The *ORIGINAL* Ultimate Chemistry O-Level Quick Revision Guide

“师傅领进门，修行在个人”

translated to: “While a teacher can guide you in your learning, it is eventually still your own effort and determination which decides the outcome”

Topics in O-Level Chemistry

	Tick off the table as you do your revision.	**1st rev**	**2nd rev**	**3rd rev**	**4th rev**	**5th rev**
1	Measurement and Experimental Techniques					
2	Purification and separation					
3	Kinetic particle theory					
4	Elements, compounds and mixtures					
5	Atomic structure					
6	Chemical Bonding and writing of equations					
7	Acids and bases					
8	Preparation of salts					
9	QA					
10	Mole concept					
11	Periodic table and periodicity					
12	Redox					
13	Metals					
14	Ammonia					
15	Energy changes					
16	Speed of reaction					
17	Air and atmosphere					
18	Electrolysis					
19	Organic chemistry					

Lecture Videos @ channel “Chem Lectures” on YouTube.

OR

Visit **www.chemlectures.sg** for Full O-Level Chemistry lecture videos AND comprehensive notes

1 Displacement of water

- For gases which are insoluble in water
- E.g., hydrogen, oxygen, carbon dioxide

2 Downward delivery

- For gases which are denser than air (and soluble in water)
- E.g., chlorine

3 Upward delivery

- For gases which are less dense than air (and soluble in water)
- E.g., hydrogen, ammonia

4 Using a gas syringe

- When there is a need to monitor the rate of evolution of a gas, i.e., how much gas was produced in the reaction within a certain duration

5 Concentrated sulfuric acid - used to dry acidic gases or gases which do not react with acids

6 Calcium oxide - used to dry alkaline gases, e.g, ammonia

7

Gas	Formula	Colour	Density (compared to air)	Solubility in water
Ammonia	NH_3	Colourless	Less dense	Extremely soluble
Carbon dioxide	CO_2	Colourless	Denser	Slightly soluble
Chlorine	Cl_2	Greenish- yellow	Denser	Soluble
Hydrogen	H_2	Colourless	Less dense	Insoluble
Hydrogen Chloride	HCl	Colourless	Denser	Very soluble
Oxygen	O_2	Colourless	Slightly denser	Slightly soluble
Carbon monoxide	CO	Colourless	Slightly denser	Slightly soluble
Nitrogen	N_2	Colourless	Slightly denser	Insoluble
Sulphur dioxide	SO_2	Colourless	Denser	Very soluble
Helium	He	Colourless	Less dense	Insoluble

8 A pure substance is a single substance not mixed with anything else, e.g., pure gold, pure silver, distilled water.

9 A mixture consists of two or more substances that are not chemically combined together, e.g., milk, Milo, Ribena, orange juice, sea water.

10 Decanting separates an insoluble solid from a liquid by pouring off the liquid from the container carefully.

11 Filtration is used to separate an insoluble solid from a liquid.

- The solid is collected in the filter paper ⇒ residue
- The liquid that passes through the filter paper and collected in the beaker ⇒ filtrate

12 Evaporation is used to separate dissolved solids from a solution.

13 Crystallisation separates a dissolved solid from a solution, forming pure crystals of the substance.

14 Sublimation is used to separate a mixture of solids, one of which sublimes. Sublimation is the process of changing from a solid to vapour on heating without going through the liquid state.

15 Simple distillation separates a pure liquid from a solution.

- The thermometer is placed near the exit of the distillation flask. This allows it to measure the temperature of the vapour which will enter the condenser and be condensed and collected as distillate.

16	Fractional distillation separates mixtures of miscible liquids with different boiling points. • Miscible liquids are completely soluble in each other to form one liquid. • The glass beads in the fractionating column provides a large surface area to volume ratio for cooling of the vapours. This helps to ensure that only one liquid of the lower boiling point distils over. This allows better separation of the two or more miscible liquids in the round-bottomed flask. • The thermometer is placed near the exit of the fractionating column. This allows it to measure the temperature of the vapour which will enter the condenser and be condensed and collected as distillate.
17	Chromatography is a method of separating and identifying mixtures. Only a very small initial sample is required to generate the chromatogram. **(1)** The dye to be chromatographed is dissolved in a suitable solvent. **(2)** A drop of the dissolved dye is placed on the pencil line near the bottom of a strip of filter paper (or chromatography paper can be used). The pencil line shows the starting position of the drop of dissolved dye. This is important in determining the R_f value. **(3)** The filter paper is dipped into a tube containing a suitable solvent, making sure the solvent level is below the spot. **(4)** The solvent travels up the filter paper and the dye in the spot dissolves and travel up the paper at different speeds. This separates the dye into its various components. **(5)** When the solvent front reaches nearly the top of the filter paper, the filter paper is removed. **(6)** The result is called a chromatogram.
18	R_f value is the distance moved by a substance relative to that moved by the solvent. $R_f = \frac{\text{distance travelled by a substance}}{\text{distance travelled by the solvent}}$
19	A locating agent is a substance that reacts with the substances on the chromatography paper to produce a coloured product.
20	When chromatographed, a pure substance only produces a single spot on the chromatogram.
21	A pure substance has a fixed melting and boiling point, e.g., pure water melts at 0 ^{o}C and boils at 100 ^{o}C.
22	Impurities affect the melting point in two ways: • Lowers the melting point. The greater the amount of impurity, the lower the melting point. • Substance melts over a range of temperatures,
23	Impurities affect the boiling point in two ways: • Raises the boiling point. The greater the amount of impurity, the higher the boiling point. • Liquid boils over a range of temperatures, i.e., no sharp boiling point.
24	The kinetic particle theory states that: • Particles are discrete and are too small to be seen directly. • There are spaces between particles of matter, and this space is different between solids, liquids and gases. • Particles are in constant random motion, and they move at different speeds in solids, liquids and gases.
25	• Diffusion is the spreading and mixing of particles in liquids and gases.

<table>
<tr><td></td><td>• The particles move from a region of higher concentration to a region of lower concentration, i.e., down a diffusion gradient.</td></tr>
<tr><td>26</td><td>Two factors affect the rate of diffusion:
• Temperature: Particles move faster at higher temperatures as they absorb the heat energy. The faster the particles move, the faster the rate of diffusion, i.e., the higher the temperature, the faster the rate of diffusion.
• Mass of particle: The greater the (molecular) mass of a particle, the slower the rate of diffusion.</td></tr>
<tr><td>27</td><td>
<table>
<tr><th>Particles in a Solid</th><th>Particles in a Liquids</th><th>Particles in a Gases</th></tr>
<tr><td>• Packed closely together in an orderly arrangement.
• Very little empty space between them.
• Cannot be compressed.
• Vibrate but cannot move freely about their fixed positions.</td><td>• Packed closely together but not in an orderly arrangement.
• Very little empty space between them, although further apart than in solids.
• Cannot be compressed.
• Vibrate and can move freely throughout the liquid.</td><td>• Far apart, and in a constant random arrangement.
• Large spaces between them.
• Can be compressed easily.
• Vibrate and are free to move throughout the container.</td></tr>
</table>
</td></tr>
<tr><td>28</td><td>Melting
(1) Particles in the solid absorb energy.
(2) This energy causes the particles to vibrate faster about their fixed positions.
(3) The vibrations of the particles overcome the forces of attraction between them.
(4) The particles are no longer in fixed positions.
(5) They can now move about freely, although still packed closely together.</td></tr>
<tr><td>29</td><td>Freezing
(1) Particles in the liquid lose energy.
(2) The particles begin to move more slowly as the temperature decreases.
(3) The particles are overcome by the forces of attraction and can no longer move freely.
(4) The particles settle into fixed positions.
(5) The liquid has now frozen to form a solid.</td></tr>
<tr><td>30</td><td>Boiling
(1) Particles in the liquid absorb energy.
(2) The particles move faster as temperature increases
(3) The particles gain enough energy to overcome the forces of attraction holding them together.
(4) The particles move further apart and more freely.
(5) When the particles gain enough energy to overcome both the forces of attraction and atmospheric pressure, the liquid changes into a gas.</td></tr>
</table>

31	An element is a substance that cannot be broken down into simpler substances by chemical methods.
32	An atom is the smallest component of an element having the chemical properties of the element.
33	A molecule is a group of two or more atoms chemically combined together.
34	A compound is a substance containing two or more elements chemically combined together.
35	An atom or a group of atoms that has an electrical charge is called an ion.
36	A mixture consists of two or more substances (can be elements, compounds, or both) that are not chemically combined together.
37	An atom is the smallest component of an element having the chemical properties of the element.
38	Isotopes are atoms of the same element (or same number of protons) with different number of neutrons. Isotopes have different physical properties but similar chemical properties.
39	Atoms lose or gain / share valence electrons to attain the stable electronic structure of a noble gas.
40	A metallic bond is the electrostatic attractive force between the positively charge metal ions and the negatively charged delocalised electrons.
41	Structure of metals: Lattice of metal cations surrounded by a “sea’ of mobile (or delocalised) valence electrons.
42	High Density: Metals consist of a giant metallic lattice of closely-packed, regular arrangement of metal cations surrounded by delocalised electrons. Hence, metals have high density as they have a high mass in a given unit volume.
43	Malleability and ductility of metals • Pure metal contains metal cations of the same size and are arranged in a regular, orderly manner, allowing the layers of cations to slide over one another easily. • Non-directional nature of the metallic bonds allows layers of metal cations to slide over one another without breaking the strong metallic bonds.
44	High melting point of metals: A large amount of energy is required to overcome the strong electrostatic attraction forces between the cations and the negative delocalised electrons.
45	Good electrical conductivity: Presence of mobile valence electrons which can migrate freely throughout the metallic structure when a potential difference is applied.
46	An ionic bond is the strong electrostatic attraction between positively charged ions (cations) and negatively charged ions (anions).
47	Ionic compounds are usually crystalline solids which are hard but brittle. • Hard: A large amount of energy is required to overcome the strong electrostatic attraction between the cations and anions. • Brittle: When a large enough force is applied, the layers of ions can slide, bringing ions of like charges next to each other. These ions of like charges repel each other and the ionic lattice shatters.
48	High melting points of ionic compounds: Large amount of energy must be supplied to overcome these strong forces of electrostatic attraction between the cations and anions. Generally, the higher the charges of the positive and negative ions the stronger the ionic bonds and the higher the melting point.
49	Ionic compounds conducts electricity in molten and aqueous states but non-conducting in solid state

	• Solid state: The ions are held in fixed positions in a giant ionic crystal lattice structure by strong electrostatic forces of attraction and thus, the ions are unable to move freely to conduct electricity. • Molten or liquid or aqueous state: The ions are mobile and free to move and will migrate to the electrodes when a potential difference is applied.
50	A covalent bond is the force of attraction between the shared electrons and the two positive nuclei of the atoms.
51	Substances with simple covalent structures • Structure consists of small, discrete covalent molecules • Atoms within a molecule are joined by strong intramolecular covalent bonds while weak intermolecular forces of attraction exists between separate molecules and holds them together. • The weak intermolecular forces are easily overcome
52	Substances with giant covalent structures All the atoms in substances with giant covalent structure are held tightly by strong covalent bonds and form a giant covalent molecule
53	Diamond • Each carbon atom is joined to 4 other carbon atoms by strong covalent bonds • Large amount of energy required to overcome the strong covalent bonds • No free electrons to conduct electricity
54	Graphite • Each carbon atom is joined to 3 other carbon atoms by strong covalent bonds • 1 unbonded valence electron per carbon atom which is delocalised between the layers and can migrate along the layers to conduct electricity • Small amount of energy is required to overcome the weak intermolecular forces between the layers, allowing the layers to slide over each other.
55	An acid is a substance which produces hydrogen ions (H^+) when it is dissolved in water.
56	Lead seemingly appears to not react with either dilute hydrochloric acid or dilute sulfuric acid. Why? This is due to the layer of lead(II) chloride or lead(II) sulfate that forms on the surface of the lead metal during its reaction with the respective acids. This layer is impervious to water and thus prevents the lead below from further reactions.
57	Most dilute acids react with metals to give a salt and hydrogen gas, e.g., $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
58	Acids react with carbonates and hydrogencarbonates to give a salt, water and carbon dioxide gas. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2 + H_2O(l) + CO_2(g)$
59	Acids react with metal oxides and hydroxides to give a salt and water. $CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$
60	Dilute acids can conduct electricity due to the present of mobile charge carriers, i.e., ions. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
61	An acid does not behave as an acid unless water is present. Water is required for the formation of the hydrogen ions.

62 Basicity (or proticity) of an acid refers to the maximum number of hydrogen ions in one molecule of an acid that can be replaced by a metal.

Type of acid	Example(s)
Monobasic (monoprotic)	HCl, HNO_3
Dibasic (diprotic)	H_2SO_4, H_2SO_3, H_2CO_3, H_3PO_3
Tribasic (triprotic)	H_3PO_4

63 A strong acid is one that is fully dissociated (ionised) in water to give H^+ ions.

64 A weak acid is one that is partially dissociated (ionised) in water to give H^+ ions, e.g., organic acids such as ethanoic acid.

$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

65 Formation of acid rain

Sulfur dioxide (SO_2) dissolves in rainwater to form sulfurous acid (H_2SO_3).

$$SO_2 + H_2O \rightarrow H_2SO_3$$

In the presence of oxygen, sulfur dioxide can be converted into sulfur trioxide (SO_3) which dissolves in water to form sulfuric acid (H_2SO_4).

$$2SO_2 + O_2 \rightarrow 2SO_3$$
$$SO_3 + H_2O \rightarrow H_2SO_4$$

Effects of acid rain includes:

- Acid rain falling on lakes kills fish and other aquatic animals.
- Acid rain falling on buildings and stone works can erode these structures.
- Acid rain falling on soil causes soil to become acidic and essential mineral salts e.g. potassium ions, calcium ions and aluminium ions are washed away.

65 Bases are the oxides or hydroxides of metals.

66 An alkali is a soluble base which produces hydroxide (OH^-) ions in water.

67 Alkalis react with acids to form salt and water only. The reaction is known as neutralisation.

68 Alkalis react with ammonium compounds on heating to give a salt, water and ammonia gas.

$Ca(OH)_2(aq) + 2NH_4Cl(s) \rightarrow CaCl_2(aq) + 2H_2O(l) + 2NH_3(g)$

69 A strong alkali undergoes complete dissociation (or ionisation) to form hydroxide ions, OH^-, in water, e.g., NaOH and KOH.

70 A weak alkali dissolves in water to form a small amount of hydroxide ions, e.g., ammonia.

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

71 An indicator is a substance that has different colours in acidic and alkaline solutions.

72 pH is a measure of the acidity of an acid or the basicity of an alkali. It is measured when a substance is dissolved in water.

73 The reaction of an acid and an alkali (or base) to form a salt and water only is called neutralisation, e.g.,

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

74 Acidic oxides are the oxides of non-metals. If they dissolve in water, they will form acids.

Acidic Oxide	Formula	Acid Produced in Water	Equation

Carbon dioxide	CO_2	Carbonic acid, H_2CO_3	$CO_2 + H_2O \rightleftharpoons H_2CO_3$
Nitrogen dioxide	NO_2	Nitric acid, HNO_3	$NO_2 + H_2O \rightarrow HNO_3$
Sulfur dioxide	SO_2	Sulfurous acid, H_2SO_3	$SO_2 + H_2O \rightarrow H_2SO_3$
Sulfur trioxide	SO_3	Sulfuric acid, H_2SO_4	$SO_3 + H_2O \rightarrow H_2SO_4$
Phosphorus(V) oxide	P_4O_{10}	Phosphoric acid, H_3PO_4	$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$
Silicon dioxide	SiO_2	Nil (oxide is insoluble in water)	-

75 Basic oxides are oxides of metals. Most basic oxides are insoluble in water, while those that dissolve in water form alkalis.

76 Amphoteric oxides are metallic oxides which can behave as an acidic oxide or a basic oxide. They react with both acids and bases, to form a salt and water.

Amphoteric oxide	Formula	Equation
Zinc oxide	ZnO	With acid: $ZnO + H^+ \rightarrow Zn^{2+} + H_2O$ With base: $ZnO + OH^- + H_2O \rightarrow [Zn(OH)_4]^{2-}$
Lead(II) oxide	PbO	With acid: $PbO + H^+ \rightarrow Pb^{2+} + H_2O$ With base: $PbO + OH^- + H_2O \rightarrow [Pb(OH)_4]^{2-}$
Aluminium oxide	Al_2O_3	With acid: $Al_2O_3 + H^+ \rightarrow Al^{3+} + H_2O$ With base: $Al_2O_3 + OH^- \rightarrow [Al(OH)_4]^-$

77 Neutral oxides are non-metallic oxides that exhibit neither basic nor acidic properties. They do not react with either acids or bases.

Neutral oxide	Formula
Water	H_2O
Carbon monoxide	CO
Nitrogen monoxide	NO

78 Step 1: Write the chemical equation, e.g., reaction of aqueous sodium hydroxide with dilute hydrochloric acid.

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Step 2: 'Ionise' all the aqueous species (i.e., those that can dissolve in water) into ions. We do this because when, for example, $NaOH(s)$ dissolves in water, it breaks down into Na^+ and OH^- ions.

Solids, liquids and gases are written in full.

$$Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

Step 3: Cancel out the 'like' ions on both sides of the equation. They are 'like' because they have not participated in the reaction. We call these spectator ions.

$$\cancel{Na^+(aq)} + OH^-(aq) + H^+(aq) + \cancel{Cl^-(aq)} \rightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$$

Step 4: This leaves us with the following equation, comprising the ions on the LHS. This is the ionic equation. We see that the hydroxide ion and the hydrogen ion reacts to form water. State symbols must always be included with ionic equations.

$$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$$

79 Solubility table

Compounds	Remarks
Na^+, K^+, NH_4^+ compounds	All are soluble
Nitrates (NO_3^-)	All are soluble
Chlorides (Cl^-)	All are soluble except $AgCl$, $PbCl_2$
Sulfates (SO_4^{2-})	All are soluble except $PbSO_4$, $BaSO_4$, $CaSO_4$
Carbonates (CO_3^{2-})	All are insoluble except those of Na^+, K^+, NH_4^+
Lead(II) compounds	All are insoluble except $Pb(NO_3)_2$

80 Salts are ionic compounds formed by replacing one or more hydrogen ions of an acid with a metallic ion or an ammonium ion (NH_4^+).

81
- Salt crystals are often formed by crystallisation from aqueous solutions, and thus often have water molecules bonded to them. This water is known as water of crystallisation.
- Salts that contain water of crystallisation are known as hydrated salts.
- Heating a hydrated salt drives away the water, leaving the anhydrous salt.

82 The method of preparation depends on three factors:

(1) Solubility of the salt.
(2) Reactivity of the metal.
(3) Chemicals being reacted together to make the salt.

83 Insoluble salts can be prepared by precipitation.

84 There are a few ways to prepare a soluble salt:
- Reacting an acid with a metal - however, note that we must never use reactive metals (e.g., Na, K) in the reaction as it will be explosive.
- Reacting an acid with an insoluble base.
- Reacting an acid with an insoluble carbonate.
- Reacting an acid and an alkali (i.e., a soluble base) via titration

85 QA – refer to QA notes. Everything is important!

<table>
<tr><td>86</td><td>Relative atomic mass, A_r, of an element is the average mass of one atom of the element when compared with $^1/_{12}$ of the mass of an atom of carbon-12.</td></tr>
<tr><td>87</td><td>Relative molecular mass, M_r, of a molecule is the average mass of one molecule of a substance when compared with $^1/_{12}$ of the mass of one atom of carbon-12.</td></tr>
<tr><td>88</td><td>Relative formula mass, M_r, is the average mass of one formula unit of an ionic compound when compared with $^1/_{12}$ of the mass of one atom of carbon-12.</td></tr>
<tr><td>89</td><td>$\% \text{ of an element} = \frac{A_r \text{ of the element}}{M_r \text{ of the compound}} \times 100\%$</td></tr>
<tr><td>90</td><td>$1 \text{ mol} = 6.02 \times 10^{23}$ of particles</td></tr>
<tr><td>91</td><td>$\text{number of moles} = \frac{\text{number of particles}}{6.02 \times 10^{23} (\text{in mol}^{-1})}$</td></tr>
<tr><td>92</td><td>$\text{n, number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g/mol)}}$</td></tr>
<tr><td>93</td><td>Empirical formula calculation sample:
<table>
<tr><th></th><th>Mg</th><th>O</th></tr>
<tr><td>Mass/ g</td><td>0.48</td><td>0.32</td></tr>
<tr><td>n/ mol</td><td>$\frac{0.48}{24.3} = 0.0198$</td><td>$\frac{0.32}{16.0} = 0.0200$</td></tr>
<tr><td>Mole ratio
(this means to divide by the smallest n throughout)</td><td>$\frac{0.0198}{0.0198} = 1$</td><td>$\frac{0.0200}{0.0198} = 1.01 \approx 1$</td></tr>
<tr><td>Simplest mole ratio</td><td>1</td><td>1</td></tr>
</table>
Empirical formula of magnesium oxide is MgO.</td></tr>
<tr><td>94</td><td>ONLY FOR GASES! NOT FOR LIQUIDS!
$\text{n, number of moles} = \frac{\text{volume of gas (in dm}^3)}{24 \text{ (in dm}^3 \text{ mol}^{-1})}$</td></tr>
<tr><td>95</td><td>$\text{concentration} = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in dm}^3)} = \frac{\text{number of moles of solute (in mol)}}{\text{volume of solution (in dm}^3)}$</td></tr>
<tr><td>96</td><td>$\text{Percentage purity} = \frac{\text{mass of pure substance}}{\text{mass of impure substance}} \times 100\%$</td></tr>
<tr><td>97</td><td>$\text{Percentage yield} = \frac{\text{actual mass of pure product}}{\text{theoretical mass of pure product}} \times 100\%$</td></tr>
<tr><td>98</td><td>The Group number indicates the number of valence electrons in an atom of an element.</td></tr>
<tr><td>99</td><td>The Period number indicates the number of shells in an atom of an element.</td></tr>
<tr><td>100</td><td>Group I
Down the Group, the atomic size increases and the distance between the nucleus and the delocalised electrons increases. The weaker attraction for these electrons results in weaker metallic bonds. Hence melting point decreases down the Group.</td></tr>
</table>

<table>
<tr><td>101</td><td>
<table>
<tr><th>Element</th><th>Observations</th></tr>
<tr><td>Lithium</td><td>Reacts quickly with water.
$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$</td></tr>
<tr><td>Sodium</td><td>Reacts rapidly with water. Sodium melts into a silvery ball and darts around the water surface. It catches fire and sometimes explodes. The hydrogen gas produced burns with a yellow flame.
$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$</td></tr>
<tr><td>Potassium</td><td>Reacts violently with water. Potassium melts into a silvery ball and darts around the water surface. It catches fire and then expodes. The hydrogen gas produced burns with a lilac flame.
$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$</td></tr>
</table>
</td></tr>
<tr><td>102</td><td>The halogens have a simple molecular structure, with the diatomic molecules held together by weak intermolecular forces which are easily broken by a small amount of energy. Hence, they have low melting and boiling points.</td></tr>
<tr><td>103</td><td>Why m.p. and b.p. increases down the Group?
The halogen molecules are held together by weak intermolecular forces. The strength of the weak intermolecular forces increases down the Group as the number of electrons increases. More energy is required to overcome these forces.</td></tr>
<tr><td>104</td><td>The more reactive halogen can displace the less reactive halogen from an aqueous solution of its ions, e.g.,
$Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$</td></tr>
<tr><td>105</td><td>Chemical properties of transition metals
• Variable valency and oxidation states
• Catalytic properties
• Formation of coloured solid compounds and coloured aqueous solutions</td></tr>
<tr><td>106</td><td>A substance is oxidised when it gains oxygen atom(s) in a reaction.

A substance is reduced when it loses oxygen atom(s) in a reaction.</td></tr>
<tr><td>107</td><td>A substance is oxidised when it loses hydrogen atom(s) in a reaction.
A substance is reduced when it gains hydrogen atom(s) in a reaction.</td></tr>
<tr><td>108</td><td>A substance is oxidised when it loses electron(s) in a reaction.
A substance is reduced when it gains electron(s) in a reaction.</td></tr>
<tr><td>109</td><td>A substance is oxidised when the oxidation number of an element in the substance increases after the reaction.
A substance is reduced when the oxidation number of an element in the substance decreases after the reaction.</td></tr>
<tr><td>110</td><td>The oxidation number of hydrogen with non-metals is +1.
The oxidation number of hydrogen with metals is -1.
The oxidation number of hydrogen in hydrogen molecule is zero</td></tr>
</table>

111 The oxidation number of oxygen as an oxide = -2.
The oxidation number of oxygen as a peroxide = -1.
The oxidation number of oxygen as a superoxide = -½.
The oxidation number of oxygen as an element is zero

112 Substances that cause oxidation of (oxidised) another substance in a reaction is called an oxidising agent. Oxidising agents themselves are reduced.

113 Substances that cause reduction of (reduced) another substance in a reaction is called a reducing agent. Reducing agents themselves are oxidised.

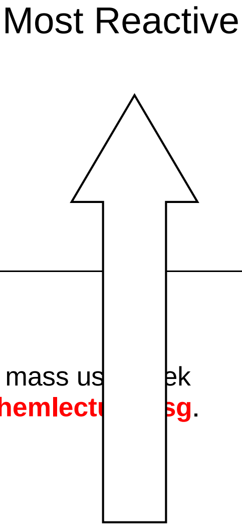
114 Test for Reducing Agents

Reagent	Observation	Explanation
aqueous potassium manganate(VII), $KMnO_4$, acidified by dilute H_2SO_4	Acidified aqueous $KMnO_4$ turned from purple to colourless.	MnO_4^- reduced to Mn^{2+} as the oxidation number of Mn decreases from +7 in MnO_4^- to +2 in Mn^{2+}. Colour of MnO_4^-(aq) is purple Colour of Mn^{2+}(aq) is pale pink (almost colourless)
aqueous potassium dichromate(VI), $K_2Cr_2O_7$, acidified by dilute H_2SO_4	Acidified aqueous $K_2Cr_2O_7$ turned from orange to green.	$Cr_2O_7^{2-}$ is reduced to Cr^{3+} as the oxidation number of Cr decreases from +6 in $Cr_2O_7^{2-}$ to +3 in Cr^{3+}. Colour of $Cr_2O_7^{2-}$(aq) is orange Colour of Cr^{3+}(aq) is green

115 Test for Oxidising Agents

Reagent	Observation	Explanation
aqueous KI, dilute H_2SO_4	Acidified aqueous. KI turned from colourless to reddish brown solution seen. [Note: May observe black solid at the bottom of the test-tube if the amount of iodine present is higher.]	I^- is oxidised to I_2 as the oxidation number of I increases from –1 in I^- to 0 in I_2.

116



Potassium	Paid
Sodium	specialists
Calcium	can
Magnesium	make

	Aluminium	a
	Zinc	zoo
	Iron	in
	Tin	this
	Lead	ludicrously
	(Hydrogen)	humid
	Copper	country
	Mercury	more
	Silver	special
Least reactive	Gold	Great!

117

Metal	Observations	Chemical Equation
Potassium	• Reacts very violently with cold water. • A lot of heat is produced and this causes the hydrogen gas produced to catch fire and explode. • Metal burns with a lilac flame. • Aqueous potassium hydroxide and hydrogen gas are produced. • Explodes with steam	$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
Sodium	• Reacts violently with cold water. • Hydrogen gas is produced and the gas may catch fire and explode. • Metal burns with a yellow flame. • Aqueous sodium hydroxide and hydrogen gas are produced. • Explodes with steam	$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

	Calcium	• Reacts readily with cold water. • Metal sinks in water. • Many bubbles of hydrogen gas are produced. • Aqueous calcium hydroxide and hydrogen gas are produced. • Water turns milky in colour as calcium hydroxide is partially soluble in water.	$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
	Magnesium	• Reacts slowly with cold water. • A few bubbles of hydrogen gas are produced. • Aqueous magnesium hydroxide and hydrogen gas are produced. • Hot magnesium reacts violently with steam to form magnesium oxide (a white powder) and hydrogen gas. A bright white glow is produced during the reaction.	$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$ $Mg(s) + 2H_2O(g) \rightarrow MgO(s) + H_2(g)$
	Zinc	• Does not react with cold water. • Hot zinc reacts readily with steam. Zinc is yellow when hot and white when cold • Solid zinc oxide and hydrogen gas are produced.	$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$
	Iron	• Does not react with cold water. • Red-hot iron reacts slowly with steam to form solid iron(II, III) oxide (tri-iron tetroxide) and hydrogen gas.	$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
	Lead, copper, silver	• No reaction with cold water or steam.	-

118

Metal	Observations	Chemical Equation
Potassium	• Explodes with dilute hydrochloric acid. • Aqueous potassium chloride and hydrogen gas are produced.	$2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$
Sodium	• Explodes with dilute hydrochloric acid. • Aqueous sodium chloride and hydrogen gas are produced.	$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
Calcium	• Reacts violently with dilute hydrochloric acid. • Many bubbles of hydrogen gas are produced. • Aqueous calcium chloride and hydrogen gas are produced.	$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
Magnesium	• Reacts readily with dilute hydrochloric acid. • Many bubbles of hydrogen gas are produced. • Aqueous magnesium chloride and hydrogen gas are produced.	$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Zinc	• Reacts readily with dilute hydrochloric acid. • Many bubbles of hydrogen gas are produced. • Aqueous zinc chloride and hydrogen gas are produced.	$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
Iron	• Reacts slowly with dilute hydrochloric acid. • Bubbles of hydrogen gas are produced. • Pale green aqueous iron(II) chloride and hydrogen gas are produced.	$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
Lead, copper, silver	• Does not react with dilute hydrochloric acid.	-

119

Metal	Observations	Chemical Equation
Potassium	• Some heat applied results in a very vigorous reaction.	$2K(s) + O_2(g) \rightarrow 2K_2O(s)$

	- A lot of heat is given out. - Metal burns with a lilac flame. - Solid potassium oxide (a white powder) is produced.	
Sodium	- Some heat applied results in a very vigorous reaction. - A lot of heat is given out. - Metal burns with a yellow flame. - Solid sodium oxide (a white powder) is produced.	$2Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
Calcium, magnesium, zinc	- Strong heat applied to start reaction. - A lot of heat is given out. - White solids are produced. - Magnesium burns with a brilliant white flame.	$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$ $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ $2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$
Iron	- Strong heat applied to start reaction. - Glows red hot and burns with white sparks.	$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$
Lead, copper	- Strong heat applied to start reaction.	$2Pb(s) + O_2(g) \rightarrow PbO(s)$ $2Cu(s) + O_2(g) \rightarrow 2CuO(s)$
Gold, silver	- No reaction occurs.	-

120

Metal	Appearance	Water on Metal	Acids on Metal	Air or Oxygen on Metal	Displacement Reactions
K	Dull, stored under oil	React with water (with decreasing ease)	Hydrogen gas produced (with decreasing ease)	Burn in air or oxygen	Each metal will displace a metal lower in the series from as solution of one of its salts.
Na					
Ca	Generally dull				
Mg		React with steam when heated			
Al					
Zn					
Fe					
Sn		No reaction with steam or water		Converted into oxide by heating in air (without burning)	
Pb					
Cu	Generally shiny		React with oxidising acids		
Hg					
Ag					
Au					

Pt			No reaction with acids	Unaffected by air or oxygen	

121 Compounds formed by reactive metals are more stable to heat than those formed by less reactive metals.

122

Metal	Carbonates
K	Stable to heat
Na	
Ca	Decompose to metal oxide and carbon dioxide gas on heating. Example: $PbCO_3(s) \rightarrow$ $PbO(s) + CO_2(g)$
Mg	
Al	
Zn	
Fe	
Sn	
Pb	
Cu	
Hg	Unstable, do not exist For silver carbonate, Ag_2CO_3, it decomposes into silver and carbon dioxide on heating.
Ag	
Au	
Pt	

123

Oxide of	Action of Heat	Action of Hydrogen	Action of Carbon
K	No reaction	No reaction	No reaction
Na			
Ca			
Mg			
Al			
Zn			Reduced to metal (with increasing ease) Example: $2CuO(s) + C(s) \rightarrow$ $2Cu(s) + CO_2(g)$
Fe		Reduced to metal (with increasing ease) Example: $CuO(s) + 3H_2(g) \rightarrow Cu(s) + H_2O(l)$	
Sn			
Pb			
Cu			
Hg	Decompose to give metal and oxygen Example: $2Ag_2O(s) \rightarrow$ $4Ag(s) + O_2(g)$		
Ag			
Au			
Pt			

124

Metal	Main Ore	Main Chemical Constituent	Extraction Method	Ease of Extraction
K	Carnallite	$KMgCl_3$	Electrolytic reduction.	Hardest ↑ Easiest
Na	Rock salt	$NaCl$		
Ca	Chalk, limestone	$CaCO_3$		
Mg	Dolomite	$CaMg(CO_3)_2$		
Al	Bauxite	Al_2O_3		
Zn	Zinc blende	ZnS	Extraction by reducing with carbon in a furnace.	
Fe	Haematite	Fe_2O_3		
Sn	Tinstone	SnO_2		
Pb	Galena	PbS		
Cu	Copper pyrites	$CuFeS_2$		
Hg	Cinnabar	HgS		
Ag	-	Ag		
Au	-	Au	Found free (unbounded) in the ground.	

125 Extraction of iron

- Main ore: Haematite, iron(III) oxide mixed with impurities such as sand and clay.
- Iron is extracted in a blast furnace. Iron ore, coke and limestone is added from the top of the furnace while hot air is pumped in near the bottom.

Stage 1

Coke reacts with oxygen from the air to form carbon dioxide in a highly exothermic reaction. Temperatures can reach 1900 ^{o}C.

$C(s) + O_2(g) \rightarrow CO_2(g)$

Stage 2

The carbon dioxide formed reacts with more coke to form carbon monoxide. This reaction is endothermic and temperature drops to about 1100 ^{o}C.

$CO_2(g) + C(s) \rightarrow 2CO(g)$

Stage 3

The carbon monoxide reduces the iron(III) oxide to produce molten iron. This collects at the bottom of the blast furnace.

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

Other reactions

	(1) Limestone added is decomposed by the heat to produce basic calcium oxide (lime) and carbon dioxide. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ Iron contains many impurities such as silicon, sulfur and phosphorus. Silicon is present as silicon dioxide (sand), an acidic oxide. SiO_2 reacts with CaO to produce slag, calcium silicate, $CaSiO_3$. Slag collects at the bottom of the furnace and floats on top of the molten iron. $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$ Hot waste gases such as carbon monoxide, carbon dioxide and nitrogen escape from the top of the furnace.
126	For rust to form, iron, water and oxygen must be present.
127	Rusting process (1) Iron reacts with water and oxygen (in the air) to form iron(II) hydroxide. In this reaction, iron is oxidised while water is reduced. $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$ (2) Iron(II) hydroxide reacts with more oxygen and water to form iron(III) hydroxide. $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$ (3) Iron(III) hydroxide loses some water to become rust. Rust is hydrated iron(III) oxide, written as Fe2O3.xH2O as they amount of water can vary. $Fe(OH)_3(s) \rightarrow Fe_2O_3.xH_2O$ (4) Overall equation $4Fe(s) + 3O_2(g) + 2xH_2O(l) \rightarrow 2Fe_2O_3.xH_2O(s)$
128	Rust is flaky and peels off easily, allowing iron underneath to rust. This is unlike aluminium oxide which forms an impervious layer on the surface of aluminium, preventing further reaction.
129	Rusting is faster when the iron is in contact with a piece of metal that is less reactive than itself.
130	• The Haber process reaction is as follows: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ Conditions: 250 atm, 450 °C, finely divided iron as catalyst Note: ▪ The Haber process reaction is reversible.

  - It is exothermic.
  - Reaction takes place under high temperature and high pressure.
  - Only about 10-15% of the reactants are converted into ammonia.

- Why high pressure (250 atm)?

  - Increasing the pressure favours the forward reaction. The speed of the forward reaction increases.
  - However, too high a pressure is costly to maintain because expensive equipment is required.
  - Hence, there is a compromise on the pressure used for the reaction. 250 atm is used.

- Why the temperature of 450 °C?

  - This is because the decomposition of ammonia through the backward reaction is reduced.
  - However, too low a temperature will reduce the rate of the forward reaction.
  - Hence, a relatively high temperature of 450 °C is used.

- Why the finely divided iron catalyst?

  - Despite the high pressure and relatively high temperature, the overall reaction is still slow.
  - A catalyst is used to speed up the reaction. It does not affect the yield.

131 Other Reactions to Obtain Ammonia

- By heating a mixture of an ammonium salt and an alkali or base.

$$NH_4Cl(s) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$$

Note: Concentrated sulfuric acid cannot be used because it reacts with ammonia gas:

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

132 Fertilisers

Nitrogen (N)	for healthy green stems and leaves; for increase in protein content of plants and size of crop
Phosphorus (P)	for good root growth
Potassium (K)	for flower and fruit formation; helps plant to photosynthesise.

133 An exothermic reaction is a reaction where energy is given out to the surroundings.

134 An endothermic reaction is a reaction where energy is absorbed from the surroundings.

135 Activation energy (E_a) is the minimum energy required to be possessed by the reactant particles in order for a chemical reaction to occur.

136 Overall enthalpy change (ΔH) = total energy absorbed for bond breaking - total energy released during bond forming

- If more energy is absorbed than released (e.g., 1000 kJ – 500 kJ), overall enthalpy change > 0, i.e., endothermic reaction.
- If more energy is released than absorbed (e.g., 500 kJ – 1000 kJ), overall enthalpy change < 0, i.e., exothermic reaction.

137 Hydrogen fuel cell

At the positive electrode	At the negative electrode
Oxygen is reduced to form hydroxide ions $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	Hydrogen is oxidised to form water $H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$

Overall reaction is the conversion of hydrogen and oxygen into water with the release of electrical energy.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

138

- Increase in concentration of reactants increases the number of reactant particles per unit volume.
- The frequency of collisions increases, and this increases the frequency of effective collisions. Hence, the speed of the chemical reaction increases.

139

- When the particle size decreases, the surface area of the reactant increases, giving a larger surface area to volume ratio.
- More surfaces are exposed for reactant particles to collide, increasing the frequency of collisions.
- This increases the frequency of effective collisions, hence increasing the speed of reaction.

140

- Increasing the pressure increases the number of particles per unit volume of the gas.
- The collisions between the gas particles become more frequent. This increases the frequency of effective collisions, increasing the speed of the chemical reaction.

141

- Increase in temperature increases the speed of a chemical reaction.
- The particles move faster and they possess more kinetic energy.
- This increases the likelihood that they collide with sufficient or more than enough energy to overcome activation energy.
- The frequency of effective collisions increases, increasing the speed of the chemical reaction.

142 A catalyst is a substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.

- Only a small amount is required as it is not chemically changed.
- A catalyst is not used up during the chemical reaction.
- A catalyst is selective and different reactions require different catalysts.
- A catalyst increases the speed of a reaction but not the yield of the reaction.

<table>
<tr><td></td><td>• A catalyst can be poisoned by impurities, and no longer works or become less effective as a catalyst.
• Catalysts are usually transition metals or their compounds.</td></tr>
<tr><td>143</td><td>Enzymes are biological catalysts
• Enzymes are proteins.
• Enzymes are substrate specific.
• Enzymes are sensitive to temperature changes.
• Enzymes are sensitive to pH changes.</td></tr>
<tr><td>144</td><td>
<table>
<tr><th>Carbon monoxide</th><th>Nitrogen oxides (NO, NO_2)</th><th>Sulfur dioxide</th></tr>
<tr>
<td>A. Source
- From incomplete combustion of fuels in vehicles and industry
B. Harmful Effects
- Colourless, odourless gas prevents easy detection.
- CO combines with haemoglobin in the blood to form stable carboxyhaemoglobin, preventing the haemoglobin from transporting oxygen to body cells, depriving them of oxygen.
C. Prevention/Removal
- Catalytic converters to remove CO in vehicles by oxidising it into CO_2.</td>
<td>A. Source
- Nitrogen and oxygen in the air combine to form nitrogen monoxide at higher temperatures in car engines and power stations/ presence of lightning.
$N_2(g) + O_2(g) \rightarrow 2NO(g)$
- NO is further oxidised to form NO_2.
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
B. Harmful Effects
- NO_2 is highly corrosive and toxic. It irritates lungs, breathing passages, and eyes.
- Oxides of nitrogen leads to the formation of acid rain and photochemical smog.</td>
<td>A. Source
- During combustion of fossil fuels such as coal and petroleum which contain sulfur as impurity
- Volcanic eruptions
- Extraction of iron in blast furnace when the sulfur impurities are oxidised off.
$S(s) + O_2(g) \rightarrow SO_2(g)$
B. Harmful Effects
- Irritates the eyes and causes breathing problems.
- Enters leaves and affects the growth of plants.
- Main cause of acid rain.
C. Prevention/Removal
- Removal of SO_2 from flue gases.
- CaO or $Ca(OH)_2$ is added to neutralise excess acidity in lakes and soil.</td>
</tr>
</table>
</td></tr>
</table>

	C. Prevention/Removal - Catalytic converters to remove NO, NO$_2$, as N$_2$.	
Unburnt Hydrocarbons	**Methane**	**Ozone**
A. Source - Come mainly from hydrocarbons in fuel that have not been burnt in vehicle engines. **B. Harmful Effects** - Some can cause cancer, they can react with sunlight and other pollutants to form photochemical smog.	**A. Source** - Produced when plant and animal matter decay. - Cows and other farm animals also produce methane. **B. Harmful Effects** - Contributes to global warming.	**A. Source** - O$_3$, a form of oxygen and colourless. - Sunlight acts on air and the pollutants such as nitrogen dioxide, carbon monoxide and hydrocarbons to produce photochemical smog. - Ozone is part of this mixture. **B. Harmful Effects** - Irritates the eyes, nose and throat and can cause asthma attacks. - It damages plant crops.

145 **Reducing Sulfur Dioxide Emissions in Power Stations**

Step 1: Limestone (calcium carbonate) added to hot flue gases. Heat causes calcium carbonate to decompose to calcium oxide and carbon dioxide.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Step 2: Calcium oxide reacts with sulfur dioxide to form calcium sulfite, which is further oxidised by oxygen in air to form calcium sulfate.

$$CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$$
$$2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$$

146 **Electrolysis of molten NaCl**

Identify electrolyte: NaC*l*(l)

In solid state, the ions are held in fixed positions in the crystal lattice by strong electrostatic forces of attraction

Identify ions present: Na$^+$(l) and C*l*$^-$(l)

When heated strongly, the electrostatic forces of attraction between the cations and anions are broken, and the ions are free to move and no longer held in fixed positions.

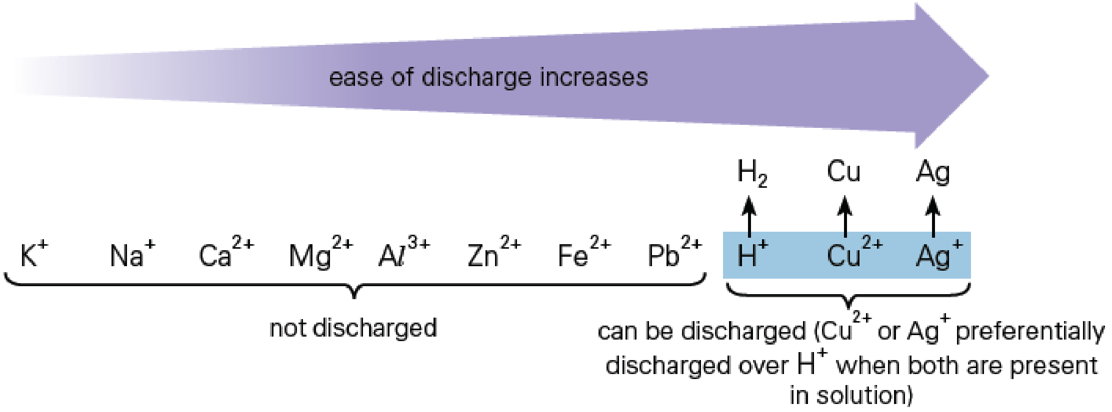
Identify electrodes used: carbon

- Carbon is an inert electrode. **Inert electrodes** do not react with the products of electrolysis or the electrolyte, and are not involved in the electrolysis process.

At the anode (electrode connected to the positive terminal of the battery)	**At the cathode** (electrode connected to the negative terminal of the battery)
• Cl^- ions are attracted to the anode. • Each Cl^- ion lose 1 electron to the anode to form a chloride atom. • 2 Cl atoms come together to form Cl_2 gas • Cl^- ions are oxidised. • Cl^- ions are discharged as chlorine gas. $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$	• Na^+ ions are attracted to the cathode. • Each Na^+ ion gain 1 electron from the cathode to form a sodium atom. • Na^+ ions are reduced. • Na^+ ions are discharged as sodium metal. $Na^+(l) + e^- \rightarrow Na(l)$

Overall reaction: $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

147



148

difficult sulfate ion, SO_4^{2-}

nitrate ion, NO_3^-

↓ chloride ion, Cl^-

bromide ion, Br^-

iodide ion, I^-

easy hydroxide ion, OH^-

149 **Electrolysis of dilute NaCl**

Step 1: Identify the ions present in the electrolyte

	Cations	**Anions**
From sodium chloride	$Na^+(aq)$	$Cl^-(aq)$
From water	$H^+(aq)$	$OH^-(aq)$

Step 2: Determine ion preferentially discharged at the cathode (electrode connected to negative terminal of battery)

- Positively charged ions will migrate to cathode: $Na^+(aq)$, $H^+(aq)$
- $H^+(aq)$ is more easily discharged over $Na^+(aq)$.
- $H^+(aq)$ is preferentially discharged as hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)\ \textbf{[reduction]}$$

- $Na^+(aq)$ remains in solution.

Step 3: Determine ion preferentially discharged at the anode (electrode connected to positive terminal of battery)

- Negatively charged ions will migrate to anode: $Cl^-(aq)$, $OH^-(aq)$
- $OH^-(aq)$ is more easily discharged compared to $Cl^-(aq)$.
- $OH^-(aq)$ is preferentially discharged as water and oxygen gas.

$$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-\ \textbf{[oxidation]}$$

- $Cl^-(aq)$ remains in solution.

Overall

- $2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$
- Equivalent to electrolysis of water.
- As electrolysis proceeds, concentration of sodium chloride will increase. This would affect the products formed during electrolysis after some time.

150 **Electrolysis of concentrated NaCl (brine)**

Step 1: Identify the ions present in the electrolyte

	Cations	**Anions**
From sodium chloride	$Na^+(aq)$	$Cl^-(aq)$
From water	$H^+(aq)$	$OH^-(aq)$

Step 2: Determine ion preferentially discharged at the cathode (electrode connected to negative terminal of battery)

- Positively charged ions will migrate to cathode: $Na^+(aq)$, $H^+(aq)$

- $H^+(aq)$ is more easily discharged over $Na^+(aq)$.
- $H^+(aq)$ is preferentially discharged as hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g) \textbf{ [reduction]}$$

- $Na^+(aq)$ remains in solution. **(Metal ions are never discharged from aqueous solutions.)**

Step 3: Determine ion preferentially discharged at the anode (electrode connected to positive terminal of battery)

- Negatively charged ions will migrate to anode: $Cl^-(aq)$, $OH^-(aq)$
- $OH^-(aq)$ is more easily discharged compared to $Cl^-(aq)$ in dilute solutions. However, in concentrated solutions, $Cl^-(aq)$ is preferentially discharged instead. **(Halide ions such as chloride, bromide and iodide ions are only discharged during the electrolysis of concentrated solutions.)**

$$2Cl^-(aq) \rightarrow Cl_2(aq) + 2e^- \textbf{ [oxidation]}$$

- $OH^-(aq)$ remains in solution.

151 **Electrolysis of dilute sulfuric acid**

Step 1: Identify the ions present in the electrolyte

	Cations	**Anions**
From sulfuric acid	$H^+(aq)$	$SO_4^{2-}(aq)$
From water	$H^+(aq)$	$OH^-(aq)$

Step 2: Determine ion preferentially discharged at the cathode (electrode connected to negative terminal of battery)

- Only positively charged ion which migrates to cathode: $H^+(aq)$
- $H^+(aq)$ is discharged as hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g) \textbf{ [reduction]}$$

Step 3: Determine ion preferentially discharged at the anode (electrode connected to positive terminal of battery)

- Negatively charged ions will migrate to anode: $SO_4^{2-}(aq)$, $OH^-(aq)$
- $OH^-(aq)$ is more easily discharged compared to $SO_4^{2-}(aq)$.
- $OH^-(aq)$ is preferentially discharged as water and oxygen gas.

$$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^- \textbf{ [oxidation]}$$

- $SO_4^{2-}(aq)$ remains in solution.

Overall

- $2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$

152 Electrolysis of aqueous copper(II) sulfate with INERT electrodes

Step 1: Identify the ions present in the electrolyte

	Cations	**Anions**
From copper(II) sulfate	Cu^{2+}(aq)	SO_4^{2-}(aq)
From water	H^+(aq)	OH^-(aq)

Step 2: Determine ion preferentially discharged at the cathode (electrode connected to negative terminal of battery)

- Positively charged ions will migrate to cathode: H^+(aq), Cu^{2+}(aq)
- Cu^{2+}(aq) is more easily discharged compared to H^+(aq).
- Cu^{2+}(aq) is preferentially discharged as Cu(s).

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)\ \textbf{[reduction]}$$

Step 3: Determine ion preferentially discharged at the anode (electrode connected to positive terminal of battery)

- Negatively charged ions will migrate to anode: SO_4^{2-}(aq), OH^-(aq)
- OH^-(aq) is more easily discharged compared to SO_4^{2-}(aq).
- OH^-(aq) is preferentially discharged as water and oxygen gas.

$$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-\ \textbf{[oxidation]}$$

- SO_4^{2-}(aq) remains in solution.

Overall

- The copper metal discharged forms a layer around the cathode.
- The concentration of Cu^{2+}(aq) in the solution decreases as the reaction proceeds.
- The intensity of the blue colour decreases.

153 Electrolysis of aqueous copper(II) sulfate using COPPER electrode

Step 1: Identify the ions present in the electrolyte

	Cations	**Anions**
From copper(II) sulfate	Cu^{2+}(aq)	SO_4^{2-}(aq)
From water	H^+(aq)	OH^-(aq)

Step 2: Determine ion preferentially discharged at the cathode (electrode connected to negative terminal of battery)

- Positively charged ions will migrate to cathode: H^+(aq), Cu^{2+}(aq)
- Cu^{2+}(aq) is more easily discharged compared to H^+(aq).

- Cu^{2+}(aq) is preferentially discharged as Cu(s).

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)\ \textbf{[reduction]}$$

Step 3: Determine ion preferentially discharged at the anode (electrode connected to positive terminal of battery)

- Negatively charged ions will migrate to anode: SO_4^{2-}(aq), OH^{-}(aq)
- However, the copper anode participates in the electrolysis process. It is known as a reactive electrode.
- The copper electrode is oxidised into Cu^{2+}(aq) ions which enter into the electrolyte solution.

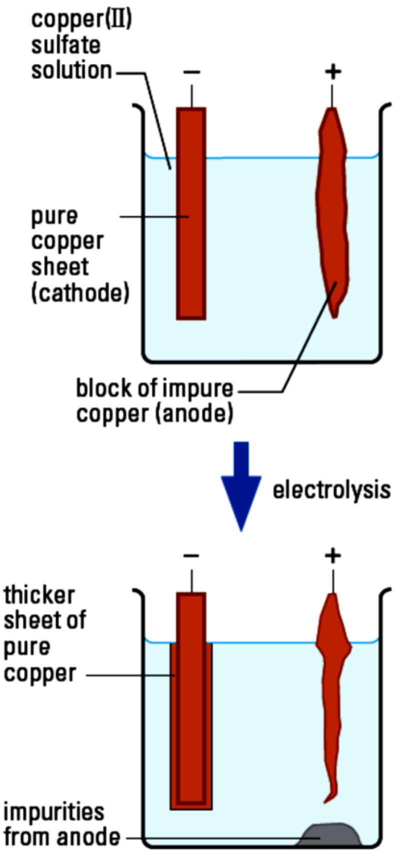
$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}\ \textbf{[oxidation]}$$

- SO_4^{2-}(aq) and OH^{-}(aq) remains in solution.

Overall

- The copper anode dissolves and decreases in mass.
- The concentration of Cu^{2+}(aq) in the solution remains unchanged as Cu^{2+}(aq) deposited as Cu(s) at the cathode is continuously replenished by the anode.
- The intensity of the blue colour remains unchanged until the anode disappears.

154 Purification of metals



- During electrolysis of copper(II) sulfate solution with copper electrodes, the anode dissolves.
- A piece of pure copper is used as the cathode.
- Impure copper is used as the anode.
- Pure copper from the anode is oxidised into Cu^{2+}(aq) which enters the electrolyte.
- Cu^{2+}(aq) is then deposited on the cathode.
- Impurities fall off the anode and collect below.
- The mass of the cathode increases.
- Impure copper can be purified in this way.

155 Electroplating

- The metal to be used for electroplating is made the anode.
- The object to be plated is made the cathode.
- For copper plating for example,
  - Copper is used as the anode.
    The electrolyte is aqueous copper(II) sulfate.
  - Copper from the electrolyte is deposited on the object, while copper from the anode replaces the $Cu^{2+}(aq)$ in the electrolyte.

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

156 Zinc-copper cell with dilute sulfuric acid as electrolyte

Step 1: Identify which electrode is oxidised. Reduction will occur at the other electrode.

The more reactive metal will be oxidised. In this case, Zn is more reactive than Cu. Hence Zn is oxidised.

Step 2: Identify the flow of electrons.

Electrons flow from the negative electrode to the positive electrode.

At the zinc electrode	At the copper electrode
• Zinc atoms are oxidised. They give up electrons to form zinc ions. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ • Because zinc gives up electrons, it behaves as the negative electrode. • The zinc ions enter into solution.	• The electrons from the zinc electrode travel to the copper electrode. • The solution contains both $Na^+(aq)$ and $H^+(aq)$ ions. • $H^+(aq)$ is more easily discharged. • $H^+(aq)$ accepts the electrons and is reduced to form hydrogen gas. $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overall redox reaction: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

157 Zinc-copper cell with aqueous copper(II) sulfate as electrolyte

Step 1: Identify which electrode is oxidised. Reduction will occur at the other electrode.

The more reactive metal will be oxidised. In this case, Zn is more reactive than Cu. Hence Zn is oxidised.

Step 2: Identify the flow of electrons.

Electrons flow from the negative electrode to the positive electrode.

At the zinc electrode	At the copper electrode
• Zinc atoms are oxidised. They give up electrons to form zinc ions. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ • Because zinc gives up electrons, it behaves as the negative electrode. • The zinc ions enter into solution.	• The electrons from the zinc electrode travel to the copper electrode. • The solution contains both $Cu^{2+}(aq)$ and $H^+(aq)$ ions. • $Cu^{2+}(aq)$ is more easily discharged. • $Cu^{2+}(aq)$ accepts the electrons and is reduced to form Cu metal. $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
Overall redox reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$	

158 The voltage reading depends on the **difference in reactivity between the two metals** used for the electrodes.

159 Hydrocarbons = compounds made up of only C and H

160 A homologous series is a group of compounds with a **general formula**, **similar chemical properties** and **showing a gradation in physical properties** as a result of increase in the size and mass of the molecules, e.g. melting and boiling points, viscosity, flammability.

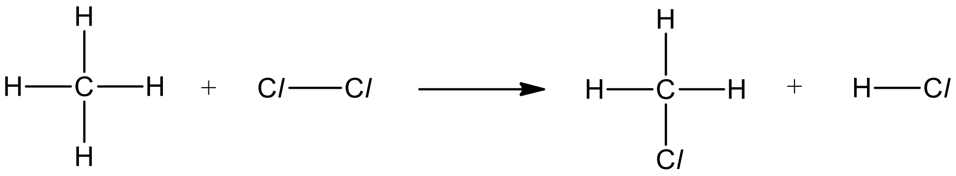
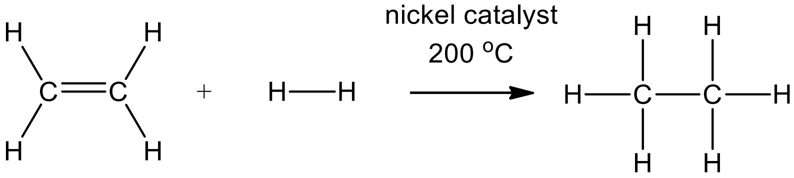
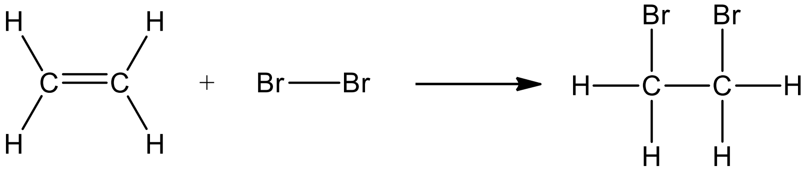
- Members have the same general formula; each member differs from the next by a $–CH_2–$ group of atoms.
- Have chemical properties which are very similar for each member (since they have the same functional group)
- Have physical properties which show a steady gradation on going down the series.
- Can be made by similar methods.

161 **Alkanes** form a homologous series consists of **saturated hydrocarbons** with the general formula $\mathbf{C_nH_{2n+2}}$. Each subsequent alkene varies by a $-CH_2-$ group. Alkanes are compounds contain only carbon-carbon single covalent bonds (C–C) and carbon-hydrogen single covalent bonds (C–H).

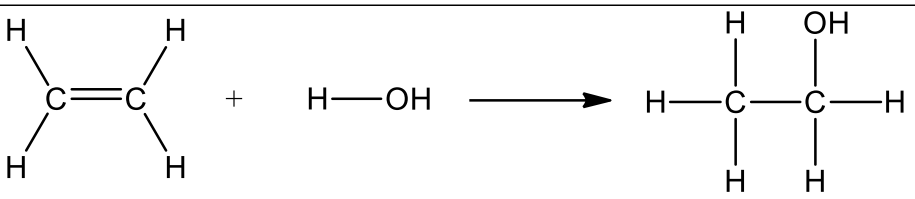
162 Isomers are compounds with the **same molecular formula** but different structural formula.

163 The **melting point and boiling point of the alkanes increases as the relative molecular mass increases**. More energy is required to overcome the stronger van der Waals forces of attraction between the alkane molecules.

164 Alkanes can react with chlorine in the presence of UV light, e.g., methane with chlorine.

	$CH_4(g) + Cl_2(g) \longrightarrow CH_3Cl(g) + HCl(g)$ 
165	**Alkenes** form a homologous series consists of **unsaturated hydrocarbons** with the general formula **C_nH_{2n}**. Each subsequent alkene varies by a -CH_2- group. Alkenes are compounds containing carbon-carbon double covalent bond (C=C) and carbon-hydrogen single covalent bond (C–H).
166	Alkenes have a higher percentage of carbon compared to alkanes, and burn with sootier flames compared to alkanes with similar number of carbon atoms, to give carbon dioxide and water.
167	Hydrogenation of alkenes $C_2H_4(g) + H_2(g) \xrightarrow{\substack{\text{nickel catalyst}\\ 200\ ^{o}C}} C_2H_6(g)$ 
168	Reddish-brown liquid bromine changes to colourless when it reacts with ethene. Useful identification test for unsaturated hydrocarbons – it decolourises reddish-brown bromine water. $C_2H_4(g) + Br_2(l) \rightarrow CH_2BrCH_2Br(l)$ 
169	Steam addition or hydration of alkenes $C_2H_4(g) + H_2O(g) \xrightarrow{\substack{\text{phosphoric acid}\\ 300\ ^{o}C,\ 60\ atm}} C_2H_5OH(l)$

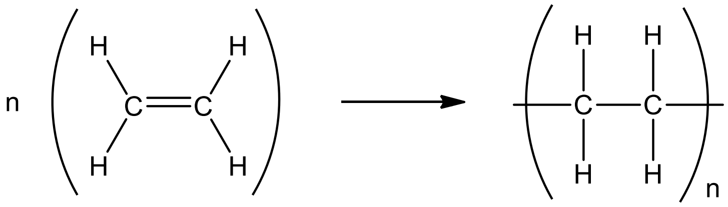
phosphoric acid
300 °C, 60 atm



170 Polymerisation of alkenes

$$\text{ethene} \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} \text{ethane}$$

$$n(CH_2{=}CH_2) \xrightarrow[\text{catalyst}]{\text{high temperature and pressure}} -(CH_2CH_2)_n-$$



171 Cracking

- Alkenes are manufactured by the cracking of hydrocarbons. Long-chain hydrocarbons are broken down into short-chain hydrocarbons.

  hexane → butane + ethane

  $C_6H_{14} \rightarrow C_4H_{10} + C_2H_4$

- A catalyst and high temperatures can be used to speed up the cracking process, i.e., catalytic cracking.

  Catalyst: aluminium oxide (Al_2O_3) and silicon dioxide (SiO_2)
  Temperature: 600 ^{o}C

  long-chain alkane → (mixture of short-chain alkenes) + (mixture of short-chain alkanes or H_2)

172

	Alkanes	**Alkenes**
Similarities	1. Contains carbon and hydrogen only 2. Flammable 3. Complete combustion produces carbon dioxide and water only	
Differences	• Only single covalent bonds between carbon atoms	• At least one double covalent bond between carbon atoms
	• Generally unreactive • Undergoes substitution reaction	• More reactive than alkanes

		• Undergoes addition reaction
	• Does not decolourise reddish-brown bromine water	• Decolourises reddish-brown bromine water.
	• Small-chain alkanes burn cleanly	• Alkenes of similar length of carbon chain burns with a smokier flame

173 **Alcohols** form a homologous series containing the –OH functional group or hydroxyl group, with the general formula $\mathbf{C_nH_{2n+1}OH}$.

174 Fermentation is a chemical reaction in which sugars are broken down into smaller molecules such as ethanol by micro-organisms such as yeast, e.g., fermentation of glucose to form ethanol.

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast, } 37\ ^oC} 2CH_3CH_2OH(aq) + CO_2(g)$$

$$\text{glucose} \xrightarrow{\text{yeast, } 37\ ^oC} \text{ethanol + carbon dioxide}$$

- The mixture of yeast and glucose is kept at 37 °C as this is the optimum temperature for enzymes in yeast to function. To high a temperature and the enzymes in yeast will be denatured and cannot function as catalysts.
- Carbon dioxide is released during fermentation. Hence frothing is observed in the flask. A while precipitate also forms in limewater.
- Fermentation is carried out in the absence of oxygen. When there is oxygen present, little to no ethanol is produced. Oxygen can also oxidise ethanol into ethanoic acid.

$$CH_3CH_2OH(aq) + O_2(g) \rightarrow CH_3COOH(aq) + H_2O(l)$$

- A dilute solution of ethanol is formed. It is only about 15% concentrated. Above this concentration, the yeast dies and fermentation stops. Ethanol is separated from the mixture by fractional distillation.

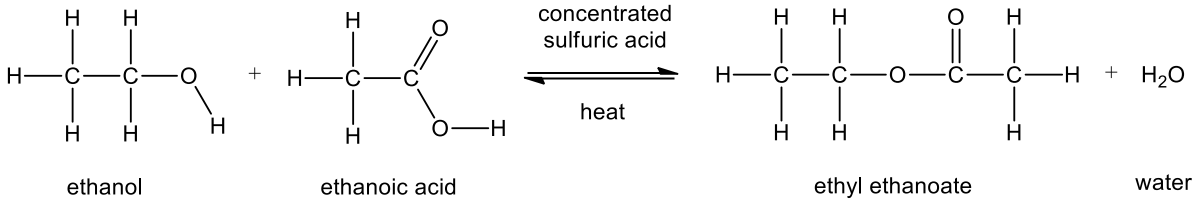
175 Main industrial method of producing ethanol.

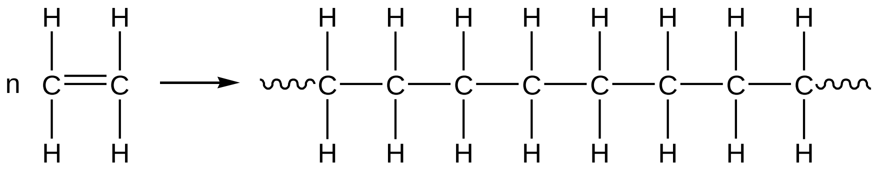
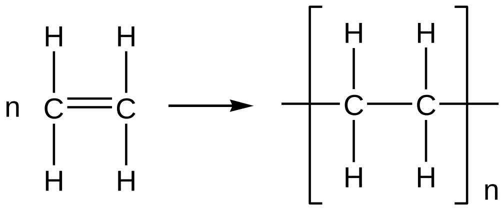
$$C_2H_4(g) + H_2O(g) \xrightarrow[300\ ^oC,\ 60\ atm]{\text{phosphoric(V) acid } (H_3PO_4)} CH_3CH_2OH(aq)$$

$$\text{ethene + steam} \xrightarrow[300\ ^oC,\ 60\ atm]{\text{phosphoric(V) acid } (H_3PO_4)} \text{ethanol}$$

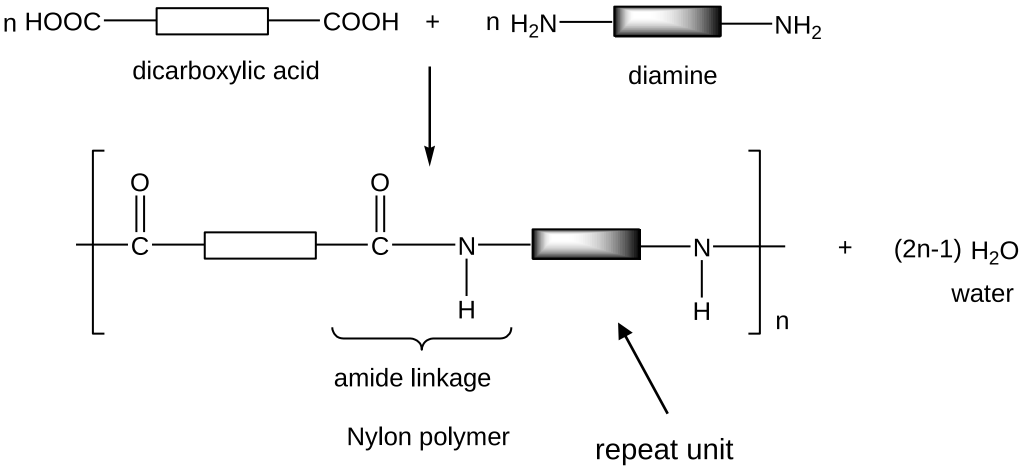
176 Alcohols can be oxidised into carboxylic acids by warming them with suitable oxidising agents such as acidified potassium manganate(VII), e.g., ethanol to ethanoic acid.

During this reaction, purple acidified potassium manganate(VII) is decolourised.

	$CH_3CH_2OH(aq) + 2[O] \xrightarrow[\text{heat}]{KMnO_4,\ \text{dil. } H_2SO_4} CH_3COOH(aq) + H_2O(l)$
177	**Carboxylic acids** form a homologous series containing the –COOH (called carboxyl group) functional group with the general formula **$C_nH_{2n+1}COOH$**.
178	Carboxylic acids behave like weak acids as they can dissociate partially in water to release hydrogen ions. The will turn blue litmus paper red (pH < 7). $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$ ethanoic acid ⇌ ethanoate ion + hydrogen ion - Reaction with relatively reactive metals – hydrogen gas is evolved $Mg(s) + 2CH_3COOH(aq) \rightarrow (CH_3COO)_2Mg(aq) + H_2(g)$ *magnesium ethanoate* - Reaction with carbonates – carbon dioxide gas is evolved $Na_2CO_3(s) + 2CH_3COOH(aq) \rightarrow 2CH_3COONa(aq) + H_2O(l) + CO_2(g)$ *sodium ethanoate* - Reaction with bases – neutralisation reaction (forms salt and water) $NaOH(aq) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$ *sodium ethanoate*
179	Esterification – alcohol + carboxylic acid, in the presence of conc sulfuric acid as catalyst and heat 
180	There are two ways for producing ethanoic acid: (1) Oxidation of ethanol by acidified potassium manganate(VII) with heating $CH_3CH_2OH(aq) + 2[O] \xrightarrow[\text{heat}]{KMnO_4,\ \text{dil. } H_2SO_4} CH_3COOH(aq) + H_2O(l)$

	(2) Oxidation of ethanol by oxygen in air $CH_3CH_2OH(aq) + O_2(g) \longrightarrow CH_3COOH(aq) + H_2O(l)$
181	A macromolecule is a large molecule made by joining together many small molecules.
182	**Polymers** (synthetic) are examples of macromolecules which are made up of large number of small molecules called **monomers** joined together. **Polymers do not have sharp melting points.** **Polymerisation** is the process of joining together these small molecules to form big molecules.
183	Addition polymerisation is a reaction where **unsaturated monomers** join together to form one large molecule as the only product. No molecules or atoms are lost during the process.
184	**<u>Polyethene</u>** n (ethene) → polyethene n = as many as 50,000  monomer  $n(CH_2{=}CH_2) \rightarrow -(CH_2CH_2)_n-$ In the presence of: 1. High temperature and pressure 2. Catalyst
185	Condensation polymerisation occurs when monomers join together to form a polymer with the elimination of small molecules.
186	**<u>Nylon</u>** • Condensation polymer • Synthetic fibre

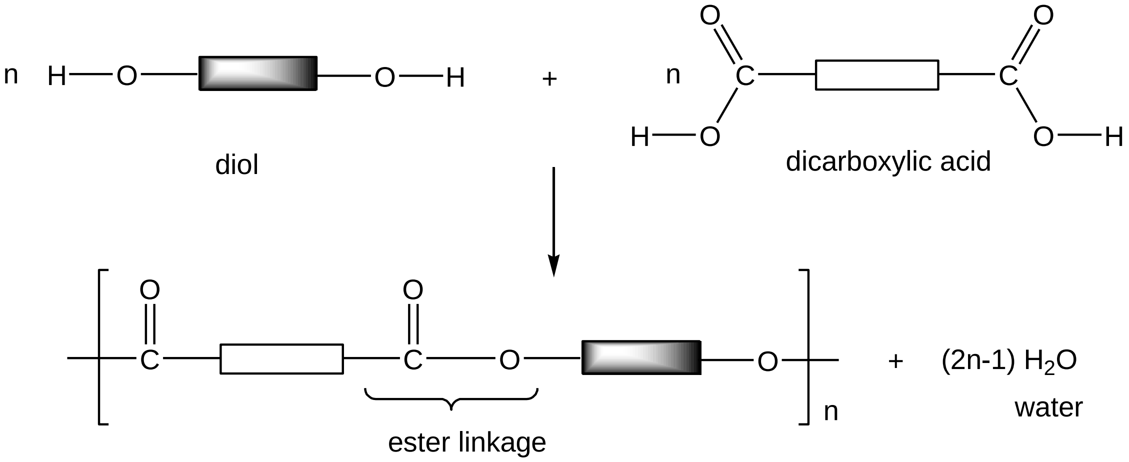
- Strong and light, stretchable without breaking
- Made from two monomers
- Also known as polyamides because the monomers are joined together by amide linkages



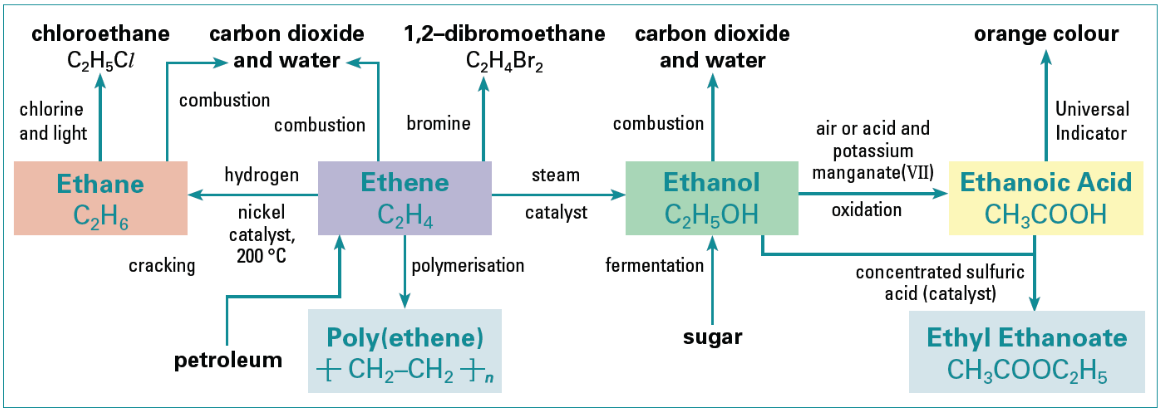
Nylon polymer

187 **Terylene**

- Condensation polymer
- Synthetic fibre
- Also known as polyester as the monomers are joined together by ester linkages



Terylene polymer

188	

chloroethane
C2H5Cl
carbon dioxide and water
1,2–dibromoethane
C2H4Br2
carbon dioxide and water
orange colour
chlorine and light
combustion
combustion
bromine
combustion
air or acid and potassium manganate(VII)
oxidation
Universal Indicator
Ethane
C2H6
hydrogen
nickel catalyst, 200 °C
Ethene
C2H4
steam
catalyst
Ethanol
C2H5OH
Ethanoic Acid
CH3COOH
cracking
polymerisation
fermentation
concentrated sulfuric acid (catalyst)
petroleum
Poly(ethene)
+ CH2–CH2 +n
sugar
Ethyl Ethanoate
CH3COOC2H5