

With 2017 NCEA  
Exam included

2018  
Version

# Chemistry AS 91166

## C2.6 Chemical Reactivity

## Achievement Criteria

**This achievement standard involves demonstrating understanding of chemical reactivity.**



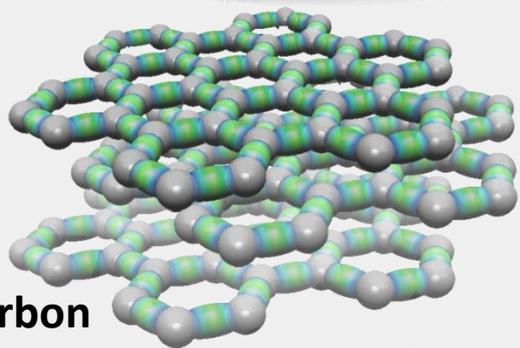
AS 91166  
C2.6

**Rates of Reaction** typically involves:

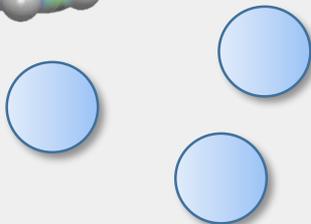
- factors affecting rates of reaction – restricted to changes in concentration, temperature, surface area, and the presence of a catalyst
- using particle theory to explain the factors (includes activation energy).

**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 ( $K_c$ ) 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  $\text{NH}_4^+$ . The properties are restricted to conductivity, rate of reaction, and pH
- calculations involving  $K_w$  and pH (restricted to strong acids and bases).



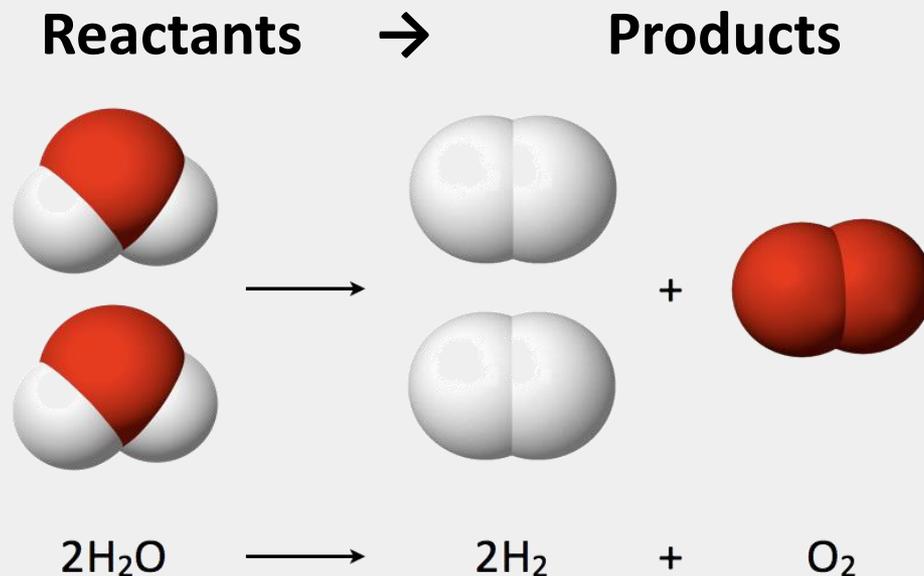
Carbon  
in the  
form of  
graphite



## 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.

## Chemical Reactions - reactants &amp; 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**.

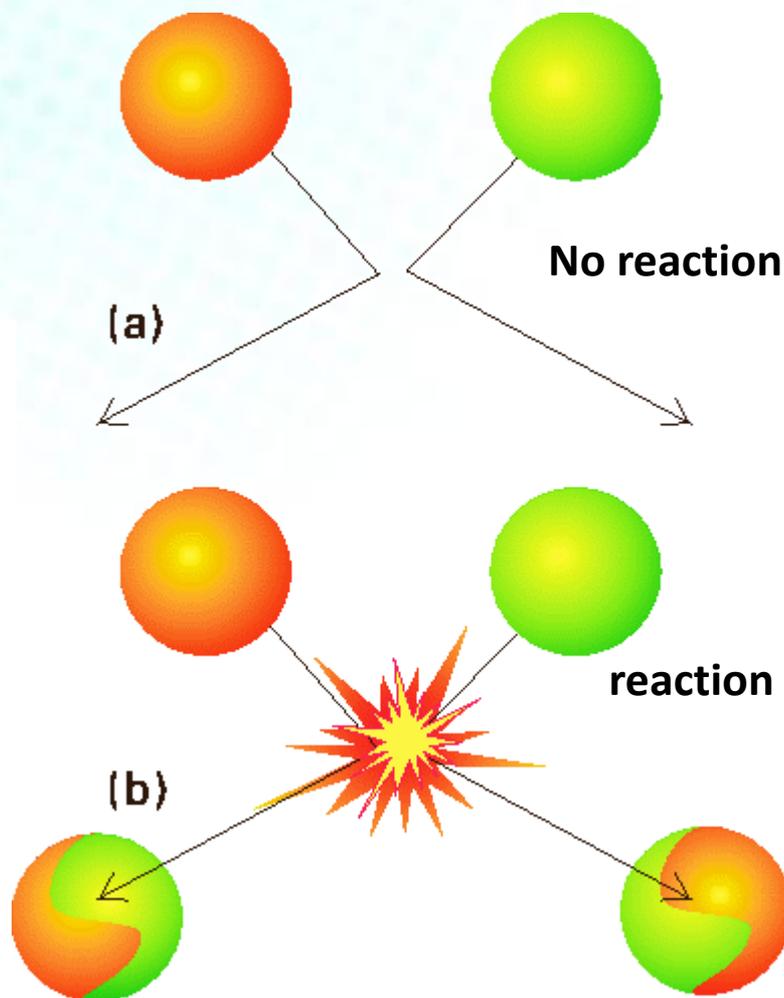
Chemical reactions between particles involve breaking bonds and forming new bonds.

## 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

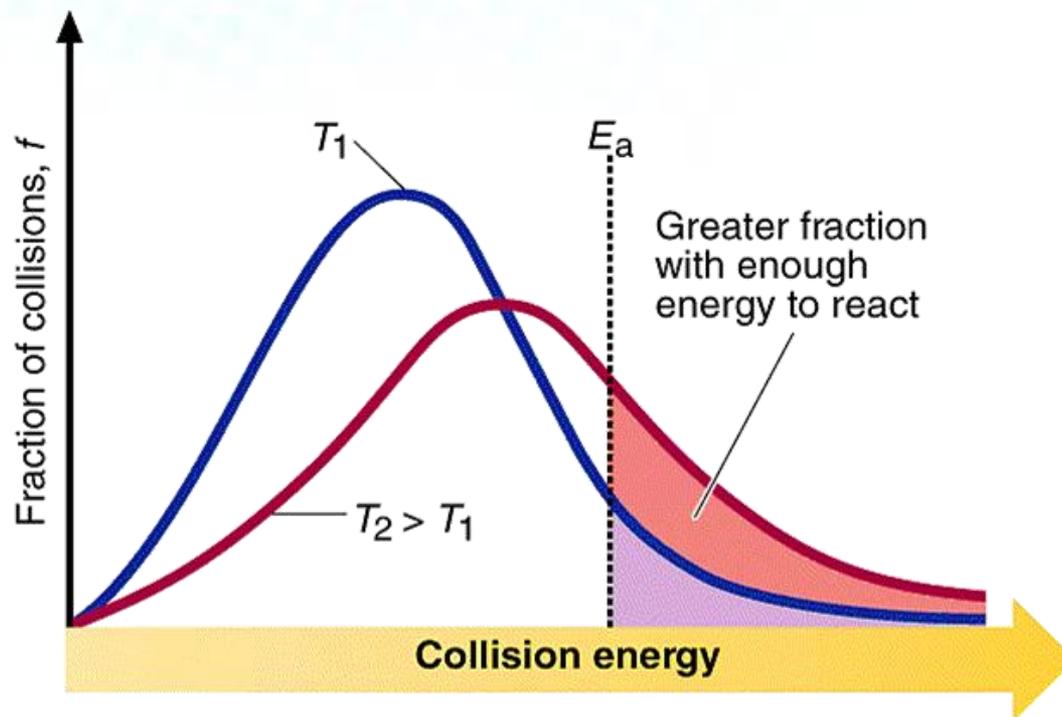
If these conditions are met the collision will be considered **successful**. (or effective)



## 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 ( $T_1$ ) most particles will have the same collision energy when they collide. As the temperature is increased ( $T_2$ ) 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.

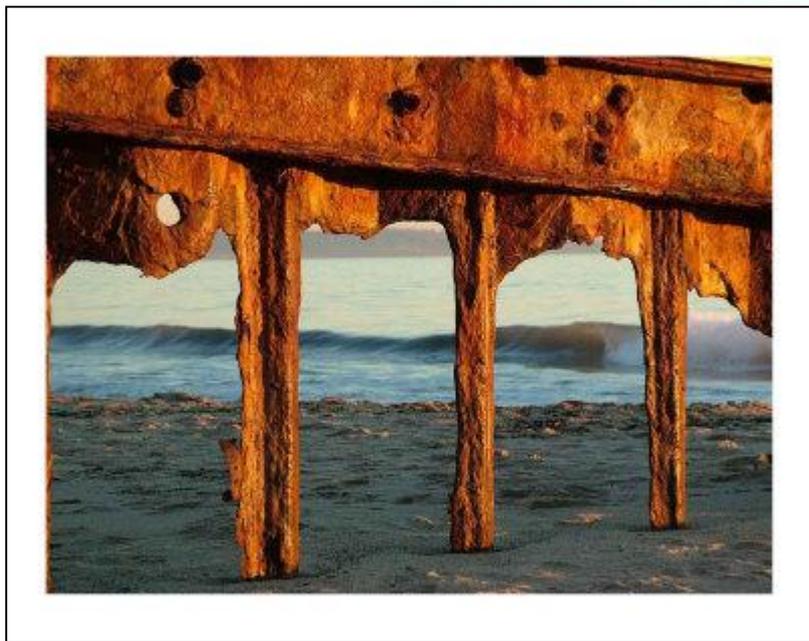


## Reaction Rate

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.



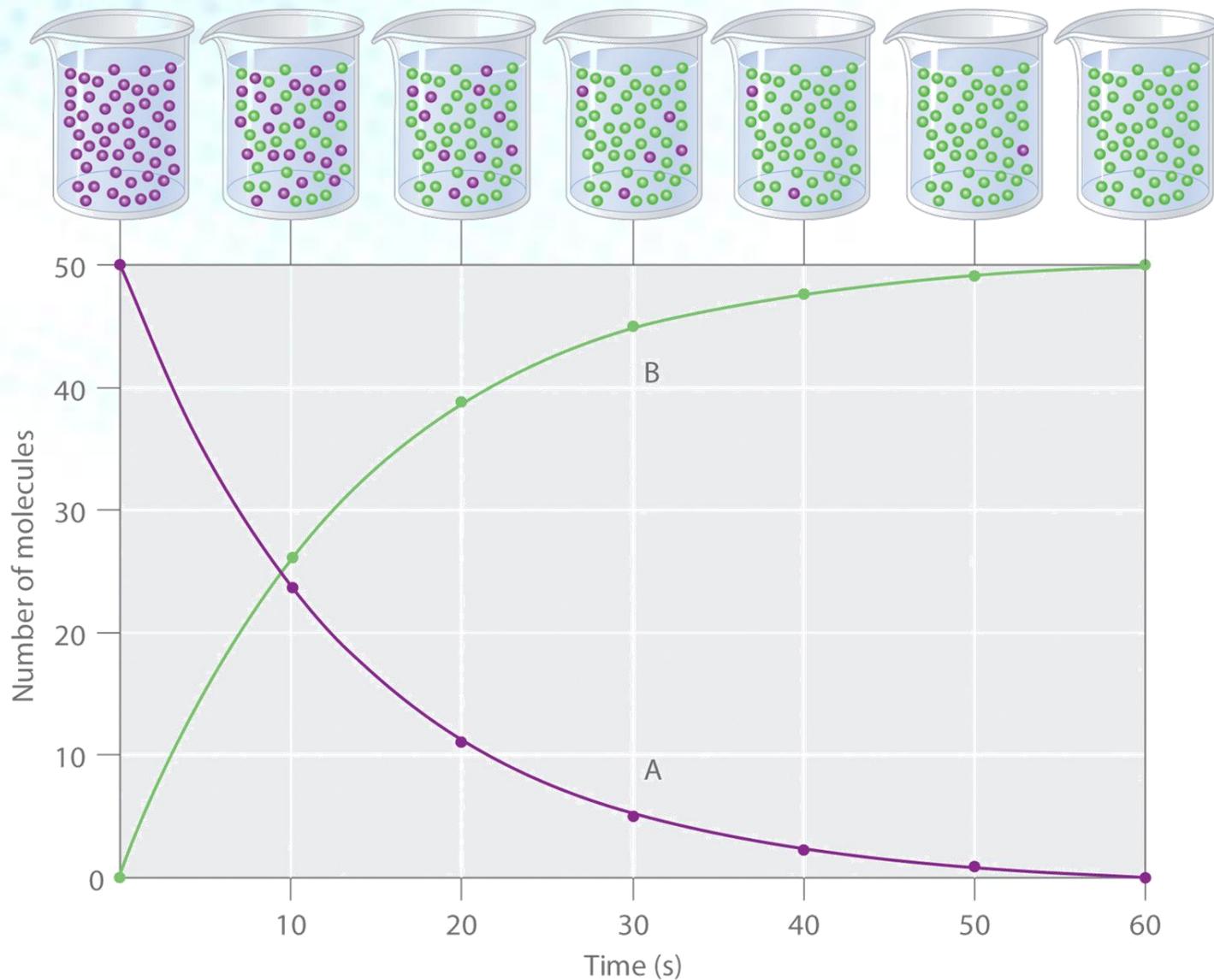
**Iron oxidising**



**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.

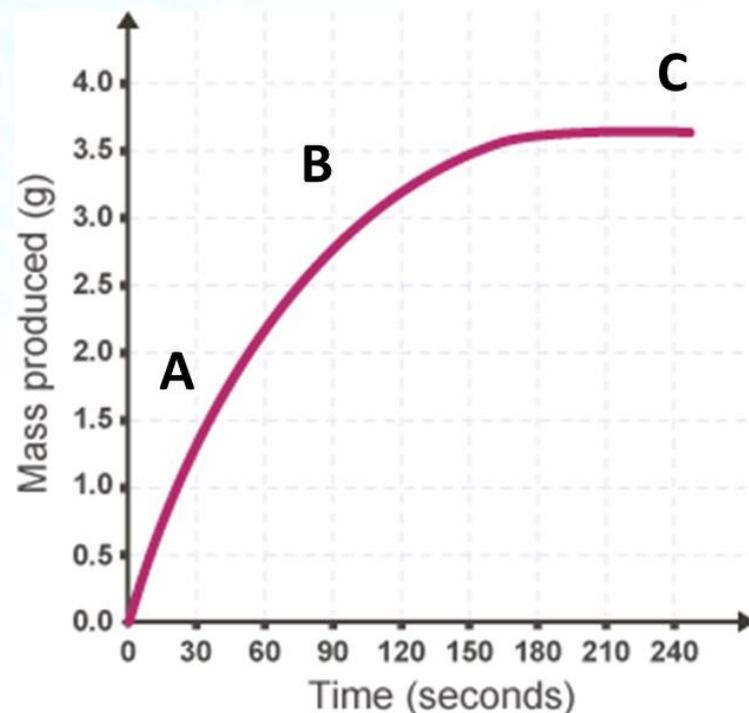


## Reaction rate is the speed at which a chemical reaction occurs

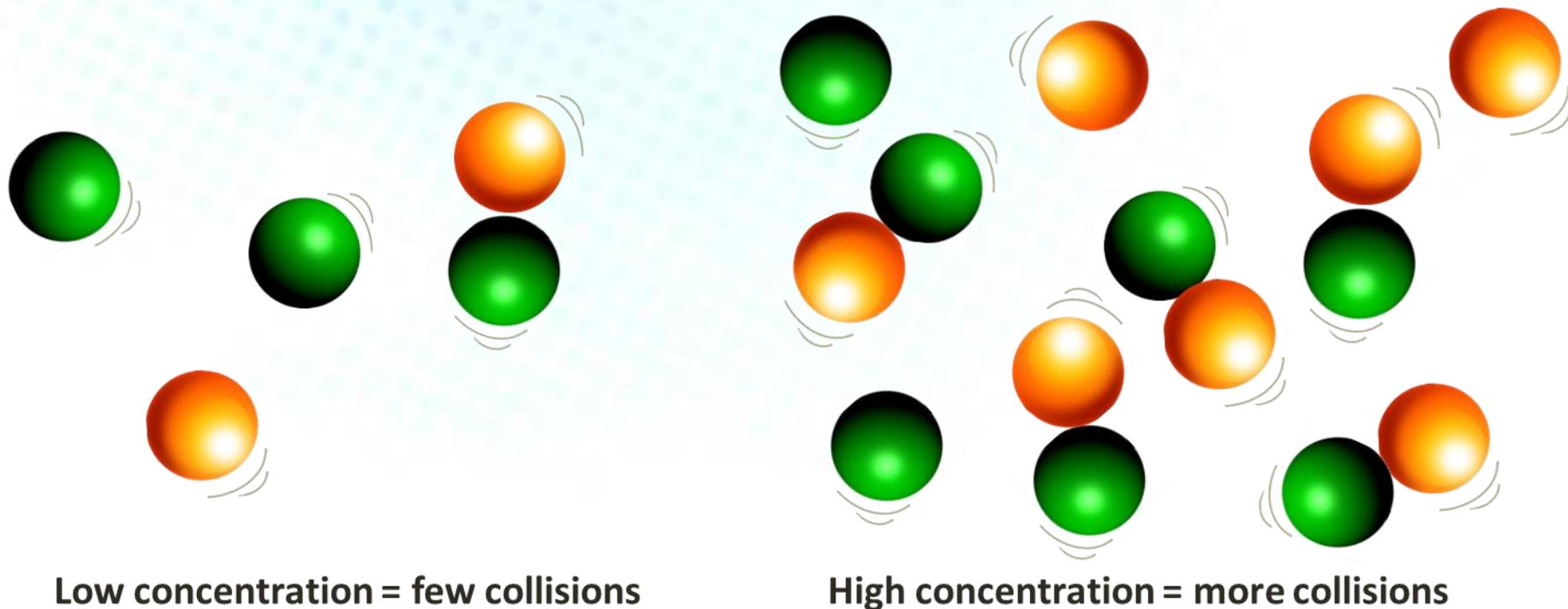
**A.** Reactions start out relatively fast because there is a much higher concentration of reactant particles available to collide and therefore the frequency of collisions will be high, resulting in more **successful collisions**; **resulting in an increase in the rate of reaction**. The gradient of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

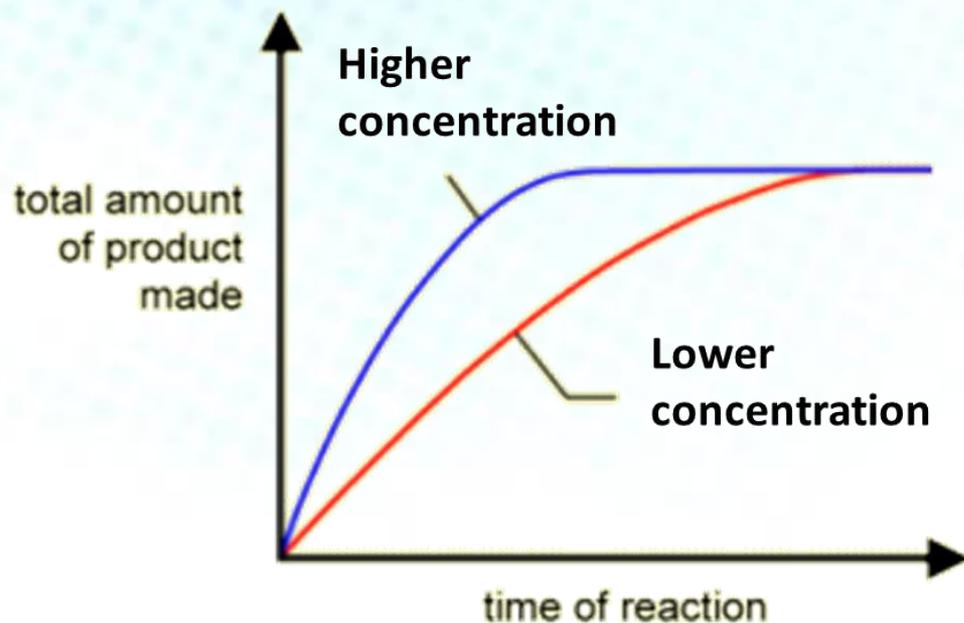


## Reaction rate can be increased by increasing the concentration



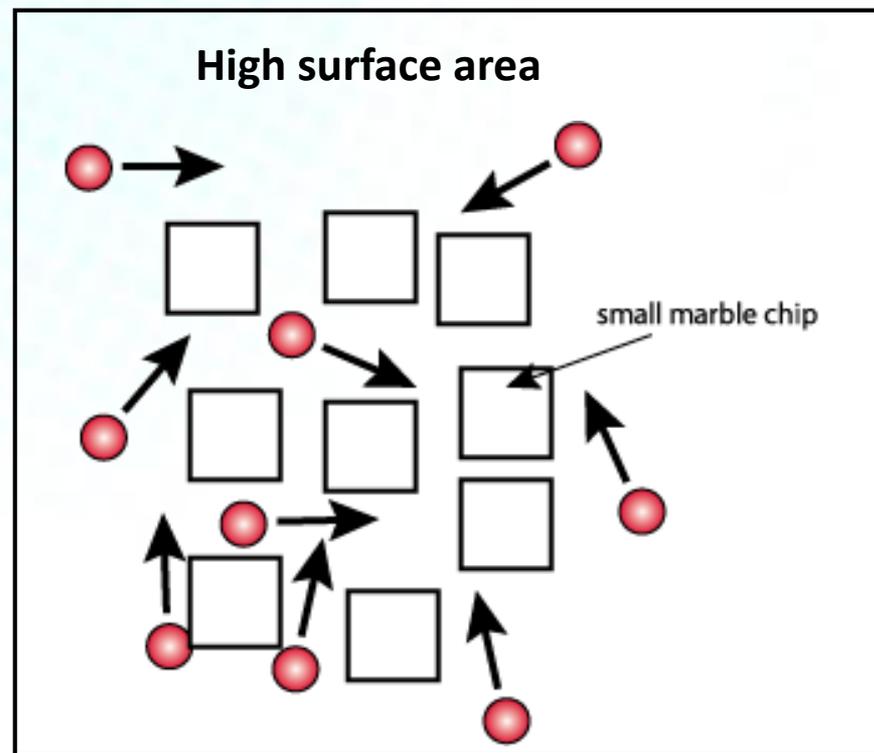
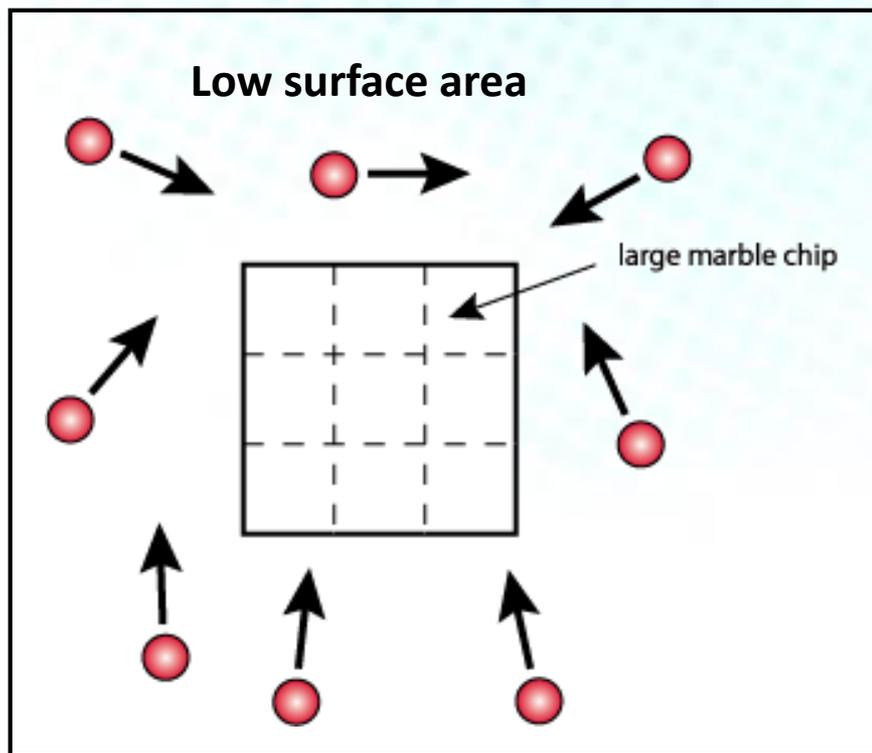
If there is a higher concentration of a substance in a system, there is a greater chance that molecules will collide, as there is less space between particles. The higher frequency of collisions means there are more effective collisions per unit of time and this will increase the rate of the reaction. 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  $1 \text{ mol L}^{-1}$  solution will contain only half the particles of a  $2 \text{ mol L}^{-1}$  so twice the volume will be required to produce the same quantity of product. Also note that the **total amount of successful collisions does not change** by increasing the concentration only the frequency (amount of collisions per unit of time) of collisions is increased.

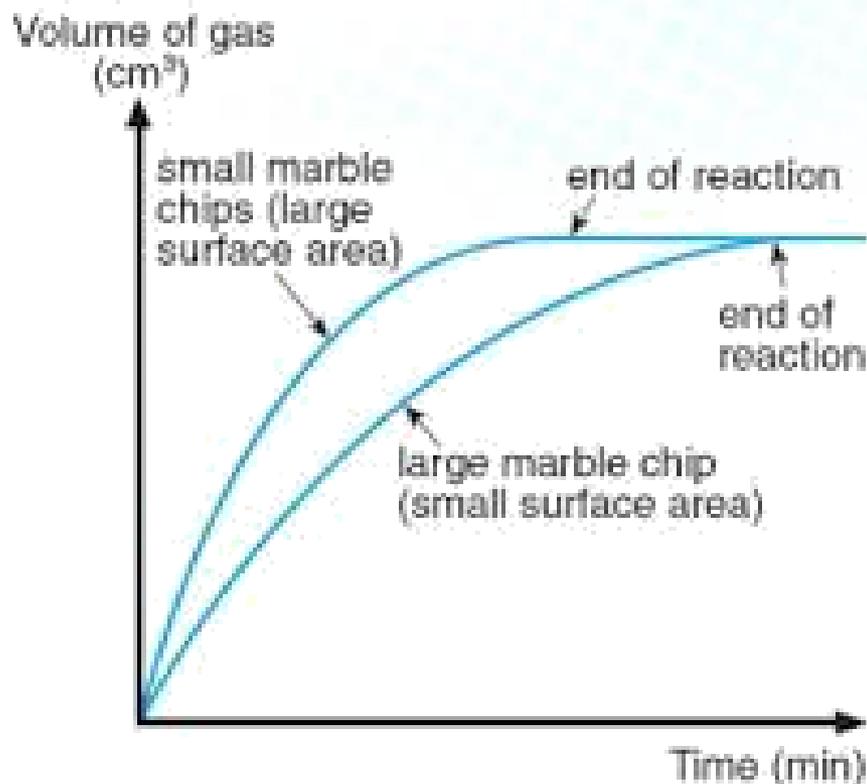
## Reaction rate can be increased by increasing the Surface Area



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

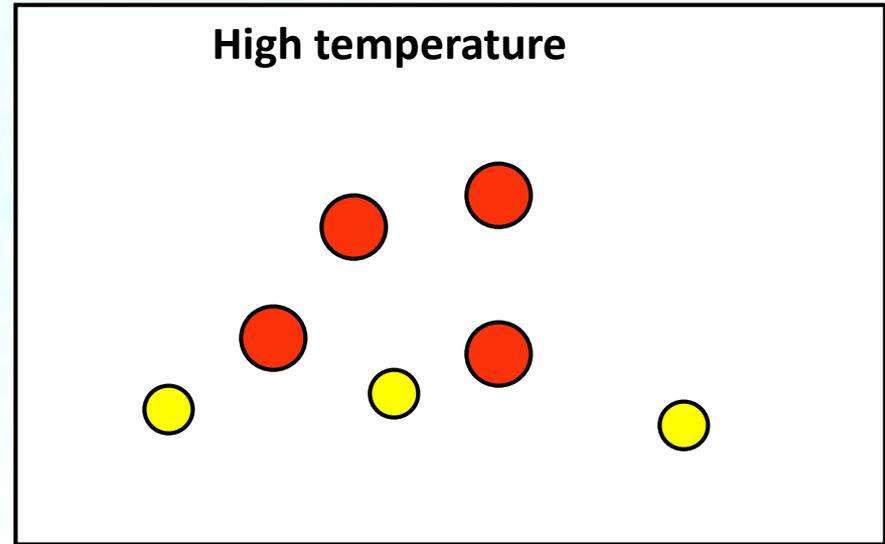
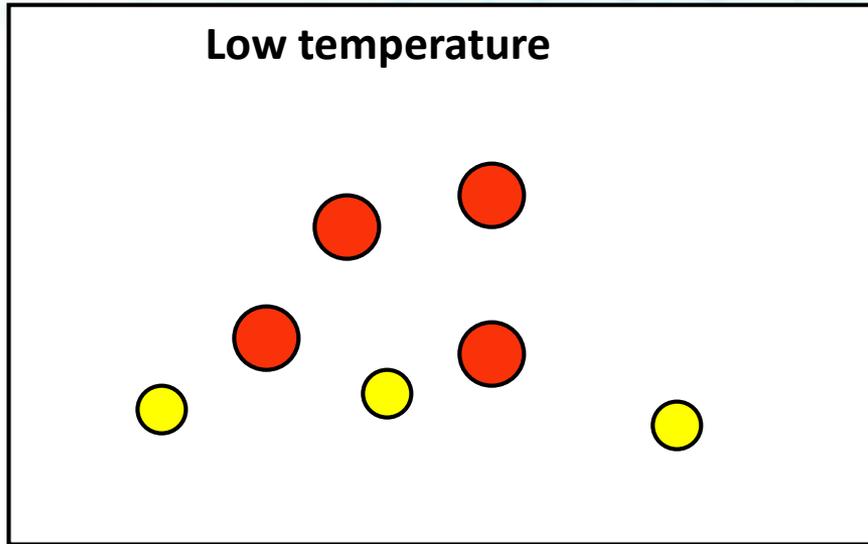
By increasing surface area a greater number of reactant particles are exposed and therefore able to collide. The frequency of collisions (number of collisions per unit of time) will increase and therefore the frequency of successful collisions so the reaction rate will also increase.



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<sub>2</sub>) 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

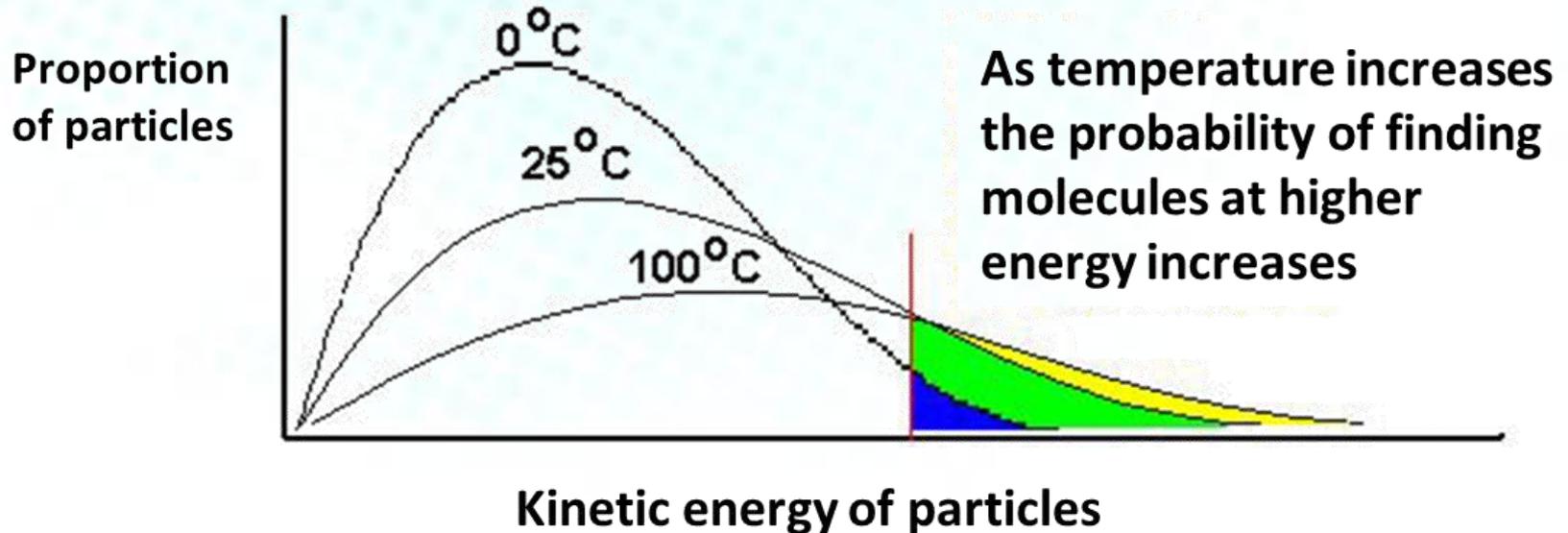


Increasing temperature effects the reaction rate in two ways.

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



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.

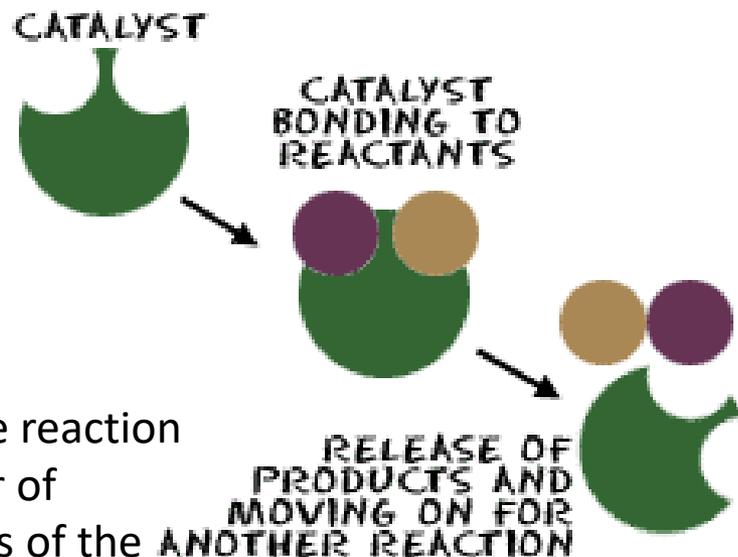
## Reaction rate can be increased by using a catalyst.

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.

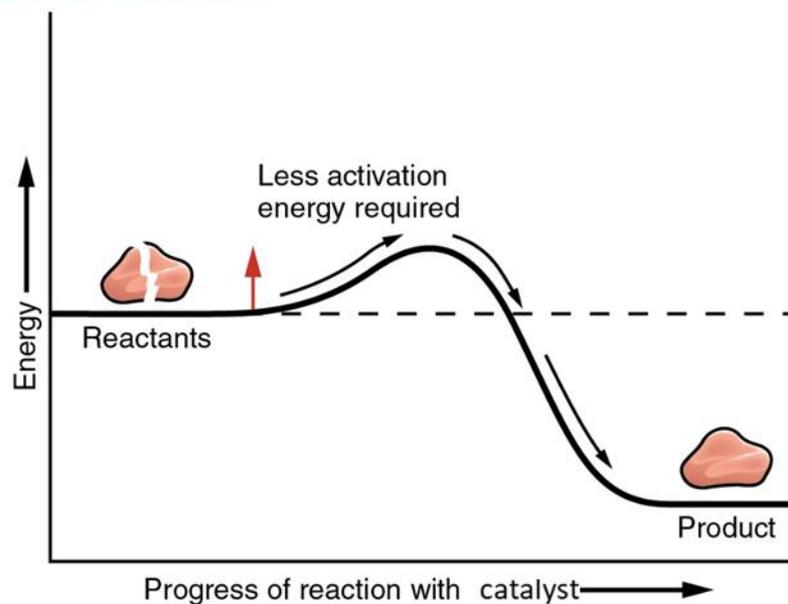
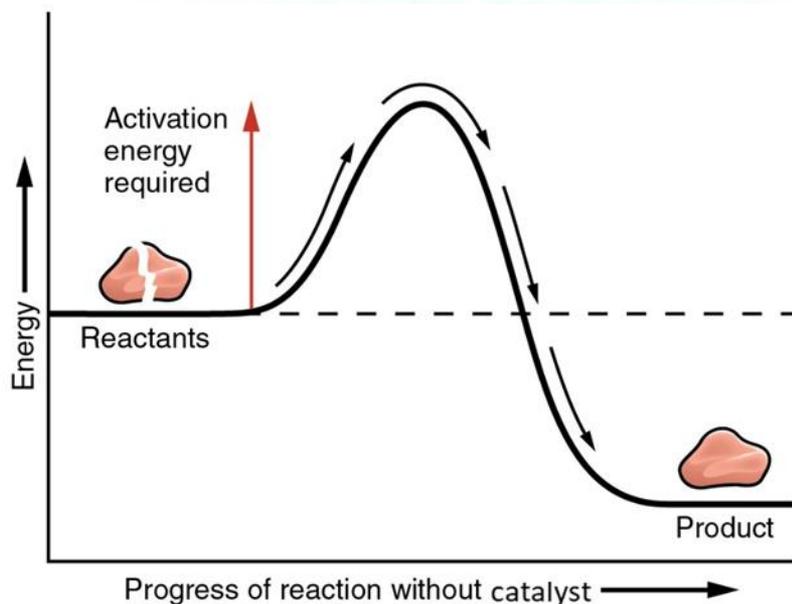


A catalyst provides a surface on which the reaction can take place. This increases the number of successful collisions between the particles of the substances as the particles come in **closer contact with each other and are more likely to be in the correct orientation.**

## Reaction rate can be increased by using a catalyst.

The reaction shown below is an exothermic reaction: the enthalpy level of the products is lower than the enthalpy of the reactants 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.



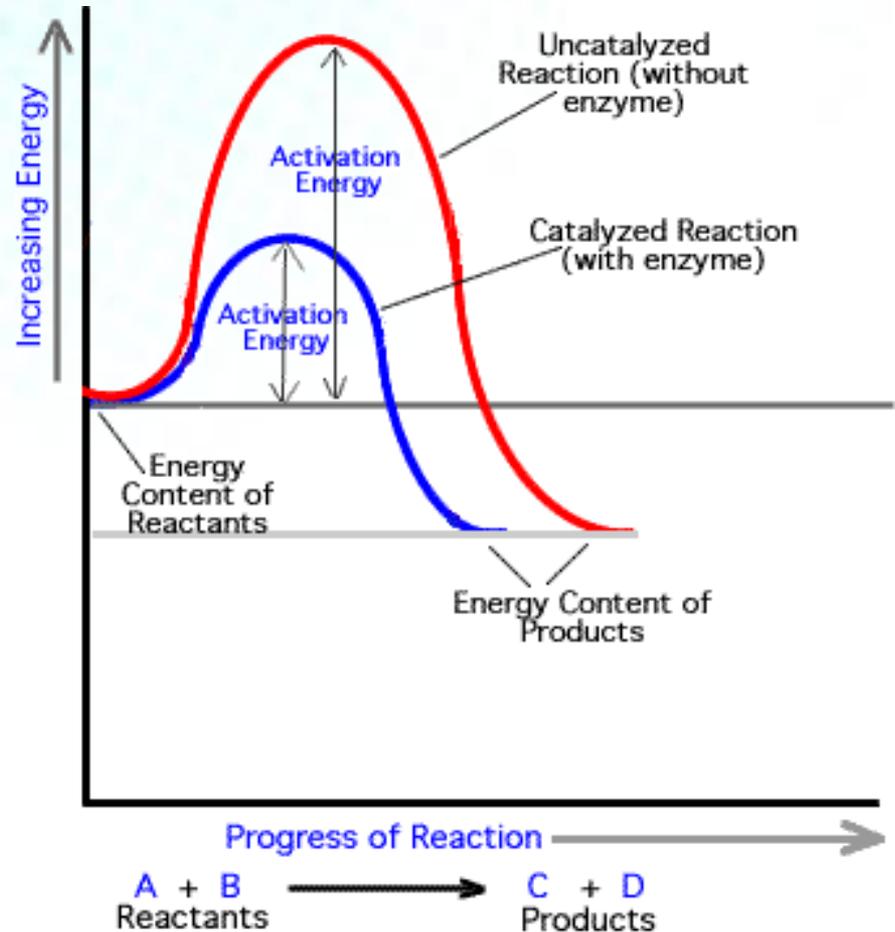
# Biological catalysts



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  $37.5^{\circ}\text{C}$  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

### Increase the frequency of collisions

➤ 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

### Increase the energy of collisions

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

Note: increasing temp also increases frequency of collisions

### Make it easier for reaction to occur

>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.**

# Reaction rates

**Factor** (if increased)

Surface Area

Increase the frequency of collisions

Concentration

Temperature

Increase the energy of particles

Catalyst

Lowers the activation energy pathway

Increase the frequency of collisions

More (%) collisions are successful

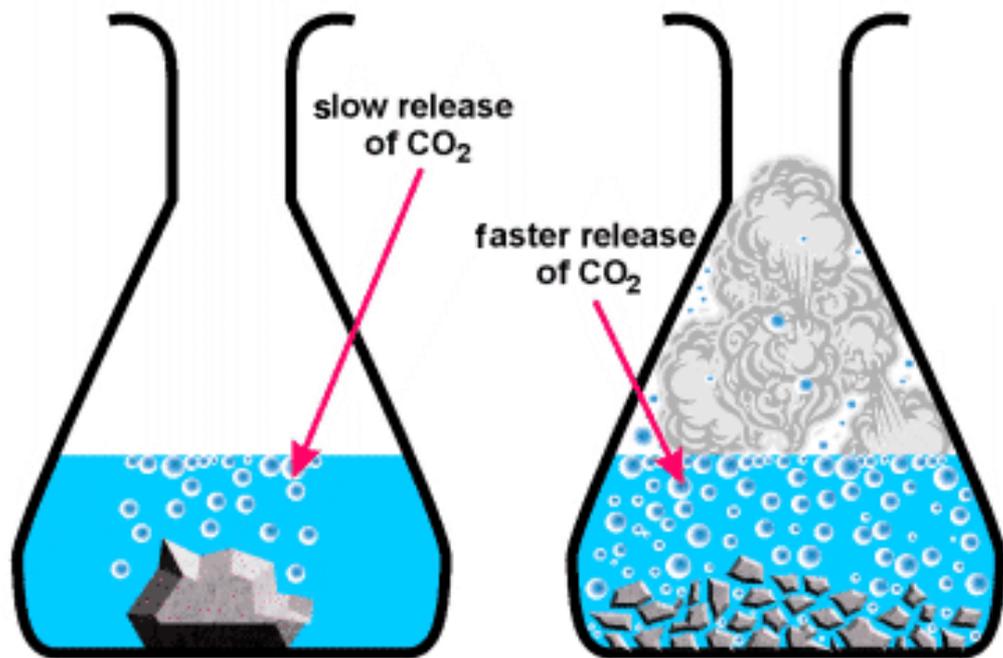
Increased Reaction Rate

## NCEA 2013 Reaction rate

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..

**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)

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 /°C	Time for cross to disappear
1	No Cu <sup>2+</sup> present	25	42
2	No Cu <sup>2+</sup> present	50	23
3	Cu <sup>2+</sup> 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 more effective collisions are occurring more frequently**.

## NCEA 2013 Reaction rate - (PART TWO)

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.

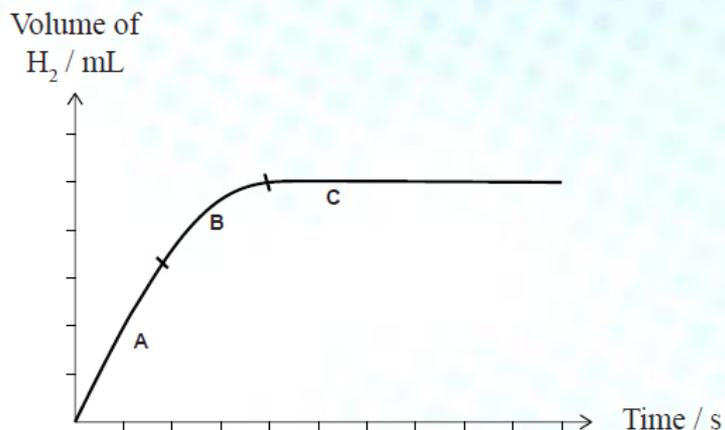
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1	No Cu <sup>2+</sup> present	25	42
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Elaborate on why the reactions in **Experiment 2** and **Experiment 3** occur faster than the reaction in **Experiment 1**.

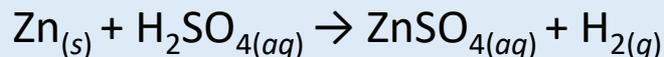
In Experiment Three, a catalyst is used (the copper ions). **Use of a catalyst speeds up the rate of chemical reaction.** For a chemical reaction to occur, the reactants must collide effectively. This means they must collide with enough energy to overcome the activation energy of the reaction. The activation 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 more effective collisions are occurring more frequently**.

## NCEA 2014 Reaction Rates - (PART ONE)

Merit  
Question



**Question 3a:** The equation for the reaction between zinc granules (lumps), Zn<sub>(s)</sub>, and sulfuric acid, H<sub>2</sub>SO<sub>4(aq)</sub>, is represented by:



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.

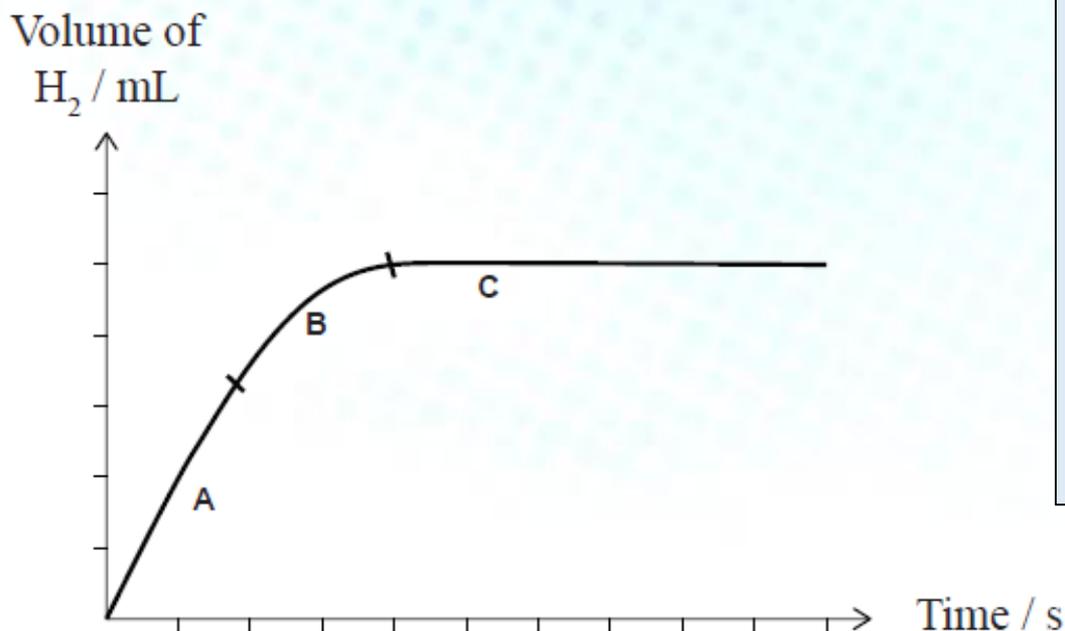
**A:** Reaction rate starts off high (steep slope) since there is a high concentration of reactants or many Zn particles at the start, so the frequency of collisions is highest, therefore the reaction rate is highest.

**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



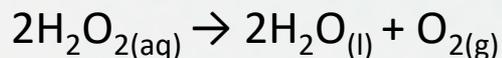
**Question 3b:** The rate of the reaction between zinc and sulfuric acid can be changed by the addition of small pieces of copper,  $\text{Cu}(s)$ , as a catalyst. Explain the role of the copper catalyst in the reaction between zinc and sulfuric acid.

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.

**Question: 1a:** The 'elephant toothpaste' demonstration shows the decomposition of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , into water and oxygen gas.



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.

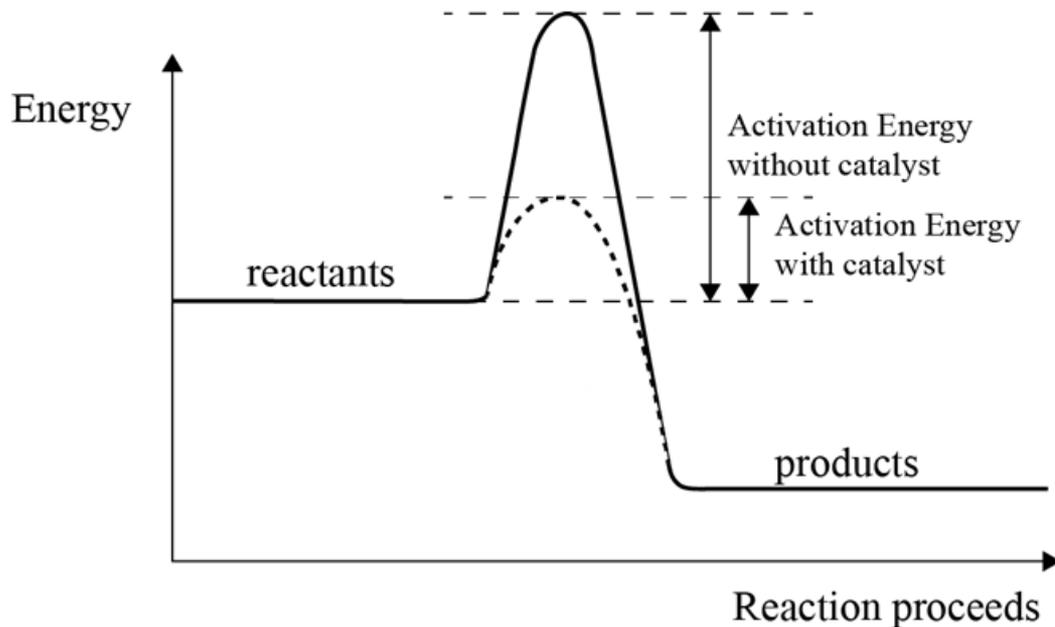
(a) The decomposition reaction of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , is very slow. By adding a small amount of powdered manganese dioxide,  $\text{MnO}_2$ , the rate of the reaction can be increased.

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

**Answer: 1a:** The added  $\text{MnO}_2$  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.

(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  $\text{MnO}_2$  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 $\text{H}_2\text{O}_2$	Temperature $^\circ\text{C}$	Presence of small amount of $\text{MnO}_2$
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 $\text{H}_2\text{O}_2$	Temperature $^\circ\text{C}$	Presence of small amount of $\text{MnO}_2$
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:

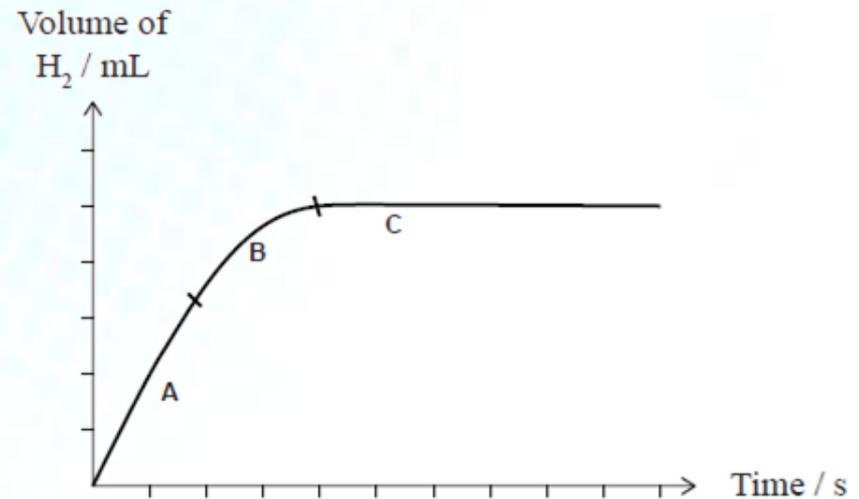
- identify the factor being changed, and the effect this will have on the rate of reaction
- explain the effect on the rate of reaction by referring to the collision of particles and activation energy, where appropriate.

**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:  
$$\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$
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<sub>2</sub> and H<sub>2</sub> 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<sub>2</sub> 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<sub>2</sub> 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<sup>-1</sup> hydrochloric acid, HCl(aq), and 0.5 g of **magnesium powder**, Mg(s), with 50.0 mL of 0.100 mol L<sup>-1</sup> 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.  
$$\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$
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 H<sub>2</sub> gas will be produced initially, resulting in a faster rate of reaction.

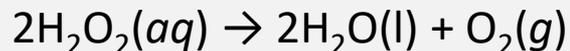
Mg ribbon will take longer to react because fewer particles are immediately available to collide, so will have a slower rate of reaction.

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

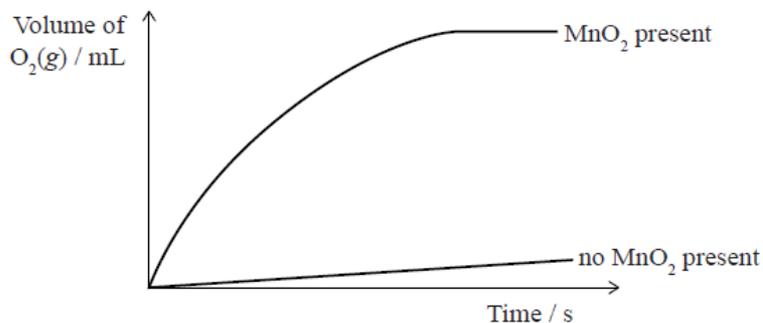
**Question 1(c):** The decomposition reaction of hydrogen peroxide solution,  $\text{H}_2\text{O}_2(aq)$ , is a slow reaction. This reaction is represented by the equation:



The rate of the decomposition reaction can be changed by adding a small amount of manganese dioxide,  $\text{MnO}_2(s)$ . The graph below shows the volume of oxygen gas formed in the reaction with and without manganese dioxide,  $\text{MnO}_2(s)$ .

- State the role of manganese dioxide,  $\text{MnO}_2(s)$ , in this reaction.
- Elaborate on how manganese dioxide,  $\text{MnO}_2(s)$ , changes the rate of the decomposition reaction of the hydrogen peroxide,  $\text{H}_2\text{O}_2(aq)$ .

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



**Answer:** the  $\text{MnO}_2$  is a catalyst.

$\text{MnO}_2$  provides an alternative pathway with lower activation energy for the decomposition of  $\text{H}_2\text{O}_2$ . 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  $\text{MnO}_2$  is required because catalysts are not used up in this reaction.)

**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.



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

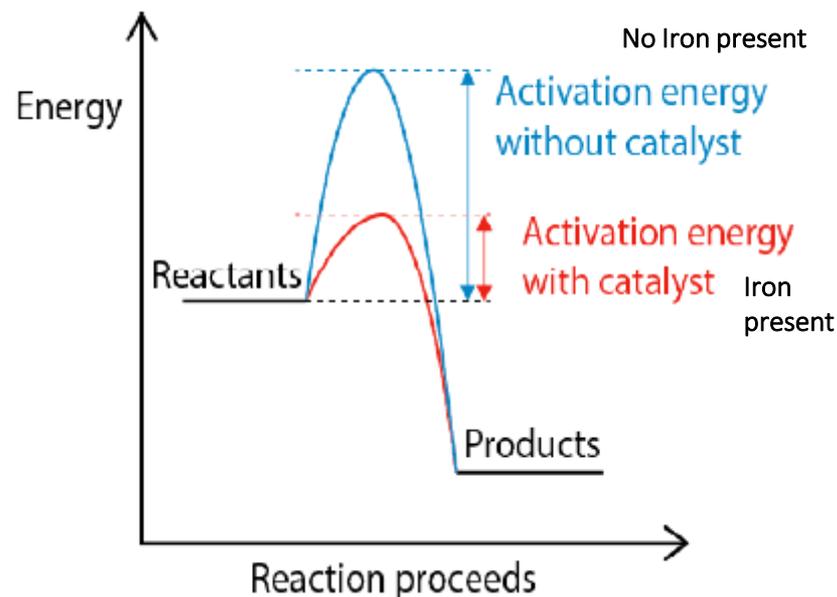
In your answer, you should refer to activation energy and collision theory.

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,  $\text{CaCO}_{3(s)}$ , and a solution of hydrochloric acid,  $\text{HCl}_{(aq)}$ .

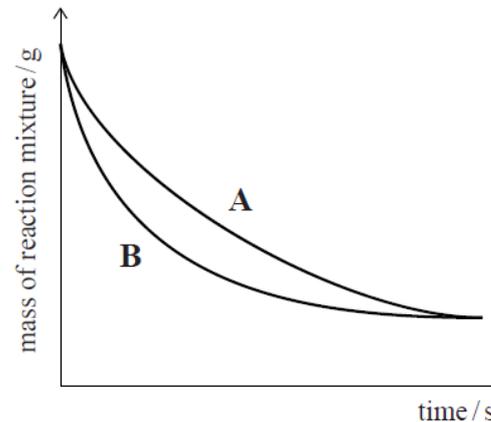
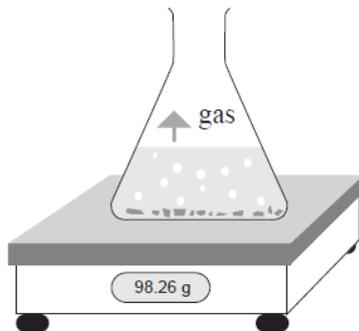
As the reaction proceeds, the mass of the reaction mixture decreases as carbon dioxide gas,  $\text{CO}_{2(g)}$ , escapes.

This is represented on the graph below.

Line A represents the reaction occurring at  $20^{\circ}\text{C}$  and line B represents the reaction occurring at  $40^{\circ}\text{C}$ .

Compare and contrast the reaction between calcium carbonate powder,  $\text{CaCO}_{3(s)}$ , and a solution of hydrochloric acid,  $\text{HCl}_{(aq)}$  at two temperatures:  $20^{\circ}\text{C}$  and  $40^{\circ}\text{C}$ , assuming all other conditions are kept the same.

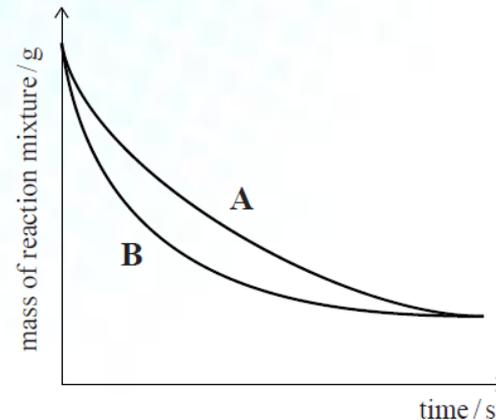
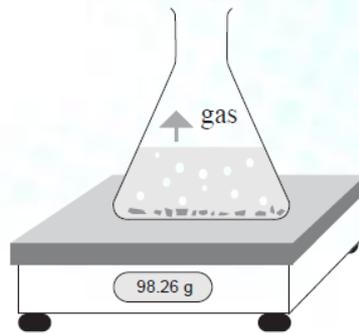
Your answer should refer to collision theory and rates of reaction.



## NCEA 2017 Reaction Rates

Excellence  
Question

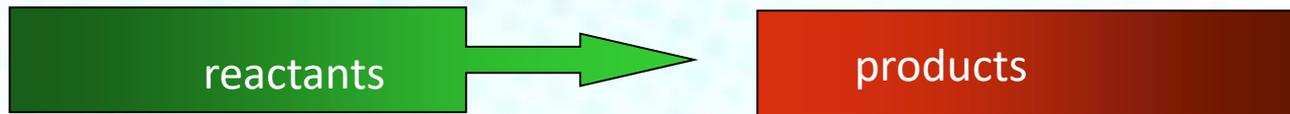
**Question 3a:** Compare and contrast the reaction between calcium carbonate powder,  $\text{CaCO}_{3(s)}$ , and a solution of hydrochloric acid,  $\text{HCl}_{(aq)}$  at two temperatures:  $20^\circ\text{C}$  and  $40^\circ\text{C}$ , assuming all other conditions are kept the same.



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  $\text{CO}_2$  gas being lost in a shorter period of time at  $40^\circ\text{C}$  than  $20^\circ\text{C}$ . 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.

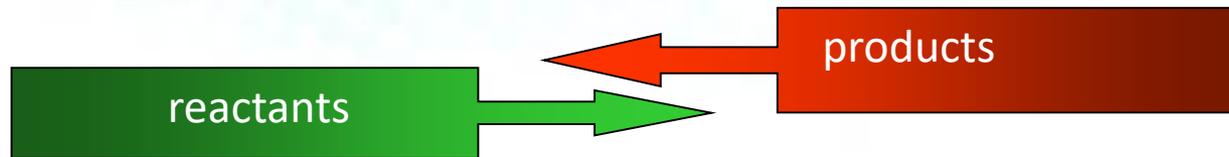
# Equilibrium

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.

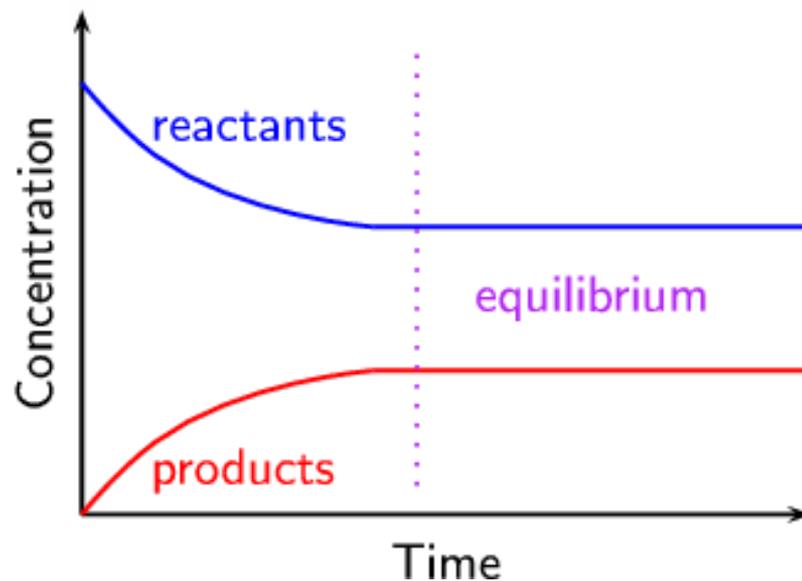
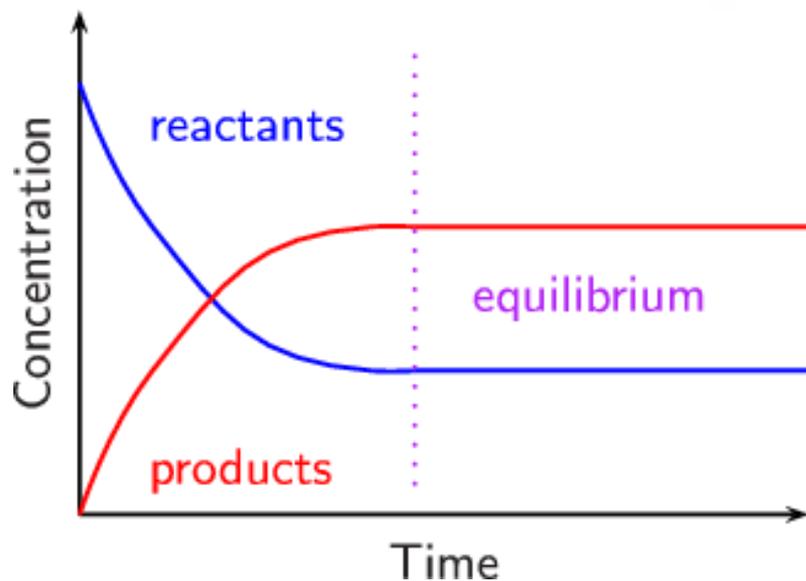
However, both the forward and reverse reactions are continuing

# 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.

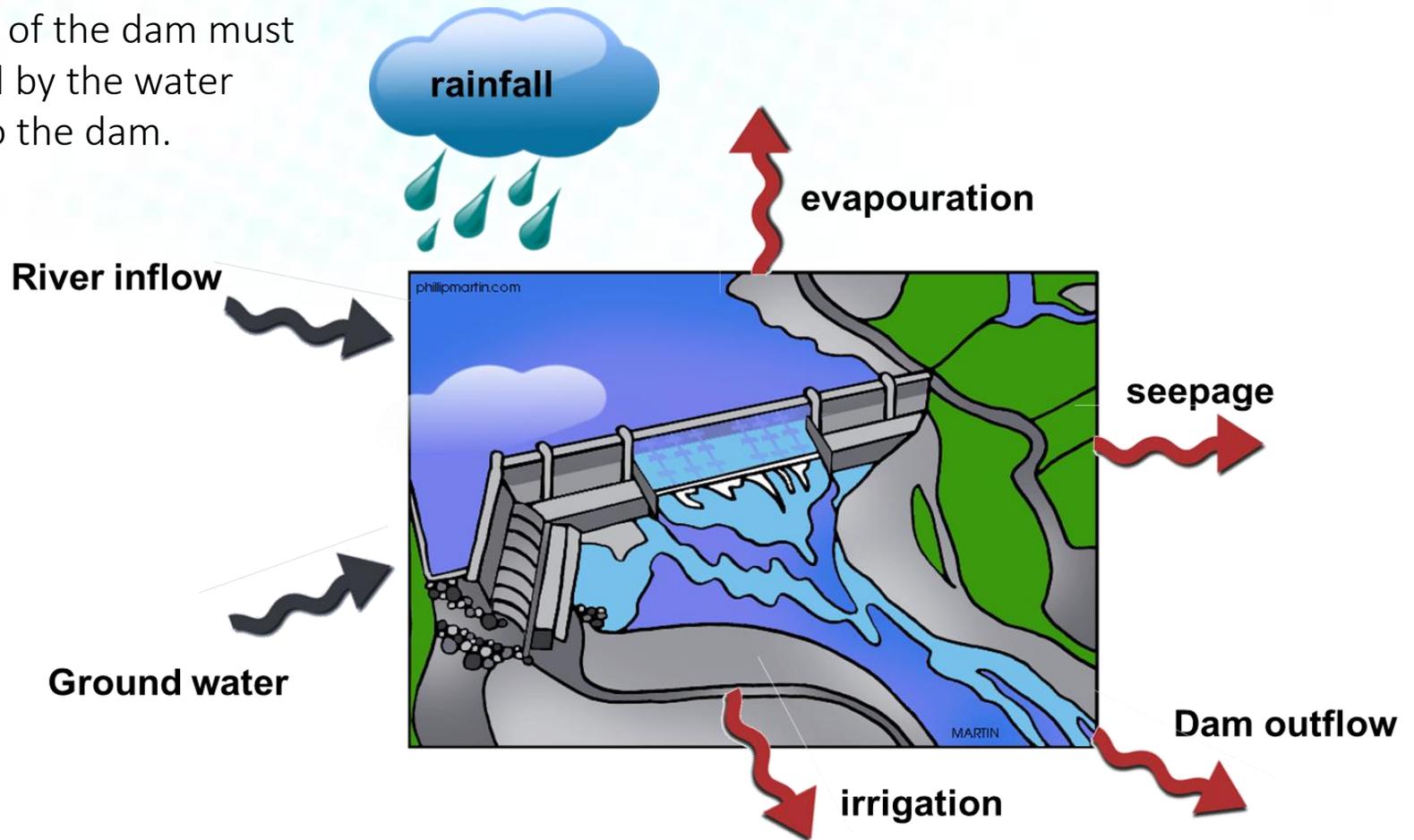
On the left hand side, the proportion of products will be higher than the reactants and on the right hand side, the proportion of reactants will be higher than products.



## 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.



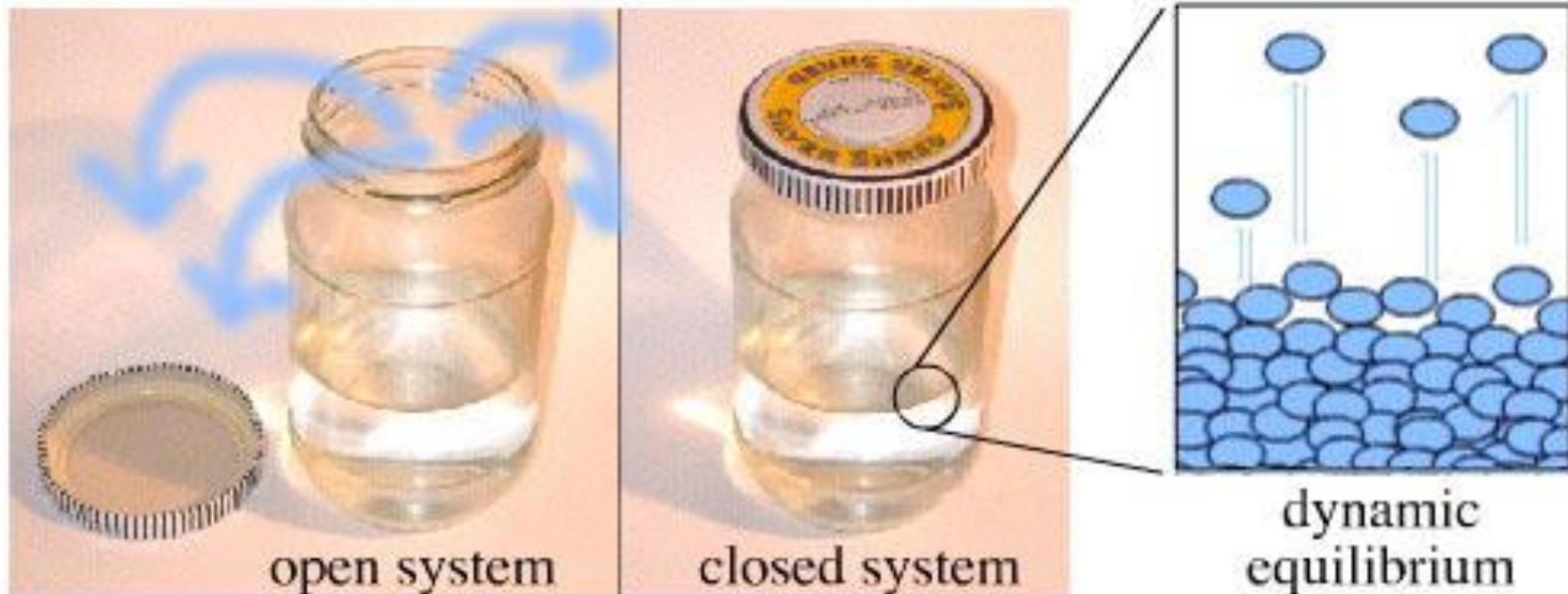
## Equilibrium – Evaporation case study

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.

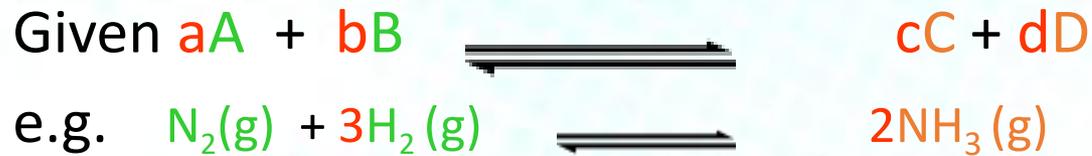
In an **open system**, the gas will escape and gradually the water level will decrease.

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 indicate 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.

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

[ ] = concentration in  $\text{molL}^{-1}$   
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.

## Equilibrium Constant: $K_c$

The size of  $K$  calculated gives information as to how far a reaction has proceeded

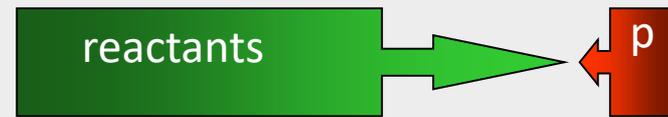
### Large $K$ value eg $K=10000$

Large amount of products produced.  
Reaction is close to completion when equilibrium was reached



### Small $K$ value eg $K=0.0001$

Small amount of products produced .  
Reaction only just underway when equilibrium was reached



**Note:** equilibrium does not necessarily mean there are equal amounts of reactants and products present

## Calculating the Equilibrium Constant: $K_c$

To calculate the equilibrium expression ( $K_c$ ) place the given concentrations in  $\text{molL}^{-1}$  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)



If the values given for concentration are  $[\text{NH}_3] = 0.325\text{molL}^{-1}$ ,  $[\text{N}_2] = 0.625\text{molL}^{-1}$  and  $[\text{H}_2] = 2.45\text{molL}^{-1}$

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

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

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

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

Step 2. expand out equation

Step 3. divide equation

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

## Changes to the Equilibrium Constant: $K_c$

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.

A temperature change to a system will permanently change the  $K_c$  value.

Questions asked about changes to a system ask for **observations immediately after a change has been made.**

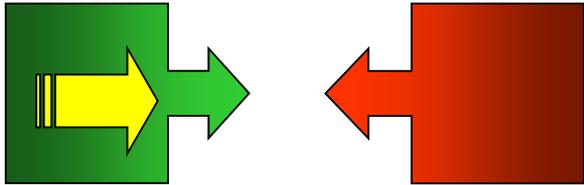
Remember, that unless a temperature change is made, the eventual observation of the system will be that it is the same as prior to the change due to the equilibrium being re-established.

## Changes in Equilibrium

A system stays in equilibrium unless a change is made

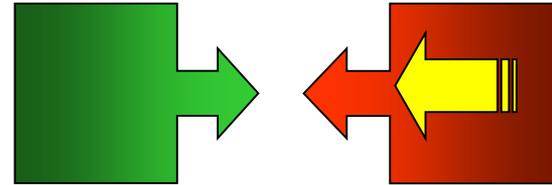
A change made to a system in equilibrium will **either**

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

## Le Chatelier's Principle

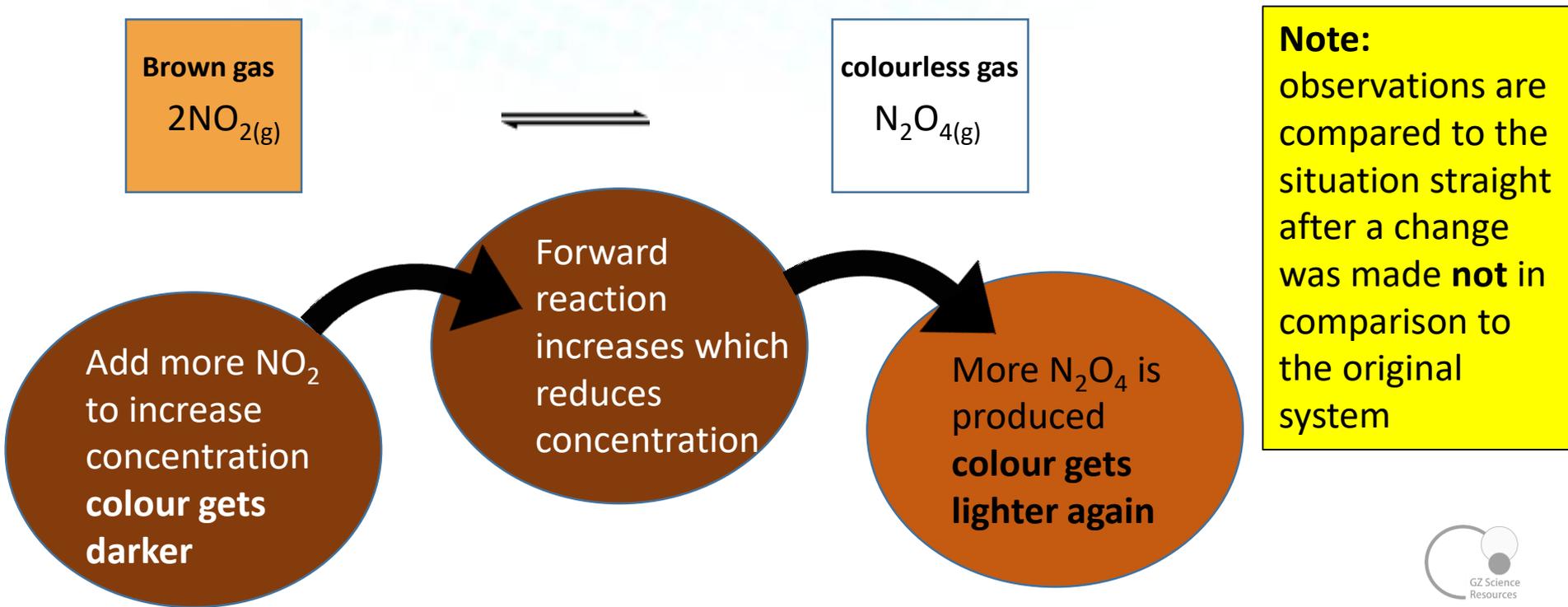
When a change is applied to a system at equilibrium, the system responds so that the effects of the change are minimised

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

## Changes in Equilibrium – Increasing Concentration

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.

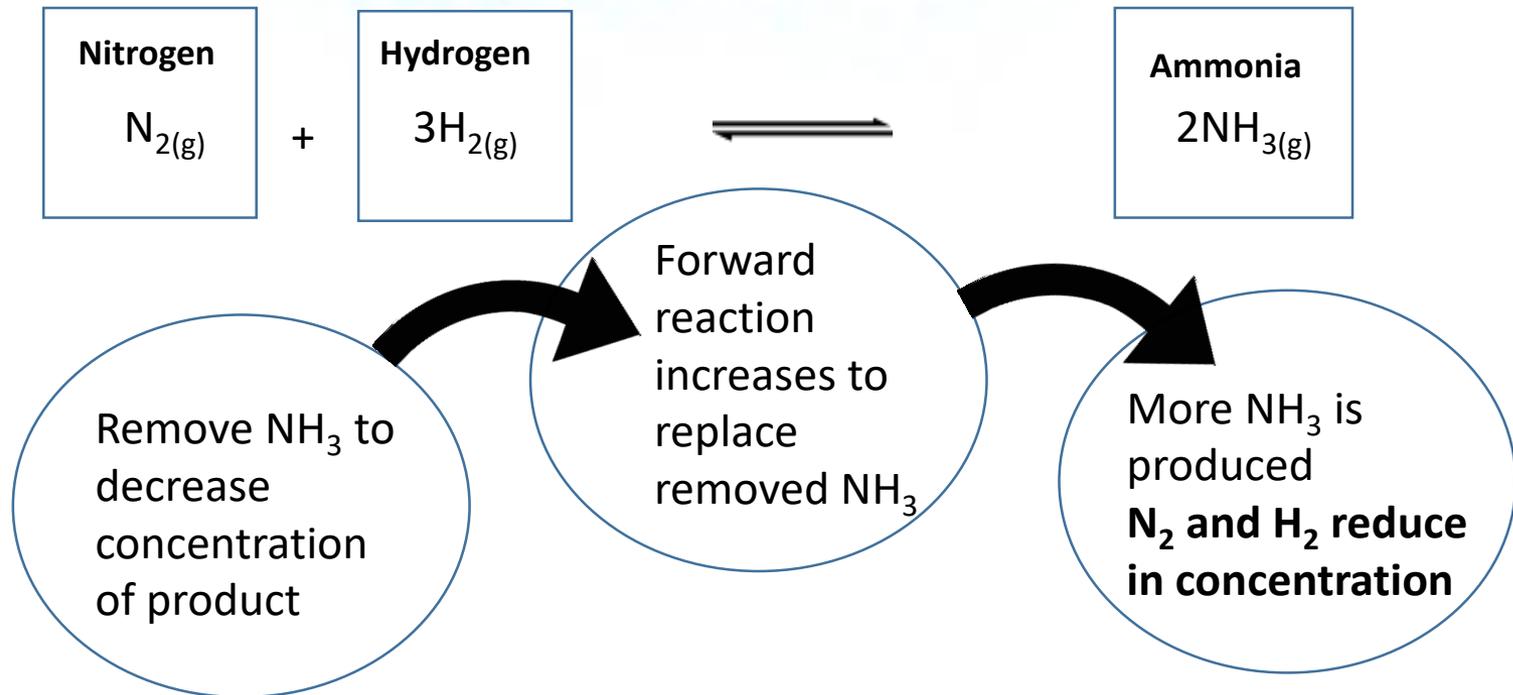
For example: if the concentration of the reactants is increased then the rate of reaction changing reactants in products will increase (due to collision theory) favouring the forward reaction. The effect of this will change more of the reactant into product – lowering the concentration of reactant and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



## 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.

For example: 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.

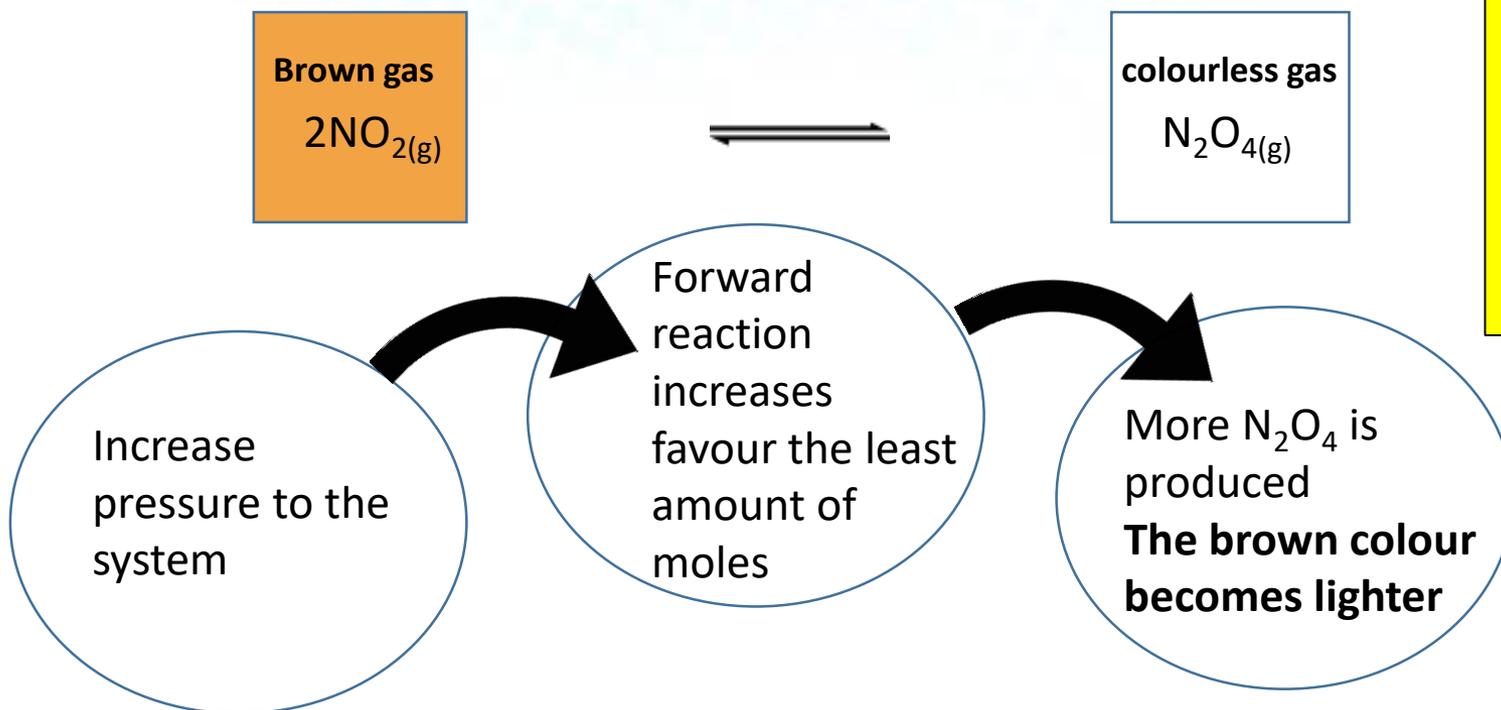


## Changes in Equilibrium – Increasing Pressure

**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.

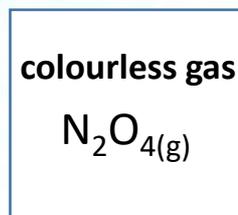
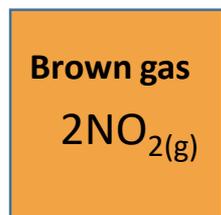


**Note:** if both sides have the same number of moles then there will be no change.

## 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  $\text{NO}_2$  has 2 moles compared to 1 mole of the  $\text{N}_2\text{O}_4$  so the reaction rate from the product to the reactant is increased.



More  $\text{NO}_2$  is produced  
**The brown colour becomes darker**

Reverse reaction increases favour the most amount of moles

Decrease pressure to the system

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

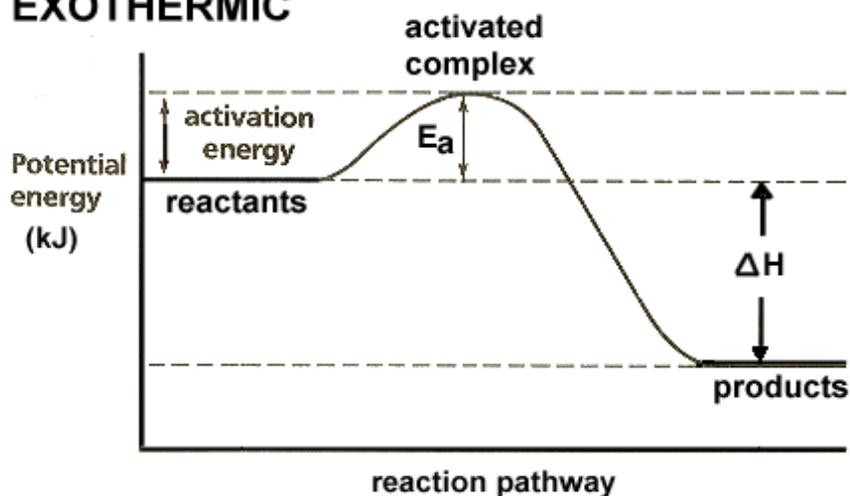
# Endothermic and Exothermic reactions

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

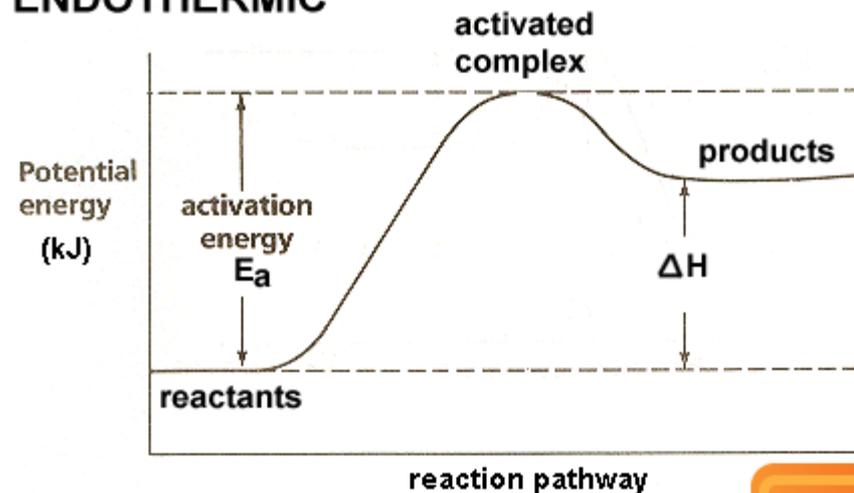
An endothermic reaction will absorb heat energy from the surrounding area as the products contain more enthalpy than the reactants. The temperature of the closed system will decrease.

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.

## EXOTHERMIC



## ENDOTHERMIC

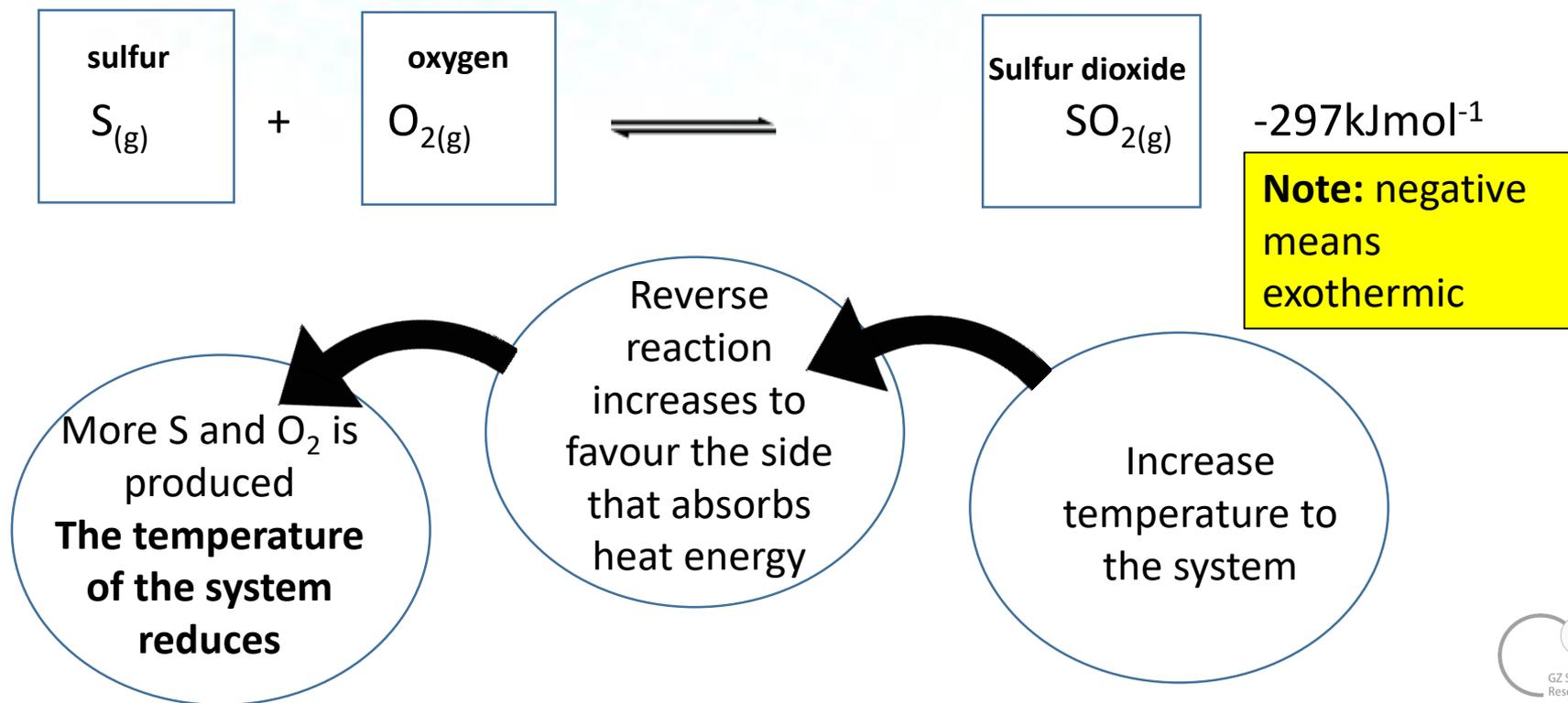


There is a difference in activation energy between the endothermic and exothermic direction. There will be uneven proportions of energy involved in activation energy and substance creation if the equilibrium is shifted. This changes the  $K_c$  value permanently.

## Changes in Equilibrium – Increasing Temperature

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

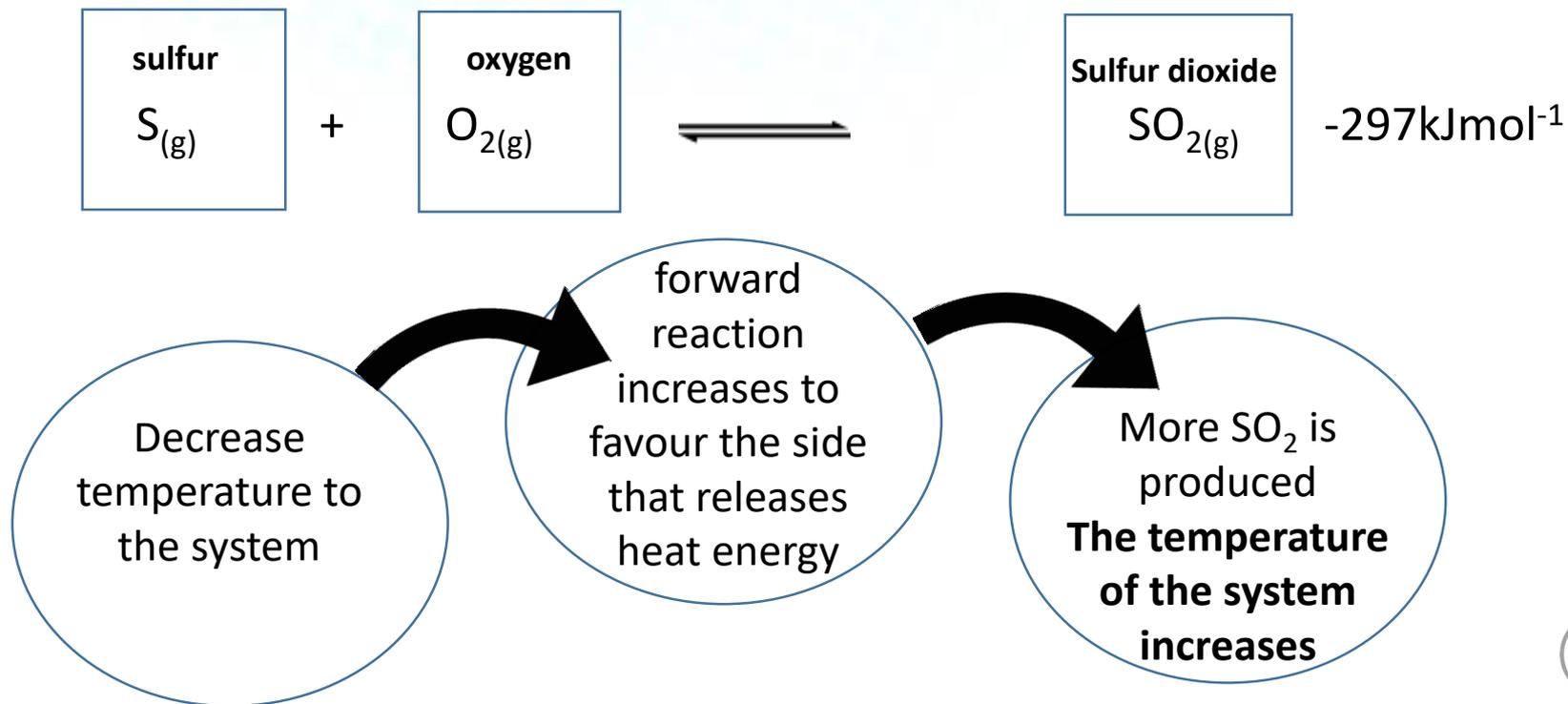
When a reaction is exothermic, it releases heat energy in the forward reaction when products are formed. An increase in temperature will favour the reverse reaction which is endothermic and more reactants will be made. (An endothermic reaction will favour the forward reaction if temperature is increased)



## 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.



## 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

The two reactions shown in the following table are both at equilibrium.

Reaction	Equation	Affected by increased pressure
One	$\text{H}_{2(g)} + \text{I}_{2(g)} \rightarrow 2\text{HI}_{(g)}$	no
Two	$\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightarrow 2\text{NH}_{3(g)}$	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.

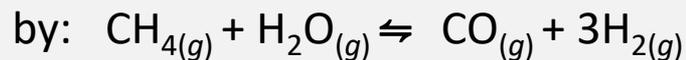
In Reaction One there are two moles of gas particles on each side of the equation. Because there are the **same numbers of gas particles on both sides** of the reaction, then a change in pressure will have **no effect** as neither reaction will be favoured. In Reaction Two however, there are four moles 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

Excellence  
Question

**Question 2a (i)** : Hydrogen can be produced industrially by reacting methane with water.

An equation for this reaction can be represented



$K_c = 4.7$  at  $1127^\circ\text{C}$

(a) (i) Complete the equilibrium constant expression for this reaction:

**Answer:**

$$K_c = \frac{[\text{H}_2]^3 [\text{CO}]}{[\text{CH}_4] [\text{H}_2\text{O}]}$$

Gas	$\text{CH}_4$	$\text{H}_2\text{O}$	$\text{CO}$	$\text{H}_2$
Concentration/mol L <sup>-1</sup>	0.0300	0.0500	0.200	0.300

**Question 2a (ii)**: The concentrations of the four gases in a reaction mixture at  $1127^\circ\text{C}$  are found to be:  
(see above)

Use these values to carry out a calculation to determine if the reaction is at equilibrium.

$K_c = 4.7$  at  $1127^\circ\text{C}$

**Answer:**

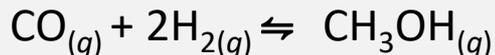
No the system is not at equilibrium.

*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  $1127^\circ\text{C}$ .

## NCEA 2014 Equilibrium

Merit  
Question

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



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<sub>3</sub>OH<sub>(g)</sub> 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<sub>2(g)</sub> is removed. Amount of CH<sub>3</sub>OH<sub>(g)</sub> would:*

**Answer: H<sub>2</sub> removed:**

Amount of methanol will **decrease**.

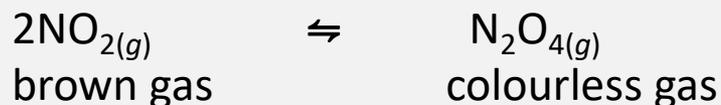
As hydrogen gas is removed, the system will oppose the change and the position of the equilibrium will shift in the reverse direction as more hydrogen is formed. This means the amount of methanol will decrease.

## NCEA 2014 Equilibrium - (PART ONE)

Merit  
Question

**Question 2c:** In a reaction, the brown gas nitrogen dioxide,  $\text{NO}_{2(g)}$ , exists in equilibrium with the colourless gas dinitrogen tetroxide,  $\text{N}_2\text{O}_{4(g)}$ .

The equation for this reaction is represented by:



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  $\text{N}_2\text{O}_4(g)$ , 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  $\text{NO}_2$ , resulting in a darker colour due to the brown colour of this gas.

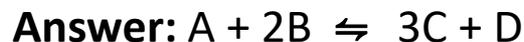
When the reaction container is placed in **hot water**, the system will adjust so that some of the heat is used up; therefore it will shift in the **endothermic direction**. In this case, the colour darkened, indicating that this favoured the reverse reaction which must be the endothermic direction. When the container was **cooled**, the colour faded indicating that the forward reaction, forming colourless  $\text{N}_2\text{O}_4$ , **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:

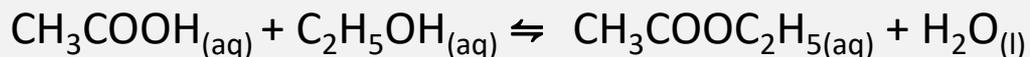
$$K_c = \frac{[C]^3[D]}{[A][B]^2}$$



Write the chemical equation for this reaction.

**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



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

Merit  
Question

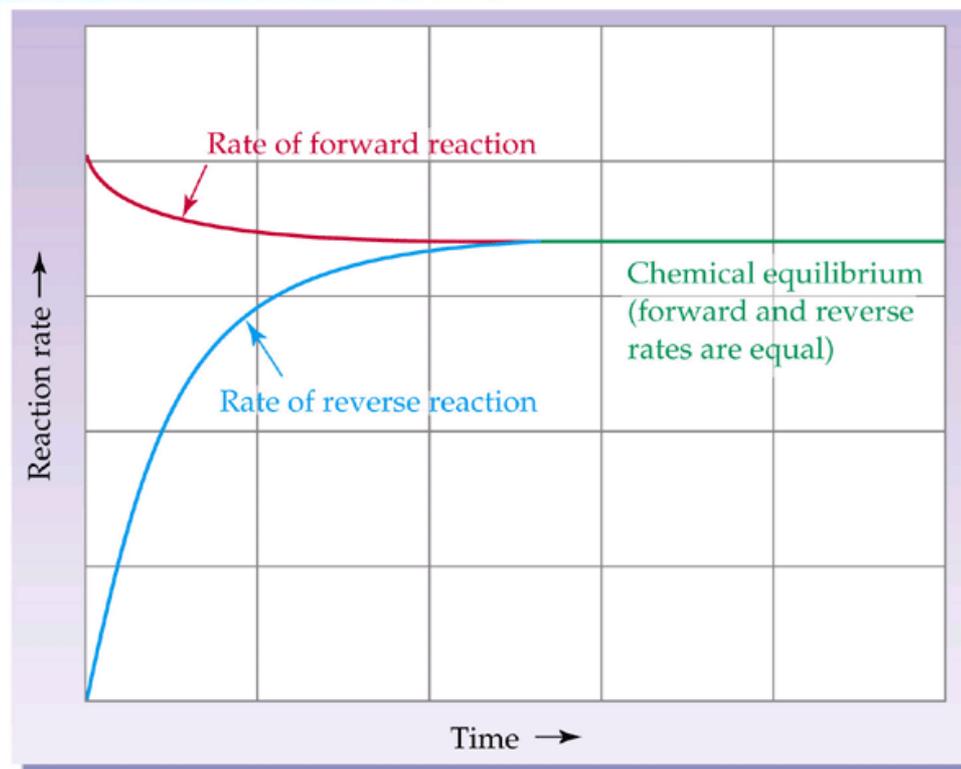
**Answer:** Adding more ethanol causes the equilibrium to move in the forward direction in order to use the extra added ethanol. This is because the equilibrium has to re-establish itself with the added reactant in order to maintain  $K_c$ .

**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.

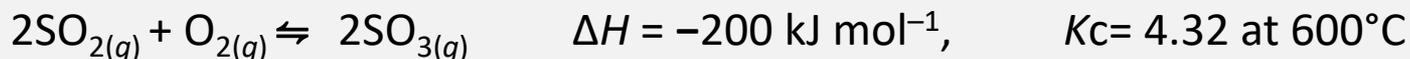
The reaction is quite slow, so a small amount of concentrated sulfuric acid is added as a catalyst.

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.



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



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

**Answer:**

$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

**Question 3(c)(ii):** A reaction mixture has the following concentration of gases at  $600^\circ\text{C}$ :

$$[\text{SO}_{2(g)}] = 0.300 \text{ mol L}^{-1}$$

$$[\text{O}_{2(g)}] = 0.100 \text{ mol L}^{-1}$$

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

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.

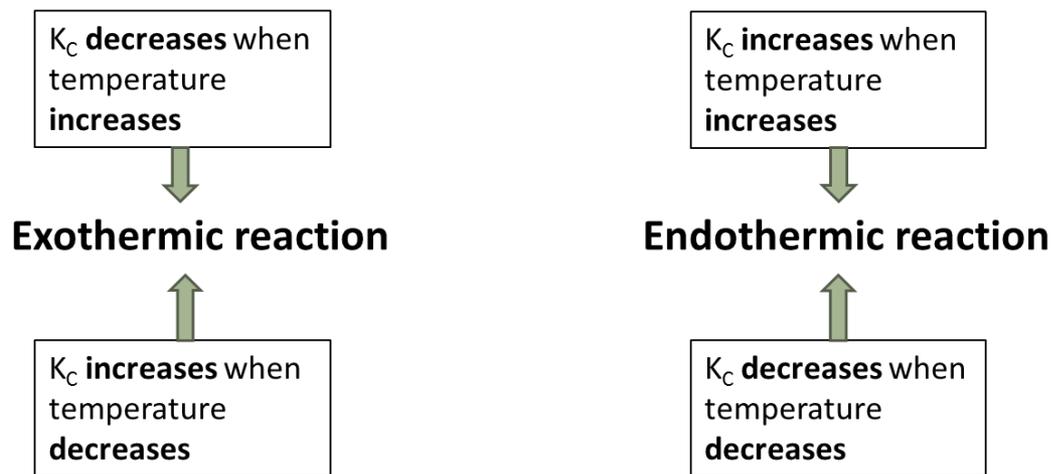
This number is greater than the  $K_c$  value, 4.32, which indicates that the reaction lies to the products side as the larger the  $K_c$  or  $Q$  value, the greater the numerator (products).

**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  $K_c$
- 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  $K_c$ .



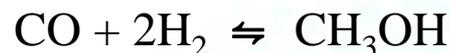
## NCEA 2016 Equilibrium

Achieved  
Question

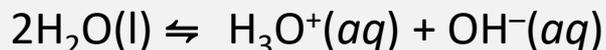
**Question 3(a):** The equilibrium constant expression for a reaction is:

$$K_e = \frac{[\text{CH}_3\text{OH}]}{[\text{CO}][\text{H}_2]^2}$$

Write the equation for this reaction.



**Question 3(b):** The ionisation of water is represented by the equation:



Give an account of the extent of ionisation of water, given  $K_w = 1 \times 10^{-14}$ .

Merit  
Question

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

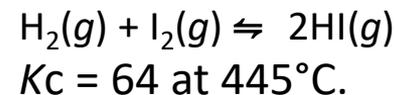


Merit  
Question

## NCEA 2016 Equilibrium (PART ONE)

### Question 3(d):

When hydrogen gas,  $\text{H}_2(g)$ , and iodine gas,  $\text{I}_2(g)$  are mixed, they react to form  $\text{HI}(g)$ , and an equilibrium is established.



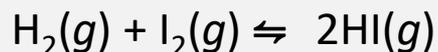
(i) Calculate the concentration of HI in an equilibrium mixture at  $445^\circ\text{C}$  when the concentrations of  $\text{H}_2(g)$  and  $\text{I}_2(g)$  are both  $0.312 \text{ mol L}^{-1}$ .

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

## NCEA 2016 Equilibrium (PART TWO)

Merit  
Question

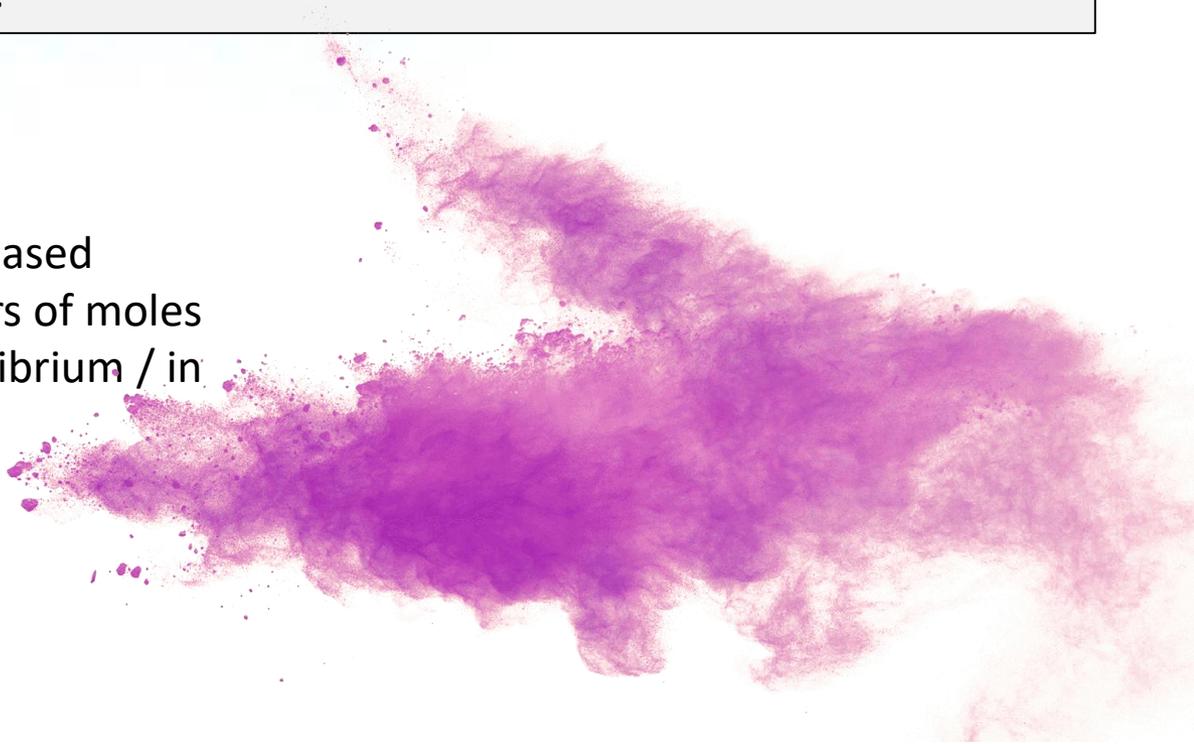
**Question 3(d):** When hydrogen gas,  $\text{H}_2(g)$ , and iodine gas,  $\text{I}_2(g)$  are mixed, they react to form  $\text{HI}(g)$ , and an equilibrium is established.



$K_c = 64$  at  $445^\circ\text{C}$ .

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

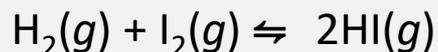
There would be no effect on the equilibrium if pressure was increased because there are equal numbers of moles of gas on either side of the equilibrium / in the reactants and products.



## NCEA 2016 Equilibrium (PART THREE)

Excellence  
Question

**Question 3(d):** When hydrogen gas,  $\text{H}_2(g)$ , and iodine gas,  $\text{I}_2(g)$  are mixed, they react to form  $\text{HI}(g)$ , and an equilibrium is established.



$K_c = 64$  at  $445^\circ\text{C}$ .

(iii) When the temperature of the equilibrium system is increased to  $510^\circ\text{C}$ , 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).

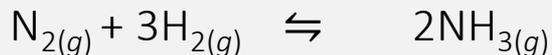
When the temperature increases, the reaction moves in the endothermic direction to absorb the added heat. In this reaction, the value of  $K$  decreased, indicating the ratio of products to reactants (numerator to denominator) decreased. Since there will be fewer products and more reactants, adding heat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

**Therefore, the forward reaction is exothermic.**

## NCEA 2017 Equilibrium (PART ONE)

Merit  
Question

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



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

$$K_c = \frac{[\text{NH}_3(aq)]^2}{[\text{N}_2(aq)][\text{H}_2(aq)]^3}$$

(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 <sub>2</sub>	H <sub>2</sub>	NH <sub>3</sub>
Concentration (mol L <sup>-1</sup> )	0.0821	0.0583	0.105

$$\begin{aligned} K_c &= \frac{0.105^2}{0.0821 \times 0.0583^3} \\ &= 677 \end{aligned}$$

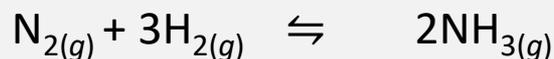
No, the reaction is not at equilibrium because  $677 > 640$  (values must be equal for a reaction to be at equilibrium). *Accept answers between 676 – 678.*

## 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,  $\text{NH}_{3(g)}$ , is an endothermic or exothermic reaction.



**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,  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$  and  $[\text{CoCl}_4]^{2-}$ , exist together in a solution in equilibrium with chloride ions,  $\text{Cl}^-_{(aq)}$ .

The forward reaction is endothermic;  $\Delta H$  is positive. The equation for this equilibrium is shown below.



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  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$  ion, so the solution will turn pink or the pink colour will intensify.

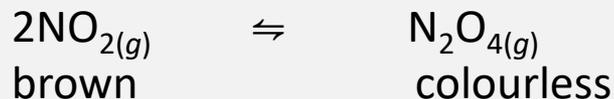
(ii) a test tube containing the reaction mixture is placed in a beaker of ice-cold water.

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  $\Delta H$  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,  $\text{NO}_{2(g)}$ , exists in equilibrium with the colourless gas, dinitrogen tetroxide,  $\text{N}_2\text{O}_{4(g)}$ .



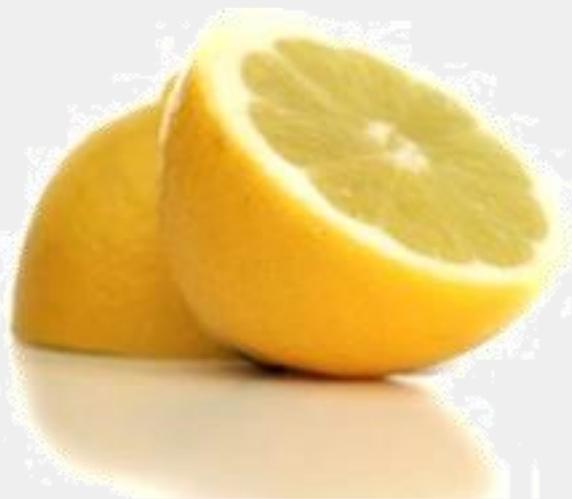
Explain using equilibrium principles, the effect of decreasing the volume of the container (therefore increasing the pressure) on the observations of this equilibrium mixture.

**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  $\text{N}_2\text{O}_4(g)$ , 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  $\text{NO}_2$ , resulting in a darker colour due to the brown colour of this gas.

## 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.

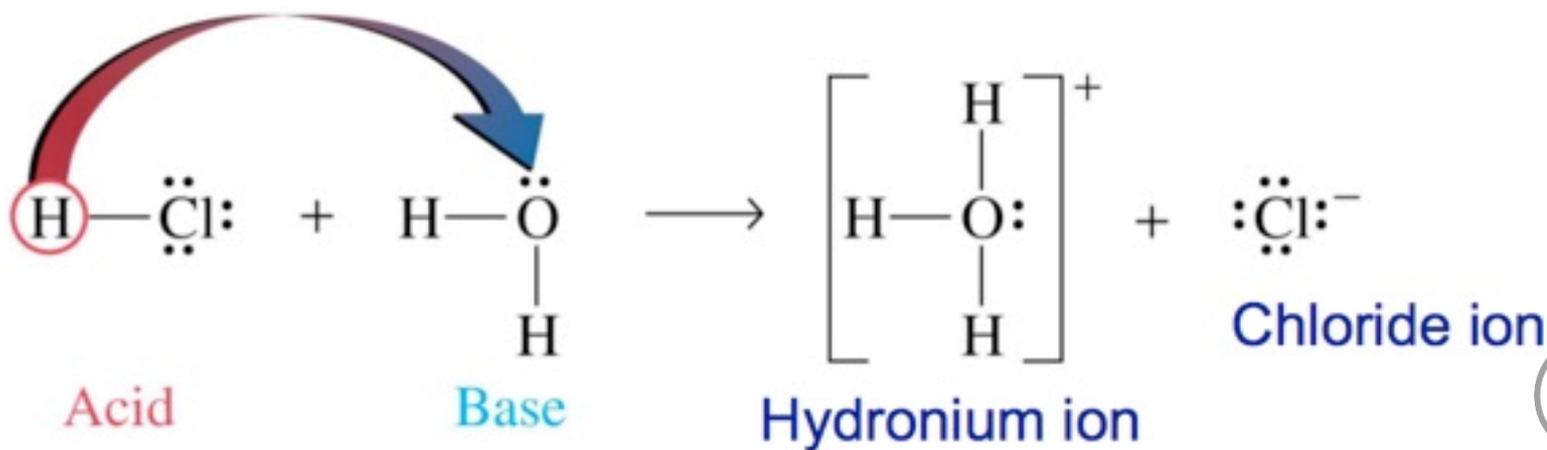
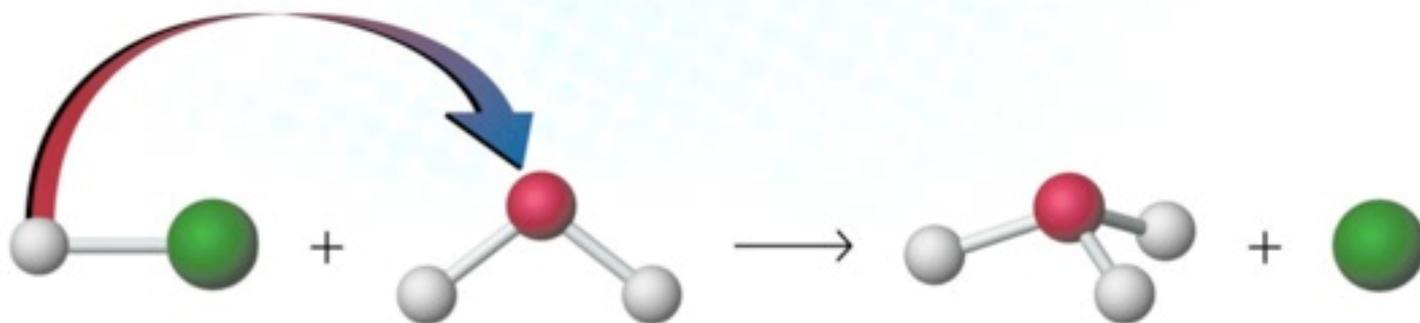
They have a **sour taste** and react with metals. Acids can be found in nature and called **organic acids** or manufactured in the laboratory and called **mineral acids**.



## Acids – their characteristics

An **Acid** donates its **Hydrogen ion** ( $\text{H}^+$ ), which is really just a proton - the electron remains behind.

Common acids include the strong acids  $\text{HNO}_3$  - nitric acid,  $\text{HCl}$  - hydrochloric acid,  $\text{H}_2\text{SO}_4$  - sulfuric acid, and the weak acid  $\text{CH}_3\text{COOH}$  - ethanoic acid.



## Bases – their characteristics

**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 tables to fix indigestion. **Bases that dissolve into water are called an alkali, and produce  $OH^-$  ions.**



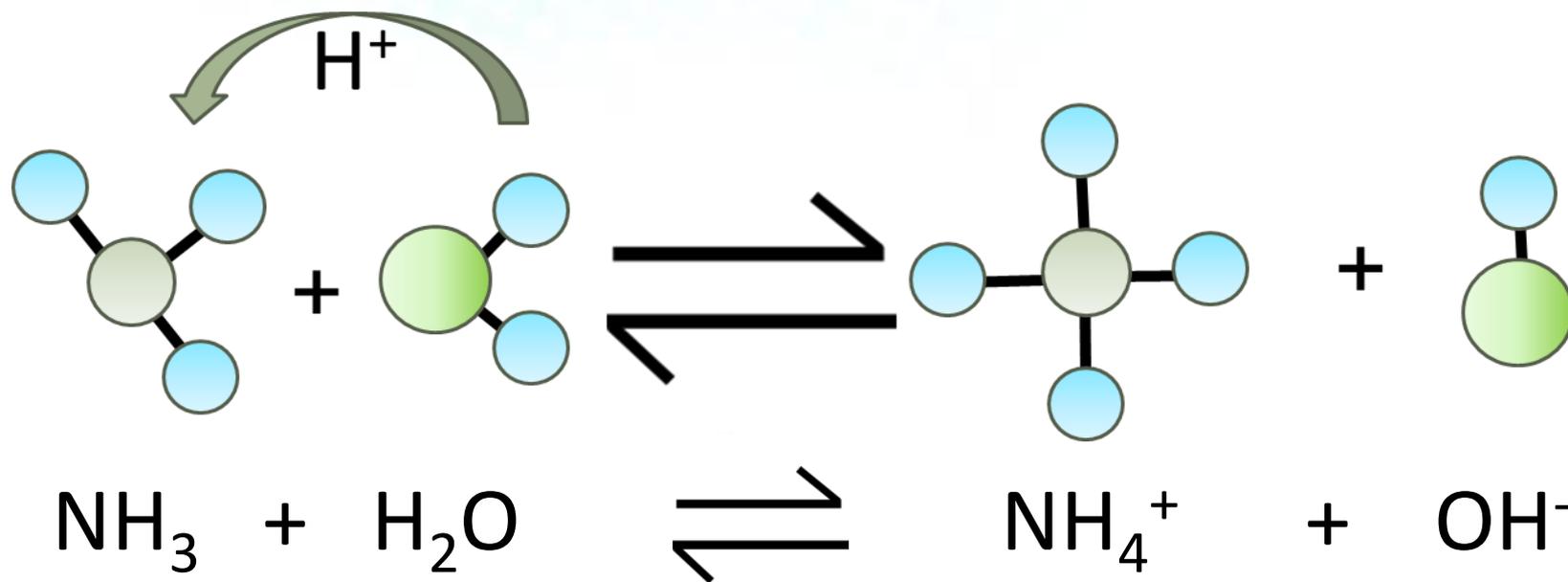
## Bases – their characteristics

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  $\text{NH}_3$  – ammonia.

An **alkali** is a base that is soluble in water and contains hydroxide ions.

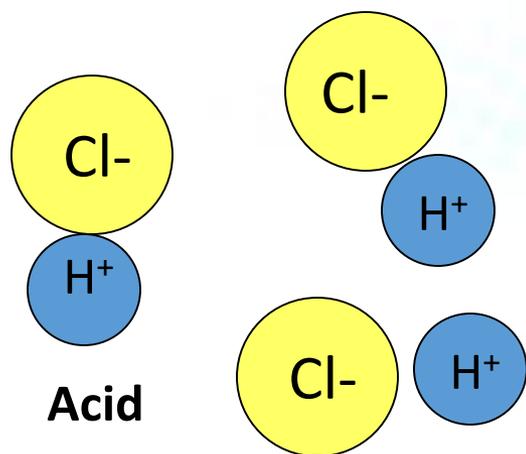
Some substances such as water are **amphiprotic** and can act as both an acid or a base depending on what other substance the water is with.



## Definition of Acid and Bases

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

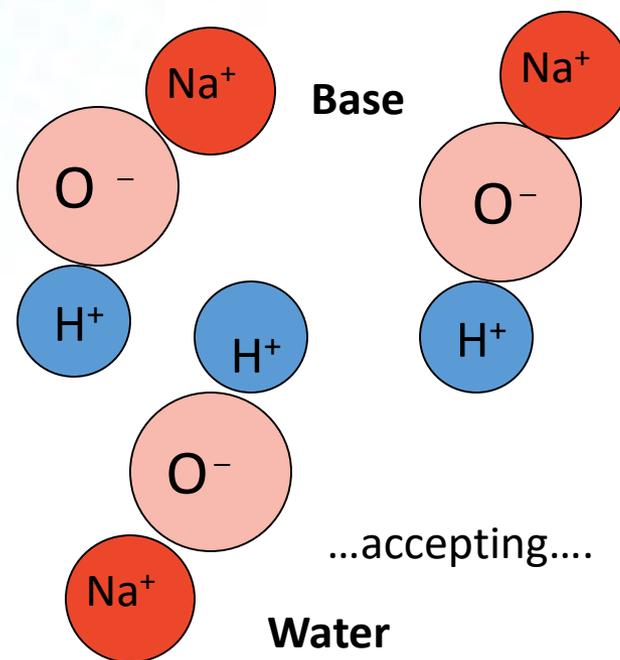
**Bases** are substances that accept hydrogen ions.



**Acid**

.....donating.....

**Salt**



**Base**

...accepting....

**Water**

**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<sup>+</sup>**

A hydrogen ion, H<sup>+</sup> is simply a lone proton (an H with the electron removed)

In water (or aqueous solutions ) H<sup>+</sup> ions exist as an H<sub>3</sub>O<sup>+</sup> ion, called **hydronium**.

Acids are substances that donate protons (H<sup>+</sup>) in solution



HCl gas dissolved in water

HCl has donated a **H<sup>+</sup>** so is acting as an acid

H<sub>2</sub>O has accepted a **H<sup>+</sup>** so it is acting as a base

Solution becomes acidic since H<sub>3</sub>O<sup>+</sup> ions form

Bases are substances that accept protons (H<sup>+</sup>) in solution



NH<sub>3</sub> gas dissolved in water

NH<sub>3</sub> has accepted a **H<sup>+</sup>** so it is acting as a base

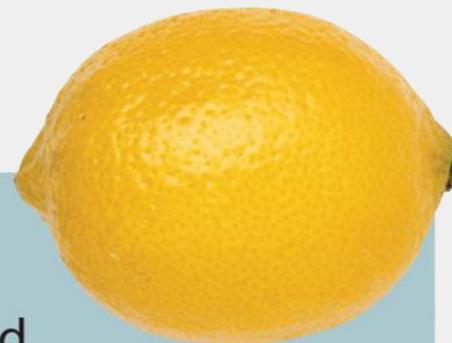
H<sub>2</sub>O has donated a **H<sup>+</sup>** so is acting as an acid

Solution becomes basic since OH<sup>-</sup> ions form.

## 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



Acid  
Sour taste  
Turns blue litmus red  
reacts with some metals to produce  $H_2$   
Dissolves carbonate salts, releasing  $CO_2$

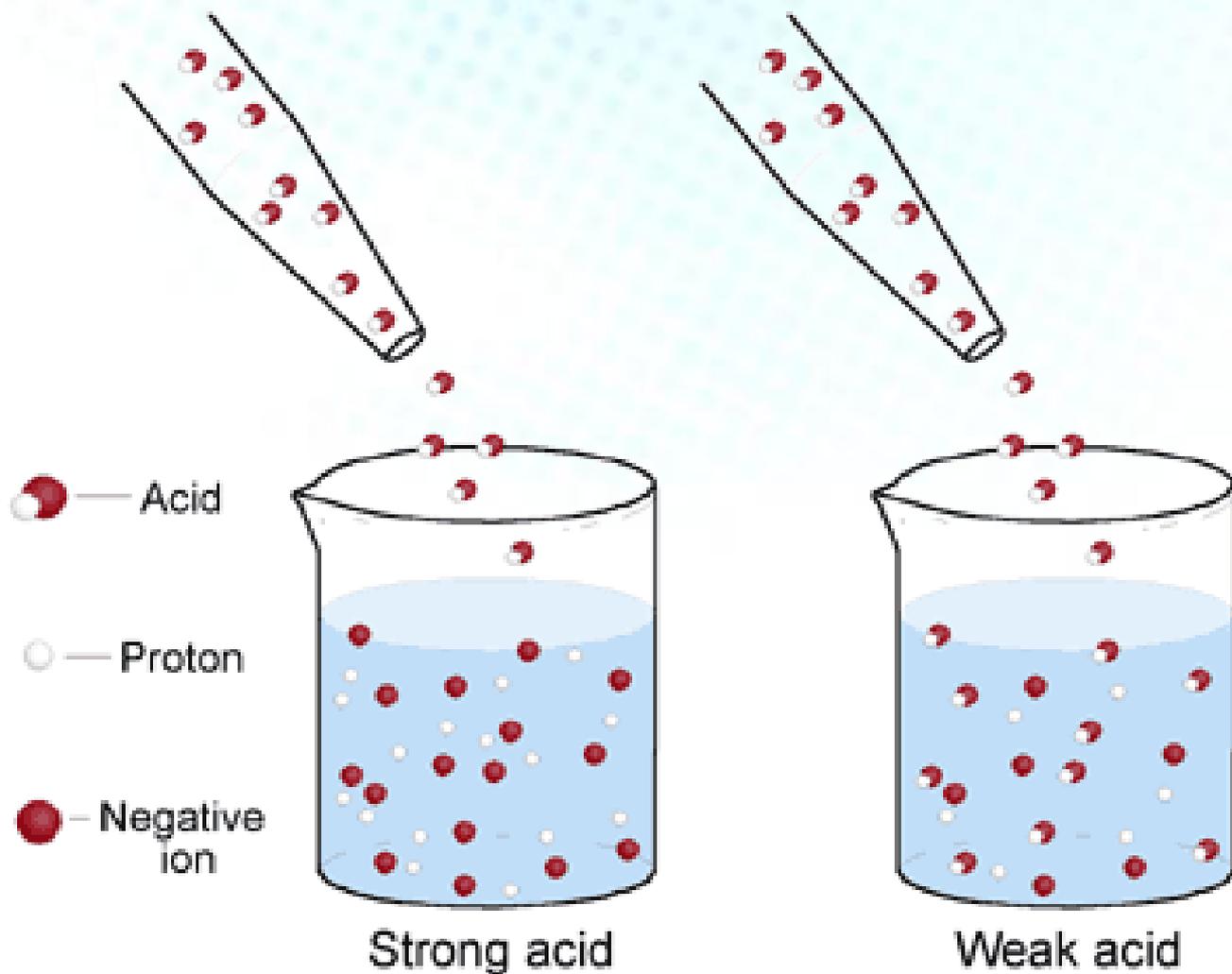
## Chemical and Physical Properties of Bases

- 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  $\text{OH}^-$  ions.

Base  
Bitter taste  
Turns red litmus blue  
Slippery to the touch

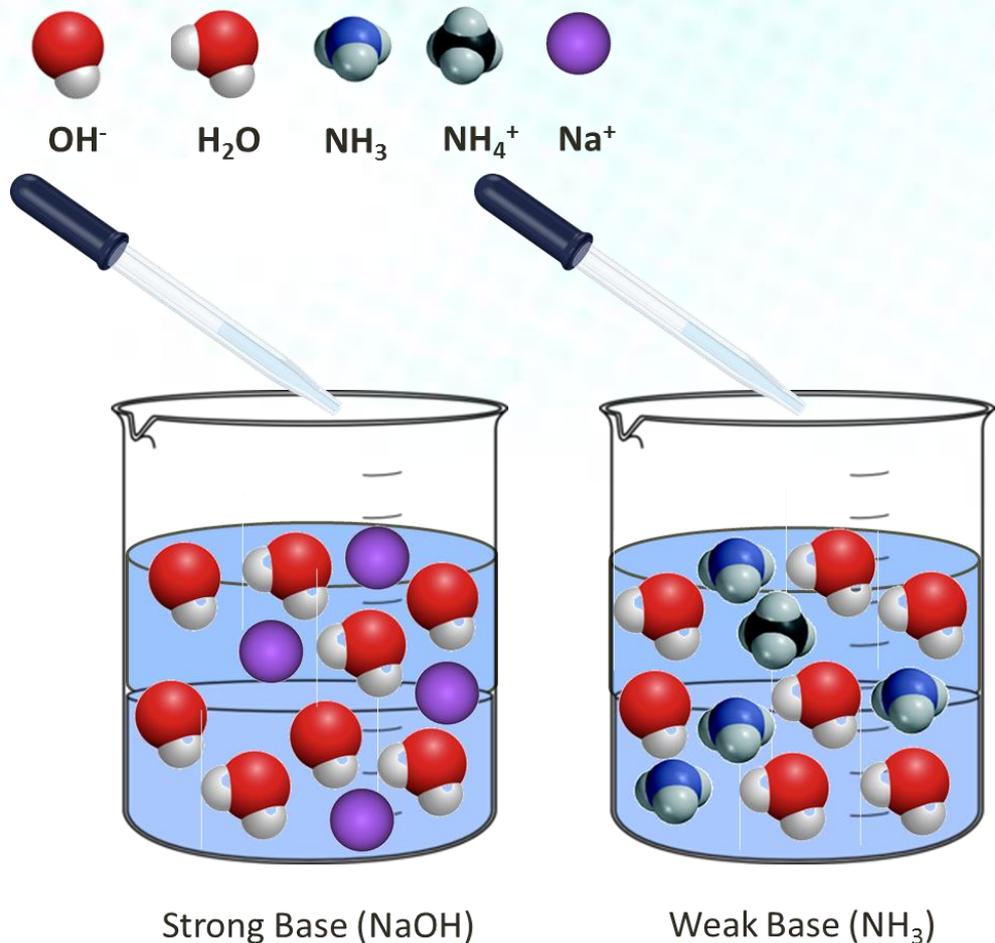


## Understand the difference between Strong and Weak acids and bases



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.

## Strong and Weak bases



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

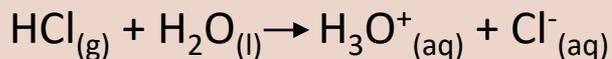
**Note:** For strong alkalis, all of the  $\text{OH}^-$  ions break away from the molecule in water.

## Strong and Weak Acids

The strength of an acid is determined by how readily it will donate its  $\text{H}^+$  ions. Strong acids will have a low pH (0-3) and include  $\text{HCl}$ ,  $\text{H}_2\text{SO}_4$  and  $\text{HNO}_3$ . Weak acids will have a higher pH (4-6). They are mostly organic acids and include  $\text{CH}_3\text{COOH}$ .

### Strong acids

Donate protons ( $\text{H}^+$ ) in aqueous solution to become completely dissociated.



HCl gas dissolved in water

HCl has donated a  $\text{H}^+$  so is acting as an acid

$\text{H}_2\text{O}$  has accepted a  $\text{H}^+$  so it is acting as a base

Solution contains virtually no intact HCl molecules after reaction.

### Weak acids

Donate protons ( $\text{H}^+$ ) in aqueous solution to become partially dissociated.



$\text{CH}_3\text{COOH}$  dissolved in water

Only some of the acetic acid molecules dissociate into acetate ions ( $\text{CH}_3\text{COO}^-$ )

Because the acetate ion is a strong base (conjugate pairs) it will readily accept  $\text{H}^+$  (from  $\text{H}_3\text{O}^+$ ) and become acetic acid.

Solution contains mostly intact  $\text{CH}_3\text{COOH}$  molecules.

## 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 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.



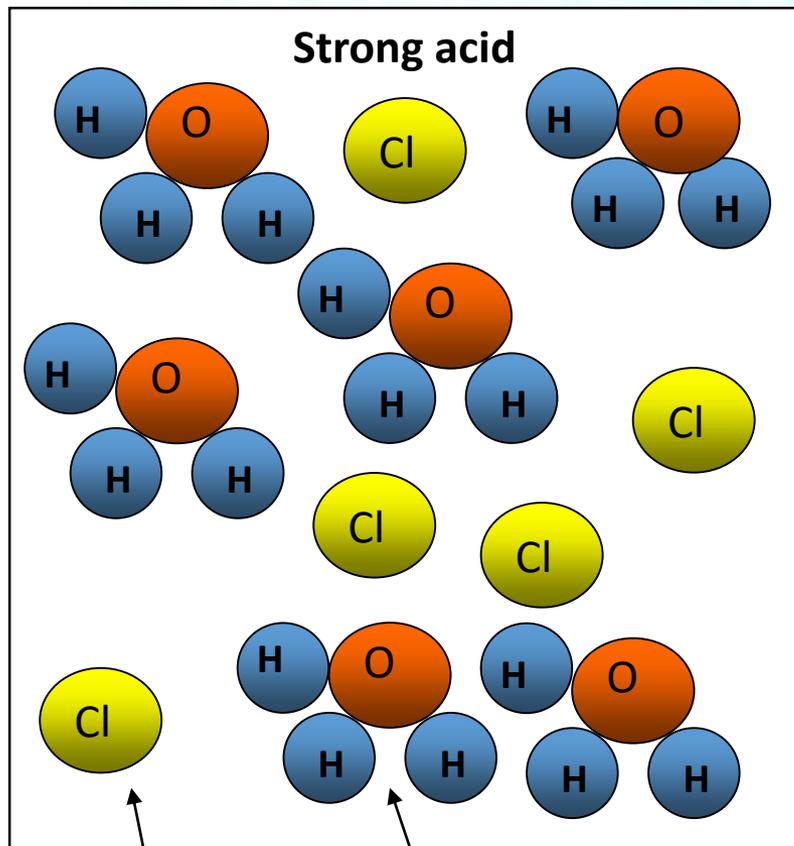
Only some of the ammonia molecules dissociate into ammonium ions ( $NH_4^+$ )

Because ammonium is a reasonably strong acid (conjugate pairs) it will readily donate  $H^+$  and become ammonia.

Solution contains mostly intact  $NH_3$  molecules.

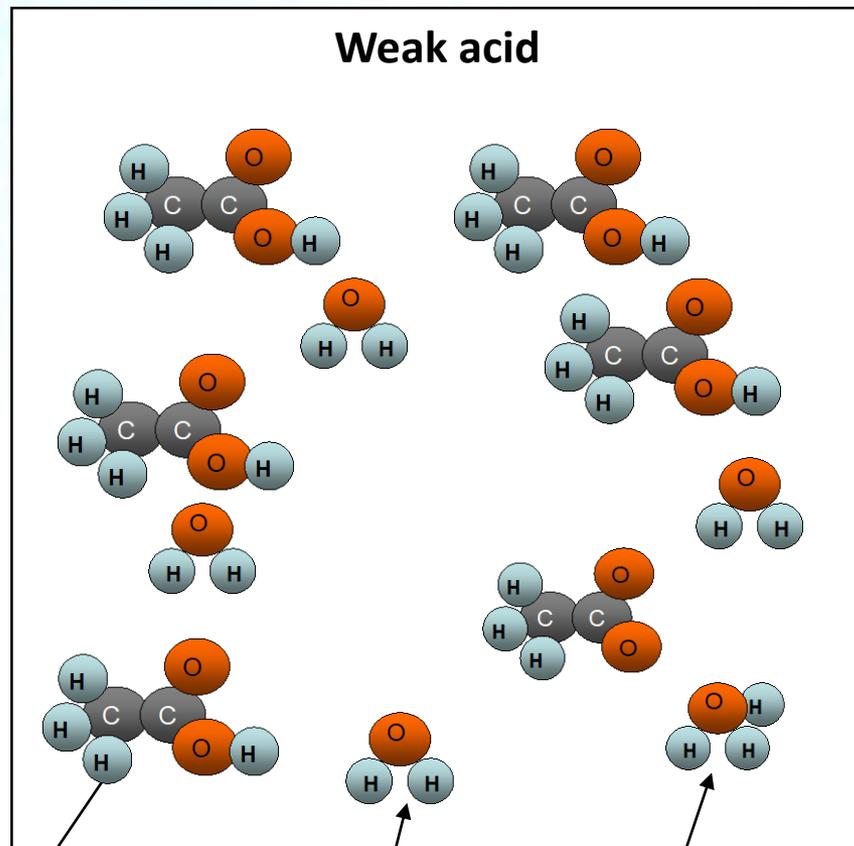
## Strong and Weak Acids

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



dissociated acid molecules

hydronium molecules



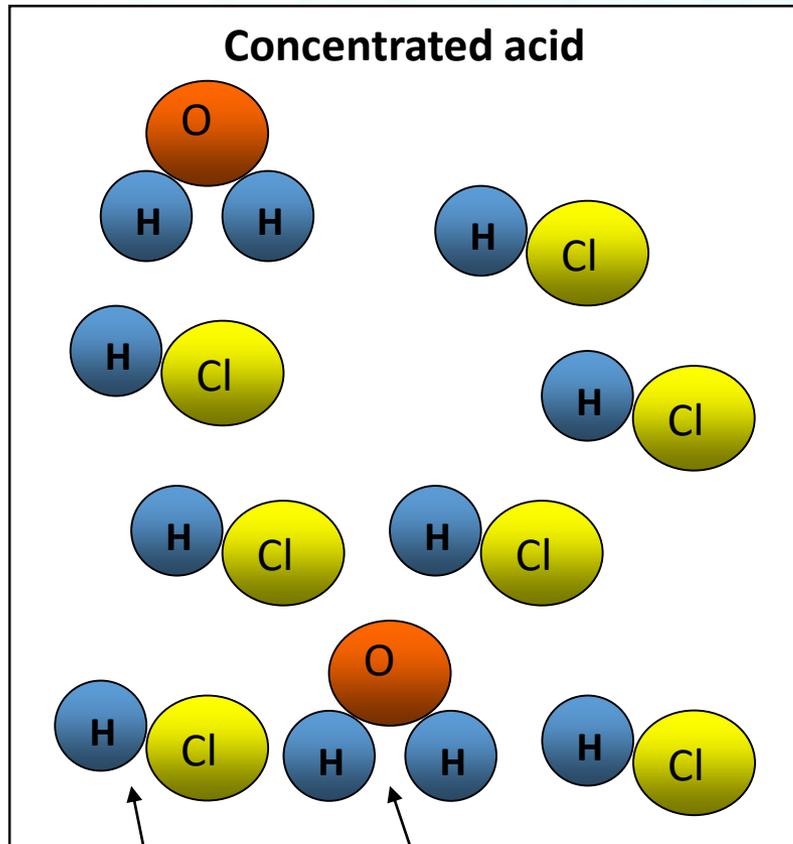
Many intact acid molecules

lots of water molecules

Few hydronium molecules

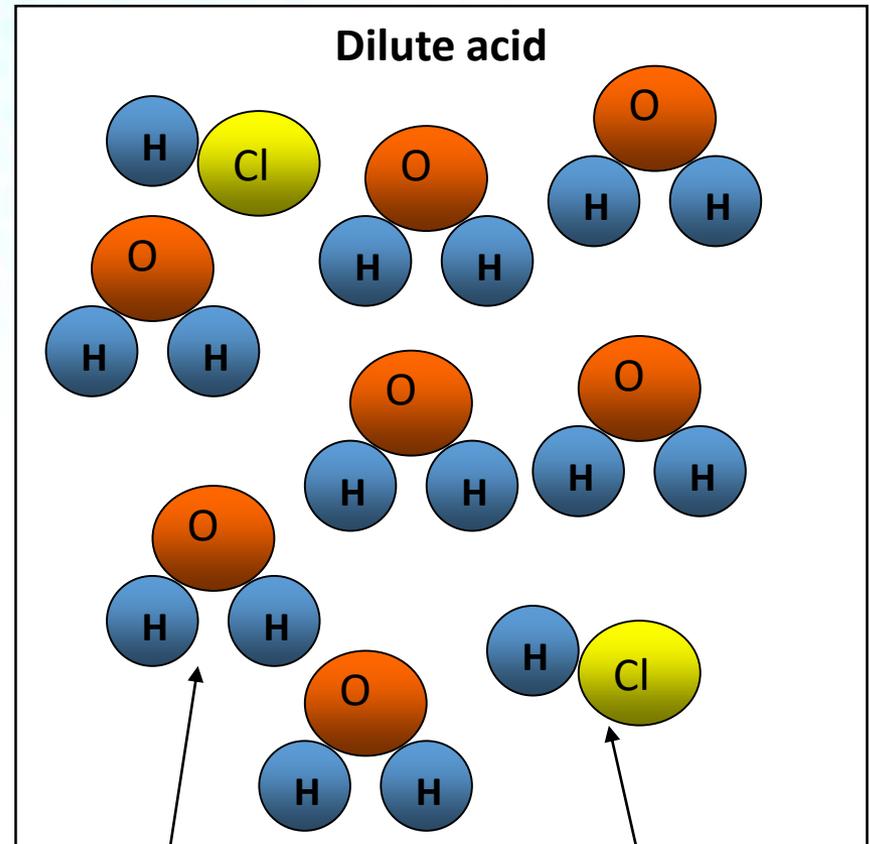
## Concentrated and Dilute Acids

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



lots of acid molecules

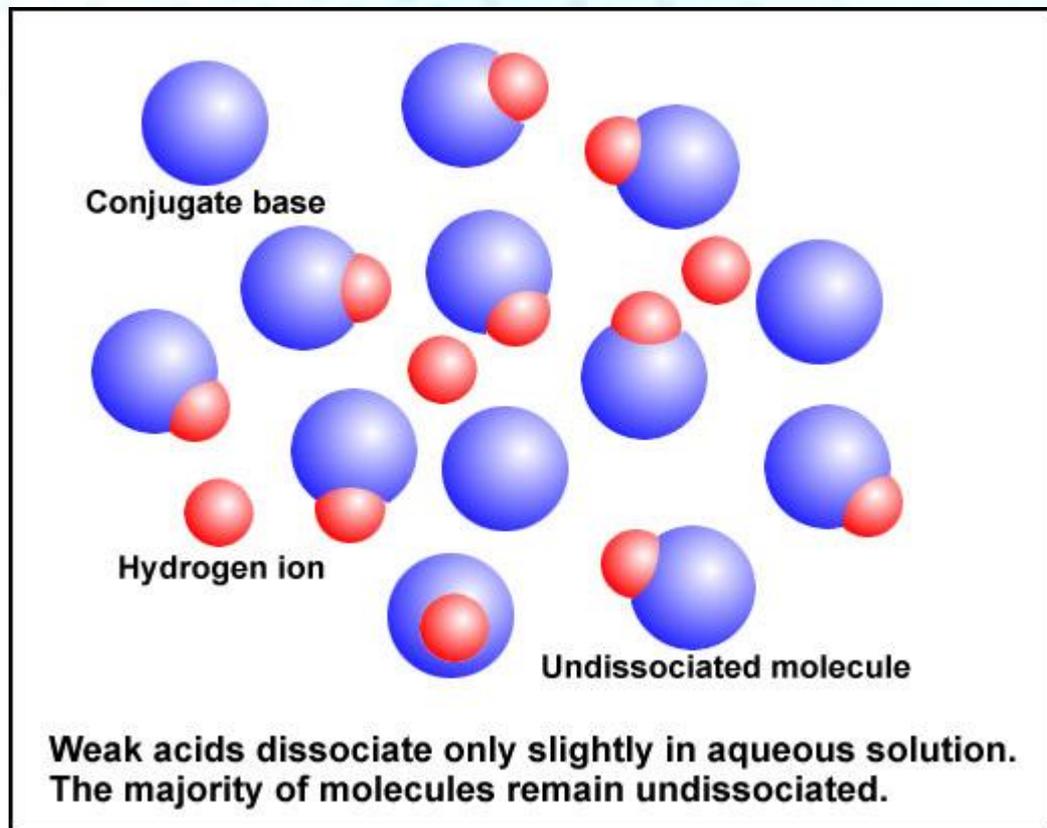
Few water molecules



lots of water molecules

few acid molecules

## Transfer of hydrogen ions in conjugate pairs



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  $\text{H}_2\text{SO}_4/\text{HSO}_4^-$ , and  $\text{NH}_4^+/\text{NH}_3$ . The acid is always the species with the additional proton. It can also be said that  $\text{NH}_3$  is the conjugate base of  $\text{NH}_4^+$ .

Acid		Conjugate Base
<b>HCl</b>	hydrochloric acid	$\text{Cl}^-$
<b><math>\text{H}_2\text{SO}_4</math></b>	sulfuric acid	$\text{HSO}_4^-$
<b><math>\text{HNO}_3</math></b>	nitric acid	$\text{NO}_3^-$
<b><math>\text{CH}_3\text{COOH}</math></b>	acetic acid	$\text{CH}_3\text{COO}^-$
<b><math>\text{NH}_4^+</math></b>	ammonium ion	$\text{NH}_3$
<b><math>\text{H}_3\text{PO}_4</math></b>	phosphoric acid	$\text{H}_2\text{PO}_4^-$

	ACID	BASE	
negligible	$\text{OH}^-$	$\text{O}^{2-}$	strong
	$\text{HS}^-$	$\text{S}^{2-}$	
← Relative acid strength increasing	$\text{H}_2\text{O}$	$\text{OH}^-$	Relative base strength increasing →
	$\text{HPO}_4^{2-}$	$\text{PO}_4^{3-}$	
	$\text{HCO}_3^-$	$\text{CO}_3^{2-}$	
	$\text{NH}_4^+$	$\text{NH}_3$	
	$\text{HCN}$	$\text{CN}^-$	
	$\text{H}_2\text{PO}_4^-$	$\text{HPO}_4^{2-}$	
	$\text{HSO}_3^-$	$\text{SO}_3^{2-}$	
	$\text{H}_2\text{S}$	$\text{HS}^-$	
	$\text{H}_2\text{CO}_3$	$\text{HCO}_3^-$	
	$\text{C}_5\text{H}_5\text{NH}^+$	$\text{C}_5\text{H}_5\text{N}$	
	$\text{CH}_3\text{CO}_2\text{H}$	$\text{CH}_3\text{CO}_2^-$	
	$\text{HF}$	$\text{F}^-$	
	$\text{H}_3\text{PO}_4$	$\text{H}_2\text{PO}_4^-$	
	$\text{H}_2\text{SO}_3$	$\text{HSO}_3^-$	
	$\text{HSO}_4^-$	$\text{SO}_4^{2-}$	
strong	$\text{H}_3\text{O}^+$	$\text{H}_2\text{O}$	
	$\text{HNO}_3$	$\text{NO}_3^-$	
	$\text{H}_2\text{SO}_4$	$\text{HSO}_4^-$	
	$\text{HCl}$	$\text{Cl}^-$	
	$\text{HBr}$	$\text{Br}^-$	

## Conjugate Acid and Base pairs

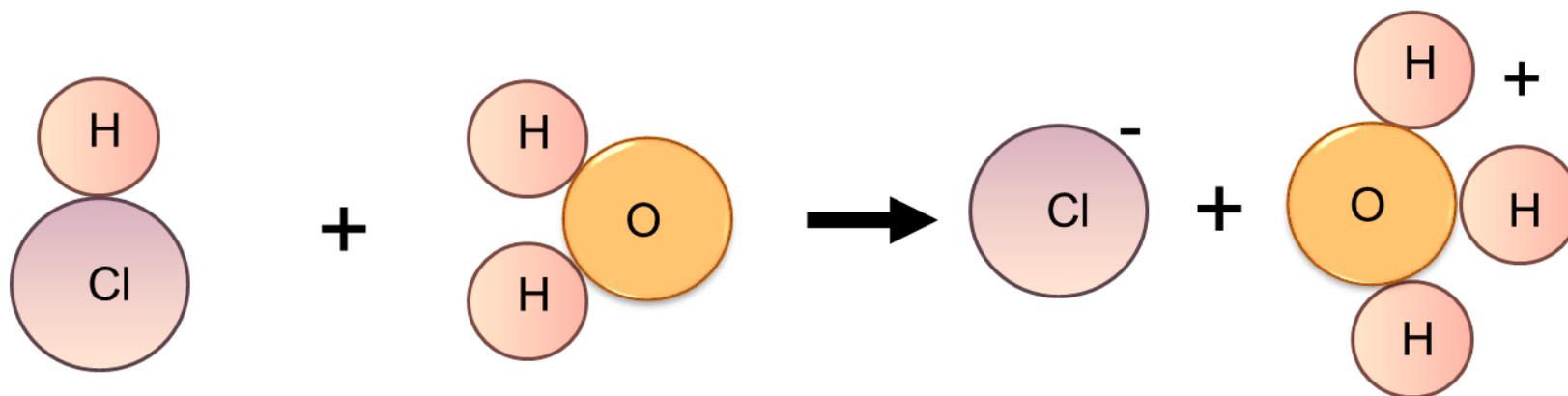
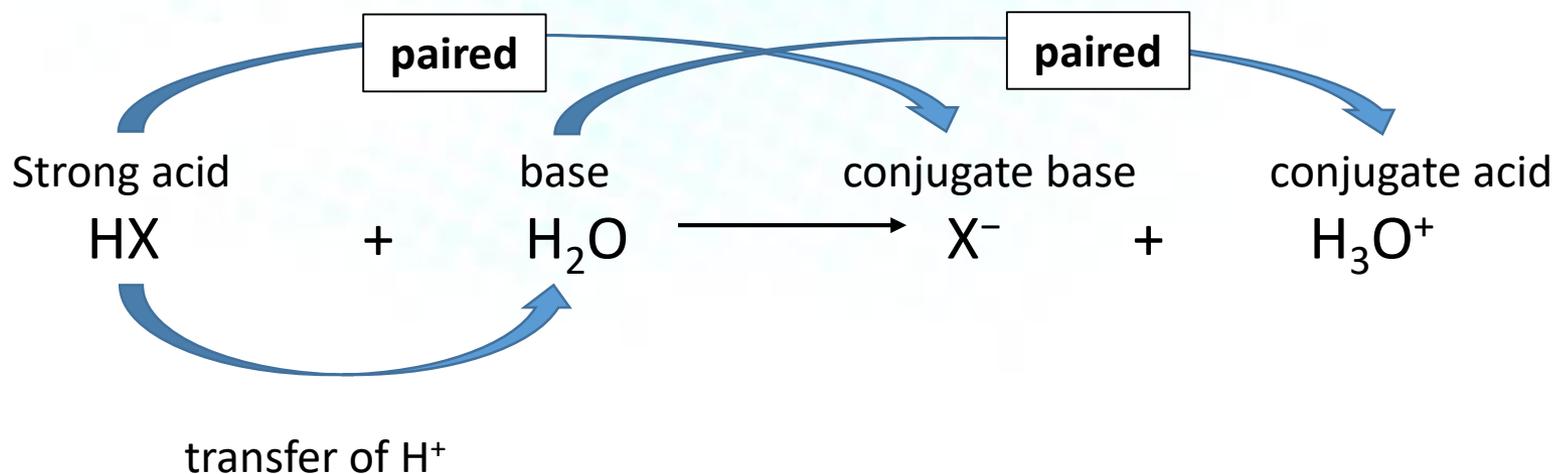
The stronger an acid, the weaker its conjugate base, and, conversely, the stronger a base, the weaker its conjugate acid.

A strong acid like HCl donates its proton so readily that there is essentially no tendency for the conjugate base  $\text{Cl}^-$  to reaccept a proton. Consequently,  $\text{Cl}^-$  is a very weak base. A strong base like the  $\text{H}^-$  ion accepts a proton and holds it so firmly that there is no tendency for the conjugate acid  $\text{H}_2$  to donate a proton. Hence,  $\text{H}_2$  is a very weak acid.

## Conjugate Acid and Base pairs (Strong Acid)

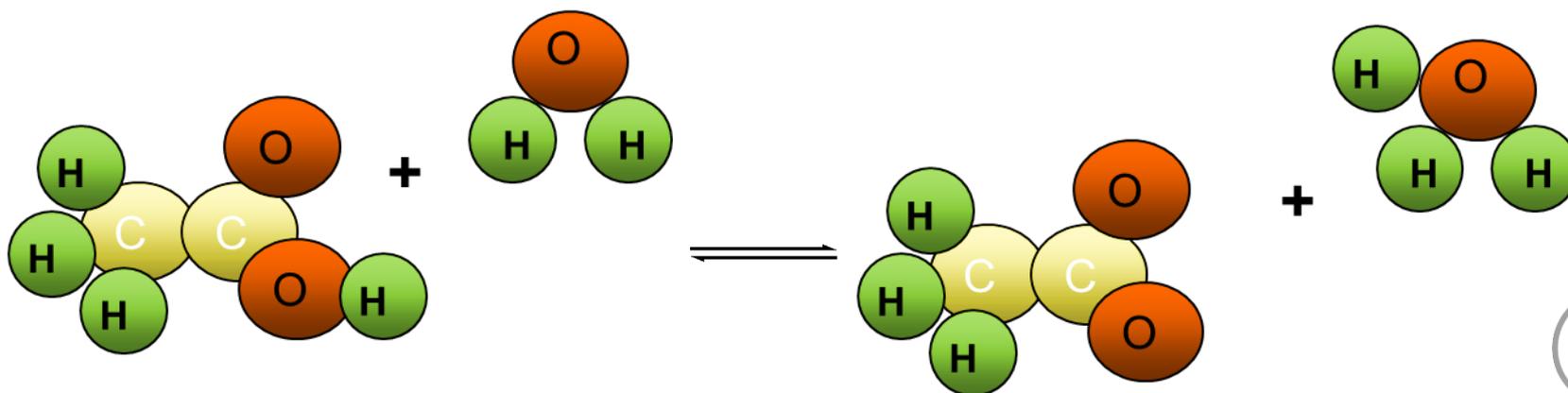
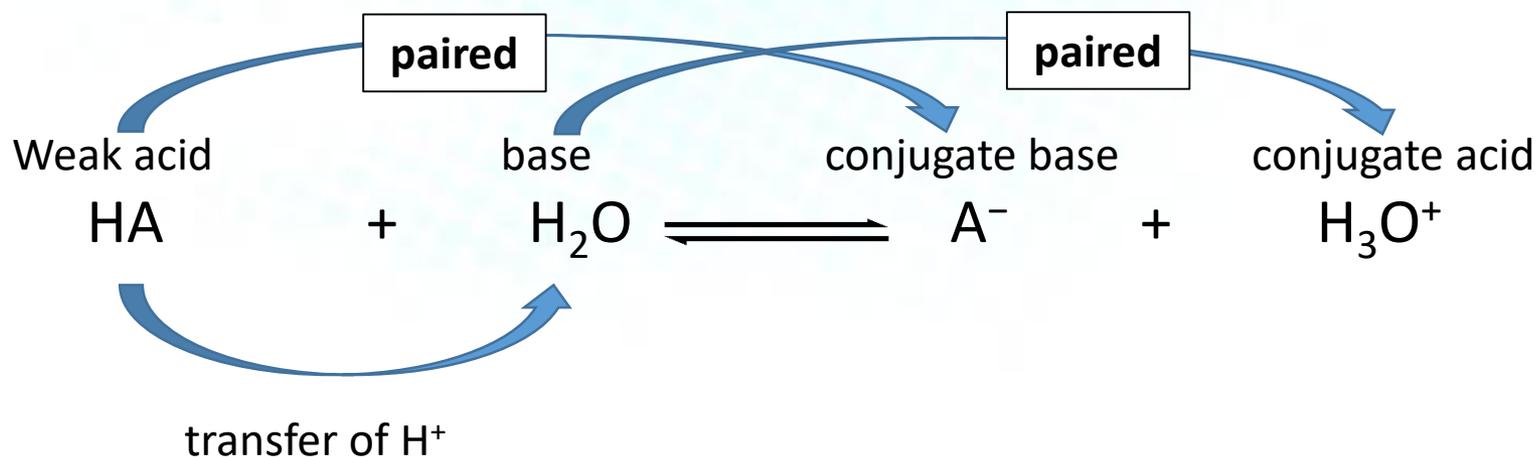
HX is a symbol used for a strong acid. A conjugate acid can be seen as the chemical substance that releases a proton in the **backward** chemical reaction.

The base produced, X<sup>-</sup>, is called the **conjugate base** and it absorbs a proton in the **backward** chemical reaction.



## 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.



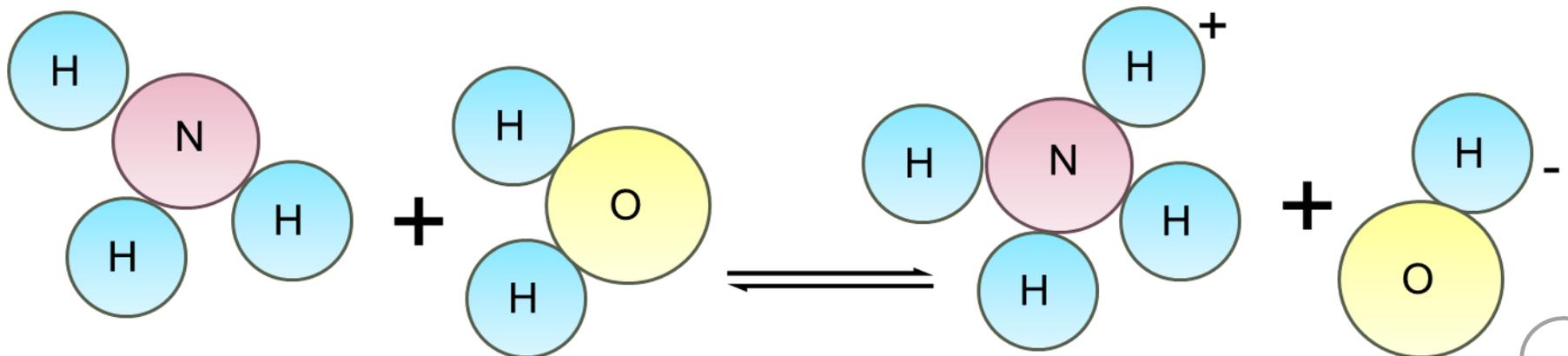
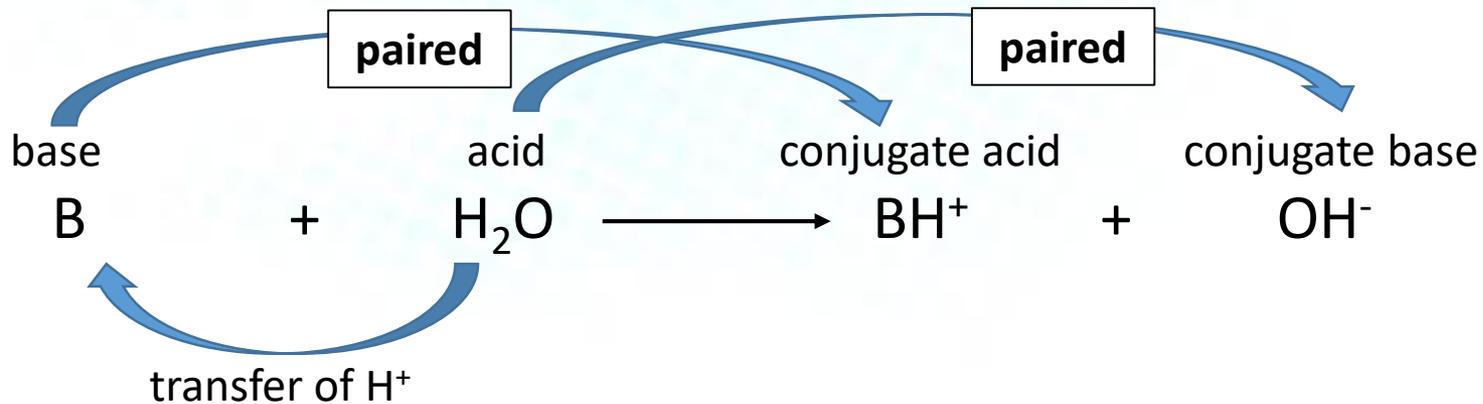
## BASE - Conjugate Acid and Base pairs

If 2 species differ by just 1 proton they are classed as a conjugate acid-base pair. Examples of acid-base pairs are  $\text{H}_2\text{SO}_4/\text{HSO}_4^-$ , and  $\text{NH}_4^+/\text{NH}_3$ . The acid is always the species with the additional proton. It can also be said that  $\text{NH}_3$  is the conjugate base of  $\text{NH}_4^+$ .

Base		Conjugate Acid
<b>H<sub>2</sub>O</b>	water	<b>H<sub>3</sub>O<sup>+</sup></b>
<b>SO<sub>4</sub><sup>2-</sup></b>	sulfate ion	<b>HSO<sub>4</sub><sup>-</sup></b>
<b>NH<sub>3</sub></b>	ammonia	<b>NH<sub>4</sub><sup>+</sup></b>
<b>OH<sup>-</sup></b>	hydroxide ion	<b>H<sub>2</sub>O</b>
<b>HCO<sub>3</sub><sup>-</sup></b>	hydrogen carbonate ion	<b>H<sub>2</sub>CO<sub>3</sub></b>
<b>CO<sub>3</sub><sup>2-</sup></b>	carbonate ion	<b>HCO<sub>3</sub><sup>-</sup></b>

## 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,  $\text{OH}^-$ , is the paired conjugate of the water once the  $\text{H}^+$  has been removed. Strong bases use a single direction arrow and weak bases use a double arrow.

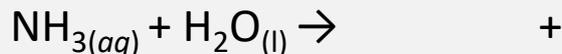


## NCEA 2014 Acids and Bases

Merit  
Question

**Question 1a:** Ammonia,  $\text{NH}_3$ , 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.



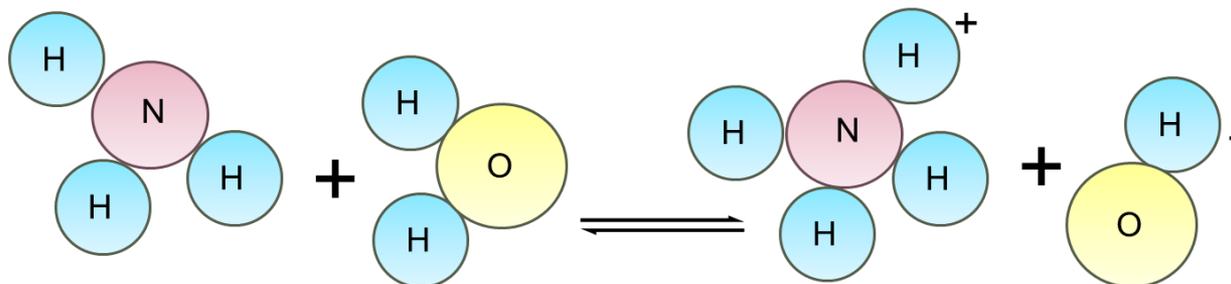
(ii) Explain what is occurring during this reaction.

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:**



Protons are being transferred from the acid,  $\text{H}_2\text{O}$  to the base,  $\text{NH}_3$ . When the  $\text{NH}_3$  gains a proton, it forms its conjugate acid,  $\text{NH}_4^+$ . When  $\text{H}_2\text{O}$  loses its proton, it forms its conjugate base,  $\text{OH}^-$ .



**Question 2(a):** Ammonia solution,  $\text{NH}_{3(aq)}$ , is a common chemical in the school laboratory.

(i) Explain, using an equation, whether ammonia solution is acidic or basic.

**Answer:** Ammonia is basic,  $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$ .

When ammonia (partially) ionises in water, it produces hydroxide ions. (When hydroxide ions are in greater concentration than  $\text{H}_3\text{O}^+$ , the pH will be above 7 and therefore basic.)

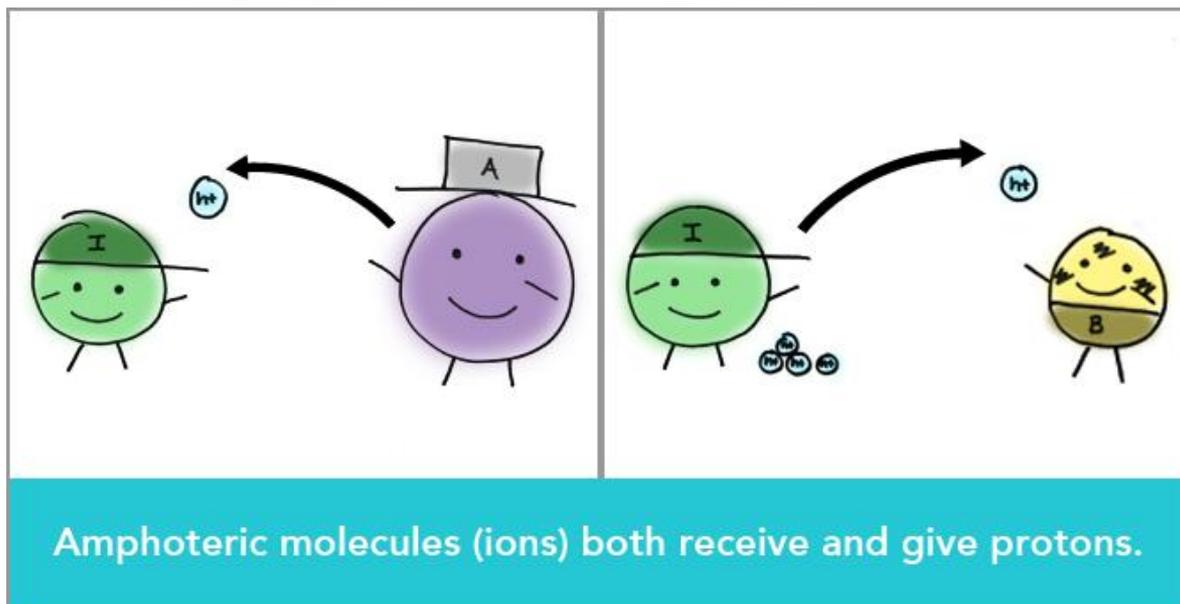
**Question 2(a):** (ii) Bottles of ammonia solution are often labelled ammonium hydroxide,  $\text{NH}_4\text{OH}_{(aq)}$ .

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.

**Question 2(b):** The hydrogen carbonate ion,  $\text{HCO}_3^-$ , 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  $\text{HCO}_3^-$  with water: one where it acts as an acid, and one where it acts as a base.

**Answer:**



## 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,  $\text{H}_2\text{O}$ , with ammonia,  $\text{NH}_3$ , and the ammonium ion,  $\text{NH}_4^+$ , in the box below.

$\text{H}_2\text{O}$ acting as	Equation
an acid	$\text{H}_2\text{O}(\ell) + \text{NH}_3(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\lambda) + \text{NH}_3(\text{aq}) \rightleftharpoons \text{OH}^-(\text{aq}) + \text{NH}_4^+(\text{aq})$
a base	$\text{H}_2\text{O}(\ell) + \text{NH}_4^+(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\lambda) + \text{NH}_4^+(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{NH}_3(\text{aq})$



## NCEA 2016 Acids and Bases

Merit  
Question

**Question 2(b):** Sodium carbonate,  $\text{Na}_2\text{CO}_3(\text{s})$ , 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:**

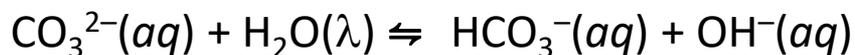
When sodium carbonate dissolves in water, it dissociates into ions.

EQUATION ONE:



The carbonate ions then react in water.

EQUATION TWO:



The  $\text{OH}^-$  ions produced make this a **basic (alkaline) solution**.

# 18 FANTASTIC USES FOR



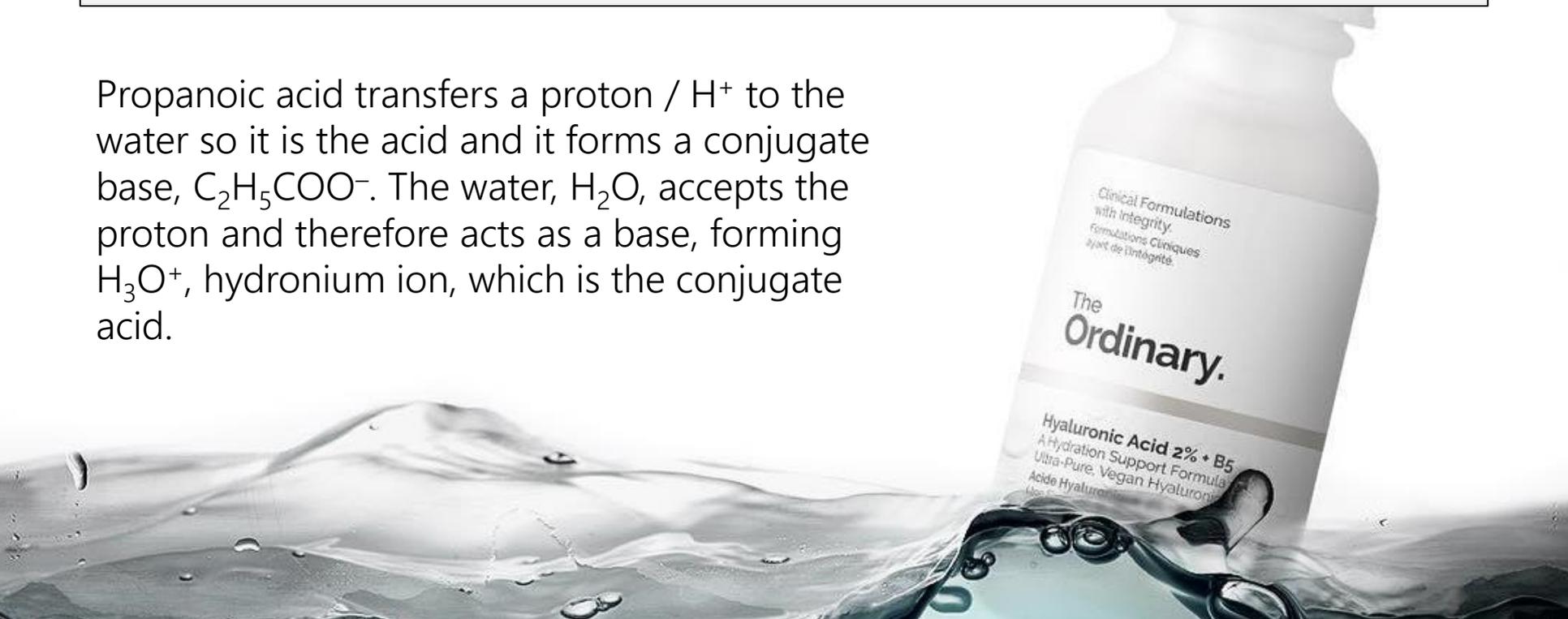
**Question 1a:** Propanoic acid,  $C_2H_5COOH$ , is dissolved in water and the resulting solution has a pH of 4.2.

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



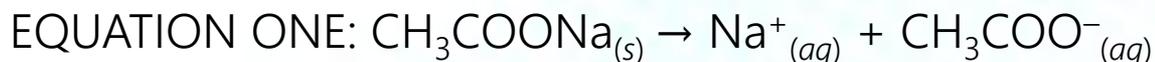
(ii) Explain the proton,  $H^+$ , transfer in this reaction, and identify the two conjugate acid-base pairs.

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.



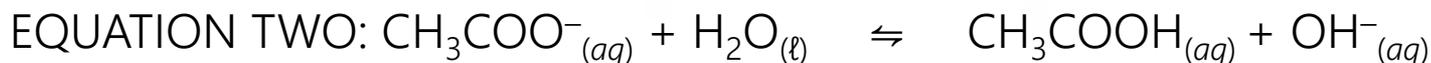
**Question 1b:** Sodium ethanoate,  $\text{CH}_3\text{COONa}_{(s)}$ , is a salt. When dissolved in water, it dissociates into ions.

Explain, including TWO relevant equations, whether a solution of sodium ethanoate is acidic or basic.



The salt dissociates / dissolves into ions in water.

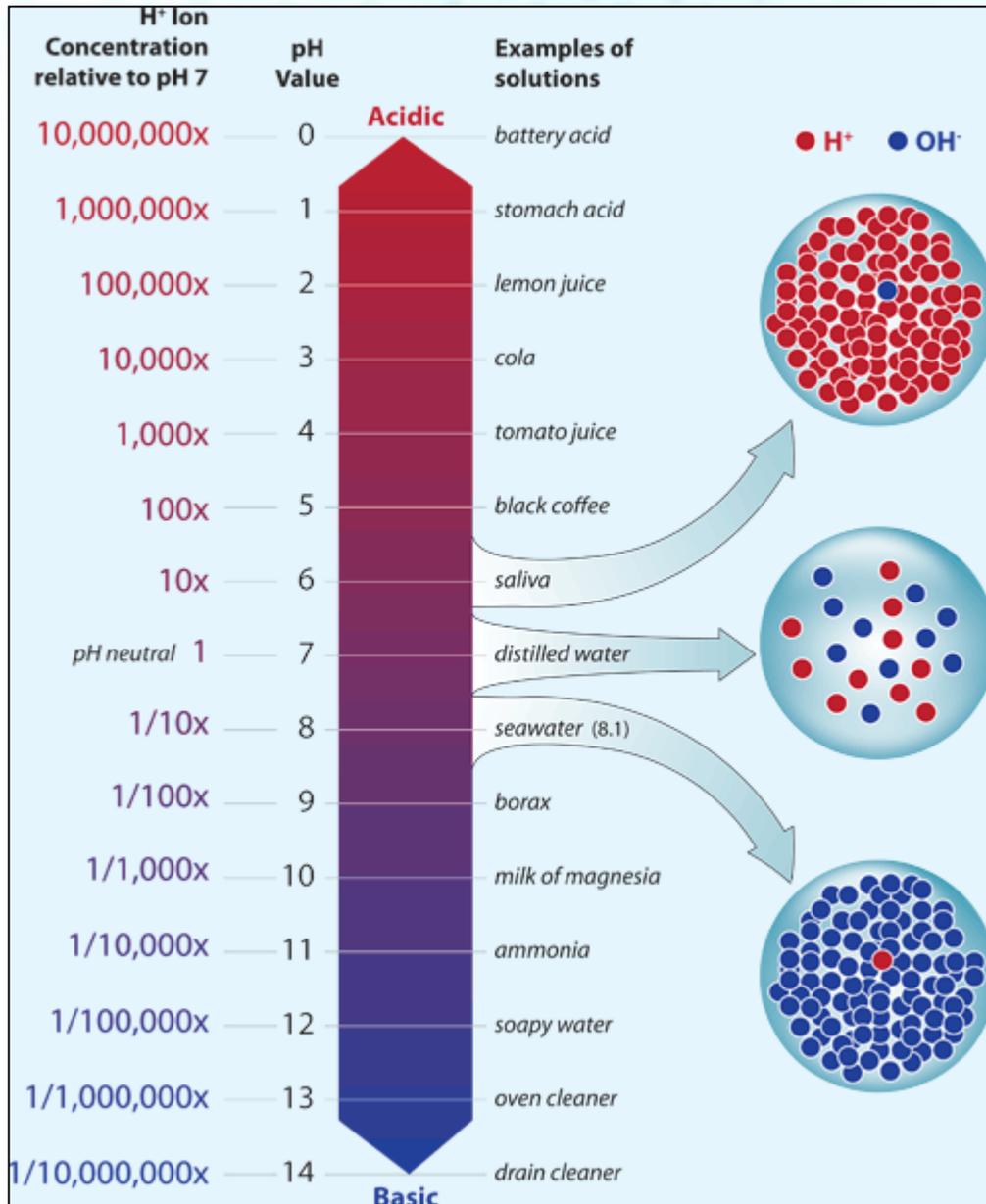
The ethanoate ions then react with water to form  $\text{CH}_3\text{COOH}_{(aq)}$  and  $\text{OH}^-_{(aq)}$ .



The  $\text{OH}^-$  ions produced make this a basic solution.



# 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

<b>Strong Acids</b> Readily donate all their protons when dissolved	<b>Weak Acids</b> donate only a small proportion of protons	<b>Neutral solution</b>	<b>Weak Bases</b> Accept only a small proportion of protons	<b>Strong Bases</b> Readily accept protons
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Concentration of $\text{H}_3\text{O}^+$ ions is <b>greater</b> than that of $\text{OH}^-$ ions	Concentration of $\text{H}_3\text{O}^+$ ions is the <b>same</b> as that of $\text{OH}^-$ ions	Concentration of $\text{H}_3\text{O}^+$ ions is <b>less</b> than that of $\text{OH}^-$ ions
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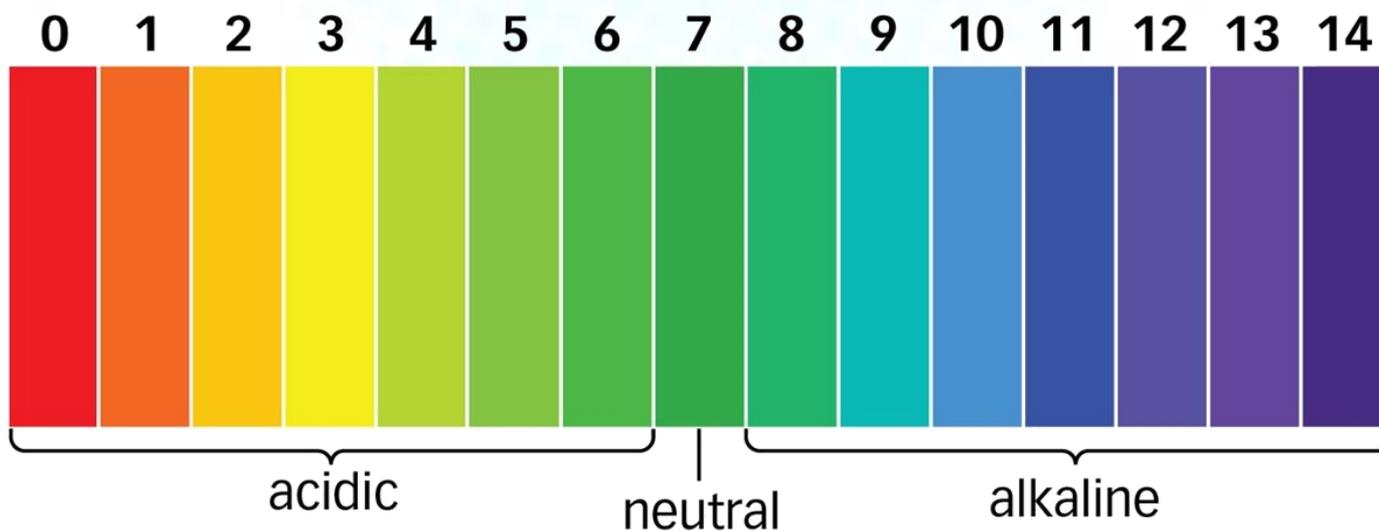
Amphiprotic Substances can act as acid or base e.g.  $\text{H}_2\text{O}$

## 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

Strong Mineral Acids

Weak Organic Acids

Water, Chlorides

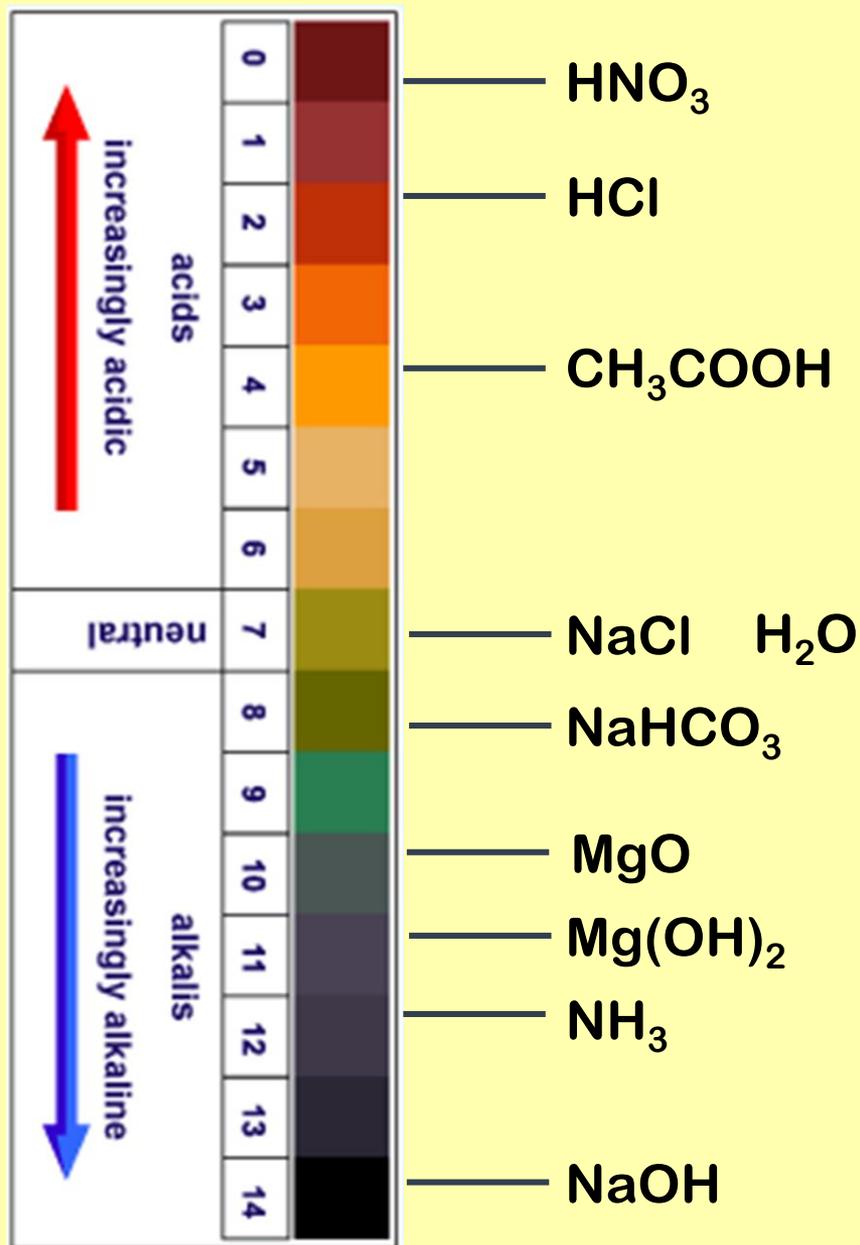
Carbonates

Oxides

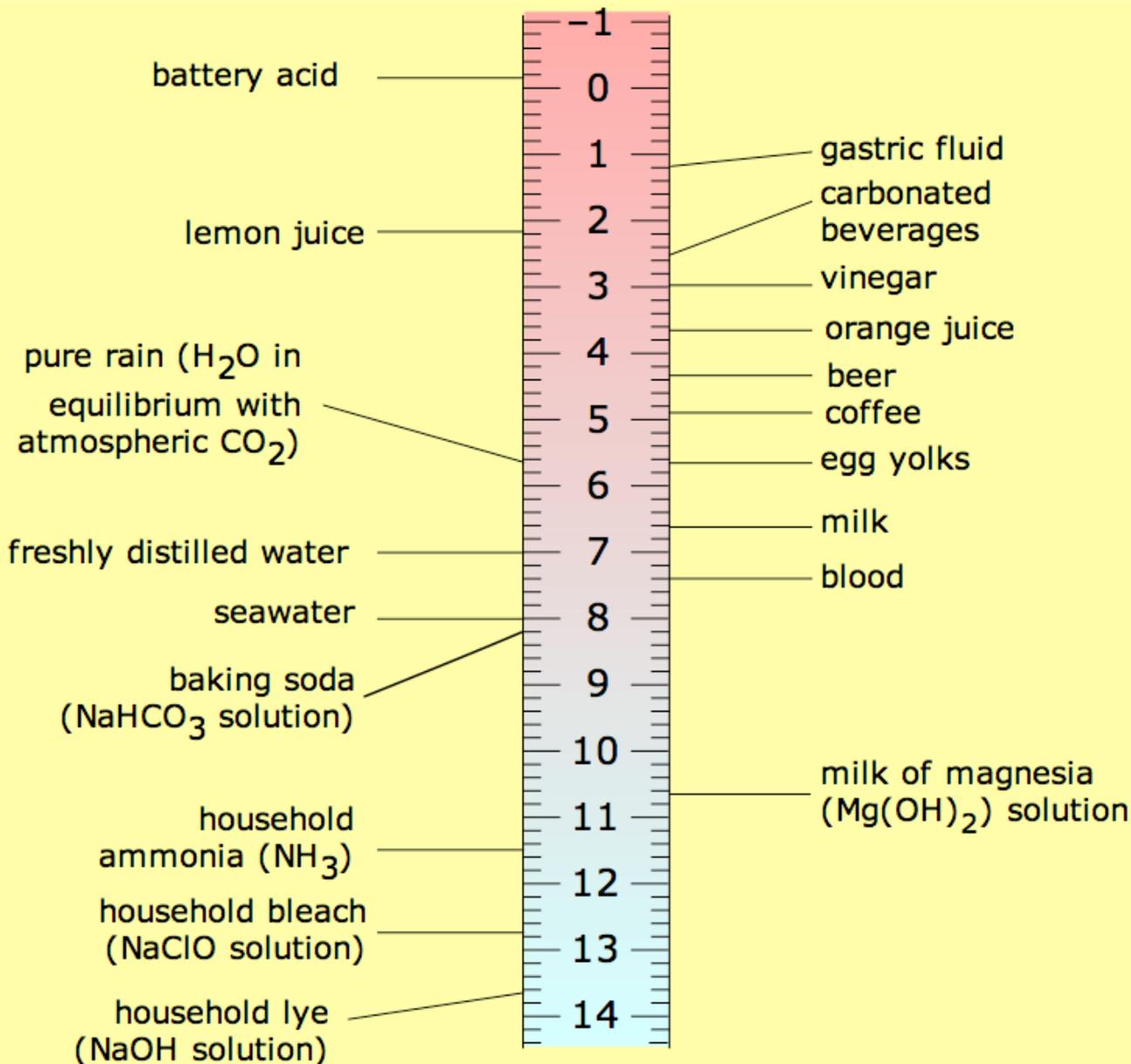
Most Hydroxides

Ammonia

Sodium Hydroxide



# 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

## Putting it all together

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



## Acid and Base Ratio of $[\text{OH}^-]$ or $[\text{H}_3\text{O}^+]$

When a strong acid dissociates into its conjugate and  $\text{H}_3\text{O}^+$  ions then 100% of its molecules will dissociate.

The concentration of  $\text{H}_3\text{O}^+$  ions produced depends upon the Strong Acid involved.

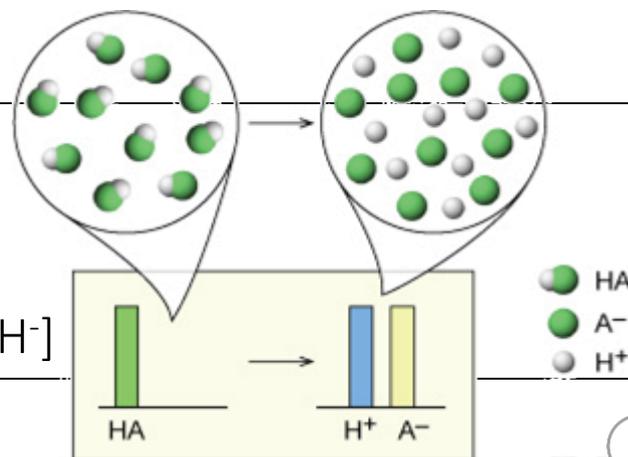
$\text{HCl}$  and  $\text{HNO}_3$  acid concentration produces the same concentration of  $\text{H}_3\text{O}^+$  ions. i.e.  $0.123\text{molL}^{-1}$  of  $\text{HCl}$  will produce  $0.123\text{molL}^{-1}$   $[\text{H}_3\text{O}^+]$  1:1 ratio as those acids have one  $\text{H}^+$  ion per molecule to donate

$\text{H}_2\text{SO}_4$  acid concentration will produce double the concentration of  $\text{H}_3\text{O}^+$  ions i.e.  $0.124\text{molL}^{-1}$  of  $\text{H}_2\text{SO}_4$  will produce  $0.248\text{molL}^{-1}$   $[\text{H}_3\text{O}^+]$  1:2 Ratio as that acid has two  $\text{H}^+$  ions per molecule to donate.

Strong Alkalis produce  $\text{OH}^-$  ions.

$\text{NaOH}$  Alkali (Base) concentration produces the same concentration of  $\text{OH}^-$  ions

i.e.  $0.267\text{molL}^{-1}$  of  $\text{NaOH}$  will produce  $0.267\text{molL}^{-1}$   $[\text{OH}^-]$



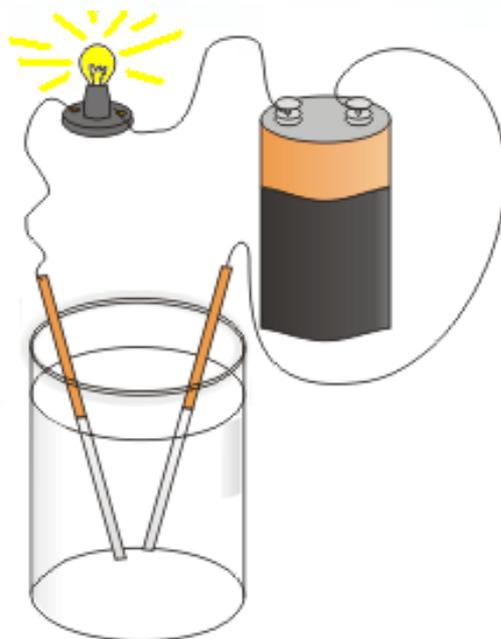
## Conductivity of Acids and Bases

Conductivity is related to the availability of free moving charged particles. The presence of  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  ions in solution and the concentration of them determine conductivity.

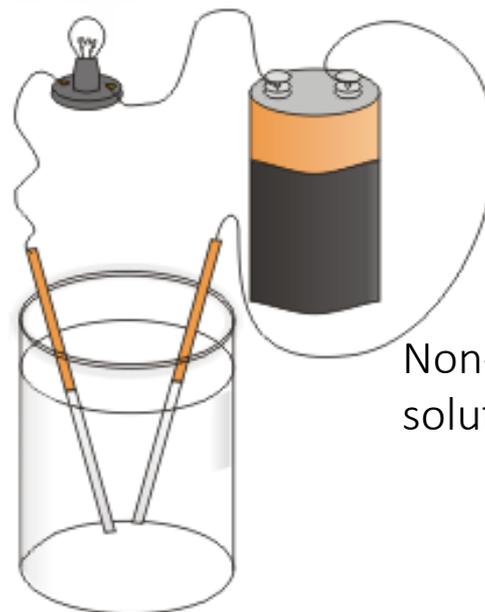
### High conductivity – strong acids and Bases

A strong electrolyte (solution containing ions) is created when a strong acid /strong base is added to water and fully dissociates. The  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  ions carry the charge

Conductive  
solution



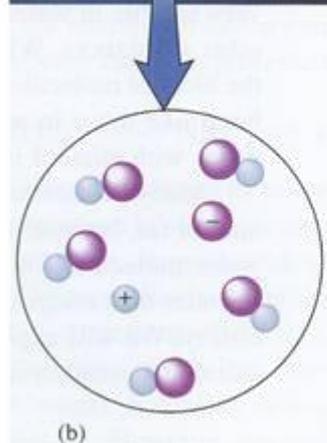
Non-conductive  
solution



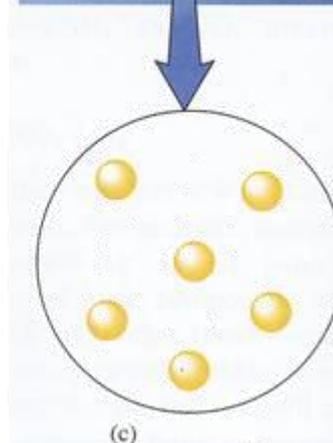
## Conductivity of Acids and Bases

### 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  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  ions are created to carry charge. (such as ethanoic acid)



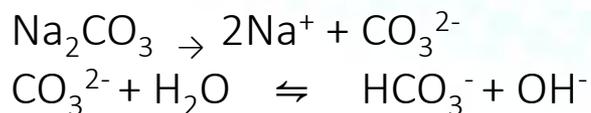
Acetic acid  
(weak acid)



Molecular substance

## Conductivity of salts (weak bases - carbonates or oxides)

If the base used is a **soluble ionic solid**, for example sodium carbonate, then before it reacts with the acid it dissolves in the water into its two ions;  $\text{Na}^+$  and  $\text{CO}_3^{2-}$ . This means that even though a carbonate is a weak base only dissociating to release a small proportion of  $\text{OH}^-$  ions / conjugate acid, it will still **be a good conductor** because of the  $\text{Na}^+$  and  $\text{CO}_3^{2-}$  ions in solution. These are free moving ions that carry charge. When drawing equations remember to show  $\text{Na}_2\text{CO}_3$  dissolving into the two ions first before  $\text{CO}_3^{2-}$  ion further reacts as a base. (2 equations required)



Conductive acids and bases	Poorly conductive acids and bases
HCl (hydrochloric acid) H <sub>2</sub> SO <sub>4</sub> (sulphuric acid) HNO <sub>3</sub> (nitric acid) NaOH (sodium hydroxide) NaCO <sub>3</sub> (sodium carbonate) Na <sub>2</sub> O (sodium oxide)	NH <sub>3</sub> (ammonia) CH <sub>3</sub> COOH (ethanoic acid)

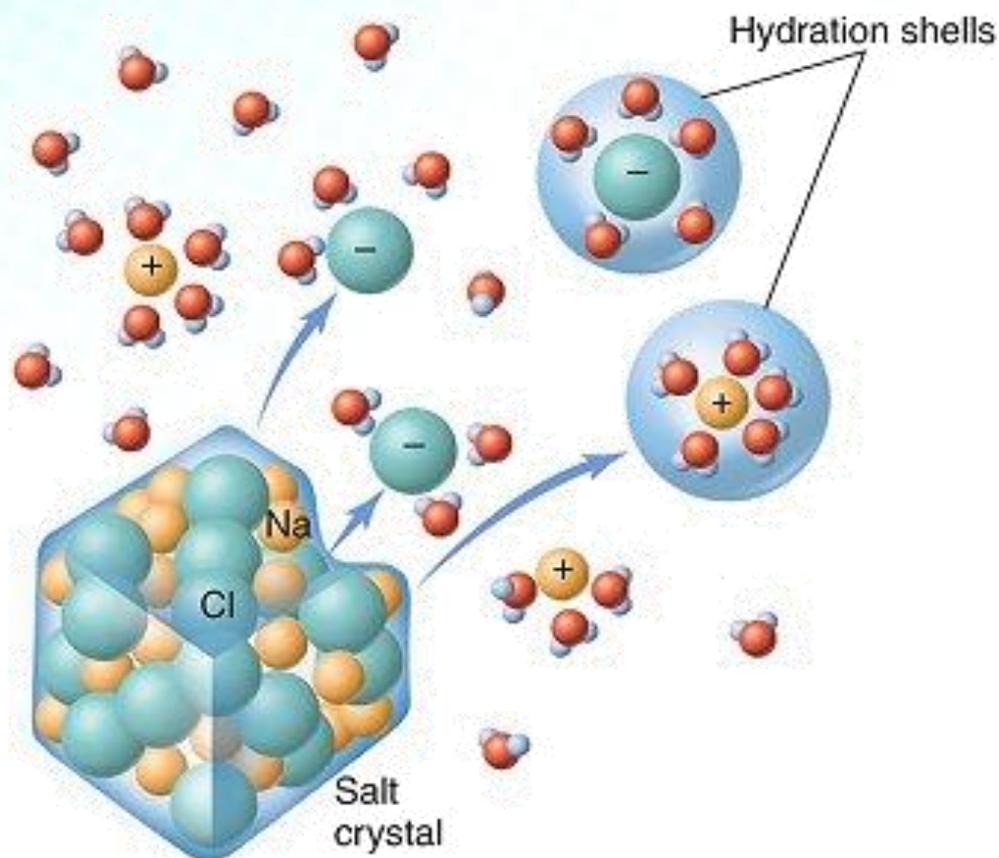
## Species in solution

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

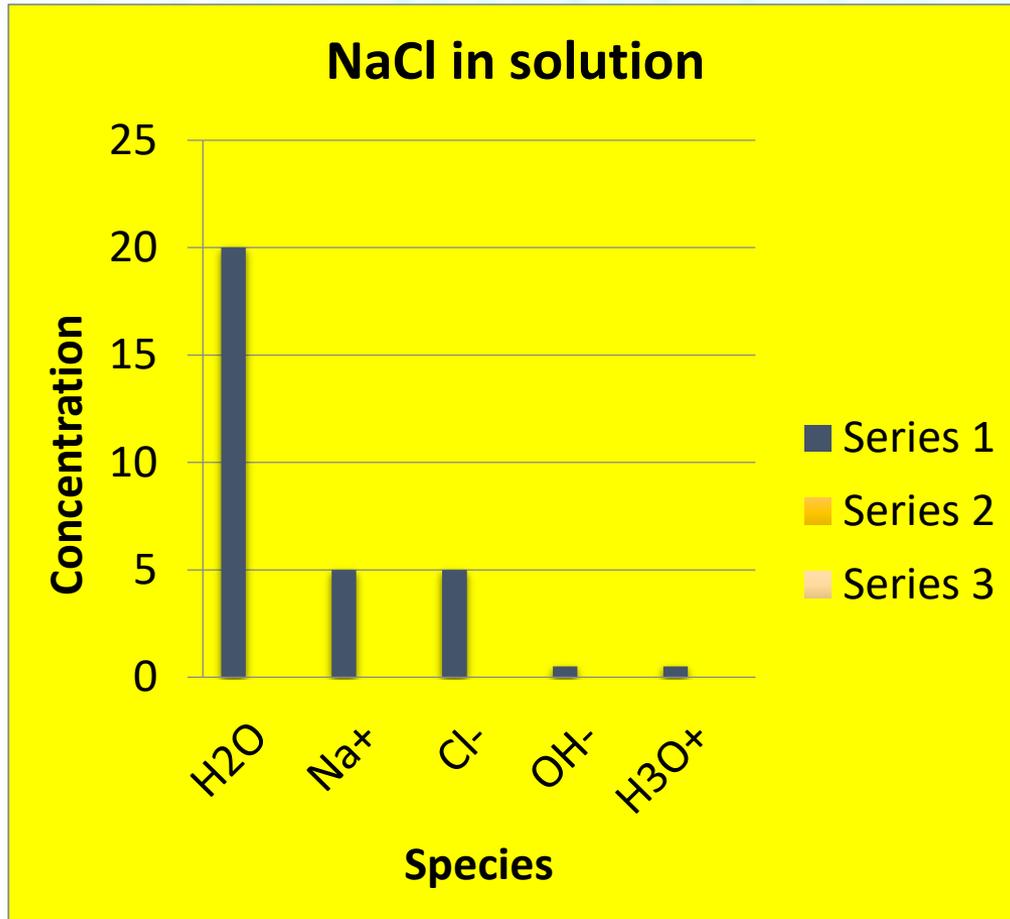
The solvent is water, and the solute can be an acid, base or ionic salt. A solute dissolves by bonds being broken between solute particles (endothermic) and new bonds being formed between solute and solvent (exothermic).

A small amount of  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  will always be present in water due to  $K_w = [\text{OH}^-] [\text{H}_3\text{O}^+] = 1 \times 10^{-14}$

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.

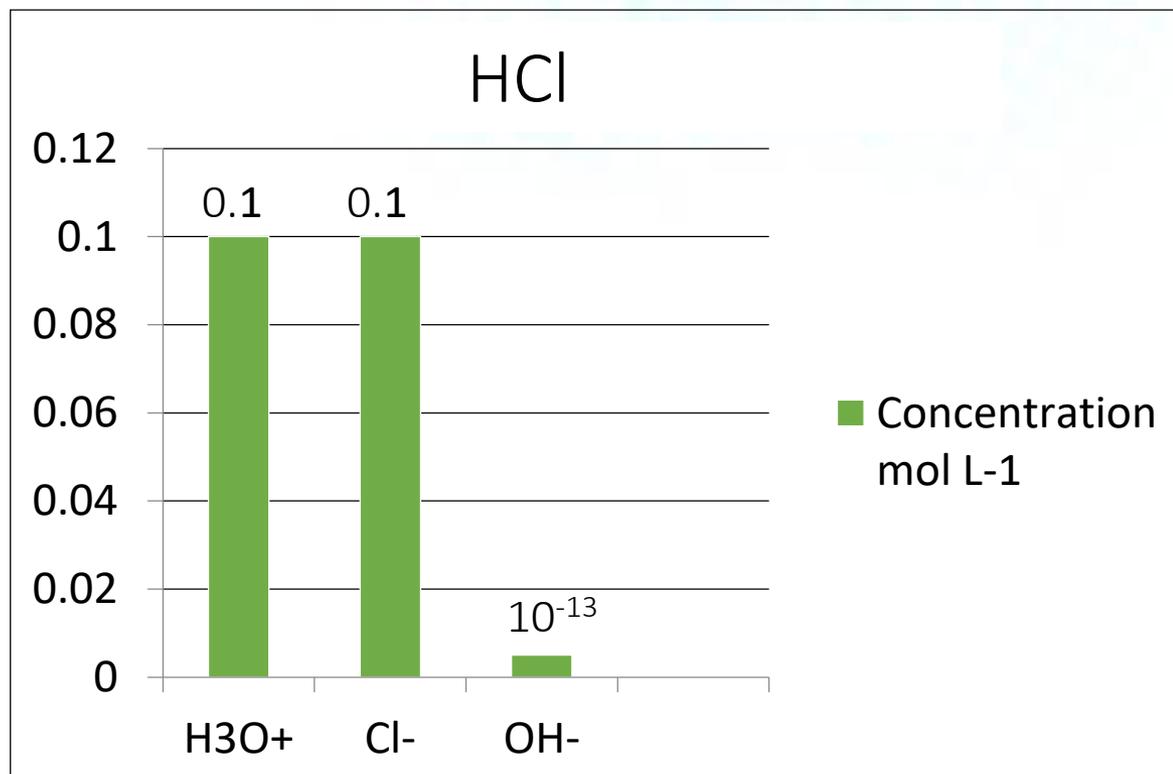
- In aqueous solutions water will almost always be present in the highest concentration.
- Small quantities of H<sub>3</sub>O<sup>+</sup> and OH<sup>-</sup> will also be present, according to the  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$
- Information on relative concentration can often be presented in a bar graph.

## Concentration of ions in solution – Strong Acid

Strong Acid i.e. HCl reacting with water



Strong acids will provide **good conductivity** and **pH 1-2** due to the high presence of  $\text{H}_3\text{O}^+$  ions



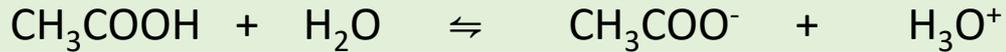
No strong acid will be left in the final mixture.

$\text{H}_3\text{O}^+$  and  $\text{Cl}^-$  are produced in equal concentrations – in the same concentration as the original strong acid.

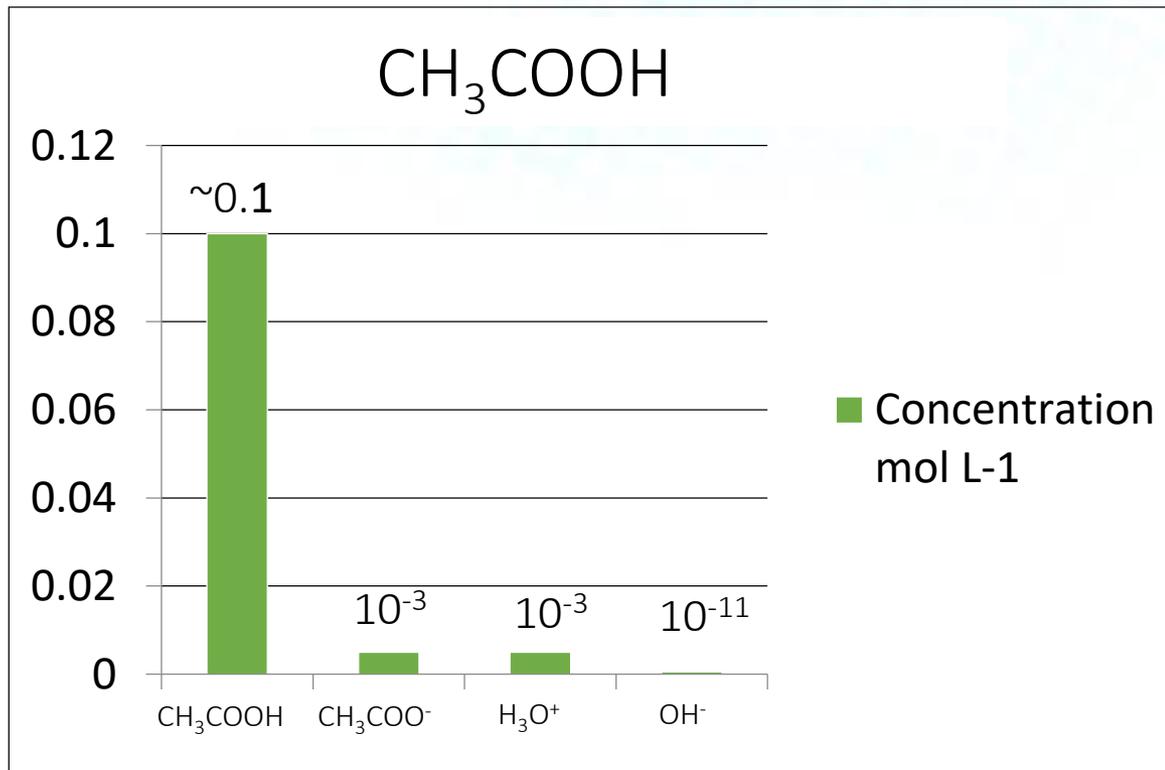
A small amount of  $\text{OH}^-$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## Concentration of ions in solution – Weak Acid

Weak Acid i.e.  $\text{CH}_3\text{COOH}$  reacting with water



Weak acids will provide **poor conductivity** and **pH 3-6** due to the low presence of  $\text{H}_3\text{O}^+$  ions



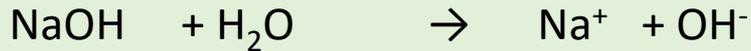
Most weak acid will be left in the final mixture.

Small amounts of  $\text{H}_3\text{O}^+$  and  $\text{CH}_3\text{COO}^-$  are produced in equal concentrations as only a small amount of the weak acid had dissociated.

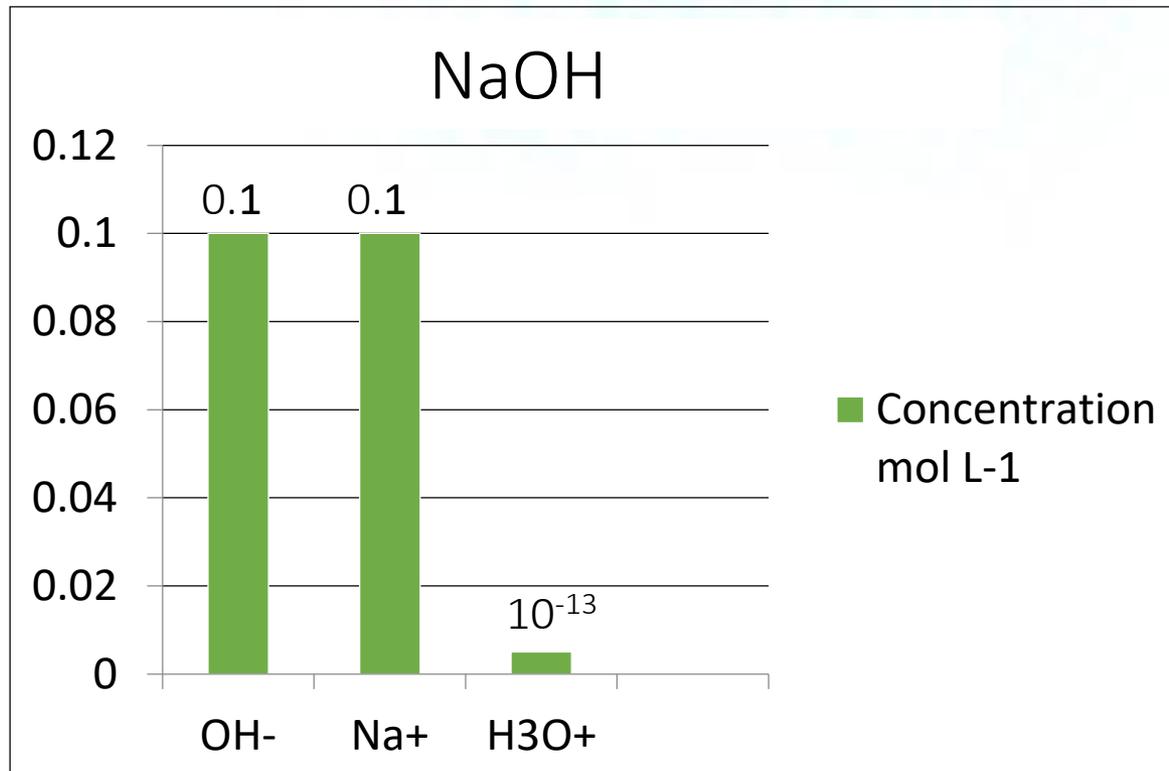
A small amount of  $\text{OH}^-$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## Concentration of ions in solution – Strong Base

Strong Base i.e. NaOH reacting with water



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



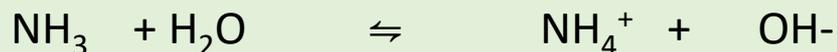
No strong base will be left in the final mixture.

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

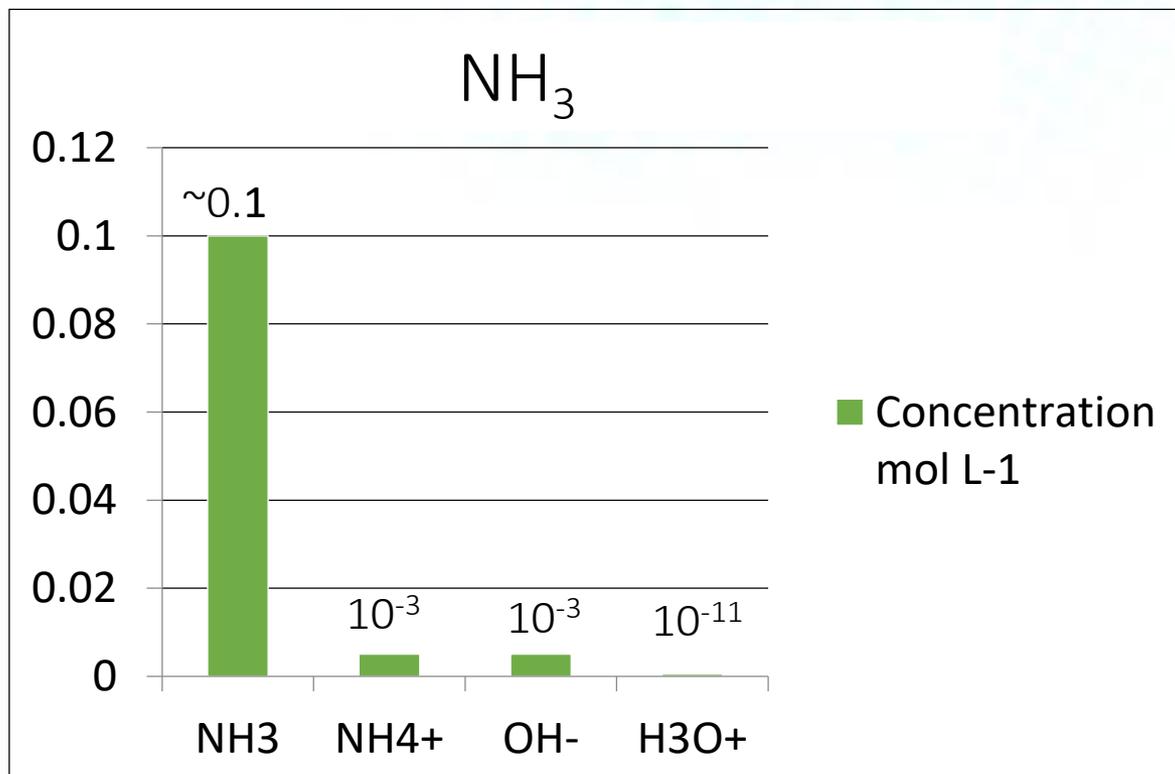
A small amount of  $\text{H}_3\text{O}^+$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## Concentration of ions in solution – Weak Base

Weak Base in water i.e.  $\text{NH}_3$



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



Most weak base will be left in the final mixture.

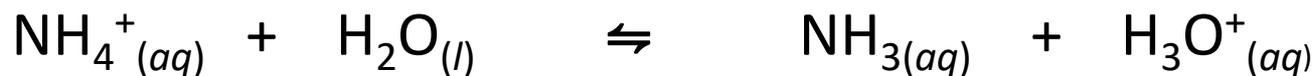
$\text{OH}^-$  and  $\text{NH}_4^+$  are produced in equal concentrations – a small amount of the weak base had dissociated.

A small amount of  $\text{H}_3\text{O}^+$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## pH of salt solutions

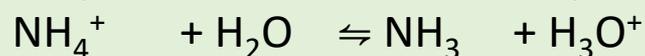
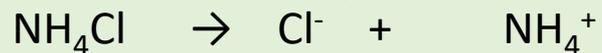
All cations that are the conjugate acids of weak bases act as weak acids and lower the pH of the solution. This means that a salt solution containing this cation could be acidic. For example, a solution of ammonium chloride,  $\text{NH}_4\text{Cl}$ , contains the cation  $\text{NH}_4^+$  and the anion  $\text{Cl}^-$ . The  $\text{Cl}^-$  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  $\text{NH}_4^+$  ion is the conjugate acid of the weak base  $\text{NH}_3$  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 an acid/base reaction with water. The non-reacting ion is left off as the spectator.

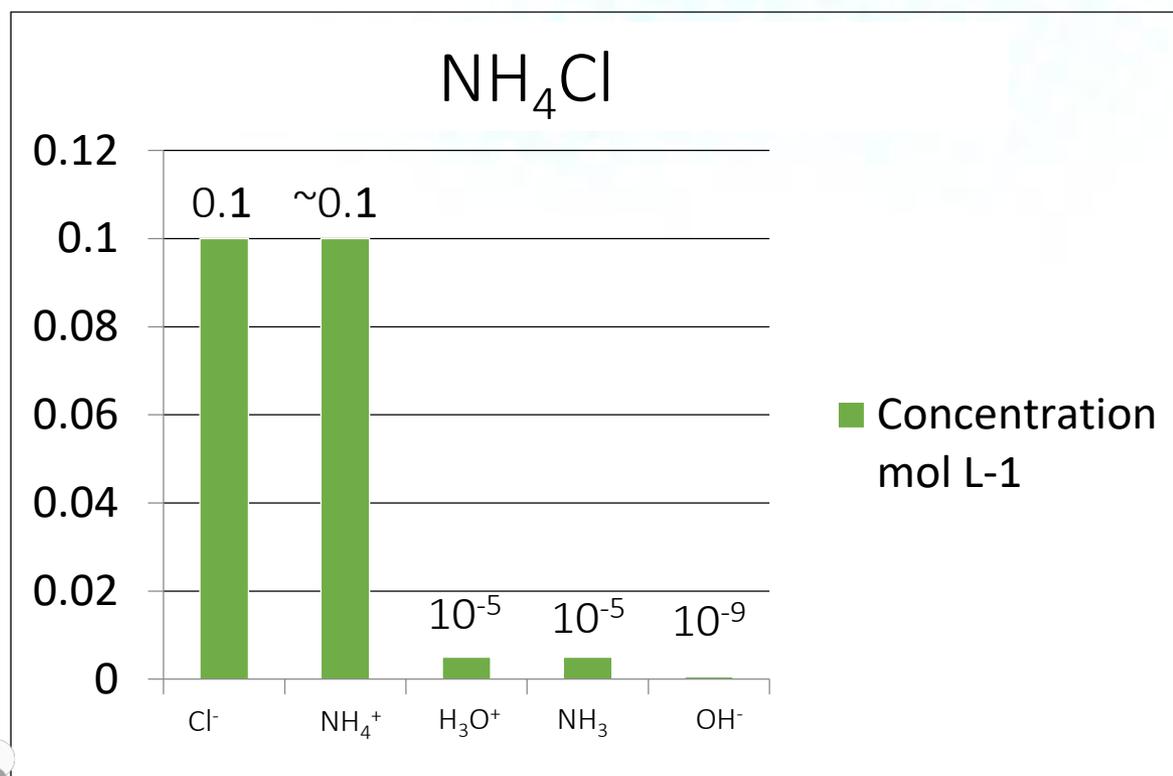


## Concentration of ions in solution – acid salt

Acid Salt i.e.  $\text{NH}_4\text{Cl}$



Acid salts will provide **good conductivity** and **pH < 7** due to the high presence of ions from dissolving and to a lesser extent  $\text{H}_3\text{O}^+$  ions



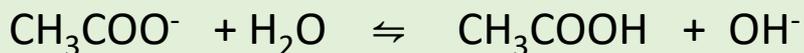
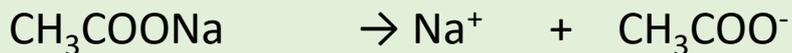
The spectator ion will be left in the highest concentration followed by the weak acid.

$\text{H}_3\text{O}^+$  and  $\text{NH}_3$  are produced in equal concentrations – a small amount of the weak acid had dissociated.

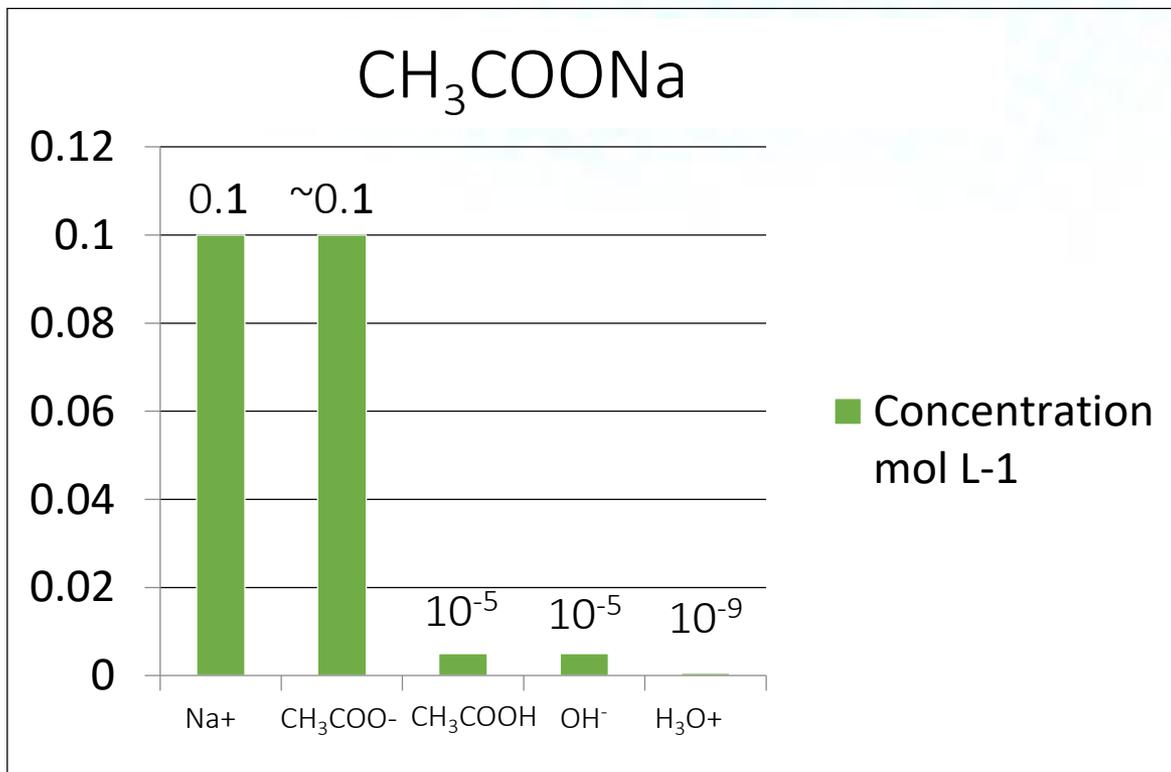
A small amount of  $\text{OH}^-$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## Concentration of ions in solution – base salt

Base Salt i.e.  $\text{CH}_3\text{COONa}$



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



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

$\text{OH}^-$  and  $\text{CH}_3\text{COOH}$  are produced in equal concentrations – a small amount of the weak base had dissociated.

A small amount of  $\text{H}_3\text{O}^+$  is present as water dissociates into  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

## NCEA 2013 Conductivity and Ions (part one)

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  $\text{NH}_3(aq)$ ,  $\text{HCl}(aq)$  and  $\text{NH}_4\text{Cl}(aq)$

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.

$\text{NH}_4\text{Cl}(aq)$  is solution **A**: good conductor of electricity – it fully dissociates in solution into ammonium and chloride ions, which conduct electricity.



Its pH is that of a weak acid, as the ammonium ion is a weak acid and partially dissociates in water, producing hydronium ions:



## NCEA 2013 Conductivity and Ions - (part two)

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  $\text{NH}_3(aq)$ ,  $\text{HCl}(aq)$  and  $\text{NH}_4\text{Cl}(aq)$

$\text{NH}_3(aq)$  is solution **B**: its pH is that of a weak base as  $\text{NH}_3$  is a weak base and it partially dissociates in water, producing hydroxide ions:



Poor conductor of electricity as it is only partially dissociated into ions in water. The remaining  $\text{NH}_3$  molecules are neutral and do not conduct electricity.

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



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 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

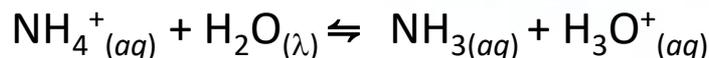
**Answer:** Strong base = **E**; red litmus paper turns blue and blue litmus remains blue indicates base; good conductivity indicates large number of charged particles to conduct current – strong base completely dissociates and produces high  $[\text{OH}^-]$ .  
Neutral salt = **C**; no colour change 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.

**Question 2(d):** Ethanoic acid solution,  $\text{CH}_3\text{COOH}_{(aq)}$ , and ammonium chloride solution,  $\text{NH}_4\text{Cl}_{(aq)}$ , are both weakly acidic. Identify and justify, using equations, which acid solution has greater electrical conductivity.

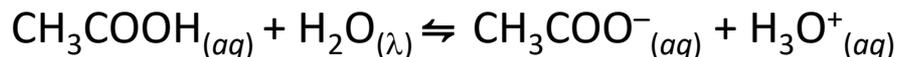
**Answer:**  $\text{NH}_4\text{Cl}$  is a better conductor of electricity because it completely dissolves / dissociates into ions:



(Then the  $\text{NH}_4^+$  ions react in water to produce hydronium ions, which makes it a weak acid:



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.



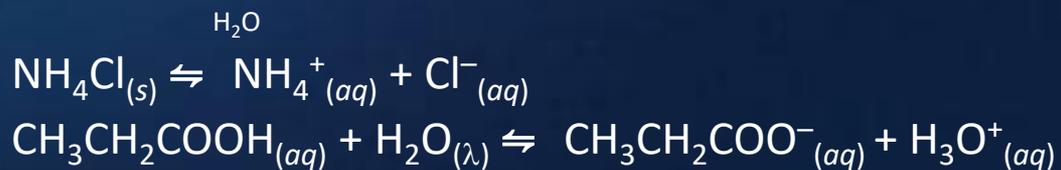
The ability to act as a conductor depends upon the concentration of ions. Ammonium chloride has a greater concentration of ions, so is a better conductor of electricity.

## NCEA 2016 Conductivity and ions

Excellence  
Question

**Question 2(d): (ii)** Explain why the solution of ammonium chloride,  $\text{NH}_4\text{Cl}(aq)$ , is a good conductor of electricity, while the solution of propanoic acid,  $\text{C}_2\text{H}_5\text{COOH}(aq)$ , is a poor conductor of electricity.

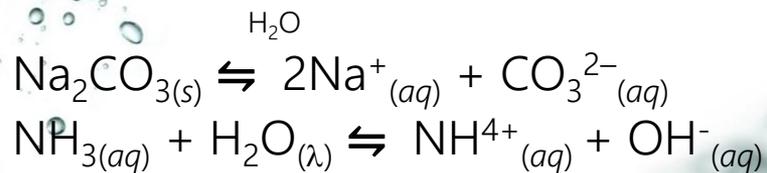
**Answer:** Charged particles that are free to move are needed to conduct electricity. Ammonium chloride is a good conductor of electricity because it is a salt that will completely dissolve / dissociate / ionise in water, releasing its ions  $\text{NH}_4^+$  and  $\text{Cl}^-$  into solution, i.e. relatively 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



## NCEA 2017 Conductivity and ions

**Question 1d:** Solutions of ammonia,  $\text{NH}_3(\text{aq})$ , and sodium carbonate,  $\text{Na}_2\text{CO}_3(\text{aq})$ , 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  $\text{CO}_3^{2-}$  and  $\text{Na}^+$  (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



## NCEA 2013 Reaction rates of Acids

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

acid	Concentration / molL <sup>-1</sup>	pH	Relative rate of reaction with Mg
HA	0.100	3.4	slow
HB	0.0100	2	fast
HC	1.00 × 10 <sup>-5</sup>	5	Very slow

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

- determine the concentration of hydronium ions, H<sub>3</sub>O<sup>+</sup>, 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

$$\text{HA } [\text{H}_3\text{O}^+] = 10^{-3.4} = 3.98 \times 10^{-4} \text{ molL}^{-1} \quad \text{HB } [\text{H}_3\text{O}^+] = 10^{-2} = 0.0100 \text{ molL}^{-1} \quad \text{HC } [\text{H}_3\text{O}^+] = 10^{-5} = 1.00 \times 10^{-5} \text{ molL}^{-1}$$

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

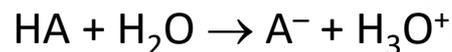
## NCEA 2014 Reaction rates of Acids (PART ONE)

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

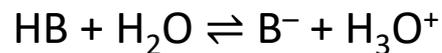
Solution	pH
0.100 mol L <sup>-1</sup> HA(aq)	1.0
0.100 mol L <sup>-1</sup> HB(aq)	2.2

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.



$$[\text{H}_3\text{O}^+] = 0.100 \text{ mol L}^{-1}$$



$$[\text{H}_3\text{O}^+] = 0.00631 \text{ mol L}^{-1}$$

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<sup>-1</sup>).

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<sup>-1</sup>.

## 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,  $\text{CaCO}_3(\text{s})$ , are reacted, separately, with excess HA and HB.

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

HA

Expect reaction to be more vigorous; rapidly produces gas / bubbles ( $\text{CO}_2$ ) – 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.

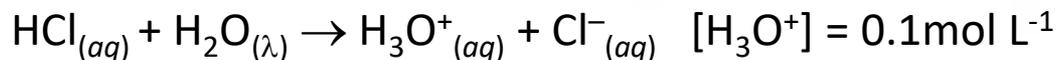
## NCEA 2015 Reaction rates of Acids

Excellence  
Question

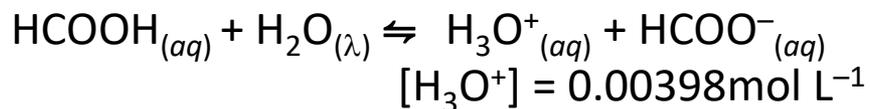
**Question 2(d):** (e) The table shows the pH of two acidic solutions, methanoic acid, HCOOH, and hydrochloric acid, HCl, which both have a concentration of 0.1 mol L<sup>-1</sup>. Compare and contrast the pH of each solution, and their expected rate of reaction with a 2 cm strip of cleaned magnesium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

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

**Answer:** HCl is a strong acid because it fully dissociates in water.



Whereas HCOOH is a weak acid, it does not readily dissociate in water.



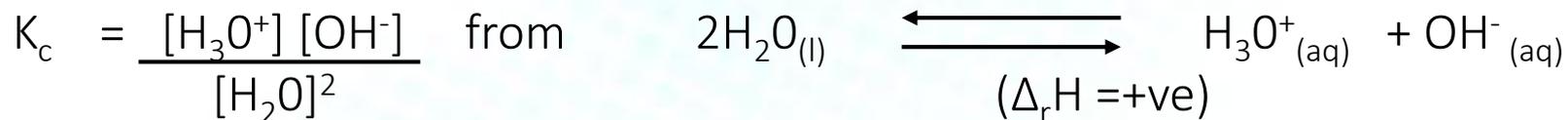
In the resulting solutions, HCl has a higher concentration of H<sub>3</sub>O<sup>+</sup>, and therefore a lower pH (1) than HCOOH, which has a lower concentration of H<sub>3</sub>O<sup>+</sup>, and therefore a higher pH (2.4).

Both acids will react with the cleaned Mg ribbon, but with a higher concentration of H<sub>3</sub>O<sup>+</sup>, HCl will have a faster rate of reaction with Mg than HCOOH, as there are more H<sub>3</sub>O<sup>+</sup> ions available to react in a given volume.

## K<sub>w</sub> – 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.



$$\text{Or } K_c \times [H_2O]^2 = [H_3O^+][OH^-]$$

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  $K_w$

As  $[H_3O^+] \times [OH^-]$  always equals  $1 \times 10^{-14}$  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 $[\text{OH}^-]$ or $[\text{H}_3\text{O}^+]$

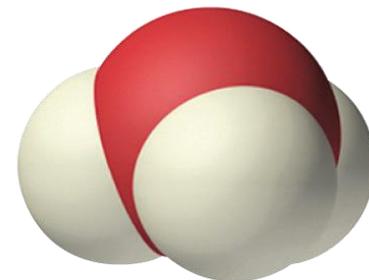
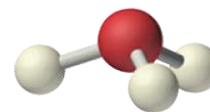
The product  $[\text{OH}^-] \times [\text{H}_3\text{O}^+]$  is a constant value (at the same temperature)

$$\text{IONIC PRODUCT} \quad K_w = [\text{OH}^-] \times [\text{H}_3\text{O}^+] = 1 \times 10^{-14}$$

e.g. If solution A has  $[\text{OH}^-] = 1 \times 10^{-1} \text{ molL}^{-1}$  find  $[\text{H}_3\text{O}^+]$

$$[\text{H}_3\text{O}^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-1}} = 1 \times 10^{-13} \text{ molL}^{-1}$$

$$[\text{OH}^-] = \frac{1 \times 10^{-14}}{[\text{H}_3\text{O}^+]} \quad \text{or} \quad [\text{H}_3\text{O}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$$



Hydronium ion,  
 $\text{H}_3\text{O}^+$

## Calculating pH given $[\text{OH}^-]$ or $[\text{H}_3\text{O}^+]$ in a strong acid or base



pH is the measure of the  $[\text{H}_3\text{O}^+]$  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.

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

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

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

Buttons on calculator

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

$$[\text{H}_3\text{O}^+] = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$$

Use this equation first if given  $[\text{OH}^-]$  instead of  $[\text{H}_3\text{O}^+]$

**Note:** the concentration of monoprotic ions, such as HCl and  $\text{HNO}_3$  that donate 1 hydrogen ion can be assumed to equal the  $[\text{H}_3\text{O}^+]$  as they fully dissociate.  $\text{H}_2\text{SO}_4$  will donate 2 hydrogen ions so the concentration of  $[\text{H}_3\text{O}^+]$  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  $[\text{OH}^-]$

These assumptions only apply for strong acids and bases.

## Calculating $[\text{OH}^-]$ or $[\text{H}_3\text{O}^+]$ given pH

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

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

e.g. calculate  $[\text{H}_3\text{O}^+]$  if solution has pH=2      $[\text{H}^+] = 10^{-2} \text{ mol L}^{-1}$

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

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

e.g. calculate  $[\text{H}_3\text{O}^+]$  if solution has pH=11.3      $[\text{H}_3\text{O}^+] = 10^{-11.3} \text{ mol L}^{-1}$

$$[\text{H}_3\text{O}^+] = 5.0 \times 10^{-12} \text{ mol L}^{-1}$$

Buttons on calculator

Shift    10<sup>x</sup>    (-)    1    1    .    3

## Summary of pH formula

Convert  $[\text{OH}^-]$  to  $[\text{H}_3\text{O}^+]$

Convert  $[\text{H}_3\text{O}^+]$  to  $[\text{OH}^-]$

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-14} / [\text{OH}^-] \quad [\text{OH}^-] = 1 \times 10^{-14} / [\text{H}_3\text{O}^+]$$

Convert  $[\text{H}_3\text{O}^+]$  to pH

Convert  $[\text{OH}^-]$  to pOH

Convert pOH to pH

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{pOH} = -\log[\text{OH}^-] \quad \text{pH} = 14 - \text{pOH}$$

Convert pH to  $[\text{H}_3\text{O}^+]$

Convert pH to  $[\text{OH}^-]$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} \quad [\text{OH}^-] = 10^{-(14-\text{pH})}$$

## 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,  $[\text{H}_3\text{O}^+]$ , and the hydroxide ion concentration,  $[\text{OH}^-]$ , in the solution.  $K_w = 1 \times 10^{-14}$

**Answer:**

$$[\text{H}_3\text{O}^+] = 1.58 \times 10^{-13} \text{ mol L}^{-1}$$

$$[\text{OH}^-] =$$

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

$$= 0.0633 \text{ mol L}^{-1}$$

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

**Answer:**

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-14} / [\text{OH}^-]$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = 10.4$$

OR

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pH} = 14 - \text{pOH}$$

**Question 2(c):** (i) A solution of nitric acid,  $\text{HNO}_{3(aq)}$ , has a hydronium ion,  $\text{H}_3\text{O}^+$ , concentration of  $0.0243 \text{ mol L}^{-1}$ .

Determine, by calculation, the pH and the concentration of hydroxide ions,  $\text{OH}^-$ , in this solution.

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

**Answer:**

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] = 1.61$$

$$\begin{aligned} [\text{OH}^-] &= \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{10^{-14}}{0.0243} \\ &= 4.12 \times 10^{-13} \text{ mol L}^{-1} \end{aligned}$$

OR

$$\begin{aligned} \text{pH} &= -\log_{10} 0.0243 \\ &= 1.61 \end{aligned}$$

$$\begin{aligned} \text{pOH} &= 14 - 1.61 \\ &= 12.4 \end{aligned}$$

$$\begin{aligned} [\text{OH}^-] &= 10^{-12.4} \\ &= 4.12 \times 10^{-13} \text{ mol L}^{-1} \end{aligned}$$

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

**Answer:**

$$\begin{aligned} [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} = 10^{-11.8} = 1.58 \times 10^{-12} \text{ mol L}^{-1} \\ [\text{OH}^-] &= \end{aligned}$$

$$\frac{K_w}{[\text{H}_3\text{O}^+]} = 6.31 \times 10^{-3} \text{ mol L}^{-1}$$

## NCEA 2016 pH Calculations

Excellence  
Question

**Question 2(c):** (i) Calculate the pH of a 0.0341 mol L<sup>-1</sup> hydrochloric acid, HCl(aq), solution.

pH =

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log 0.0341 = 1.47$$

**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<sub>3</sub>O<sup>+</sup>, and hydroxide ions, OH<sup>-</sup>, in this solution.

[H<sub>3</sub>O<sup>+</sup>] =

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-12.4} = 3.98 \times 10^{-13} \text{ mol L}^{-1}$$

[OH<sup>-</sup>] =

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = 0.0251 \text{ mol L}^{-1}$$

## NCEA 2016 pH Calculations

Excellence  
Question

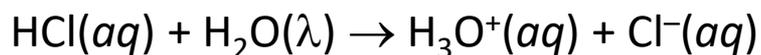
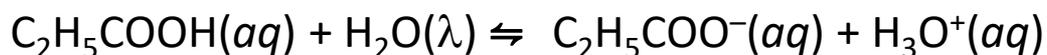
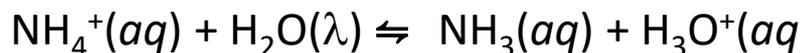
**Question 2(d):** The table shows the pH of three acidic solutions, ammonium chloride,  $\text{NH}_4\text{Cl}$ , propanoic acid,  $\text{C}_2\text{H}_5\text{COOH}$ , and hydrogen chloride,  $\text{HCl}$ .

(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.

	$\text{NH}_4\text{Cl}(aq)$	$\text{C}_2\text{H}_5\text{COOH}(aq)$	$\text{HCl}(aq)$
Concentration/mol L <sup>-1</sup>	0.1	0.1	0.1
pH	5.62	3.44	1.0

pH reflects the concentration of  $\text{H}_3\text{O}^+$  ions. The higher the pH, the lower the  $[\text{H}_3\text{O}^+]$ . Both ammonium ions,  $\text{NH}_4^+$ , and propanoic acid are weak acids, so both only partially dissociate in water and produce fewer  $\text{H}_3\text{O}^+$  ions in water than  $\text{HCl}$ , which is a strong acid. As a strong acid,  $\text{HCl}$  completely dissociates in water, resulting in a high concentration of  $\text{H}^+$  ( $\text{H}_3\text{O}^+$ ) ions and a low pH.



$\text{NH}_4^+$  has a higher pH than propanoic acid, which means that it has a lower  $[\text{H}_3\text{O}^+]$  in solution.

## NCEA 2017 pH Calculations

Excellence  
Question

**Question 1c:** (i) A solution of sodium hydroxide,  $\text{NaOH}_{(aq)}$ , has a pH of 11.6. Calculate the hydronium ion concentration  $[\text{H}_3\text{O}^+]$ , and the hydroxide ion concentration,  $[\text{OH}^-]$ , in the solution.

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

$$\begin{aligned} [\text{H}_3\text{O}^+] &= 10^{-\text{pH}} \\ &= 10^{-11.6} \\ &= 2.51 \times 10^{-12} \text{ mol L}^{-1} \end{aligned}$$

$$\begin{aligned} [\text{OH}^-] &= \frac{K_w}{[\text{H}_3\text{O}^+]} \\ &= 3.98 \times 10^{-3} \text{ mol L}^{-1} \end{aligned}$$

(ii) Calculate the pH of a  $2.96 \times 10^{-4} \text{ mol L}^{-1}$  solution of potassium hydroxide,  $\text{KOH}_{(aq)}$ .  
pH =

$$\begin{aligned} \text{pOH} &= -\log(2.96 \times 10^{-4}) = 3.53, \\ \text{so pH} &= 14.0 - 3.53, \text{ pH} = 10.5. \end{aligned}$$

OR

$$\begin{aligned} [\text{H}_3\text{O}^+] &= \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{2.96 \times 10^{-4}} = 3.38 \times 10^{-11} \end{aligned}$$

$$\text{So, pH} = -\log(3.38 \times 10^{-11}) = 10.5.$$

## Definitions

**Acid:** A solution that has an excess of  $H^+$  ions.

**Alkali:** A base in solution that has an excess of  $OH^-$  ions.

**Amphiprotic:** A substance that can act as either an acid or a base.

**Aqueous:** A solution that is mainly water.

**Base:** A substance that accepts  $H^+$  ions.

**Neutral:** A solution that has a pH of 7. It is neither acidic nor basic.

**Strong Acid:** An acid that has a very low pH (0-4). The  $H^+$  ions completely disassociate in solution

**Strong Base:** A base that has a very high pH (10-14). 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).