

FACTORS AFFECTING THE RATE OF A CHEMICAL REACTION

TEMPERATURE

- THE HIGHER THE TEMPERATURE THE FASTER THE RATE OF REACTION
 - PARTICLES ARE MOVING FASTER - HAVE GREATER KINETIC ENERGY.
 - SO THERE ARE MORE COLLISIONS PER SECOND
- AND**
- MORE OF THE PARTICLES HAVE ENERGY GREATER OR EQUAL TO E_a AND SO MORE OF THE COLLISIONS THAT ARE OCCURRING LEAD TO A SUCCESSFUL REACTION.

When explaining temperature you need BOTH of these important ideas!!

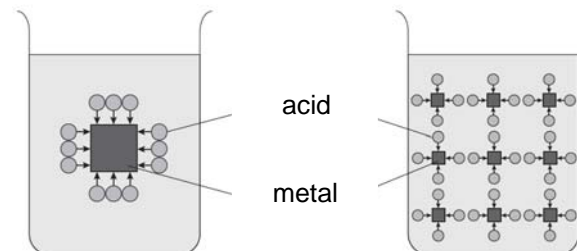
Measuring rates of reaction.

- Collect gas in a syringe or water filled measuring cylinder: record volume of gas at time intervals
- Measure loss of mass over time eg CaCO_3 & acid in a flask standing on a digital balance
- Time how long for a coloured chemical to appear or disappear

RATES OF REACTION

SURFACE AREA

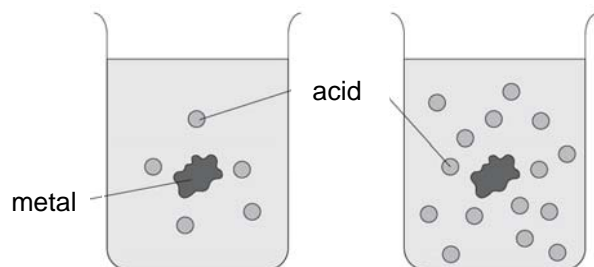
Consider a solid substance in a solution, e.g. magnesium metal in hydrochloric acid. When the solid is broken up, more ACID particles can collide with each piece in the same time.



Increase in collision frequency.....

CONCENTRATION

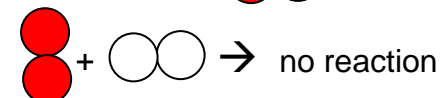
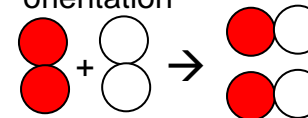
The higher the concentration of a reactant (or reactants), the more particles per unit volume will collide per second.



Collision theory:

For a reaction to occur

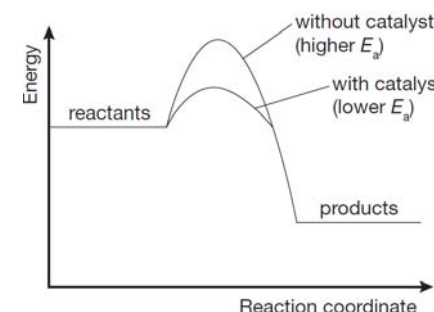
- The reactants must collide!!!!
- They must collide with the correct orientation



- They must collide with energy greater or equal to E_a (the activation energy) for that reaction

Anything that brings about more effective collisions per unit time will increase the rate of the reaction. "effective" – leading to a reaction. "collisions per unit time"? This just means how FREQUENTLY the particles collide, eg the number of collisions per second.

Catalysts Provide an alternative pathway for a reaction with a lower E_a (so more molecules have the required energy to react and so the rate is faster). Are NOT used up in the process. Don't alter ΔH for a reaction.



... EQUILIBRIUM ...

LE CHATELIER'S PRINCIPLE

"When a change is applied to a system in dynamic equilibrium, the system reacts in such a way as to oppose the effect of the change."

When a chemical equilibrium is REACHED:

- reaction is dynamic - it is moving forwards & backwards
- the rates of forward & backward reactions are equal
- both reactants & products are present
- concentrations of reactants & products remain constant

Equilibrium can only occur in a closed system, one in which nothing can be added or removed. For gases, the container *must* be sealed.

The equilibrium sign \rightleftharpoons does *not* imply equal proportions of reactants & products, but that rates of forward & reverse reactions are equal.

K_c is known as the Equilibrium Constant

Consider: $aA + bB \rightleftharpoons cC + dD$ eg $N_2 + 3H_2 \rightleftharpoons 2NH_3$

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

Product(s)
is/are ON
TOP!!!

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

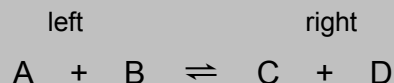
[] denotes the equilibrium concentration in mol L⁻¹
K_c

- is **AFFECTED** by a change of temperature
- but is **NOT AFFECTED** by changes in concentration, a change of pressure OR adding a catalyst.

Altering the temperature affects the rate of both backward & forward reactions but to different extents. The equilibrium moves producing a new equilibrium constant. The direction of movement depends on the sign of ΔH .

CONCENTRATION

K_c is not affected by changes in any concentration (at constant temperature). To maintain the constant the composition of the equilibrium mixture changes



- increase [A] or [B] equilibrium moves to right
- increase [C] or [D] equilibrium moves to left
- decrease [A] or [B] equilibrium moves to left
- decrease [C] or [D] equilibrium moves to right

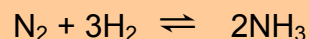
CATALYSTS

Do not affect the position of equilibrium but equilibrium is reached quicker.

PRESSURE

Increase pressure (decrease volume) – equilibrium moves to side with **FEWER GASEOUS MOLECULES**

Decrease pressure (increase volume) – equilibrium moves to side with **MORE GASEOUS MOLECULES**

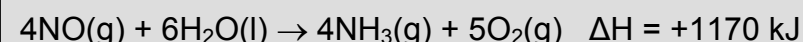


Increase in pressure favours NH₃ production (4 mol of gas → 2 mol of gas will decrease pressure of system).
No change will occur when equal numbers of gaseous molecules appear on both sides.

TEMPERATURE

An increase in temperature favours the reaction that absorbs heat (endothermic reaction). A decrease in temperature favours the reaction that releases heat (exothermic reaction).

Look @ a thermochemical equation eg

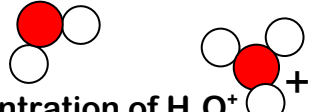

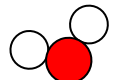

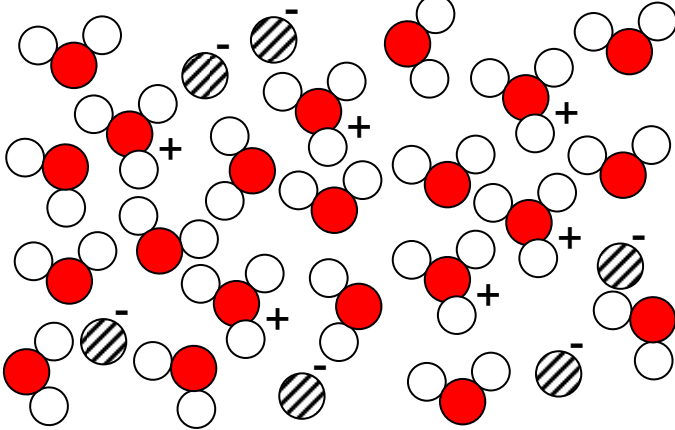
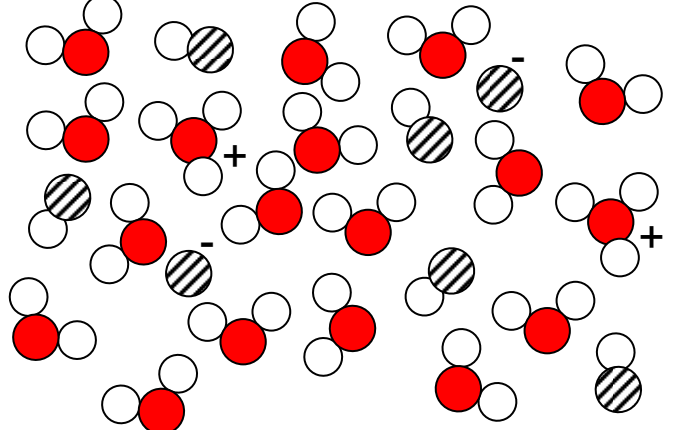


The reaction producing NH₃ and O₂ is endothermic and so it would be favoured by an increase in temperature.

K_c is affected by temperature. If K_c increases (bigger products ÷ smaller reactants) when the temperature is increased then the forward reaction is endothermic. + ΔH .

Determining the position of equilibrium from equilibrium constants

Value of K	Reactants or products favoured
< 0.001	reactant
between 0.001 and 1000	mixed, comparable amounts of reactants & product
> 1000	product

<h2>ACIDS & BASES</h2> <p>Brønsted-Lowry definitions</p> <p>Acids are proton donors</p> <p>Bases are proton acceptors</p> <p>Acid-base conjugate pairs differ by a proton eg HA / A^-, $\text{H}_3\text{O}^+ / \text{H}_2\text{O}$, $\text{NH}_4^+ / \text{NH}_3$, $\text{H}_2\text{O} / \text{OH}^-$, $\text{HCO}_3^- / \text{CO}_3^{2-}$</p> <p>Species that can act as acid or base are called amphiprotic eg H_2O and HCO_3^-</p>	<p>pH</p> <p>$\text{pH} = -\log [\text{H}_3\text{O}^+]$ (often write $\text{pH} = -\log [\text{H}^+]$)</p> <p>$[\text{H}_3\text{O}^+] = \text{invlog} (-\text{pH})$</p> <p>[] = concentration in mol L^{-1}</p> <p>$K_w = \text{ionic product of water} = 1 \times 10^{-14}$</p> <p>$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$ (or $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$)</p> <p>$[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$</p> <p>NB when entering 2.34×10^{-5} into calculator, do 2.34 EXP -5 (don't x 10 as well)</p>	<p>pH of some salt solutions</p> <p>NH_4Cl, ammonium chloride, dissolved in water has a $\text{pH} < 7$. It dissolves $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$. The NH_4^+ ion then reacts with water</p> <p>$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$: H_3O^+ makes $\text{pH} < 7$</p> <p>CH_3COONa, sodium ethanoate, dissolved in water has a $\text{pH} > 7$. $\text{CH}_3\text{COONa} \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+$. The CH_3COO^- then reacts with water.</p> <p>$\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$; OH^- makes it alkaline.</p>
<p>Strong and weak acids and bases</p> <p>Strong – fully ionised / dissociated OR react completely with water; use \rightarrow in equations eg $\text{HA} + \text{aq} \rightarrow \text{H}^+ + \text{A}^-$</p> <p>Weak – partially ionised or dissociated OR incomplete reaction with water; use \rightleftharpoons in equations eg $\text{HA} + \text{aq} \rightleftharpoons \text{H}^+ + \text{A}^-$</p>	<p>Strong Acids</p> <p>Fully ionised in water </p> <p>Low pH as high concentration of H_3O^+</p> <p>React faster than same concentration of weak acid as more H_3O^+ available for collision at any one time.</p> <p>Good electrical conductors as a high conc. of ions (BOTH H^+ and A^-) to carry current</p> <p>In diagram below is <u>no</u> HA</p> <p>– is 100% ionised </p>	<p>Weak Acids</p> <p>Partially ionised in water </p> <p>Not such a low pH as lower concentration of H_3O^+ than the strong acid</p> <p>React slower than same concentration of strong acid as less H_3O^+ available for collision at any one time.</p> <p>Poor electrical conductors as very low conc. of ions (H^+ and A^-) to carry current</p> <p>In diagram below it is mainly HA </p>
<p>Stuff you should know from level 1...</p> <p>$\text{pH} < 7$ = acid, $\text{pH} 7$ = neutral, $\text{pH} > 7$ = basic or alkaline. An alkali is a soluble base.</p> <p>Litmus: Red litmus: stays red in acid, turns blue in alkali. Blue litmus: turned red in acid, stays blue in alkali</p> <p>Universal Indicator: Red – yellow (acid), green (neutral), blue-purple (alkali)</p> <ul style="list-style-type: none"> ○ Acid + metal \rightarrow salt + hydrogen ○ Acid + base \rightarrow salt + water ○ Acid + carbonate \rightarrow salt + water + carbon dioxide 		

pH in (much) more detail...

pH calculations

Remember that [] = concentration, and has the units mol L⁻¹. pH of course has NO units!

Turning H⁺ into pH

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

Eg 1. what is the pH if $[\text{H}^+] = 3.45 \times 10^{-3} \text{ mol L}^{-1}$

- $\text{pH} = -\log 3.45 \times 10^{-3}$
- Enter as: (-) log 3 . 4 5 EXP (-) 3, and then EXE

Turning pH into H⁺

$[\text{H}_3\text{O}^+] = \text{invlog} (-\text{pH})$ (OR simply $[\text{H}^+] = \text{invlog} (-\text{pH})$)

Eg 2. what is the $[\text{H}^+]$ if the pH is 4.55

- $[\text{H}^+] = \text{invlog} (-\text{pH})$
- Enter as: SHIFT LOG (-) 4 . 5 5, and then EXE

K_w, the ionic product for water

Water self ionises to a very small extent - $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$

(or $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$)

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

$K_w = 1 \times 10^{-14}$ In the exam you will be told **$K_w = 10^{-14}$**

This is useful when you are asked to calculate $[\text{OH}^-]$ when you know the $[\text{H}^+]$ or $[\text{H}^+]$ when you know the $[\text{OH}^-]$

$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$, so $[\text{H}_3\text{O}^+] = K_w / [\text{OH}^-]$, and of course $[\text{OH}^-] = K_w / [\text{H}_3\text{O}^+]$

Eg 3. what is the $[\text{OH}^-]$ when $[\text{H}_3\text{O}^+]$ is 1.76×10^{-8}

- $[\text{OH}^-] = K_w / [\text{H}_3\text{O}^+]$
- Enter as: 1 EXP (-) 14 ÷ 1 . 7 6 EXP (-) 8, and then EXE



But what about $[\text{OH}^-]$ to pH? Or pH to $[\text{OH}^-]$?

You need to do it in 2 steps. There are TWO different ways! They (should) give you the same answer! You only need to know one way.

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

Eg 4. what is the pH if $[\text{OH}^-] = 6.35 \times 10^{-3} \text{ mol L}^{-1}$

Method 1.

Find $[\text{H}^+]$, then find pH

- $[\text{H}^+] = K_w / [\text{OH}^-]$
- $\text{pH} = -\log [\text{H}^+]$

Method 2.

Find pOH, then find pH

- $\text{pOH} = -\log [\text{OH}^-]$
- $\text{pH} = 14 - \text{pOH}$

Don't round numbers along the way.... Keep all in calculator until the final answer and then write it to 3 s.f. (if requested in the question)

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

This is merely the reverse. Either

- first calculate $[\text{H}^+]$, then $[\text{OH}^-]$, OR
- first calculate pOH, then $[\text{OH}^-]$

Answers to examples

1. 2.46
2. $2.82 \times 10^{-5} \text{ mol L}^{-1}$
3. $5.68 \times 10^{-7} \text{ mol L}^{-1}$
4. 11.8