

Chemical Reactivity AS 91166

Achievement Criteria

This achievement standard involves demonstrating understanding of chemical reactivity.

Rates of Reaction typically involves:

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

Equilibrium principles is limited to:

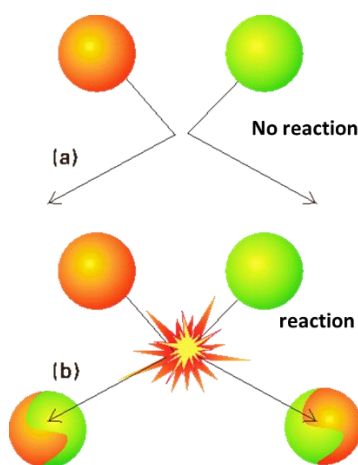
- the dynamic nature of equilibrium
- the effect of changes in temperature, concentration, pressure, or addition of a catalyst on equilibrium systems
- the significance of the equilibrium constant (K_c) for homogeneous systems. This may involve calculations.
- the nature of acids and bases in terms of proton transfer
- properties of aqueous solutions of strong and weak acids and bases including ionic species such as 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).

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

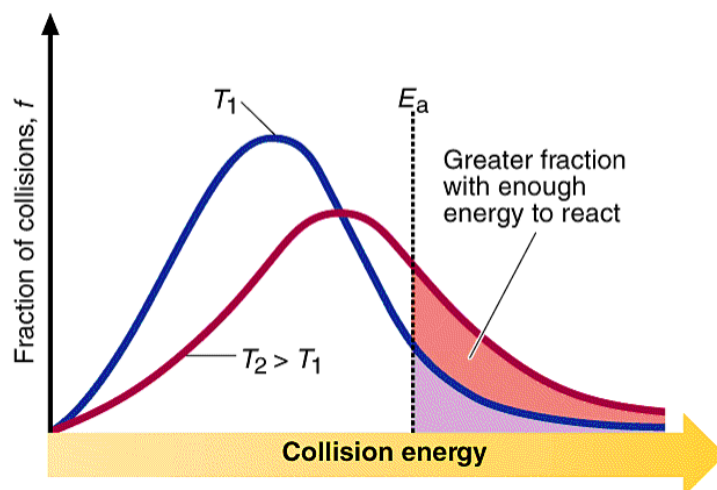
If these conditions are met, the collision will be considered successful. (Or effective)



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.

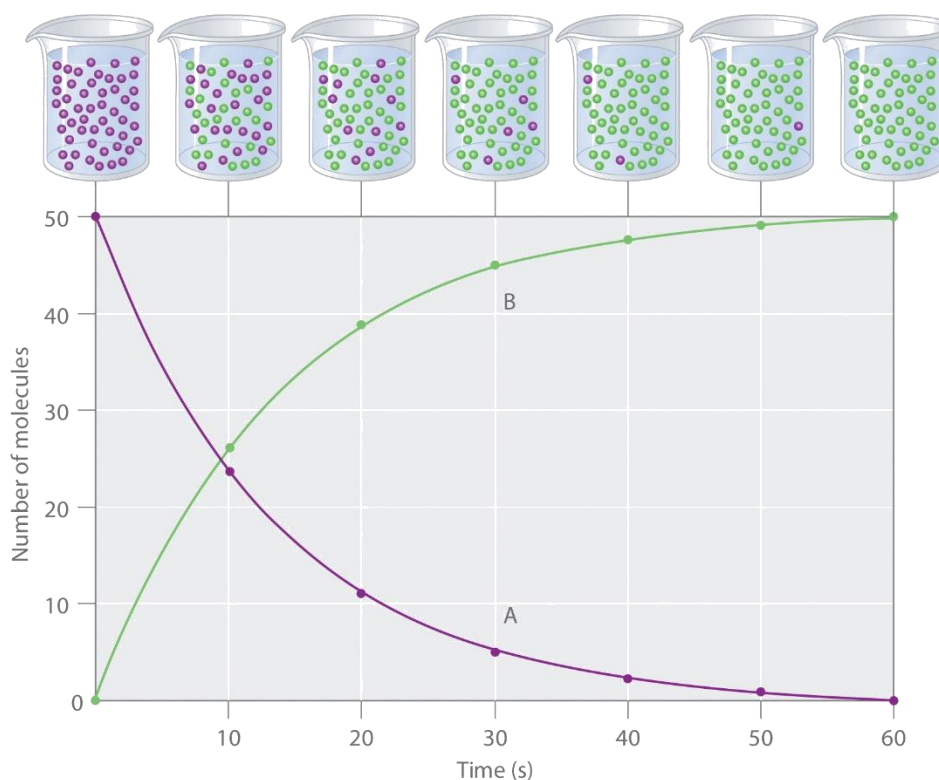
At lower temperatures (T_1) most particles will have the same collision energy when they collide. As the temperature is increased (T_2) the range of collision energy is more spread out and a greater proportion of particles will have enough energy to cross the activation energy threshold and therefore 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.

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.

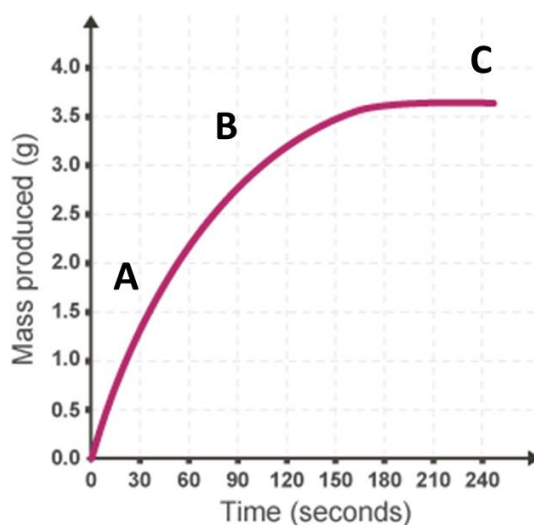


Reaction rate is the speed at which a chemical reaction occurs

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.

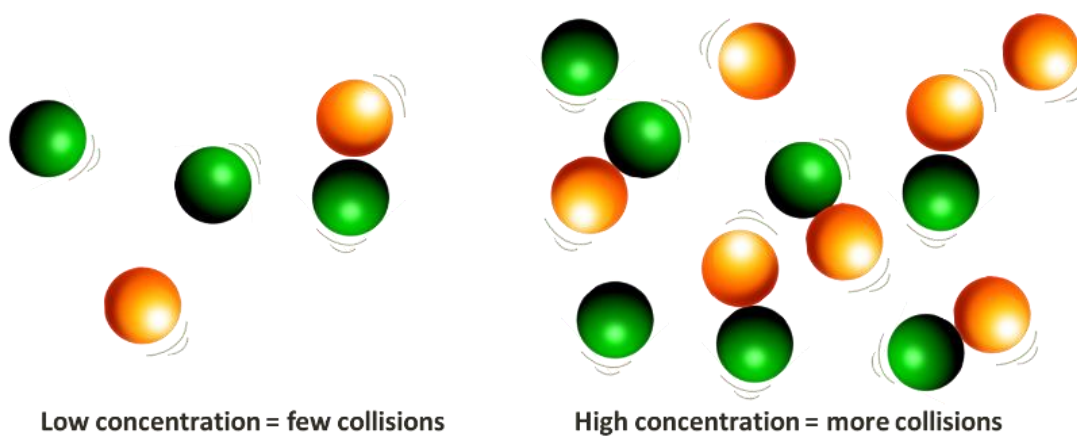
B. As the reaction proceeds there will be less reactant particles available to collide as many have already reacted to form products. The gradient of the line will be lower, and therefore the frequency of collisions will be less, resulting in less successful collisions; resulting in the slowing down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



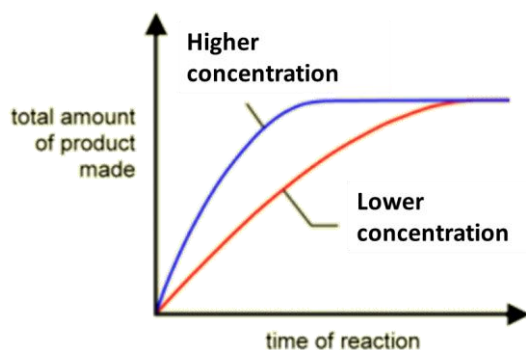
Reaction rate can be increased by increasing the concentration

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.



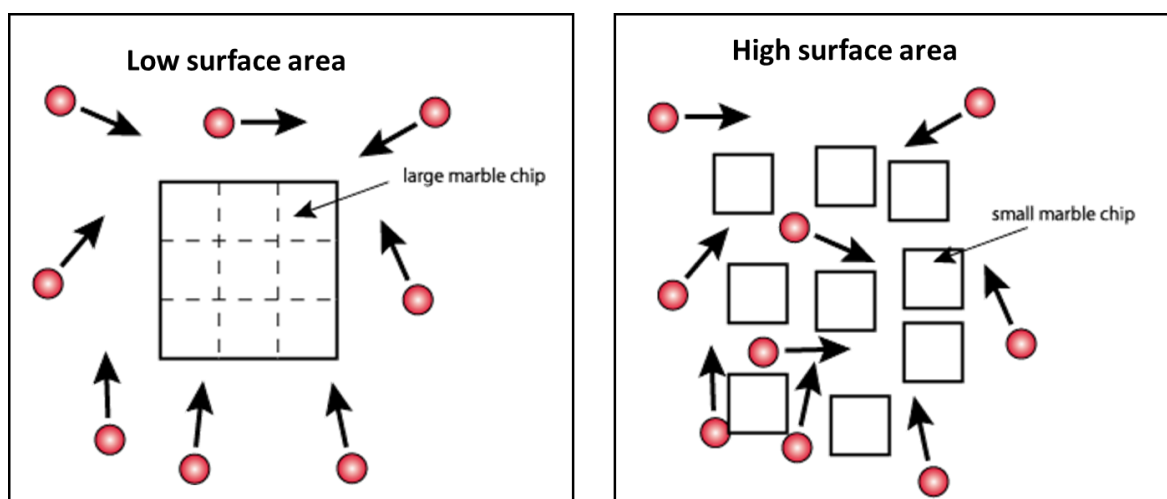
It is important to note that the total amount of product made depends upon the total amount of reactants at the start. A 1 molL^{-1} solution will contain only half the particles of a 2 molL^{-1} so twice the volume will be required to produce the same quantity of product.

Also, note that 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

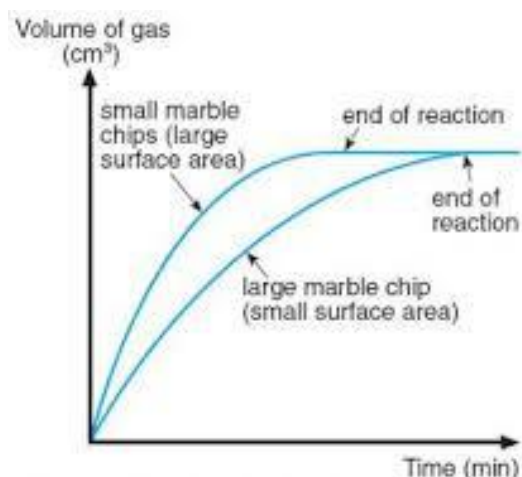
Surface area can be increased by grinding and crushing large lumps into a finer powder. The smaller the pieces, the greater the surface area. The reactant(s) with a greater surface area will have a faster rate of reaction than the same amount of reactants with a smaller surface area.



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

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

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



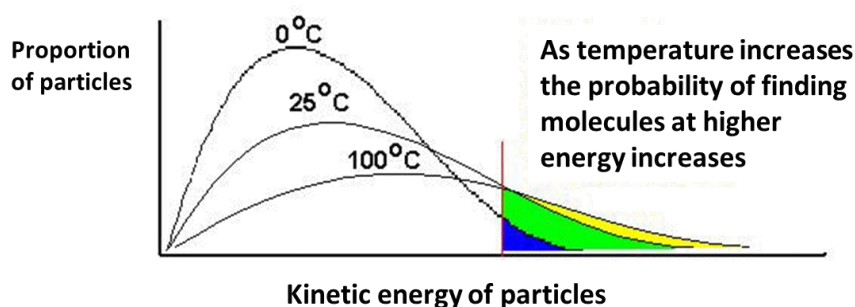
Reaction rate can be increased by increasing the Temperature

Increasing temperature effects the reaction rate in two ways.

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

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

Reaction rate can be increased by increasing the Temperature



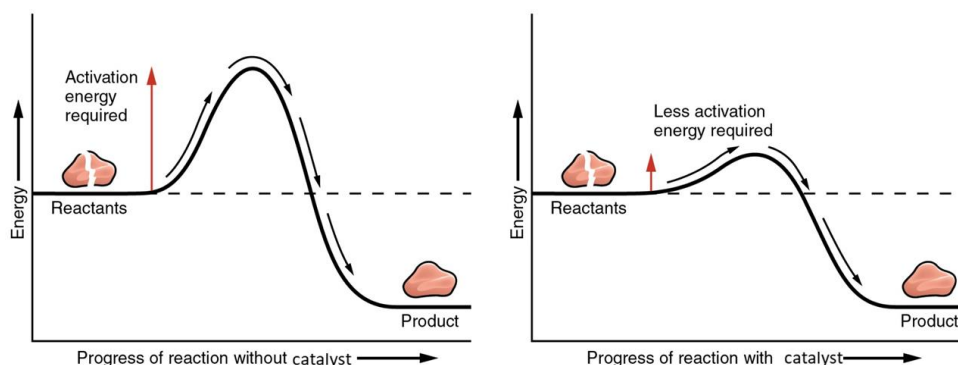
At any given temperature, there will be a range in the kinetic energy of particles. At a lower temperature a greater proportion of particles are likely to have insufficient kinetic energy during a collision in order for a successful collision, and therefore a reaction, to take place. Increasing the temperature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

Reaction rate can be increased by using a catalyst.

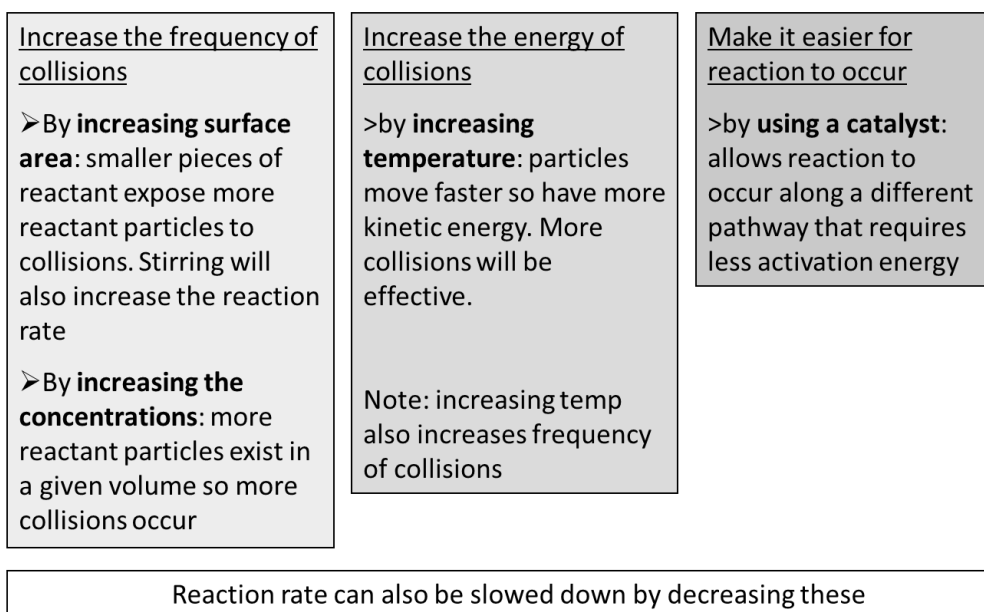
A catalyst is a substance that increases the reaction rate without being used up or forming part of the products. Only some reactions have catalysts that are effective, but for many reactions, there is no catalyst that works.

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.

The reaction shown below is an exothermic reaction: the enthalpy level of the products is lower than the enthalpy of the reactions therefore energy is released during this reaction. Activation energy is required before a collision is successful between particles. A catalyst lowers the activation energy pathway and therefore a greater proportion of particles will have sufficient energy during collision for it to be successful.

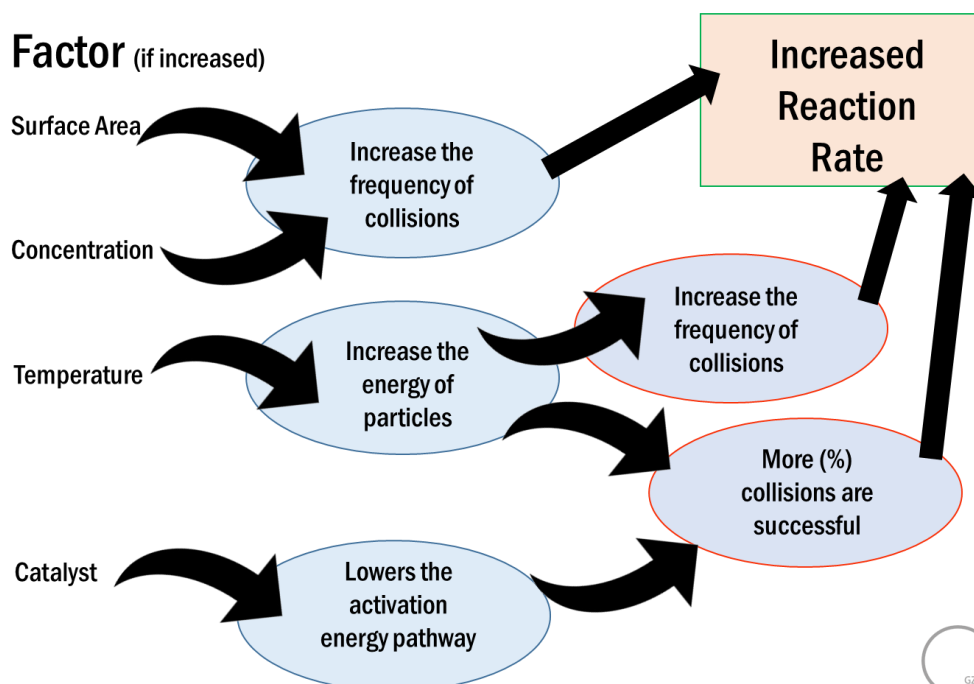


Summary of Factors affecting Reaction rate



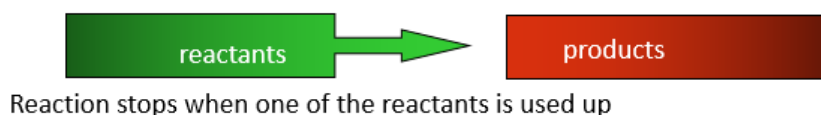
Writing Reaction rate Answers

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.

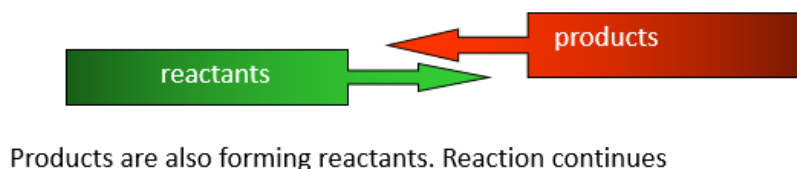


Equilibrium

Some reactions go to completion



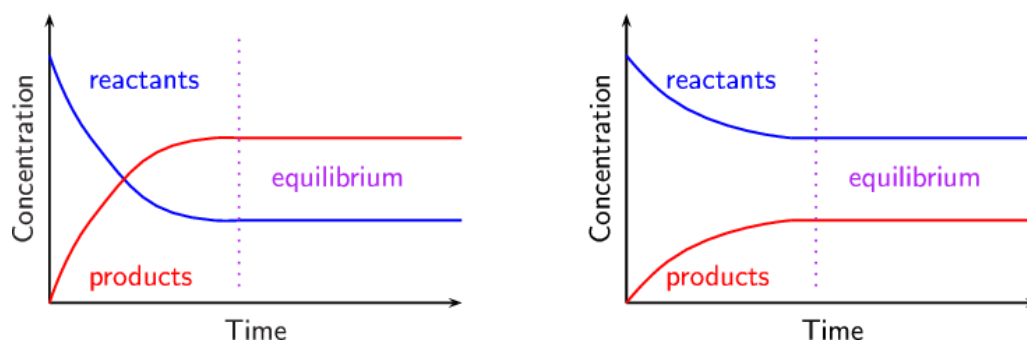
Other reactions are reversible



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

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.

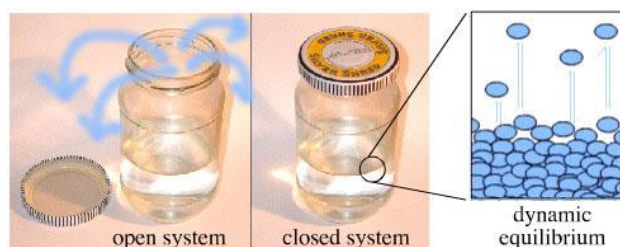


On the left hand side, the proportion of products will be higher than the reactants and on the right hand side, the proportion of reactants will be higher than products.

Equilibrium – Evaporation case study

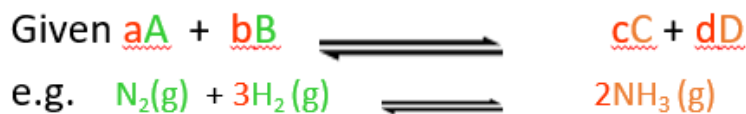
A dynamic equilibrium must occur in a closed system where all reactants and products are retained in an area where particles can collide with each other.

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.



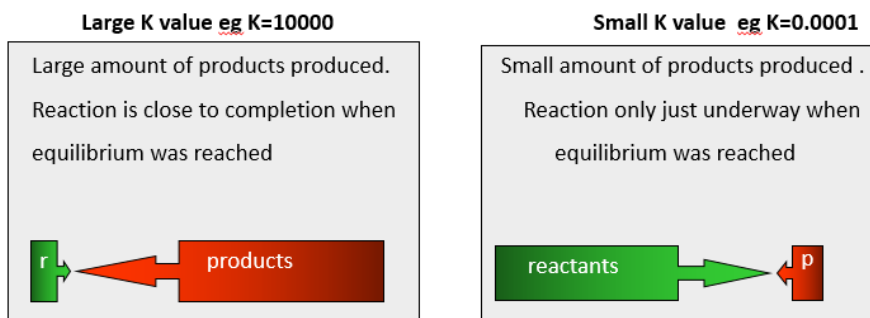
$$K_c = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

[] = concentration in mol L^{-1}
at equilibrium

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

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

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



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 mol L^{-1} into the correct position in the expression. (Remember do not put solids in as they have a fixed density and their concentration does not change)



If the values given for concentration are: $[NH_3] = 0.325 \text{ mol L}^{-1}$, $[N_2] = 0.625 \text{ mol L}^{-1}$ and $[H_2] = 2.45 \text{ mol L}^{-1}$

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

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

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

Step 2. expand out equation

Step 3. divide equation

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

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

Changes to the Equilibrium Constant: K_c

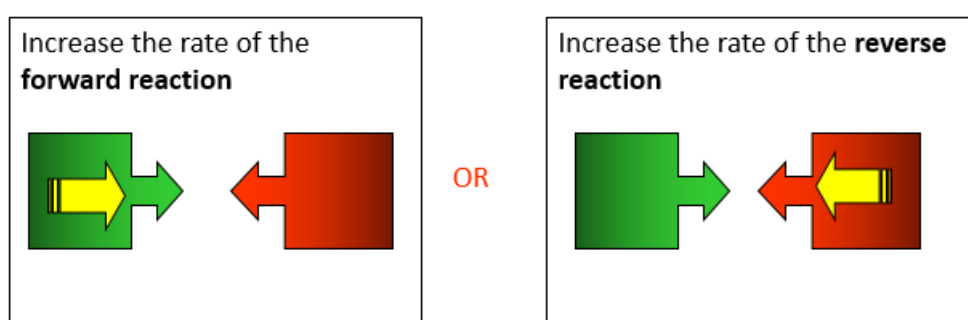
At a particular temperature, the equilibrium constant always has the same value for a set reaction. It is unaffected by a change in concentration, pressure or whether or not you are using a catalyst.

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:



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

Le Chatelier's Principle

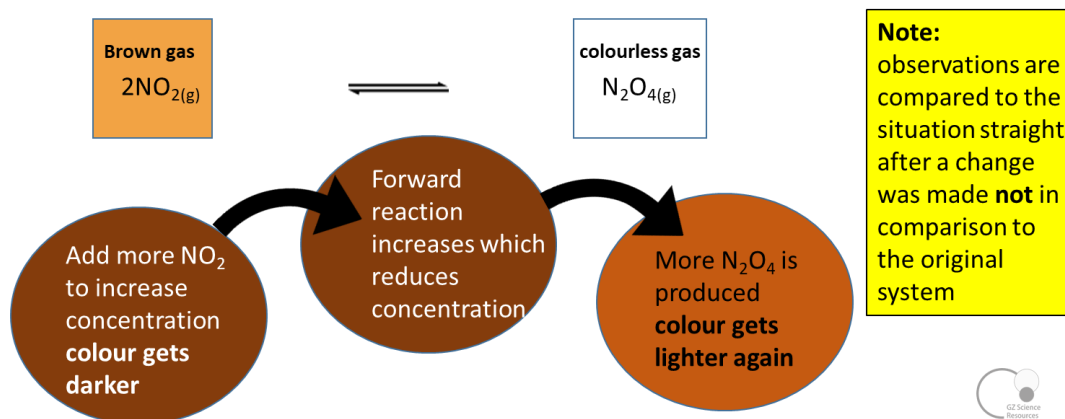
When a change is applied to a system at equilibrium, the system responds so that the effects of the change are minimised

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

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.

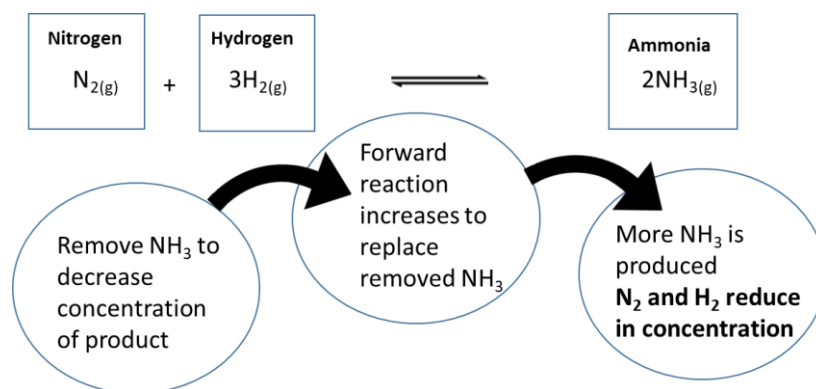
For example: if the concentration of the reactants is increased then the rate of reaction changing reactants in products will increase (due to collision theory) favouring the forward reaction. The effect of this will change more of the reactant into product – lowering the concentration of reactant and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

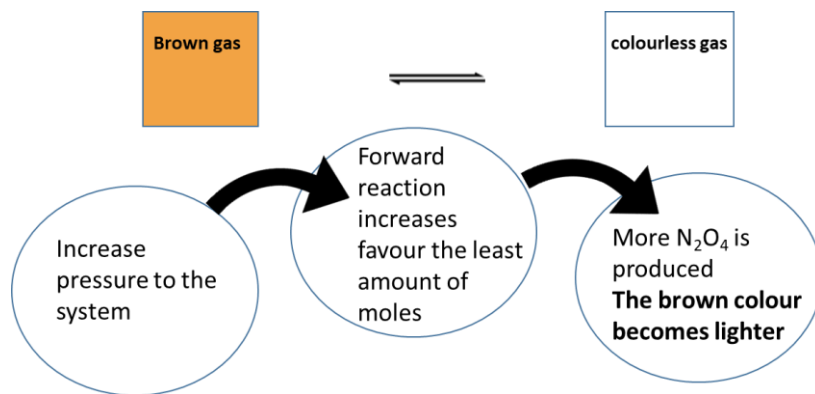


Changes in Equilibrium – Increasing Pressure

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.

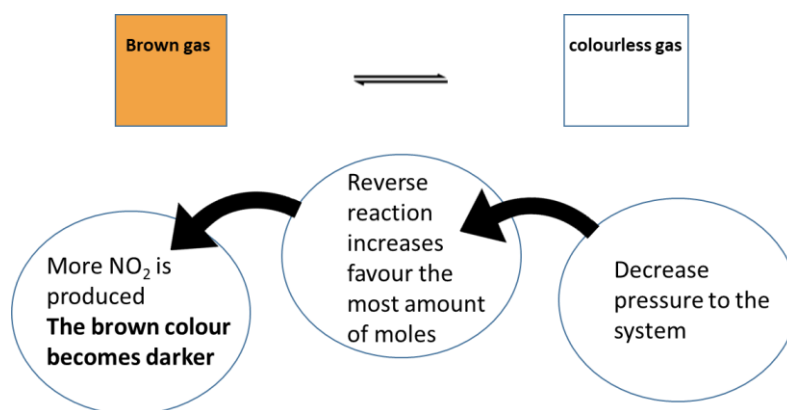


Note: if both sides have the same number of moles then there will be no change.

Changes in Equilibrium – Decreasing Pressure

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

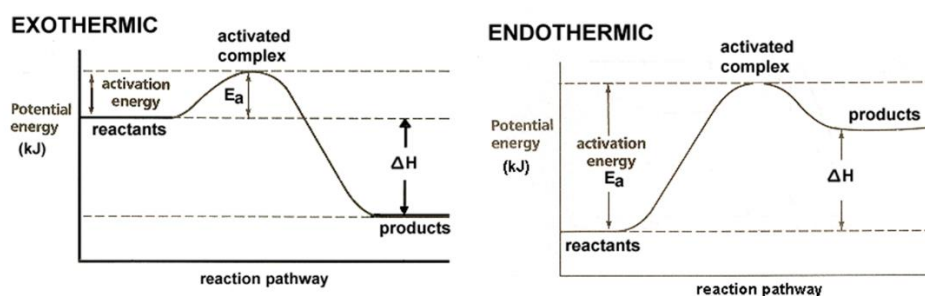
For example: In this reaction, the NO_2 has 2 moles compared to 1 mole of the N_2O_4 so the reaction rate from the product to the reactant is increased.



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.

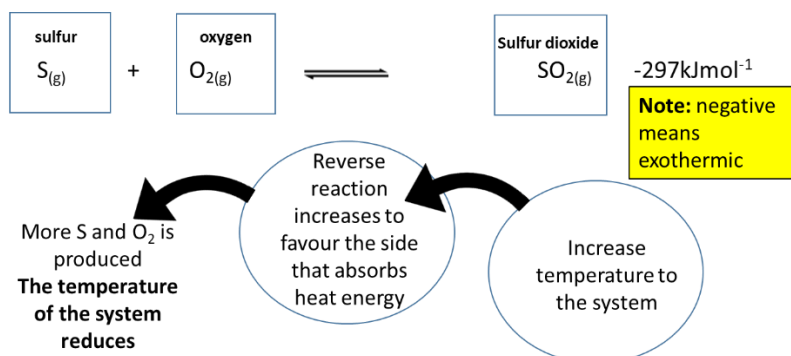


There is a difference in activation energy between the endothermic and exothermic direction. There will be uneven proportions of energy involved in activation energy and substance creation if the equilibrium is shifted. This changes the K_c value permanently.

Changes in Equilibrium – Increasing Temperature

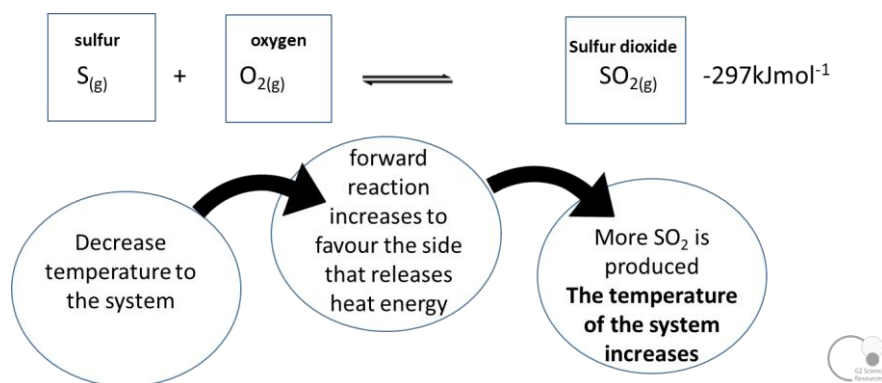
A change in temperature will permanently change the equilibrium, and K_c , of the system

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)



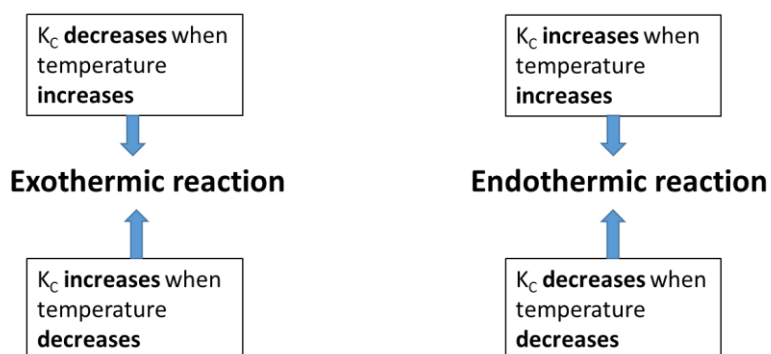
Changes in Equilibrium – Decreasing Temperature

When a reaction is exothermic, it releases heat energy in the forward reaction when products are formed. A 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.



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.

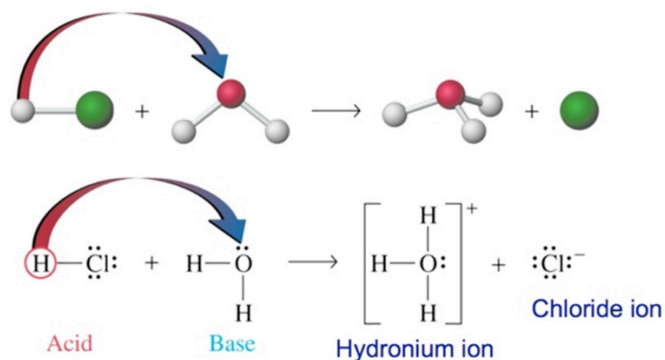
Acids – their characteristics

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

Common acids (that you need to know) include:

Strong acids: HNO_3 - nitric acid, HCl - hydrochloric acid, H_2SO_4 – sulfuric acid, (HBr - Hydrobromic acid)

Weak acids: CH_3COOH – ethanoic acid, NH_4^+ - ammonium and (HF - Hydrofluoric acid)



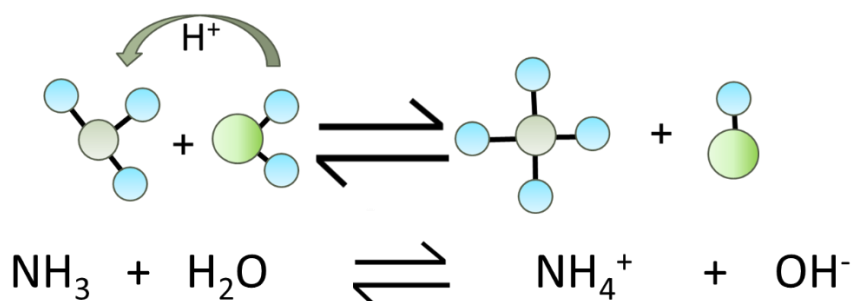
Bases – their characteristics

A Base accepts a Hydrogen ion that have been donated from an Acid.

Common bases (that you need to know) include:

Strong bases: $NaOH$ – sodium hydroxide, KOH – potassium hydroxide

Weak bases: NH_3 – ammonia, CH_3COO^- - ethanoate ion and (CH_3NH_2 – methylamine)

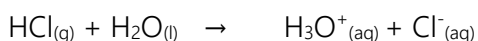


Some substances such as water are amphoteric and can act as both an acid and a base depending on what other substance the water is with.

Brønsted–Lowry theory of Acid and Base reactions

Acid-Base reactions involve the transfer of Hydrogen ions, H^+ . A hydrogen ion, H^+ is simply a lone proton (an H with the electron removed). In water (or aqueous solutions), H^+ ions exist as an H_3O^+ ion, called hydronium.

Acids are substances that donate protons (H^+) in solution



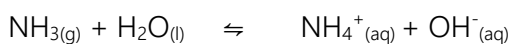
HCl gas dissolved in water

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

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

Solution becomes acidic since H_3O^+ ions form

Bases are substances that accept protons (H⁺) in solution



NH₃ gas dissolved in water

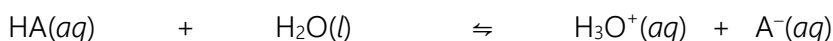
NH₃ has accepted an H⁺ so it is acting as a base

H₂O has donated an H⁺ so is acting as an acid

Solution becomes basic since OH⁻ ions form.

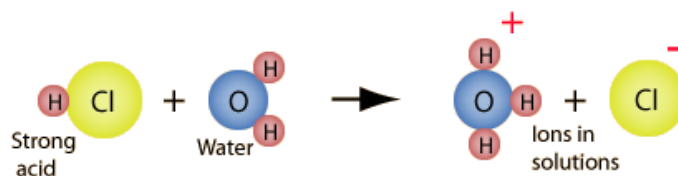
Brønsted–Lowry acids and bases summary

Proton donation to a water molecule forms H₃O⁺ (hydronium) ions.



acid base acid base

proton donor proton acceptor

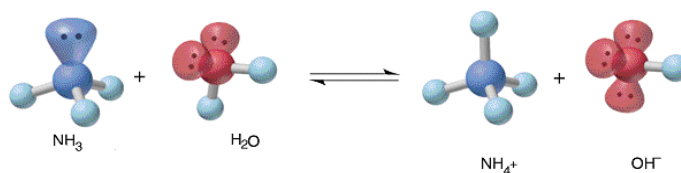


Similarly, proton donation from water to a base produces OH⁻ ions.



base acid acid base

proton acceptor proton donor



Amphiprotic substances

An amphiprotic substance is a substance that can donate or accept a proton, H⁺

For a substance to be amphiprotic it must

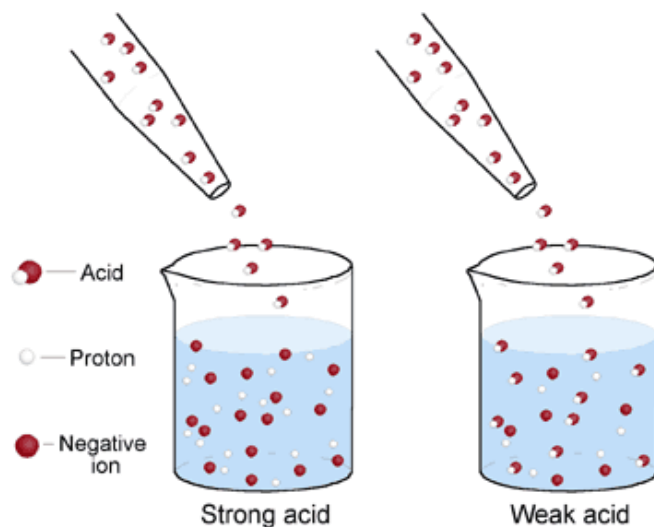
1. Contain a hydrogen atom, which is able to be donated to another chemical species.
2. Be able to accept a hydrogen ion from another species.

Examples of amphiprotic species include, H₂O, HCO₃⁻, HSO₄⁻, HPO₄²⁻ and H₂PO₄⁻

Name of amphiprotic species	Chemical formula	Able to donate a proton, H ⁺	Able to accept a proton, H ⁺
Water	H ₂ O	H ₂ O(l) → H ⁺ (aq) + OH ⁻ (aq)	H ₂ O(l) + H ⁺ (aq) → H ₃ O ⁺ (aq)
Hydrogen carbonate ion	HCO ₃ ⁻	HCO ₃ ⁻ (aq) → H ⁺ (aq) + CO ₃ ²⁻ (aq)	HCO ₃ ⁻ (aq) + H ⁺ (aq) → H ₂ CO ₃ (aq)
Hydrogen sulfate ion	HSO ₄ ⁻	HSO ₄ ⁻ (aq) → H ⁺ (aq) + SO ₄ ²⁻ (aq)	HSO ₄ ⁻ (aq) + H ⁺ (aq) → H ₂ SO ₄ (aq)
Dihydrogen phosphate ion	H ₂ PO ₄ ⁻	H ₂ PO ₄ ⁻ (aq) → H ⁺ (aq) + HPO ₄ ²⁻ (aq)	H ₂ PO ₄ ⁻ (aq) + H ⁺ (aq) → H ₃ PO ₄ (aq)

Understand the difference between Strong and Weak acids and bases

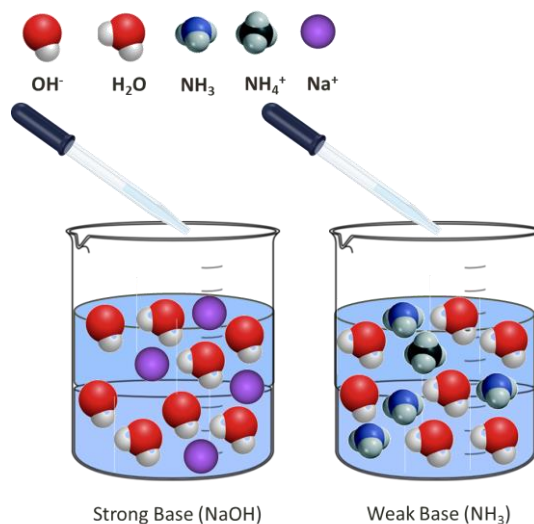
You can define acids and bases as being "strong" or "weak". Strong acids are compounds that completely dissociate (break up) in water. All of the H⁺ ions (protons) break away from the original acid molecule in water. A weak acid only partially dissociates and loses just some of its H⁺ ions (protons)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



Strong and Weak bases

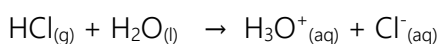
You can define bases as being "strong" or "weak". Strong bases are compounds where each molecule will accept an H⁺ ion. A weak base is a compound where only some of the molecules will accept an H⁺ ion. Most weak base molecules remain unreacted.

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



Strong acids

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



HCl gas dissolved in water

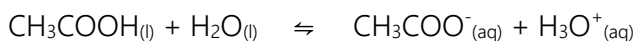
HCl has donated an H⁺ so is acting as an acid

H₂O has accepted an H⁺ so it is acting as a base

Solution contains virtually no intact HCl molecules after reaction.

Weak acids

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

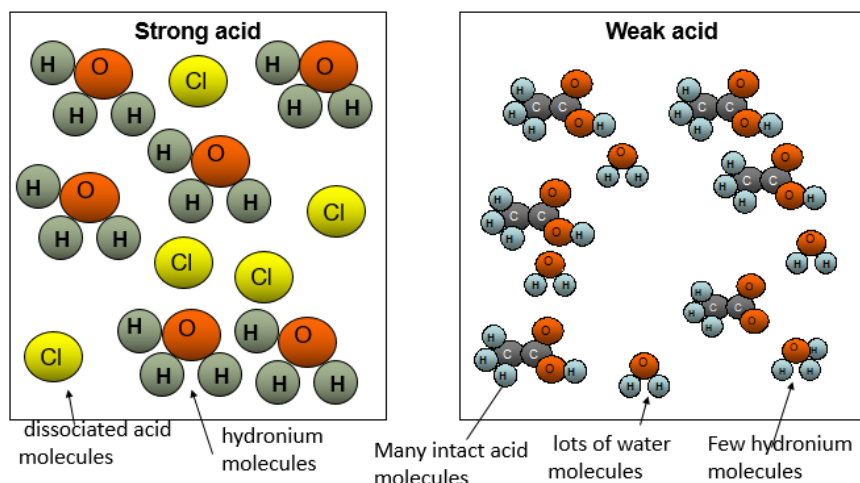


CH_3COOH dissolved in water

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

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

Solution contains mostly intact CH_3COOH molecules.



Strong and Weak Bases

The strength of a 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 , CH_3NH_2 and CH_3COO^-

Strong Bases

Completely accept protons (H^+) in aqueous solution



NaOH completely dissociates

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

Solution contains very few intact NaOH molecules after reaction.

Weak Bases

Partially accept protons (H^+) in aqueous solution



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.

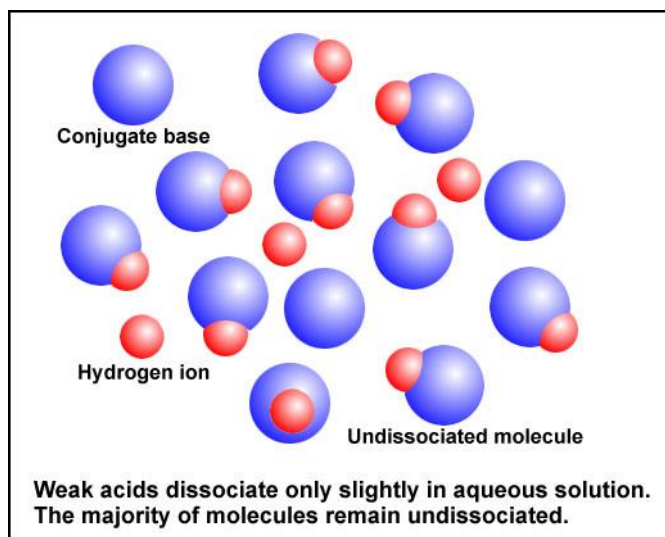
Solution contains mostly intact NH_3 molecules.

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. Moreover, 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 NH_3 is the conjugate base of NH_4^+ .

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

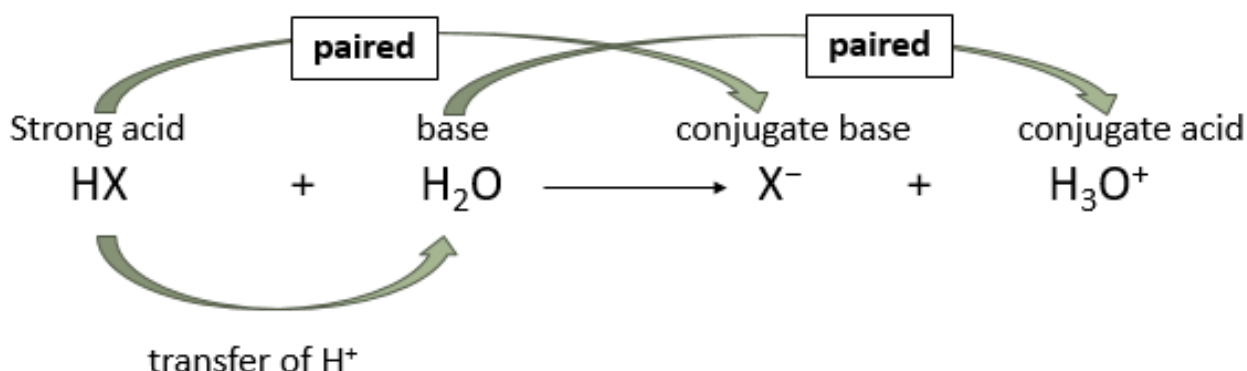
The stronger an acid, the weaker its conjugate base, and, conversely, the stronger a base, the weaker its conjugate acid.

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

Conjugate Acid and Base pairs (Strong Acid)

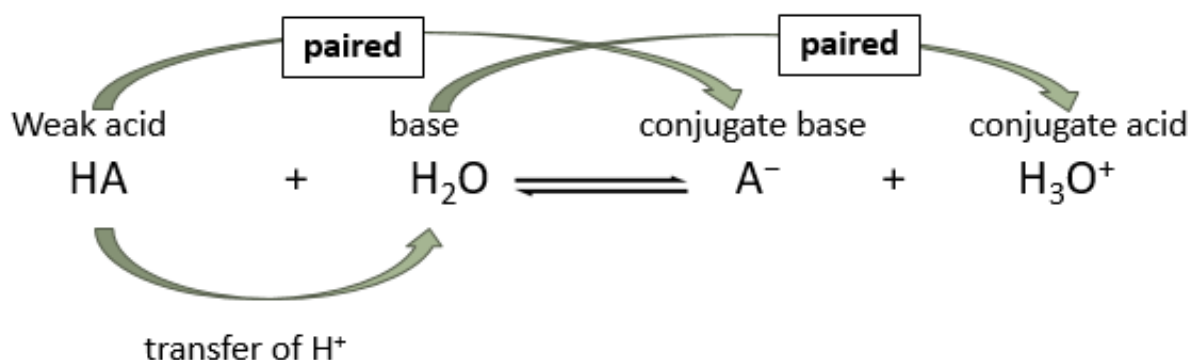
HX is a symbol used for a strong acid. A conjugate acid can be seen as the chemical substance that releases a proton in the backward chemical reaction.

The base produced, X⁻, is called the conjugate base and it absorbs a proton in the backward chemical reaction.



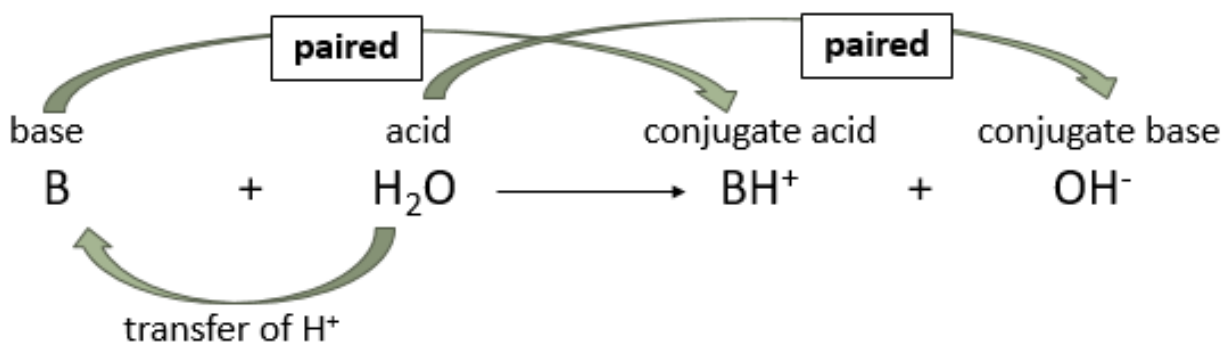
Conjugate Acid and Base pairs (weak acid)

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



Conjugate Acid and Base pairs (Base)

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



BASE - Conjugate Acid and Base pairs

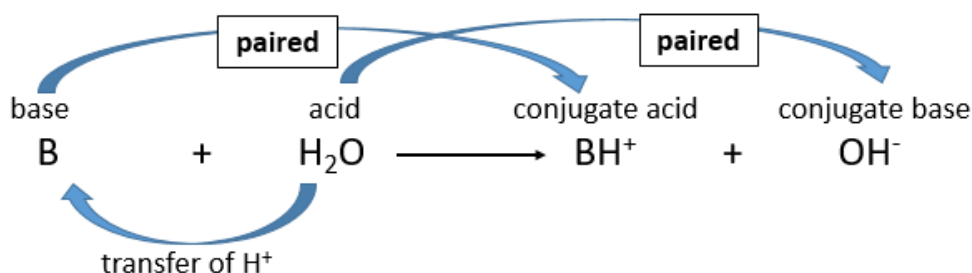
If 2 species differ by just 1 proton they are classed as a conjugate acid-base pair.

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 NH_3 is the conjugate base of NH_4^+ .

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

Conjugate Acid and Base pairs (Base)

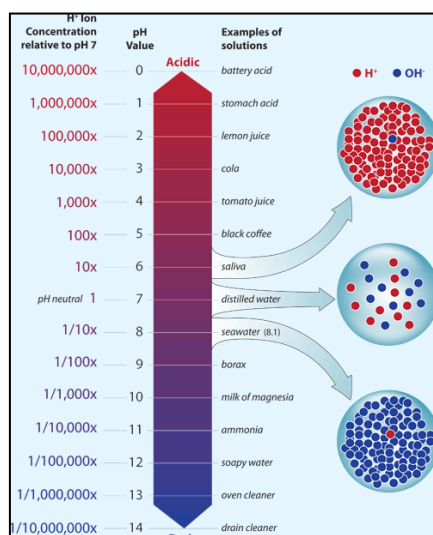
B is a symbol used for a base. The base now accepts the hydrogen ion from the water. The hydroxide ion, OH^- , is the paired conjugate of the water once the H^+ has been removed. Strong bases use a single direction arrow and weak bases use a double arrow.



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.

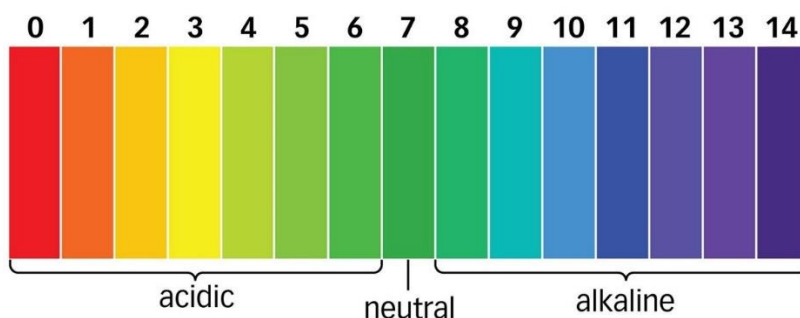


Describe solutions as acidic, alkaline or neutral in terms of the pH scale.

Acids have a pH less than 7

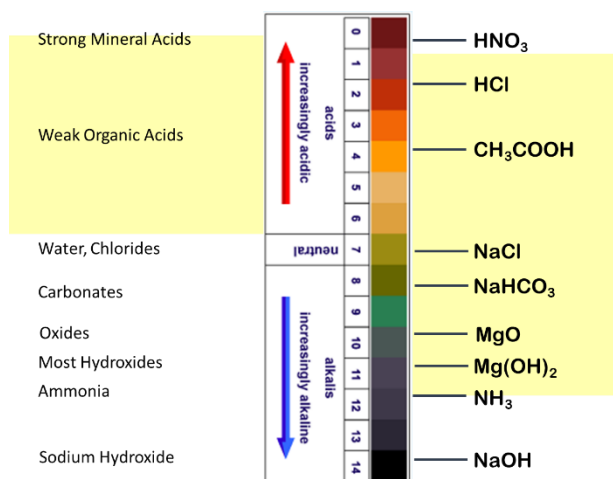
Neutral substances have a pH of 7

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



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

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

When a strong acid dissociates into its conjugate and H_3O^+ ions then 100% of its molecules will dissociate. The concentration of H_3O^+ ions produced depends upon the Strong Acid involved.

HCl and HNO_3 acid concentration produces the same concentration of H_3O^+ ions i.e. 0.123molL^{-1} of HCl will produce 0.123molL^{-1} $[\text{H}_3\text{O}^+]$ 1:1 ratio as those acids have one H^+ ion per molecule to donate

Note: 1 mol of H_2SO_4 acid produces 1 mol of H^+ ions and 1 mol of HSO_4^- ions. The HSO_4^- acts as a weak acid and further dissociation only produces a small amount of H^+ ions. However, at the completion of the reaction with NaOH the ratio will be 2:1, acid:base

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

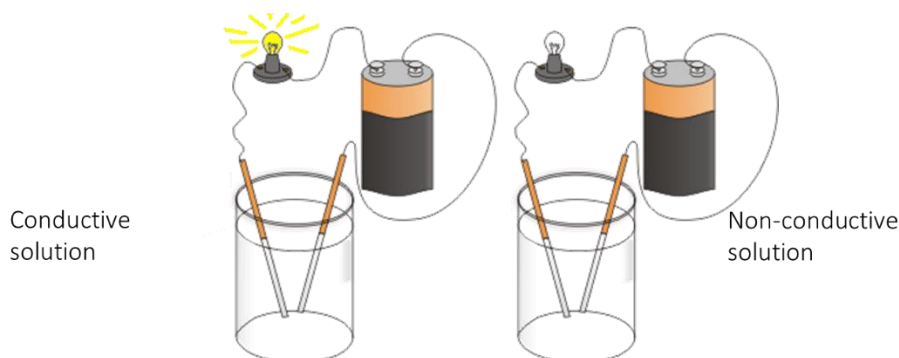
Conductivity of Acids and Bases

Conductivity is related to the availability of free moving charged particles.

The presence of H_3O^+ or OH^- ions in solution and the concentration of them determine conductivity.

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 H_3O^+ or OH^- ions carry the charge



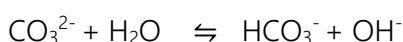
Low Conductivity – Weak acids and Bases

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

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; Na^+ and CO_3^{2-} . This means that even though a carbonate is a weak base only dissociating to release a small proportion of OH^- ions / conjugate acid, it will still be a good conductor because of the Na^+ and CO_3^{2-} ions in solution. These are free moving ions that carry charge.

When drawing equations remember to show Na_2CO_3 dissolving into the two ions first before 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_2SO_4 (sulphuric acid) HNO_3 (nitric acid) NaOH (sodium hydroxide) NaCO_3 (sodium carbonate) Na_2O (sodium oxide)	NH_3 (ammonia) CH_3COOH (ethanoic acid)

Species in solution

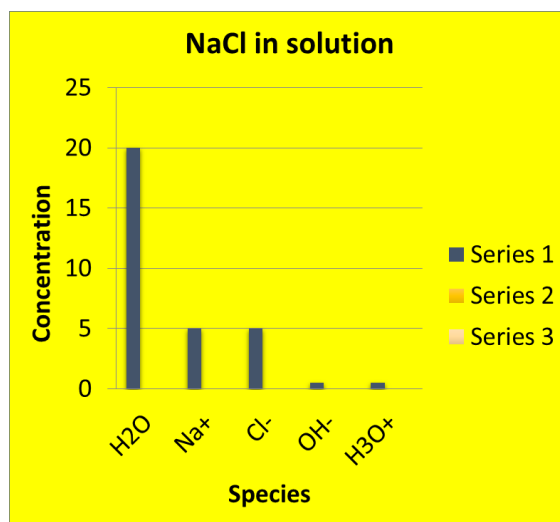
A solution is formed by mixing a solute (a dissolved substance) into a solvent (the solution that dissolves the solute). The solvent is water, and the solute can be an acid, base or ionic salt.

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 H_3O^+ and 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

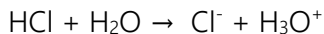
The relative concentration of the species in solution at equilibrium will depend upon the type of substances dissolved into water initially.

- In aqueous solutions, water will almost always be present in the highest concentration.
- Small quantities of H_3O^+ and OH^- 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

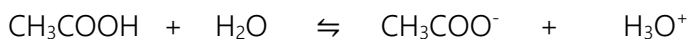


Strong acids will provide good conductivity and pH 1-2 due to the high presence of H_3O^+ ions

No strong acid will be left in the final mixture. H_3O^+ and Cl^- are produced in equal concentrations – in the same concentration as the original strong acid. A small amount of OH^- is present as water dissociates into H_3O^+ and OH^-

Concentration of ions in solution – Weak Acid

Weak Acid i.e. CH_3COOH reacting with water



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

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

Concentration of ions in solution – Strong Base

Strong Base i.e. NaOH reacting with water

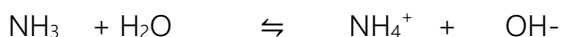


Strong bases will provide good conductivity and pH 12 – 14 due to the high presence of OH^- ions and Na^+ ions

No strong base will be left in the final mixture. OH^- and Na^+ are produced in equal concentrations – in the same concentration as the original strong base.

Concentration of ions in solution – Weak Base

Weak Base in water i.e. NH_3



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

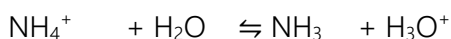
Most weak base will be left in the final mixture. OH^- and NH_4^+ are produced in equal concentrations, only a small amount of the weak base had dissociated.

pH of salt solutions

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

Concentration of ions in solution – acid salt

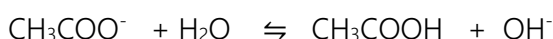
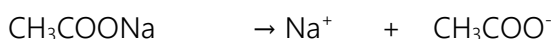
Acid Salt i.e. NH_4Cl



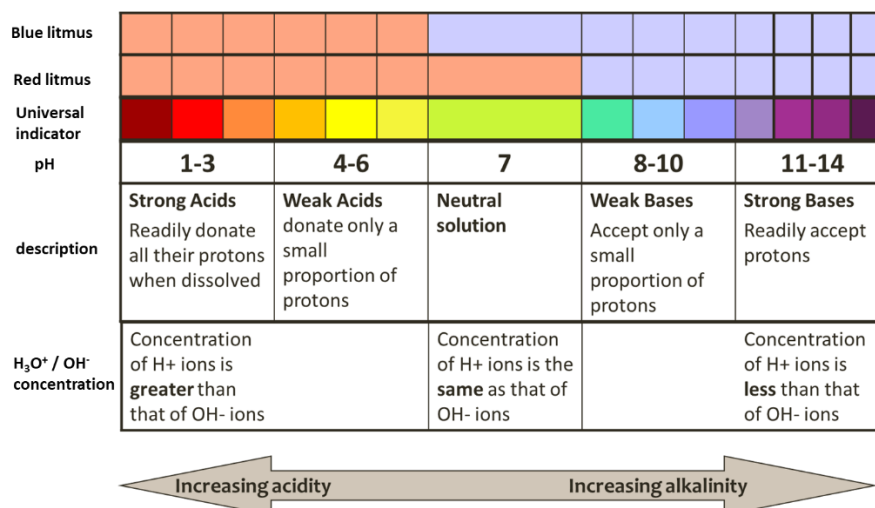
Acid salts will provide good conductivity and $\text{pH} < 7$ due to the high presence of ions from dissolving and to a lesser extent H_3O^+ ions. The spectator ion will be left in the highest concentration followed by the weak acid. H_3O^+ and NH_3 are produced in equal concentrations – a small amount of the weak acid had dissociated.

Concentration of ions in solution – base salt

Base Salt i.e. CH_3COONa

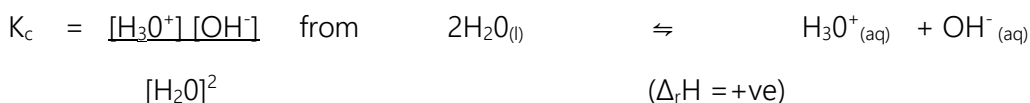


Base salts will provide good conductivity and $\text{pH} > 7$ due to the high presence of ions from dissolving and to a lesser extent OH^- ions. The spectator ion will be left in the highest concentration followed by the weak base. OH^- and CH_3COOH are produced in equal concentrations – a small amount of the weak base had dissociated.



K_w – 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 1×10^{-14} then so does K_w

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

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

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

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

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

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

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

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

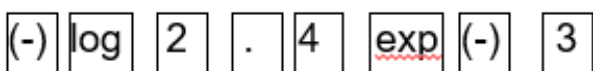
$$pH = -\log [H_3O^+]$$

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

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

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

Buttons on calculator



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

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

Note: the concentration of monoprotic ions, such as HCl and HNO_3 that donate 1 hydrogen ion can be assumed to equal the $[H_3O^+]$ as they fully dissociate. H_2SO_4 will donate 2 hydrogen ions so the concentration of $[H_3O^+]$ will be twice that of the original starting concentration of acid.

The starting concentrations of strong bases such as NaOH can be assumed to be the same as $[OH^-]$. These assumptions only apply for strong acids and bases.

Calculating $[\text{OH}^-]$ 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 $[\text{H}^+] = 10^{-2} \text{ mol L}^{-1}$

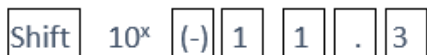
calculate $[\text{OH}^-]$ if solution has pH=2 $[\text{OH}^-] = 10^{(14-2)}$

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

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



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 $[\text{OH}^-]$

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

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 dissociate 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).