# Combined Science - Chemistry

## CC13-15 Knowledge organiser

## C13 to C15: Groups, rates and heat changes

### Lesson sequence

1. Group 1
2. Group 7
3. Reactivity of halogens
4. Group 0
5. Rates of reaction
6. Collision theory
7. Core practical – rates of reaction (CP11)
8. Catalysts
9. Exothermic and endothermic reactions
10. Explaining energy changes

### 1. Group 1

| | |
|---|---|
| **Alkali metals** | The name of the metals in group 1 – lithium, sodium, potassium etc. |
| **Group 1 symbols** | Li – lithium<br>Na – sodium<br>K – potassium |
| **Reaction of alkali metals with water** | Metal + water → metal hydroxide + hydrogen<br>E.g: sodium + water → sodium hydroxide + hydrogen<br>$2Na + 2H_2O \rightarrow 2NaOH + H_2$ |
| **Lithium and water** | Lithium floats and bubble vigorously |
| **Sodium and water** | Sodium melts into a ball and moves around the surface bubbling vigorously. |
| **Potassium and water** | Potassium melts into a ball, catches fire (lilac) and moves around the surface bubbling vigorously. |
| **Group 1 reactivity** | Reactivity increases as you move down the group. |
| **Explaining group 1 reactivity** | When metals react they lose their outer electrons. Further down the group there are more shells of electrons so the outer electrons are less attracted to the nucleus and easier to remove. |

### 2. Group 7

| | |
|---|---|
| **Halogens** | The names given to the non-metals in group 7 – fluorine, chlorine, bromine and iodine. |
| **Chlorine** | $Cl_2$ - A pale green gas. |
| **Bromine** | $Br_2$ - A red-brown liquid. |
| **Iodine** | $I_2$ - A shiny purple-black solid. |
| **Reaction of halogens with metals** | Halogen + metal → metal halide<br>E.g:<br>Bromine + sodium → sodium bromide<br>$Br_2 + 2Na \rightarrow 2NaBr$ |
| **Reaction of halogens with hydrogen** | Halogen + hydrogen → hydrogen halide<br>E.g: Chlorine + hydrogen → hydrogen chloride<br>$Cl_2 + H_2 \rightarrow 2HCl$ |
| **Hydrogen halides** | Hydrogen halides dissolve in water to form acids, for example hydrogen chloride makes hydrochloric acid. |
| **Chlorine test** | Chlorine gas turns damp blue litmus red then quickly bleaches it white. |

### 3. Reactivity of halogens

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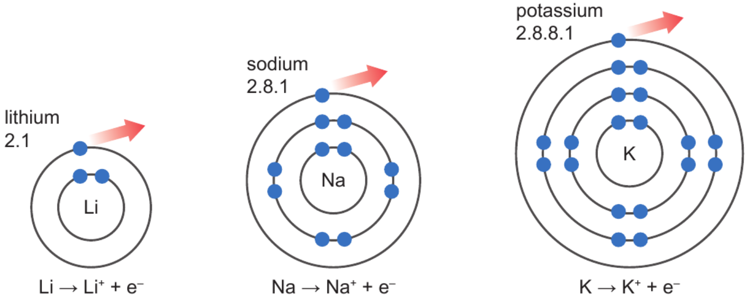
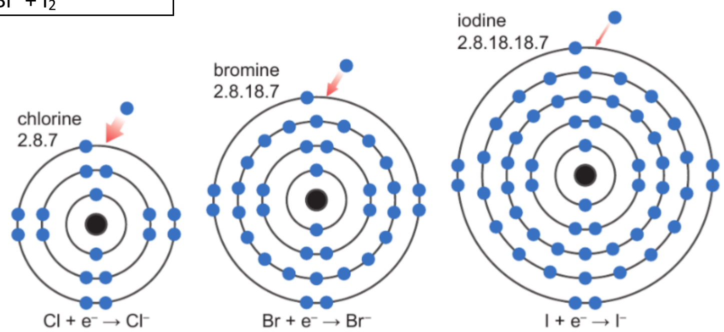
|---|---|
| **Group 7 reactivity** | Reactivity increases as you go up the group. |
| **Explaining group 7 reactivity** | When non-metals react they complete their outer shells. Further up the group the elements have fewer shells so the nucleus attracts electrons more strongly. |
| **Displacement reactions** | Reactions in which a more reactive metal displaces a less reactive metal from a salt eg:<br>*copper sulfate + zinc → zinc sulfate + copper*<br>Does not work backwards as copper is less reactive than zinc. |
| **Displacement reactions of halogens** | A more reactive halogen displaces a less reactive halide ion by taking its electrons.<br>E.g: bromine + sodium iodide → iodine + sodium bromide |
| **Redox reactions of halogens** | The more reactive halogen oxidises the less reactive halide by taking its electrons. The more reactive halogen is reduced.<br>E.g: $Br_2 + 2I^- \rightarrow 2Br^- + I_2$ |

### 4. Group 0

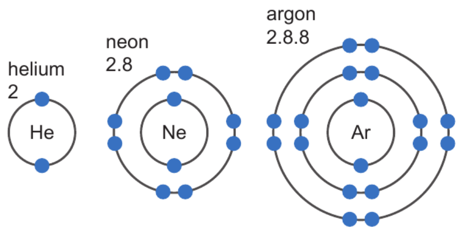
| | |
|---|---|
| **Noble gases** | The name given to the non-metals in group 0 – helium, neon, argon, krypton and xenon. |
| **Melting point of noble gases** | They are all gases at room temperature but the melting and boiling point increase down the group. |
| **Reactivity of group 0** | The noble gases do not (easily) do any reactions – they are inert. |
| **Explaining reactivity of group 0** | When elements react they try to complete their outer shells. Because group 0's outer shells are already complete, they do not react. |
| **Uses of noble gases** | -Helium is used in airships because it is inert and has low density<br>- Argon is used in fire extinguishers because it is inert and denser than air.<br>- Neon is used in lighting because it glows red when electricity is passed through it. |




**D** As the distance between the outer electron and the nucleus increases, the alkali metals get more reactive.




**D** Going down group 7, the outermost electron shell gets further from the nucleus and the ions are less readily formed.




**F** Noble gases do not react as they already have a complete outer shell of electrons.

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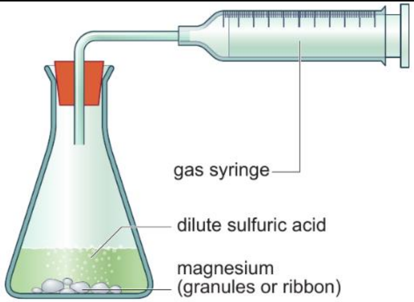
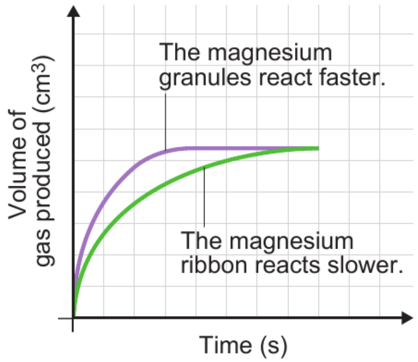
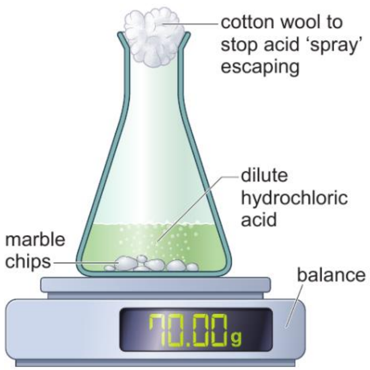
| 5. Rates of reaction | |
|---|---|
| **Rate of reaction** | The rate at which reactants are used up or products are made. |
| **Reactants vs time graph** | Starts high and curves downward, decreasing rapidly at first and then more gently. Steeper line = faster rate. |
| **Products vs time graph** | Starts low and curves upwards, increasing rapidly at first and then more gently. Steeper line = faster rate. |
| **Measuring rates – reactions that produce gas** | - Collect gas in a gas syringe and measure the volume every 30 secs.<br>- Collect gas over water (up-turned measuring cylinder full of water) and measure volume every 30 secs.<br>- Do reaction on a balance and record the change in mass every 30 secs. |
| **Measuring rates – reactions that go cloudy** | Do the reaction in a beaker placed on piece of paper with a cross marked on it. Looking down through the beaker, time how it takes for the cross to disappear. |



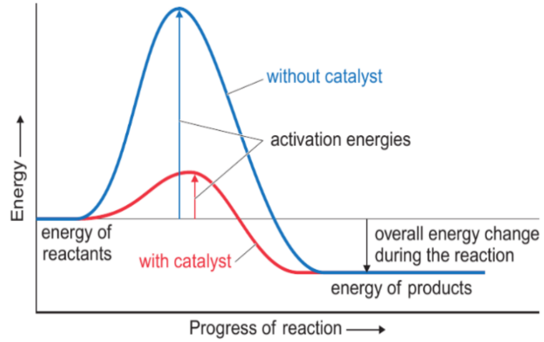

| 6. Collision theory | |
|---|---|
| **Collision theory** | States that for two particles to react they must:<br>- Collide with each other<br>- Collide with enough energy to react |
| **Activation energy** | The minimum energy that two particles must have when they collide in order to react. |
| **Effect of concentration on rate** | Increasing the concentration increases the rate because there are more particles so there are more collisions and more reactions. |
| **Effect of surface area on rate** | Increasing the surface area (by decreasing particle sizes) in creases the rate by exposing more particles to collisions leading to more collisions and more reactions. |
| **Effect of pressure on rate** | Increasing the pressure increases the rate because particles are pushed closer together so they collide more often. |
| **Effect of temperature on rate** | Increasing the temperature increases the rate because particles move faster so they collide more, and collide with more energy to a greater proportion of collisions lead to reactions. |




**D** As the reaction proceeds, the mass of the flask and contents will decrease.

| 7. Core practical – rates of reaction (CP11) | |
|---|---|
| **CP11 – Aim** | To explore the rate of two reactions by collecting gas and observing a colour change. |
| **CP11 – Gas collection – setup** | Place a measuring cylinder full of water upside down in a basin of water. Place 5 g of marble chips in a conical flask with 40 $cm^3$ hydrochloric acid. Insert a bung with delivery tube and insert the delivery tube into the measuring cylinder. |
| **CP11 – Gas collection – measurements** | Record the volume of gas collected every 15 seconds until it stops. |
| **CP11 – Gas collection – variations** | Repeat with a different size of marble chips. |
| **CP11 – Gas collection – results** | The amount of gas collected increases quickly at first and then more slowly. The smaller marble chips produce gas more quickly, but the same amount in total. |
| **CP11 – Colour change – setup** | Draw a cross on a piece of paper and place a beaker on it. Measure out 50 $cm^3$ of sodium thiosulfate solution and 5 $cm^3$ of hydrochloric acid into two test tubes and leave to warm in a water bath at 30°C. |
| **CP11 – Colour change – run the experiment** | Quickly pour both test tubes into the beaker, mix and start the stopwatch. Looking down through the beaker, stop when you can no longer see the cross. |
| **CP11 – Colour change – variations** | Repeat with water baths set to 35°C, 40°C, 45°C and 50°C. |
| **CP11 – Colour change – results** | The cross disappears most quickly at 50°C and least quickly at 30°C. |

| 8. Catalyst | |
|---|---|
| **Catalyst** | A substance that speeds up a chemical reaction without being used up. |
| **Effect of catalysts on rate** | Catalysts increase the rate of reaction by reducing the activation energy so that a greater proportion of collisions lead to reactions. |
| **Reaction profile** | A graph that shows the changes in energy during a reaction. Starts with large 'hump' that represents the activation energy. |
| **Effect of catalysts on reaction profiles** | The 'hump' representing the activation energy is smaller. |
| **Enzyme** | A protein that works as a catalyst to speed up the reactions in our cells. |
| **Enzymes in alcohol production** | Alcoholic drinks are produced using enzymes found in yeast which catalyse a reaction that turns glucose into ethanol. |




**C** This reaction profile shows that a catalyst lowers the activation energy.

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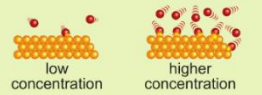
### Concentration and reaction rate

*Change*: Increasing the concentration of solutions increases the rate of reaction.

*Explanation*: There are more reacting particles in the same volume so collisions occur more often.

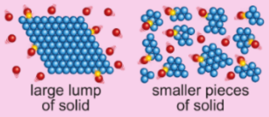



C

### Surface area and reaction rate

*Change*: Increasing the surface area to volume ratio, by decreasing the size of solid pieces while keeping the total volume of solid the same, increases the rate of the reaction.

*Explanation*: There is more surface for collisions to occur on, so collisions occur more often.

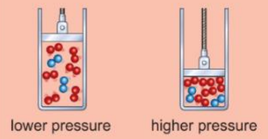



D

### Pressure of gases and reaction rate

*Change*: Increasing the pressure of gases increases the rate of reaction.

*Explanation*: The reactant particles are squeezed closer together so collisions occur more often.

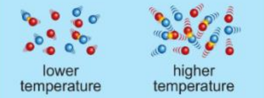



E

### Temperature and reaction rate

*Change*: Increasing the temperature increases the rate of reaction.

*Explanation*: The reactant particles speed up and have more energy. They therefore collide more often and more particles have enough energy to react when they collide.

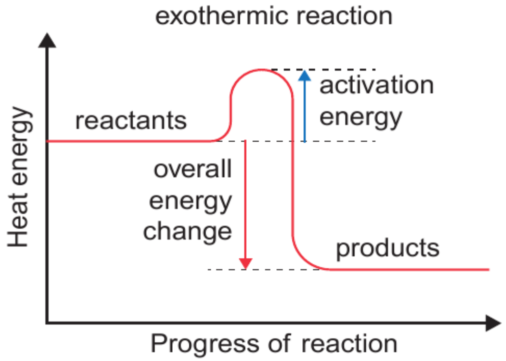


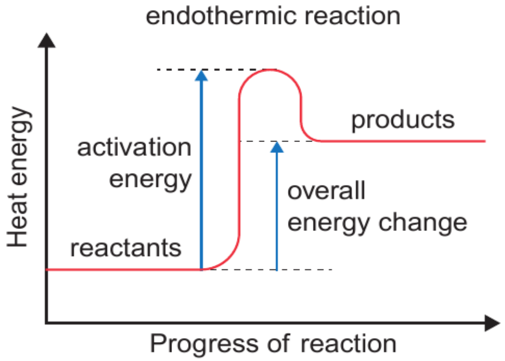

F

| 9. Endothermic and exothermic reactions | |
|---|---|
| **Exothermic reaction** | A reaction that transfers energy to the surroundings (gets hotter). |
| **Exothermic reaction examples** | - Neutralisation<br>- Displacement<br>- Combustion<br>- Some precipitation<br>- Respiration |
| **Endothermic reaction** | A reaction that absorbs energy from the surroundings (gets colder) |
| **Endothermic reaction examples** | - Dissolving (most) salts<br>- Some precipitation<br>- Photosynthesis |
| **Exothermic reaction profile** | The reactants have more energy than the products, so their line on the graph is higher. |
| **Endothermic reaction profile** | The reactants have less energy than the products, so their line on the graph is lower. |
| **Measuring energy changes** | -Sit a polystyrene beaker inside a glass beaker (insulation)<br>- Measure the starting temperature of the reactants.<br>- Mix the reactants in the polystyrene beaker<br>- Cover with lid fitted with a thermometer<br>- Monitor and record the lowest temperature. |

| 10. Explaining energy changes | |
|---|---|
| **Chemical bonds in reactions** | During chemical reactions, old chemical bonds are broken and new ones are formed. |
| **Breaking bonds** | Breaking bonds absorbs energy, breaking stronger bonds absorbs more energy. |
| **Making bonds** | Making bonds releases energy, making stronger bonds releases more energy. |
| **Energy changes and bond formation** | The energy change in a reaction is the difference between the energy required to break the old bonds and the energy released by making the new ones. |
| **Exothermic reactions and bonds** | Exothermic reactions break weaker bonds and make stronger ones. |
| **Endothermic reactions and bonds** | Endothermic reactions break stronger bonds and make weaker ones. |
| **Bond strength** | The energy required to break one mole of a particular covalent bond in kJ/mol. |
| **Calculating energy changes from bond strengths** | Add up the total strength of old bonds broken and subtract the total strength of new bonds made. A negative answer is exothermic. |

| Covalent bond | Bond energy (kJ $mol^{-1}$) |
|---|---|
| C–O | 358 |
| C–H | 413 |
| H–H | 436 |
| O–H | 464 |
| O=O | 498 |
| C=O | 805 |




**B** The activation energy is the difference in energy between the reactants and the top of the 'hump'.

**Energy change example:**

Hydrogen and chlorine react to form hydrogen chloride. The bond strengths are as follows: H-H = 436 kJ/mol, Cl-Cl = 240 kJ/mol , H-Cl = 428 kJ/mol. Calculate the energy change of the reaction

$$H\text{-}H + Cl\text{-}Cl \rightarrow 2(H\text{-}Cl)$$

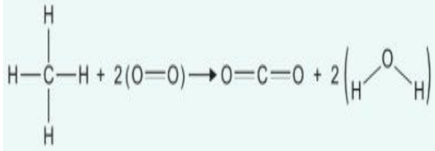
**Bonds broken** = 436 + 240 = 676

**Bonds made** = 2 x 428 = 856

Reaction energy = 676 – 856 = -180 kJ/mol, the reaction is exothermic because the answer is negative.

### Worked example

Methane burns completely in oxygen to form carbon dioxide and water:



D

Calculate the energy change during this reaction.

**Step 1 Calculate energy in (bonds broken)**

| | | |
|---|---|---|
| 4 × (C–H) | = 4 × 413 | = 1652 kJ $mol^{-1}$ |
| 2 × (O=O) | = 2 × 498 | = 996 kJ $mol^{-1}$ |
| Total in | = 1652 + 996 | = 2648 kJ $mol^{-1}$ |

**Step 2 Calculate energy out (bonds made)**

| | | |
|---|---|---|
| 2 × (C=O) | = 2 × 805 | = 1610 kJ $mol^{-1}$ |
| 4 × (O–H) | = 4 × 464 | = 1856 kJ $mol^{-1}$ |
| Total out | = 1610 + 1856 | = 3466 kJ $mol^{-1}$ |

**Step 3 Energy change = energy in – energy out**

= 2648 – 3466 = –818 kJ $mol^{-1}$

The negative sign shows that the reaction is exothermic (endothermic reactions have a positive sign).