate, $CH_3COONa_{(aq)}$, ammonia, $NH_{3(aq)}$, and sodium hydroxide, $NaOH_{(aq)}$.

	$CH_3COONa(aq)$	$NH_3(aq)$	$NaOH(aq)$
Concentration ($mol\ L^{-1}$)	0.1	0.1	0.1
pH	8.88	10.6	13.0

Explain why each of these solutions has a different pH value, yet they are the same concentration. Use equations to support your answer.

The higher the pH, the more hydroxide ions present and the fewer hydronium ions, because the product of hydronium and hydroxide ions is 1×10^{-14}, and pH is the negative log of the hydronium ion concentration. The more hydroxide ions, the greater the dissociation or ionisation in water due to more ions being present.
NaOH is a strong base, so fully dissociates releasing many OH^- ions giving a very high $[OH^-]$, a very low $[H_3O^+]$ and in turn a high pH.
$NaOH \rightarrow Na^+ + OH^-$

NH_3 is a weak base, so only partially ionises, releasing fewer hydroxide ions than NaOH; giving a high $[OH^-]$,a low $[H_3O^+]$, in turn a high (10.6) pH, but still lower NaOH.
$NH_3 + H_2O \leftrightharpoons NH_4^+ + OH^-$

NCEA 2020 pH and ions (Part TWO)

Excellence Question

Question 1c: The table below shows the concentration and pH of three basic solutions, sodium ethanoate, $CH_3COONa_{(aq)}$, ammonia, $NH_{3(aq)}$, and sodium hydroxide, $NaOH_{(aq)}$.

	$CH_3COONa(aq)$	$NH_3(aq)$	$NaOH(aq)$
Concentration ($mol\ L^{-1}$)	0.1	0.1	0.1
pH	8.88	10.6	13.0

Explain why each of these solutions has a different pH value, yet they are the same concentration. Use equations to support your answer.

CH_3COONa dissolves into ions and in turn the released weak base CH_3COO^- partially dissociates, releasing even fewer hydroxide ions than ammonia, so an increased $[OH^-]$,a lowered $[H_3O^+]$, in turn a pH above 7 (8.88), but lower than ammonia.

$CH_3COONa \rightarrow CH_3COO^- + Na^+$

$CH_3COO^- + H_2O \leftrightharpoons CH_3COOH + OH^-$

Excellence Question

NCEA 2020 Reaction Rates of Acids

Question 2b: (ii) 2.0 g of powdered calcium carbonate, $CaCO_{3(s)}$, is added to each of the three solutions, A, B, and C, below. The volume of acid in each solution is the same. Identify which solution would have the highest rate of reaction with $CaCO_{3(s)}$. Explain your answer, with reference to collision theory.

Solution	Acid	pH	$[H_3O^+]$ mol L^{-1}
A	HCl(aq)	0.89	0.129 mol L^{-1}
B	HCl(aq)	1.80	0.0158
C	HCl(aq)	2.94	0.00115 mol L^{-1} or 1.15×10^{-3} mol L^{-1}

Solution A will have the highest rate of reaction. This is because the lowest pH (solution A) has the highest concentration of H_3O^+, therefore there are more acid particles in the same volume, so there will be more frequent collisions leading to more effective collisions per second.

NCEA 2020 Conductivity and Ions

Excellence Question

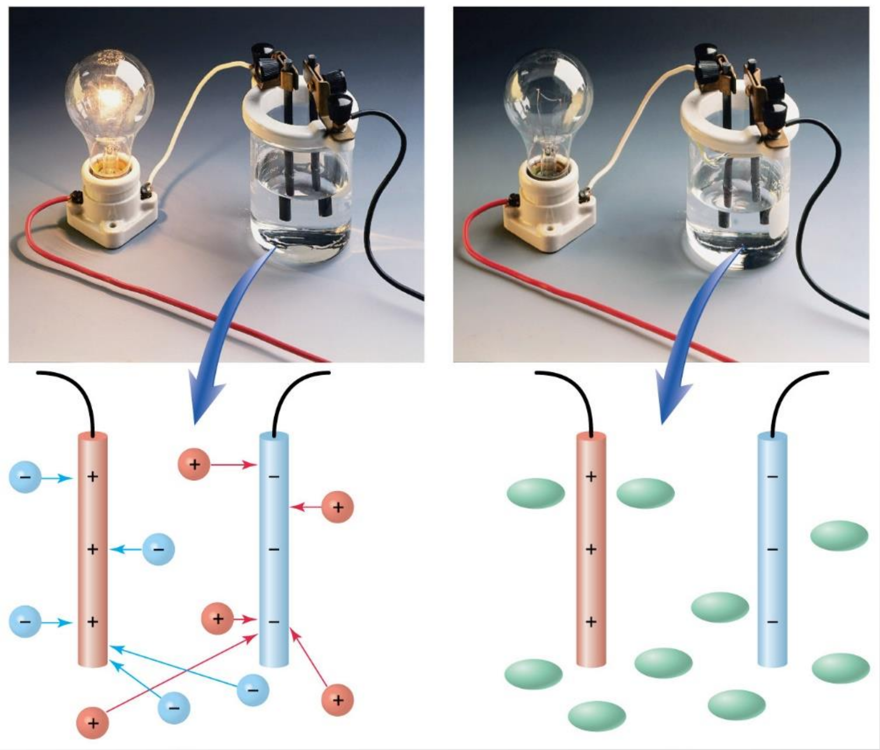
Question 2c: Compare the electrical conductivity of a hydrochloric acid solution, $HCl_{(aq)}$, with a solution of ethanoic acid, $CH_3COOH_{(aq)}$, of the same concentration.
Use equations to support your answer.

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

HCl is a strong acid, so it ionises completely in solution to produce a lot of ions. Conductivity depends on the number of mobile charged particles, in this instance ions, to conduct.

Ethanoic is a weak acid, so partially ionises in solution, therefore not producing many ions. So ethanoic acid is a poor conductor with HCl is a good conductor.

$CH_3COOH + H_2O \leftrightharpoons CH_3COO^- + H_3O^+$



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Kw – the ionic product for water

K_w is ionic product for water and an equilibrium constant based on the reaction of water molecules transferring H^+ in an acid base reaction to create OH^- and H_3O^+ in equal quantities. The rate of reaction from reactants to products is the same as products to reactants once equilibrium is reached.

$$K_c = \frac{[H_3O^+]\,[OH^-]}{[H_2O]^2} \quad \text{from} \quad 2H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + OH^-{}_{(aq)} \; (\Delta_r H = +ve)$$

Equilibrium constant

Or $K_c \times [H_2O]^2 = [H_3O^+]\,[OH^-]$

Because the concentration of water is so large it doesn't change → considered constant
So $K_c \times [H_2O]^2$ is also constant – called $\mathbf{K_w}$

As $[H_3O^+] \times [OH^-]$ always equals 1×10^{-14} then so does K_w

Temperature increase causes an increase in K_w as the reaction is endothermic this favours the forward reaction (Le Chatelier's Principle)

Using K_w to Calculate $[OH^-]$ or $[H_3O^+]$

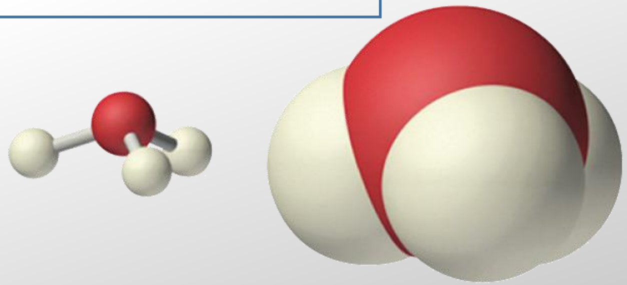
IONIC PRODUCT $K_W = [OH^-] \times [H_3O^+] = 1 \times 10^{-14}$

e.g. If solution A has $[OH-] = 1 \times 10^{-1}$ $molL^{-1}$ find $[H_3O^+]$

$$[H_3O^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-1}} = 1 \times 10^{-13} \quad molL^{-1}$$

$$[OH^-] = \frac{1 \times 10^{-14}}{[H_3O^+]} \quad \text{or} \quad [H_3O^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

You will be provided this value of K_w in Examinations



Hydronium ion, H_3O^+

Calculating pH given $[OH^-]$ or $[H_3O^+]$ in a strong acid or base

pH is the measure of the $[H_3O^+]$ concentration in a solution

It is an **inverse logarithmic scale** from 1-14. As the number gets lower then the concentration of hydronium ions increases exponentially.

$$pH = -\log [H_3O^+]$$

$pH = -\log [H_3O^+]$

e.g. $pH = -\log(10^{-5}) = 5$

e.g. $pH = -\log(2.4 \times 10^{-3}) = 2.6$

Buttons on calculator

(-)	log	2	.	4	exp	(-)	3

$$[H_3O^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

Use this equation first if given $[OH^-]$ instead of $[H_3O^+]$

Note: the concentration of monoprotic ions, such as HCl and HNO_3 that donate 1 hydrogen ion can be assumed to equal the $[H_3O^+]$ as they fully dissociate. H_2SO_4 will donate 2 hydrogen ions so the concentration of $[H_3O^+]$ will be twice that of the original starting concentration of acid.
The starting concentrations of strong bases such as NaOH can be assumed to be the same as $[OH^-]$
These assumptions only apply for strong acids and bases.

Calculating $[OH^-]$ or $[H_3O^+]$ given pH

$[H_3O^+] = 10^{-pH}$ $[OH^-] = 10^{-(14-pH)}$

e.g. calculate $[H_3O^+]$ if solution has pH=2 $[H^+] = 10^{-2}$ mol L^{-1}

calculate $[OH^-]$ if solution has pH=2 $[OH^-] = 10^{(14-2)}$

$[OH^-] = 10^{-12}$ mol L^{-1}

e.g. calculate $[H_3O^+]$ if solution has pH=11.3 $[H_3O^+] = 10=^{-11.3}$ mol L^{-1}

$[H_3O^+] = 5.0 \times 10^{-12}$ mol L^{-1}

Buttons on calculator

Shift	10^x	(-)	1	1	.	3

Calculating pOH given [OH^-]

$$pOH = -\log[OH^-] \qquad pH = 14 - pOH$$

Another method to calculate pH if you are provided the concentration of OH= ions [OH-], is to calculate pOH, and then subtract that value from 14.

[H^+]	pH	Color	[OH^-]	pOH
10^0	0		10^{-14}	14
10^{-1}	1		10^{-13}	13
10^{-2}	2		10^{-12}	12
10^{-3}	3		10^{-11}	11
10^{-4}	4		10^{-10}	10
10^{-5}	5		10^{-9}	9
10^{-6}	6		10^{-8}	8
10^{-7}	7		10^{-7}	7
10^{-8}	8		10^{-6}	6
10^{-9}	9		10^{-5}	5
10^{-10}	10		10^{-4}	4
10^{-11}	11		10^{-3}	3
10^{-12}	12		10^{-2}	2
10^{-13}	13		10^{-1}	1
10^{-14}	14		10^0	0

GZ Science Resources

Summary of pH formula

Calculate pH if given **Acid** Concentration

$$pH = -\log[H_3O^+]$$

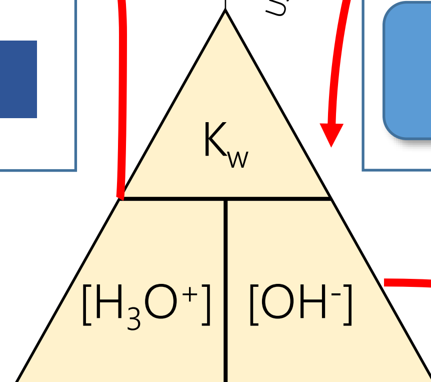
(—) log $[H_3O^+]$

Calculate **Acid** concentration if given pH

$$[H_3O^+] = 10^{-pH}$$

Use triangle to convert

shift 10^x (—) pH



$$K_w = 1 \times 10^{-14}$$

Use triangle to convert

$[OH^-]$

$[OH^-]$

Calculate **$[H_3O^+]$** if given **$[OH^-]$**

Calculate **$[OH^-]$** if given **$[H_3O^+]$**

Summary of pH formula

Convert $[OH^-]$ to $[H_3O^+]$

$[H_3O^+] = 1 \times 10^{-14} / [OH-]$

Convert $[H_3O^+]$ to $[OH^-]$

$[OH^-] = 1 \times 10^{-14} / [H_3O^+]$

Convert $[H_3O^+]$ to pH

$pH = -\log[H_3O^+]$

Convert $[OH^-]$ to pOH

$pOH = -\log[OH^-]$

Convert pOH to pH

$pH = 14 - pOH$

Convert pH to $[H_3O^+]$

$[H_3O^+] = 10^{-pH}$

Convert pH to $[OH^-]$

$[OH^-] = 10^{-(14-pH)}$

NCEA 2014 pH Calculations

Excellence Question

Question 1b: (i) In a solution of potassium hydroxide, KOH, the pH is found to be 12.8.
Calculate the hydronium ion concentration, $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, in the solution. $K_w = 1 \times 10^{-14}$

Answer:

$[H_3O^+] = 1.58 \times 10^{-13}$ mol L^{-1}
$[OH^-] =$

$$\frac{1 \times 10^{-14}}{1.58 \times 10^{-13}}$$

$= 0.0633$ mol L^{-1}

Question 1b: (ii) Calculate the pH of a 2.25×10^{-4} mol L^{-1} sodium hydroxide, NaOH, solution.

Answer:

$[H_3O^+] = 1 \times 10^{-14} / [OH^-]$
pH= -log$[H_3O^+]$
pH = 10.4

OR
pOH= -log$[OH^-]$
pH = 14 - pOH

NCEA 2015 pH Calculations

Excellence Question

Question 2(c): (i) A solution of nitric acid, $HNO_{3(aq)}$, has a hydronium ion, H_3O^+, concentration of 0.0243 mol L^{-1}.
Determine, by calculation, the pH and the concentration of hydroxide ions, OH^-, in this solution.
$K_w = 1 \times 10^{-14}$

Answer:

$$\mathrm{pH} = -\log_{10}[H_3O^+] = 1.61$$

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{10^{-14}}{0.0243}$$

$$= 4.12 \times 10^{-13} \text{ mol L}^{-1}$$

OR

$$\mathrm{pH} = -\log_{10} 0.0243$$

$$= 1.61$$

$$\mathrm{pOH} = 14 - 1.61$$

$$= 12.4$$

$$[OH^-] = 10^{-12.4}$$

$$= 4.12 \times 10^{-13} \text{ mol L}^{-1}$$

Question 2(c): (ii) (ii) Determine the hydroxide ion concentration, $[OH^-]$, of a solution of potassium hydroxide, $KOH_{(aq)}$, with a pH of 11.8.

Answer:

$[H_3O^+] = 10^{-pH} = 10^{-11.8} = 1.58 \times 10^{-12} \text{ mol L}^{-1}$

$[OH^-] =$

$$\frac{K_w}{[H_3O^+]} = 6.31 \times 10^{-3} \text{ mol L}^{-1}$$

NCEA 2016 pH Calculations

Excellence Question

Question 2(c): (i) Calculate the pH of a 0.0341 mol L–1 hydrochloric acid, HCl(*aq*), solution.
pH =

$$pH = -\log[H_3O^+] = -\log 0.0341 = 1.47$$

Question 2(c): (ii) A solution of sodium hydroxide, NaOH(*aq*), has a pH of 12.4. Calculate the concentrations of both hydronium ions, H_3O+, and hydroxide ions, OH^-, in this solution.

$[H_3O+]$ =

$$[H_3O^+] = 10^{-pH} = 10^{-12.4} = 3.98 \times 10^{-13} \text{ mol L}^{-1}$$

[OH–] =

$$\left[OH^-\right] = \frac{K_w}{\left[H_3O^+\right]} = 0.0251 \text{ mol L}^{-1}$$

NCEA 2017 pH Calculations

Excellence Question

Question 1c: (i) A solution of sodium hydroxide, $NaOH_{(aq)}$, has a pH of 11.6. Calculate the hydronium ion concentration $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, in the solution.
*K*w = 1 × 10–14

$[H_3O^+] = 10^{-pH}$
$= 10^{-11.6}$
$= 2.51 \times 10^{-12}$ mol L^{-1}
$[OH-] = \frac{Kw}{[H_3O]^+}$
$= 3.98 \times 10^{-3}$ mol L^{-1}

(ii) Calculate the pH of a 2.96×10^{-4} mol L^{-1} solution of potassium hydroxide, $KOH_{(aq)}$.
pH =

pOH = $-\log(2.96 \times 10^{-4}) = 3.53$, so pH = 14.0 – 3.53, pH = 10.5.

OR
$$[H_3O]^+ = \frac{Kw}{[OH]^-} = \frac{1\times10^{-14}}{2.96\times10^{-4}} = 3.38\times10^{-11}$$
So, pH = $-\log(3.38 \times 10^{-11}) = 10.5$.

NCEA 2018 pH Calculations

Excellence Question

Question 1b (ii) : The pH of the original solution **X** is 10.8.
Calculate the hydronium ion concentration, $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, in the solution.

$[H_3O^+] = 10^{-pH} = 1.58 \times 10^{-11}$ mol L^{-1}
$[OH]- = Kw / [H_3O^+] = 6.31 \times 10^{-4}$ mol L^{-1}

Question 1b (iii) : The sodium hydroxide solution, $NaOH_{(aq)}$, used to prepare solution **X** has a concentration of 0.0125 mol L^{-1}.
Calculate the pH of the sodium hydroxide solution.

$[H_3O^+] = Kw / [OH] = 1 \times 10^{-14} / 0.0125$
$= 8.00 \times 10^{-13}$ mol L^{-1}

$pH = -\log_{10} [H_3O^+] = 12.1$

OR

$pOH = -\log_{10} 0.0125 = 1.90$
$pH = 14 - 1.9 = 12.1$

NCEA 2018 pH Calculations

Excellence Question

Question 3c (i) : The table below gives the pH of solutions of ethanoic acid, $CH_3COOH_{(aq)}$, and nitric acid, $HNO_{3(aq)}$, of concentrations of 0.200 mol L^{-1}.
Use the pH values to analyse the strength of the acids by calculating the concentration of their H_3O^+ ions.

Solution	**$CH_3COOH(aq)$**	**$HNO_3(aq)$**
pH	2.73	0.70

$[H_3O^+]$ of HNO_3 = $10^{-0.7}$ = 0.200 mol L^{-1}
This is the same as the concentration of acid, so it is a strong acid that fully dissociates, producing lots of H_3O^+.
$[H_3O^+]$ of CH_3COOH = $10^{-2.73}$ = 0.00186 mol L^{-1}
This is less than the concentration of acid, so it is a weak acid that only partially ionises in water, producing fewer H_3O^+ ions than HNO_3.

NCEA 2019 pH Calculations

Excellence Question

Question 3b: (i) A solution of hydrochloric acid, $HCl_{(aq)}$, has a hydronium ion concentration, $[H_3O^+]$, of 0.0164 mol L^{-1}.
Calculate the pH and hydroxide ion concentration, $[OH^-]$, of the solution.

$$pH = -\log\left[H_3O^+\right] = 1.79$$

$$\left[OH^-\right] = \frac{K_w}{\left[H_3O^+\right]} = \frac{1\times10^{-14}}{0.0164} = 6.10\times10^{-13} \text{ mol L}^{-1}$$

(ii) Calculate the hydroxide ion concentration, $[OH^-]$, of a solution of potassium hydroxide, $KOH_{(aq)}$, with a pH of 9.4.

$$\left[H_3O^+\right] = 10^{-pH} = 10^{-9.4} = 3.98\times10^{-10} \text{ mol L} - 1$$

$$\left[OH^-\right] = \frac{K_w}{\left[H_3O^+\right]} = 2.51\times10^{-5} \text{ mol L}^{-1}$$

NCEA 2020 pH Calculations

Excellence Question

Question 1b: (i) A solution of sodium hydroxide, $NaOH_{(aq)}$, has a pH of 11.8. Calculate the concentration of hydroxide ions, OH^-, in this solution.

$[H_3O^+] = 10^{-11.8} = 1.58 \times 10^{-12}$ mol L^{-1}

$$[OH^-] = \frac{1 \times 10^{-14}}{1.58 \times 10^{-12}} = 6.31 \times 10^{-3} \text{ mol L}^{-1}$$

(or 0.00631 mol L^{-1})

(ii) The ionisation constant of water, K_w , like all equilibrium constants, varies with temperature.
Calculate the pH of pure water at 0 °C when $K_w = 0.114 \times 10^{-14}$
$K_w = [H_3O^+][OH^-]$

$$[H_3O^+] = \sqrt{0.114 \times 10^{-14}} = 3.38 \times 10^{-8} \text{ mol L}^{-1}$$

$$pH = -\log[H_3O^+] = 7.47$$

NCEA 2020 pH Calculations

Achieved Question

Question 2b: (i) Complete the table below by calculating either the pH or the hydronium ion concentration, $[H_3O^+]$, for the three hydrochloric acid solutions, $HCl_{(aq)}$.

Solution	Acid	pH	$[H_3O^+]$ mol L^{-1}
A	HCl(aq)	0.89	0.129 mol L^{-1}
B	HCl(aq)	1.80	0.0158
C	HCl(aq)	2.94	0.00115 mol L^{-1} or 1.15×10^{-3} mol L^{-1}

Definitions

Acid: A substance donates H^+ ions – producing H_3O^+ ions
Alkali: A base in solution that has an excess of OH^- ions, and accepts H^+ ions
Amphiprotic: A substance that can act as either a H^+ acceptor or donator depending on the other substance it reacts with
Aqueous: A solution that is mainly water.
Base: A substance that accepts H^+ ions.
Neutral: A solution that has equal quantities of H^+ and OH^- ions. It is neither acidic nor basic.
Strong Acid: An acid that has a very low pH (0-3).The H^+ ions completely disassociates in solution
Strong Base: A base that has a very high pH (11-14). A substance that readily accepts all H^+ ions.
Weak Acid: An acid that only partially ionizes in an aqueous solution. That means not every molecule breaks apart. They usually have a pH close to 7 (3-6).
Weak Base: A base that only partially ionizes in an aqueous solution. That means not every molecule breaks apart. They usually have a pH close to 7 (8-10).