Figure 3 If the $6.02 \times 10^{23}$ atoms in 12 g of carbon were turned into marbles, the marbles could cover Great Britain to a depth of 1500 km!
Figure 3 The mass spectrometer.

Figure 7 Variation of first ionisation enthalpy with atomic number for elements with atomic numbers 1 to 56.

Figure 8 Successive ionisation enthalpies for aluminium.
Figure 9  Successive ionisation enthalpies for phosphorus.

Figure 10  Energies of electron sub-shells from \( n = 1 \) to \( n = 4 \) in a typical many-electron atom. The energy of a sub-shell is not fixed, but falls as the charge on the nucleus increases as you go from one element to the next in the Periodic Table. The order shown in the diagram is correct for the elements in Period 3 and up to nickel in Period 4. After nickel the 3d sub-shell has lower energy than 4s.

Figure 13  Arrangement of electrons in atomic orbitals in a ground state sodium atom.
Figure 14  Building up the Periodic Table.
Figure 15  Dividing up the Periodic Table.
Figure 9: In a hydrogen molecule, the atoms are held together because their nuclei are both attracted to the shared electrons.

Figure 13: Pauling electronegativity values for some main group elements in the Periodic Table.

Figure 2: The sodium chloride lattice, built up from oppositely charged sodium ions and chloride ions.

Figure 15: A model of metallic bonding.
Figure 17 The relative sizes of atoms and ions.

Li atom
$r_{atom} = 0.152$ nm
Li$^+$ ion
$r_{ion} = 0.078$ nm

Na atom
$r_{atom} = 0.186$ nm
Na$^+$ ion
$r_{ion} = 0.098$ nm

Mg atom
$r_{atom} = 0.160$ nm
Mg$^{2+}$ ion
$r_{ion} = 0.078$ nm

F atom
$r_{atom} = 0.071$ nm
F$^-$ ion
$r_{ion} = 0.133$ nm

O atom
$r_{atom} = 0.073$ nm
O$^{2-}$ ion
$r_{ion} = 0.132$ nm

Figure 18 Ions with higher charge densities attract more water molecules (ions are only shown in two dimensions).

Na$^+$ ion
(charge 1+, radius 0.098nm)
On average, each Na$^+$ ion is surrounded by 5 water molecules

Mg$^{2+}$ ion
(charge 2+, radius 0.078nm)
On average, each Mg$^{2+}$ ion has 15 water molecules around it

Figure 19 Hydrated ions are much bigger than isolated ions.

Li$^+$ (g)
hydrated Li$^+$ ion (radius = 1.00nm)

Li$^+$ (aq)

Na$^+$ (g)
hydrated Na$^+$ ion (radius = 0.79nm)

Na$^+$ (aq)

Figure 36 Two isomers of alanine.

Figure 38 The CORN rule. Look down the $H-C$ bond from hydrogen towards the central carbon atom.

L isomer

R isomer

H$_2$N

C

COOH

H

C

H$_2$N

C

NH$_2$
Enthalpy

**Reactants**

**Products**

**Progress of reaction**

- Energy given out to surroundings: $\Delta H$ negative
- Energy taken in from surroundings: $\Delta H$ positive

**Figure 1** Enthalpy level diagram for an exothermic reaction, e.g., burning methane: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$.

**Figure 2** Enthalpy level diagram for an endothermic reaction, e.g., decomposing calcium carbonate: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$.

**Figure 3** A bomb calorimeter for making accurate measurements of energy changes. The fuel is ignited electrically and burns in the oxygen inside the pressurised vessel. Energy is transferred to the surrounding water, whose temperature rise is measured.

Note that the experiment is done at constant volume in a closed container. Enthalpy changes are for reactions carried out at constant pressure, so the result needs to be modified accordingly.

**Figure 4** An enthalpy cycle for finding the enthalpy change of formation of methane, $\text{CH}_4$.

$\Delta H$ going this way... is the same as $\Delta H$ going this way.
Figure 9  Breaking and making bonds in the reaction between methane and oxygen.

Figure 14 A comparison of solids, liquids and gases.

Figure 15 Each electronic energy level has within it several vibrational, rotational and translational energy levels. Note that the levels are not to scale.
Figure 21 An enthalpy cycle to show the dissolving of an ionic solid.

Figure 26 Born–Haber cycle for sodium chloride.
Figure 27 Born–Haber cycles for sodium chloride and potassium chloride.
Figure 27  Born–Haber cycles for sodium chloride and potassium chloride.
Figure 2 What happens when an ionic substance such as sodium chloride dissolves in water.

Figure 5 Polar water molecules attract the ions in a solid lattice.

Figure 11 (a) The structure of diamond and (b) the structure of graphite.

Figure 13 The fullerenes are a recently discovered molecular form of carbon. (a) shows C_{60}, named buckminsterfullerene. It is made up of a mixture of 5-membered and 6-membered rings and looks like a football. (b) is another way of presenting C_{60}, showing the positions of the carbon atoms. (c) shows C_{70}, which is shaped like a rugby ball.

Figure 12 Imagine you are inside a diamond. The regular network structure would repeat in all directions – as far as the edge of the diamond.
Figure 14  On heating, the molecular substance (represented by ) changes from a solid to a liquid and then to a gas. Energy must be supplied to overcome the intermolecular forces. Note that the covalent bonds within the molecules remain intact.

Electron cloud evenly distributed; no dipole. At some instant, more of the electron cloud happens to be at one end of the molecule than the other; molecule has an instantaneous dipole.

Figure 15  How a dipole forms in a chlorine molecule.

This atom is not yet polarised, but its electrons are repelled by the dipole next to it ...

Figure 18  How an induced dipole is formed in a Xe atom.

Figure 20  Variation in the boiling points of the hydrides of some Group 4, 5, 6 and 7 elements.

Figure 21  Variation in the enthalpy changes of vaporisation of some Group 4, 5, 6 and 7 elements.
Figure 23  The positively charged H atom lines up with the lone pair on an F atom.

Figure 25  The positively charged H atoms line up with the lone pairs on the O atoms.

Figure 28  The arrangement of water molecules in ice.
HEAT

(a) **Thermoplastic**: no cross-linking

Weak forces between polymer chains easily broken by heating; polymer can be moulded into new shape.

(b) **Thermoset**: extensive cross-linking

Strong covalent bonds between polymer chains cannot be easily broken; polymer keeps shape on heating.

**Figure 29** Thermoplastics and thermosets.

**Figure 32** Crystalline and amorphous regions of a polymer.

**Figure 36** A summary of the structures of substances.
Figure 3 Obtaining a line emission spectrum.

Figure 4 Obtaining a line absorption spectrum.

Figure 6 How the Lyman series in the emission spectrum is related to energy levels in the H atom.
Figure 7. An HCl molecule has energy associated with different aspects of its behaviour.

Figure 16. The basic parts of a double beam infrared spectrometer.

Figure 21. Infrared spectrum of butane.
Figure 22 Infrared spectrum of methylbenzene.

Figure 23 Infrared spectrum of benzoic acid.

Figure 36 The principle of n.m.r.: a small magnet in a strong magnetic field can have two different energies.

Figure 37 A simplified diagram of an n.m.r. spectrometer.
**Figure 43** N.m.r. spectrum of hexane.

**Figure 44** N.m.r. spectrum of trans-but-2-ene.

**Figure 45** N.m.r. spectrum of propan-1-ol.

**Figure 52** How light behaves with transparent and opaque objects.
transparent object absorbs red
transmits wavelengths corresponding to other colours: **appears green**

opaque object absorbs blue
reflects wavelengths corresponding to other colours: **appears orange**

**Figure 53** How colours arise from absorption of light.

**Figure 57** The absorption spectrum of carotene (in solution in hexane).

**Figure 59** Absorption and reflectance spectra of Monastral Blue.
Figure 63 The energy needed to excite an electron in a coloured compound and in a colourless compound.
Figure 10 Ion-exchange columns are used to soften water. The ion-exchange resin removes calcium ions from the ‘hard’ water, and replaces them with sodium ions to form ‘soft’ water.

Figure 11 Thin-layer chromatography.

Figure 15 A gas–liquid chromatograph.
Figure 5 The pH scale.

Figure 6 How a buffer solution keeps the pH constant.

- Plenty of HA to make more H⁺(aq) if some used up by alkali that gets added.
- Plenty of A⁻ to combine with any H⁺(aq) that gets added.
- H⁺(aq) stays roughly constant because there is plenty of HA to make more H⁺(aq) if some used up by alkali that gets added.
Figure 2. The reaction of chlorine with potassium iodide solution.

Figure 4. The reaction of zinc with copper(II) sulphate solution:
\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Cu}(s) + \text{Zn}^{2+}(aq) \]

Figure 5. The general arrangement for an electrochemical cell.

Figure 6. An ordinary dry cell – the kind you use in a torch.
Figure 9 A copper–zinc cell.

Figure 11 The standard hydrogen half-cell (sometimes called a standard hydrogen electrode).
solution containing equal concentrations of Fe^{3+} (aq) and Fe^{2+} (aq)
platinum electrode

Figure 12 A standard half-cell for the Fe^{3+} (aq)/Fe^{2+} (aq) half-reaction.
Figure 2  Enthalpy profile for an exothermic reaction.

Figure 4  Distribution curves for molecular kinetic energies in a gas at 300 K and 310 K.

Figure 5  Distribution curve showing collisions with energy 50 kJ mol⁻¹ and above.

Figure 6  Distribution curves showing the effect on the proportion of collisions with energy 50 kJ mol⁻¹ and above of changing the temperature from 300 K to 310 K.
Figure 8  Apparatus for investigating the rate of decomposition of hydrogen peroxide. The yeast provides the enzyme catalase.

Figure 9  The decomposition of hydrogen peroxide solutions of differing concentrations.

Figure 10  The initial rate of decomposition of hydrogen peroxide plotted against concentration of hydrogen peroxide.
Figure 13 Finding half-lives, $t_{\frac{1}{2}}$, for the decomposition of hydrogen peroxide.

Figure 14 Using a progress curve to find the rate of a reaction at different concentrations, by drawing tangents.

Figure 17 For a reaction involving two steps, there are two activation enthalpies.
Reactants get adsorbed onto catalyst surface. Bonds are weakened.

New bond forms.

Second bond forms, and product diffuses away from catalyst surface, leaving it free to adsorb fresh reactants.

Figure 18 An example of heterogeneous catalysis. The diagrams show a possible mechanism for nickel catalysing the reaction between ethene and hydrogen to form ethane.

Figure 19 The effect of a catalyst on the enthalpy profile for a reaction.
Figure 3 Densities of elements in Periods 2 and 3.

Figure 4 Melting and boiling points of elements in Period 3.

Figure 5 Variation in atomic size across Period 3.

Figure 6 First ionisation enthalpies of elements 1–20.
<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic number</th>
<th>Electronic arrangement</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sc</td>
<td>21</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Ti</td>
<td>22</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>V</td>
<td>23</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Cr</td>
<td>24</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Mn</td>
<td>25</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
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<tr>
<td>Fe</td>
<td>26</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
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<tr>
<td>Co</td>
<td>27</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Ni</td>
<td>28</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Cu</td>
<td>29</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>Zn</td>
<td>30</td>
<td>[Ar] $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$ $\uparrow$</td>
</tr>
</tbody>
</table>

Figure 20  Arrangement of electrons in the ground state of elements of the first row of the d block. [Ar] represents the electronic configuration of argon.

Figure 22  The arrangement of metal atoms in a crystal (a) before and (b) after slip has taken place. The shaded circles represent the end row of atoms.

Figure 23  The arrangement of metal atoms in an alloy. The open circles represent atoms of iron. The black circles are the larger atoms of a metal added to make an alloy.

<table>
<thead>
<tr>
<th>Sc</th>
<th>Ti</th>
<th>V</th>
<th>Cr</th>
<th>Mn</th>
<th>Fe</th>
<th>Co</th>
<th>Ni</th>
<th>Cu</th>
<th>Zn</th>
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<td>+1</td>
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<td>+2</td>
<td>+2</td>
<td>+2</td>
<td>+2</td>
<td>+2</td>
<td>+2 only</td>
</tr>
<tr>
<td>+3 only</td>
<td>+3</td>
<td>+3</td>
<td>+3</td>
<td>+3</td>
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<td>+3</td>
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<td></td>
</tr>
</tbody>
</table>

Figure 24  Oxidation states shown by elements in the first row of the d block. The most important oxidation states are in boxes.
Figure 27 An octahedral complex of Fe(III). Coordination number 6.

Figure 28 A tetrahedral complex of Ni(II). Coordination number 4.

Figure 29 A square planar complex of Ni(II). Coordination number 4.

Figure 30 A linear complex of Ag(I). Coordination number 2.

Figure 32
(a) Edta$^{4-}$, a hexadentate ligand.
(b) The nickel-edta complex ion [Ni(edta)]$^{2-}$. 
Figure 33 Relative energy levels for the five 3d orbitals of the hydrated Ti$^{3+}$ ion.

Figure 36 The cis and trans isomers of $[\text{Co(NH}_3)_4\text{Cl}_2]^+$ have different colours.

- **cis isomer:** violet
- **trans isomer:** green
<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular formula</th>
<th>Full structural formula</th>
<th>Shortened structural formula</th>
<th>Further shortened to</th>
</tr>
</thead>
<tbody>
<tr>
<td>methane</td>
<td>CH₄</td>
<td>H:H—C:H:H</td>
<td>CH₄</td>
<td></td>
</tr>
<tr>
<td>ethane</td>
<td>C₂H₆</td>
<td>H:H:C:H</td>
<td>CH₃—CH₃</td>
<td>CH₃CH₃</td>
</tr>
<tr>
<td>propane</td>
<td>C₃H₈</td>
<td>H:H:C:H:C:H</td>
<td>CH₃—CH₂—CH₃</td>
<td>CH₃CH₂CH₃</td>
</tr>
</tbody>
</table>

Table 2  Structural formulae of alkanes.

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**Figure 5**  The three-dimensional shape of methane.

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Figure 6  The three-dimensional shape of ethane.

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Figure 7  The three-dimensional shape of butane.
### ALKANES

<table>
<thead>
<tr>
<th>Isomerisation</th>
<th>Reforming</th>
<th>Steam cracking</th>
<th>Cracking</th>
<th>Catalytic cracking</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Feedstock</strong></td>
<td><strong>naphtha</strong></td>
<td><strong>naphtha/kerosene</strong></td>
<td><strong>gas oil</strong></td>
<td></td>
</tr>
<tr>
<td>C&lt;sub&gt;4&lt;/sub&gt; – C&lt;sub&gt;6&lt;/sub&gt; alkanes</td>
<td>naphtha</td>
<td>naphtha/kerosene</td>
<td>gas oil</td>
<td></td>
</tr>
<tr>
<td><strong>Product molecules</strong></td>
<td>same number of C atoms as reactants; cyclic</td>
<td>more molecules; fewer C atoms than reactants; some unsaturated; some branched; some cyclic</td>
<td>more molecules; fewer C atoms than reactants; cyclic</td>
<td></td>
</tr>
<tr>
<td>Feedstock</td>
<td>naphtha</td>
<td>naphtha/kerosene</td>
<td>gas oil</td>
<td></td>
</tr>
<tr>
<td><strong>Conditions</strong></td>
<td>Pt/Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;; 150 °C</td>
<td>Pt/Al&lt;sub&gt;2&lt;/sub&gt;O&lt;sub&gt;3&lt;/sub&gt;; 500 °C; hydrogen is recycled through mixture to reduce 'coking'</td>
<td>no catalyst; 900 °C; short residence time; steam is a diluent to prevent 'coking'</td>
<td>zeolite; 500 °C</td>
</tr>
<tr>
<td><strong>Example</strong></td>
<td><img src="example1.png" alt="Example" /></td>
<td><img src="example2.png" alt="Example" /></td>
<td><img src="example3.png" alt="Example" /></td>
<td><img src="example4.png" alt="Example" /></td>
</tr>
<tr>
<td><strong>Uses of products</strong></td>
<td>to improve octane rating of petrol</td>
<td>to improve octane rating of petrol</td>
<td>to manufacture polymers</td>
<td>to improve octane rating of petrol</td>
</tr>
<tr>
<td><strong>Further details</strong></td>
<td>Developing Fuels storyline, Section DF4</td>
<td>Developing Fuels storyline, Section DF4</td>
<td>The Polymer Revolution storyline, Sections PR3 and PR4</td>
<td>Developing Fuels storyline, Section DF4</td>
</tr>
</tbody>
</table>

Table 5 The action of heat on alkanes. Skeletal formulae are used in the examples.

[<](https://wwwимерной.jpg) is the skeletal formula for a benzene ring.
The C–C bond length is 0.139 nm.

Figure 11 The structure of the benzene ring.

Figure 14 An electron density map for benzene at –3 °C. The lines are like contour lines on a map: they show parts of the molecule with equal electron density.

Figure 15 Enthalpy changes for the hydrogenation of benzene and the hypothetical Kekulé structure.

\[ \Delta H^\circ = -360 \text{ kJ mol}^{-1} \] (estimated enthalpy change for Kekulé’s benzene)

\[ \Delta H^\circ = -208 \text{ kJ mol}^{-1} \] (measured enthalpy change for benzene)

Figure 16 The regions of higher electron density above and below the benzene ring.
Table 1  Naming halogenoalkanes.

<table>
<thead>
<tr>
<th>Full structural formula</th>
<th>Skeletal formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—C—C—C—C—Cl</td>
<td>Cl</td>
<td>1-chloropropane</td>
</tr>
<tr>
<td>H—C—C—C—Cl</td>
<td>Cl</td>
<td>1,2-dichloropropane</td>
</tr>
<tr>
<td>H—C—C—C—Cl</td>
<td>Br</td>
<td>3-bromo-1-chlorobutane</td>
</tr>
</tbody>
</table>

Table 3  Some common nucleophiles.

<table>
<thead>
<tr>
<th>Name and formula</th>
<th>Structure, showing lone pairs</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydroxide ion, OH⁻</td>
<td>H—O⁻</td>
</tr>
<tr>
<td>cyanide ion, CN⁻</td>
<td>N≡C⁻</td>
</tr>
<tr>
<td>ethanoate ion, CH₃COO⁻</td>
<td>CH₃C—O⁻</td>
</tr>
<tr>
<td>ethoxide ion, C₂H₅O⁻</td>
<td>CH₂CH₂—O⁻</td>
</tr>
<tr>
<td>water molecule, H₂O</td>
<td>H—O⁻</td>
</tr>
<tr>
<td>ammonia molecule, NH₃</td>
<td>N—H</td>
</tr>
</tbody>
</table>

Table 4  Naming alcohols.

<table>
<thead>
<tr>
<th>Full structural formula</th>
<th>Skeletal formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—C—C—C—C—OH</td>
<td>OH</td>
<td>propan-1-ol</td>
</tr>
<tr>
<td>H—C—C—C—OH</td>
<td>OH</td>
<td>propan-2-ol</td>
</tr>
<tr>
<td>H—C—C—C—OH</td>
<td>OH</td>
<td>pentan-3-ol</td>
</tr>
</tbody>
</table>
Table 6 Some examples of acid derivatives.

<table>
<thead>
<tr>
<th>Acid derivative</th>
<th>Dealt with in Section(s)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>ester</td>
<td>13.5 and 14.2</td>
<td>ethyl ethanoate</td>
</tr>
<tr>
<td>acyl chloride</td>
<td>13.5 and 14.2</td>
<td>ethanoyl chloride</td>
</tr>
<tr>
<td>amide</td>
<td>13.8</td>
<td>ethanamide</td>
</tr>
<tr>
<td>acid anhydride</td>
<td>13.5 and 14.2</td>
<td>ethanoic anhydride</td>
</tr>
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</table>

Table 7 Primary, secondary and tertiary alcohols.

<table>
<thead>
<tr>
<th>Type of alcohol</th>
<th>Position of —OH group</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>primary</td>
<td>at end of chain:</td>
<td>propan-1-ol</td>
</tr>
<tr>
<td></td>
<td>H</td>
<td>CH₃CH₂—C—OH</td>
</tr>
<tr>
<td></td>
<td>R—C—OH</td>
<td></td>
</tr>
<tr>
<td></td>
<td>H</td>
<td></td>
</tr>
<tr>
<td>secondary</td>
<td>in middle of chain:</td>
<td>propan-2-ol</td>
</tr>
<tr>
<td></td>
<td>H</td>
<td>CH₃—C—CH₃</td>
</tr>
<tr>
<td></td>
<td>R—C—R'</td>
<td></td>
</tr>
<tr>
<td></td>
<td>OH</td>
<td></td>
</tr>
<tr>
<td>tertiary</td>
<td>attached to a carbon</td>
<td>2-methylpropan-2-ol</td>
</tr>
<tr>
<td></td>
<td>atom which carries</td>
<td></td>
</tr>
<tr>
<td></td>
<td>no H atoms:</td>
<td></td>
</tr>
<tr>
<td></td>
<td>R'</td>
<td>CH₃</td>
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<tr>
<td></td>
<td>R—C—R''</td>
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<td>CH₃—C—CH₃</td>
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</table>

Figure 6 How to name an ester.
Figure 7  The general structure of the triesters found in fats and oils.

Figure 9  Representations of triglycerides.
Figure 18  Making a secondary amide.

Figure 20  The generalised structure of an α-amino acid.

Figure 21  How an amino acid forms a zwitterion.

Figure 22  A generalised coupling reaction.
Figure 2 Reactions of alkenes.

$\text{CH}_2=\text{CH}_2 \xrightarrow{\text{Cl}_2(g) \text{ sunlight}} \text{CH}_2=\text{CH}_2 + \text{HCl(g)}$

$\text{H}_2(g)$

finely divided Ni
at 150 °C and 5 atm.
(or Pt at room temp
and 1 atm.)

$\text{Cl}_2(g)$

$\text{Cl}_2(g)$

$\text{H}_2O(g)$

$\text{NaBH}_4$

$\text{Cr}_2\text{O}_7^{2-}/\text{H}^+(aq)$

$\text{NaOH(aq)}$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

Figure 3 Reactions of halogenoalkanes.

$\text{R} \xrightarrow{\text{NH}_3(aq)} \text{R} \xrightarrow{\text{NaOH(aq)}} \text{R} \xrightarrow{\text{H}_2\text{O(l)}} \text{R} \xrightarrow{\text{c. NH}_3(aq)} \text{R} \xrightarrow{\text{H}_2\text{O(l)}} \text{R} \xrightarrow{\text{c. NH}_3(aq)} \text{R}$

$\text{R} \xrightarrow{\text{H}_2\text{SO}_4} \text{R} \xrightarrow{\text{H}_2\text{SO}_4} \text{R} \xrightarrow{\text{H}_2\text{SO}_4} \text{R}

Figure 4 Reactions of primary alcohols.

$\text{R} \xrightarrow{\text{H}_2\text{O(l)}} \text{R} \xrightarrow{\text{c. NH}_3(aq)} \text{R} \xrightarrow{\text{NaOH(aq)}} \text{R} \xrightarrow{\text{c. H}_2\text{SO}_4} \text{R} \xrightarrow{\text{c. H}_2\text{SO}_4} \text{R} \xrightarrow{\text{c. H}_2\text{SO}_4} \text{R}$
Note: Esters and amides are both hydrolysed by heating with aqueous acid or aqueous alkali. Alkaline hydrolysis gives the salt of the corresponding carboxylic acid. The free carboxylic acid is formed on acidification of the solution.

Figure 5 Reactions of aldehydes.

Figure 6 Reactions of ketones.

Figure 7 The reactions of carboxylic acids and some related compounds.
Figure 8 Reactions of arenes.
Figure 1 Sequence of unit operations in a chemical plant.

Figure 2 Comparison of (a) batch and (b) continuous stirred tank reactors.

Figure 3 Pipelines from BP, Grangemouth, and Exxon, Mossmorran, for the distribution of ethene.
Figure 4 Feedstocks from natural gas.

Figure 5 Feedstocks from oil.

Figure 6 (a) An example of how energy can be used in a chemical process. (b) A diagram to illustrate a heat exchanger.