Chem Factsheet

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Key Definitions in AS Chemistry

To succeed in this topic you need to understand:
- the basics of AS level Chemistry

After working through this Factsheet you will be able to:
- define the key terms you will encounter in AS Chemistry
- understand and use the terms in their correct context

A word of warning

Make sure you have a copy of your particular AS specification, because it will tell you which definitions you need. If you don't know which specification you are following, ask your chemistry teacher.

Physical Chemistry Definitions

Physical chemistry deals with the physical properties of elements and compounds. At AS, this covers the topics of atomic structure, calculations, energetics, equilibria and kinetics.

Atomic Structure

Atomic number (Z) The number of protons in an atom. This is shown in the periodic table.

Isotopes Atoms of an element which have the same number of protons, but a different number of neutrons.

Mass (or nucleon) number (A) The number of protons plus the number of neutrons in an atom
.

Mass spectrometry A technique used to determine the relative atomic mass of elements and the relative molecular mass and structure of compounds.

Relative atomic mass (A_r) The average mass of all the isotopes of an element, relative to carbon-12. This is shown in the periodic table.

Relative formula mass (F_r) The sum of the relative atomic masses of the atoms in one formula unit of an ionic compound. The formula unit of sodium chloride is NaCl, so its relative formula mass is 23+35.5 = 58.5.

Relative isotopic mass
The mass of an isotope, relative to carbon-12.

Relative molecular mass (M_r) The sum of the relative atomic masses of all the atoms in a molecule. For example, the relative molecular mass of H_2O is 1+1+16 = 18.

Chemical Calculations

Avogadro's constant Avogadro's constant is 6.022×10^{23} mole^{-1}. It is the number of particles in one mole.

Empirical formula The simplest ratio of atoms in a compound. For example, the empirical formula of butane (C_4H_8) is CH_2.

Molar mass The mass in grams of one mole of a substance. The molar mass of carbon-12 is 12 g.

Mole One mole of any substance contains 6.022×10^{23} particles. This is the same as the number of atoms in exactly 12 g of carbon-12.

Molecular formula Formula which shows the actual number of atoms in one molecule of a compound.

Percentage atom economy

$$\frac{\text{Mass of useful product in balanced equation}}{\text{Total mass of reactants in balanced equation}} \times 100$$

Percentage yield

$$\frac{\text{Actual mass of product}}{\text{Theoretical maximum mass of product}} \times 100$$

Equilibria and Kinetics

Activation energy The minimum amount of energy two particles must have for a reaction between them to occur when they collide.

Catalyst A substance that increases the rate of a chemical reaction but remains chemically unaltered at the end. A catalyst increases the rate of a reaction by providing an alternative route through the reaction with lower activation energy.

Dynamic equilibrium A reversible reaction in which forward and reverse reactions are happening at the same rate, so that the concentrations of reactants and products do not change.

Heterogeneous catalyst A catalyst in a different phase from the reaction mixture. For example, nickel (solid phase) in the hydrogenation of alkenes (gaseous phase) to make alkanes.

Homogeneous catalyst A catalyst in the same phase as the reaction mixture. For example, aqueous H^+ ions in the reaction between ethanol and ethanoic acid to form the ester, ethyl ethanoate.

Le Chatelier's principle If a system at dynamic equilibrium is subjected to a change, the position of equilibrium will shift so as to counteract the change. For example, if the temperature of a reaction at equilibrium is increased, the position of equilibrium will shift in the endothermic direction so as to reduce the temperature again.

Energetics

Endothermic A process which absorbs heat energy. Endothermic processes have a positive enthalpy change, ΔH.

Enthalpy Heat content of a substance or system measured at constant pressure.

Enthalpy change, ΔH The change in heat energy at constant pressure.

Exothermic A process which gives out heat energy. Exothermic processes have a negative enthalpy change, ΔH.

Hess's Law The total energy change for a process is independent of the route taken.

Standard conditions Atmospheric pressure (101kPa) and a specified temperature, normally 298K (25°C). Under standard conditions, solutions have a concentration of 1 moldm^{-3}, and all substances are in their normal physical state at the specified temperature.

Standard enthalpy of atomisation ($\Delta H^{\ominus}_{at}$) The enthalpy change when one mole of atoms is formed from the element in its standard state, measured under conditions of standard temperature and pressure. e.g. the standard enthalpy change for $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

Standard enthalpy of combustion ($\Delta H^{\ominus}_{c}$) The enthalpy change when one mole of a substance burns completely in oxygen, measured under conditions of standard temperature and pressure. e.g. the standard enthalpy of combustion of methane is the standard enthalpy change for $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

Standard enthalpy of formation ($\Delta H^{\ominus}_{f}$) The enthalpy change when one mole of a substance is formed from its elements in their standard states, measured under conditions of standard temperature and pressure. e.g. the standard enthalpy of formation of sodium chloride is the standard enthalpy change for:

$$Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$$

Standard enthalpy of neutralisation ($\Delta H^{\ominus}_{neut}$) The enthalpy change when an acid and a base react to form one mole of water, measured under conditions of standard temperature and pressure. e.g. the standard enthalpy change for :

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

Inorganic Chemistry Definitions

Inorganic chemistry covers the reactions and properties of elements other than carbon. At AS, this includes types of bonding, oxidation and reduction, and periodic trends across and down the periodic table.

Types of Bonding and Intermolecular Forces

Covalent bond A pair of electrons shared between two atoms.

Dative covalent (co-ordinate) bond A covalent bond in which both electrons come from the same atom.

Dipole-dipole force Intermolecular forces resulting from the electrostatic attractions between molecules possessing permanent dipoles.

Hydrogen bonds Intermolecular forces between a δ+ hydrogen atom bonded to a N, O or F atom and the lone pair of an atom such as oxygen, nitrogen or fluorine in a neighbouring molecule. For molecules of similar size, these are the strongest of the three types of intermolecular force.

Ionic bonding The strong electrostatic attraction between oppositely charged ions.

Metallic bonding The attraction between a lattice of positive ions and a sea of delocalised outer electrons.

Permanent dipole An uneven sharing of the electrons in a covalent bond, brought about by one atom in the bond being more electronegative than the other. For example, in a C-O bond the oxygen has the greater share of the electrons, and is therefore slightly negative (δ-) since it is more electronegative than carbon, which is therefore slightly positive (δ+). i.e. $^{ä+}$C-O$^{ä-}$

Polar bond A covalent bond which contains a permanent dipole.

Van der Waal's forces Induced dipole-dipole forces between **all** covalent molecules. For molecules of similar size, these are the weakest of the three types of intermolecular force.

Redox

Disproportionation A redox reaction in which the same element is both oxidised and reduced.

$$\text{e.g. } Cl_2\,(g) + H_2O\,(l) \rightarrow HOCl\,(aq) + HCl\,(aq)$$

Ore A rock which contains a metal compound, usually a sulphide or oxide, in high enough proportions to make it worth mining and extracting.

Oxidation Loss of electrons (see also Types of Reaction).

Oxidation state (or number) A numerical way of expressing the relative degrees of oxidation of atoms. More positive numbers indicate a greater state of oxidation and vice versa.

For example, the oxidation state of Na in NaCl is +1, while the oxidation state of the Cl is –1. The sum of the oxidation states of the atoms in a compound is zero.

Oxidising agent An electron acceptor.

Reducing agent An electron donor.

Reduction Gain of electrons (see also Types of Reaction).

Periodic trends

Alkali metal A Group 1 element. e.g. lithium, sodium and potassium.

d Block The elements, including the transition metals, in the middle of the periodic table. The outer electrons of the d block elements are in a d subshell.

Electronegativity The ability of an atom to attract the bonding electrons in a covalent bond.

First electron affinity The standard enthalpy change when one mole of electrons is added to one mole of gaseous atoms. e.g. the first electron affinity of O is the standard enthalpy change for

$$O(g) + e^- \rightarrow O^-(g)$$

First ionisation energy The standard enthalpy change to remove one mole of electrons from one mole of gaseous atoms. e.g. the first ionisation energy of Na is the standard enthalpy change for

$$Na(g) \rightarrow Na^+(g) + e^-$$

Halogen A Group 7 element. e.g. fluorine, chlorine and bromine.

Noble gas A Group 0 element. e.g. helium, neon and argon.

p Block Groups 3-0 of the periodic table. The outer electrons of the p block elements are in a p subshell.

s Block Groups 1 and 2 of the periodic table. The outer electrons of the s block elements are in an s subshell.

Second electron affinity The standard enthalpy change when one mole of electrons is added to one mole of gaseous negatively charged ions. e.g. the second electron affinity of O is the standard enthalpy change for $O^-(g) + e^- \rightarrow O^{2-}(g)$.

Second ionisation energy The standard enthalpy change to remove one mole of electrons from one mole of gaseous singly positively charged ions. e.g. the second ionisation energy of Na is the standard enthalpy change for $Na^+(g) \rightarrow Na^{2+}(g) + e^-$.

Transition metal An element which forms at least one stable ion with an incompletely filled d subshell.

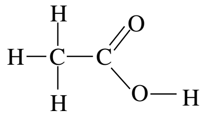
Organic Chemistry

Organic chemistry is the chemistry of covalently bonded carbon compounds excluding carbon dioxide, carbon monoxide and their derivatives such as carbonates.

At AS level, this covers structures, isomerism, basic reaction types, and the physical and chemical properties of some simple compounds, usually alkanes, alkenes, alcohols and halogenoalkanes.

Structures

Displayed formula The displayed formula shows the positions of all atoms and all of bonds between them. For example, the displayed formula of ethanoic acid is:



Functional group The part of a molecule which is involved in most of its reactions. In an alcohol, for example, the functional group is an OH group.

General formula The general formula is used to represent any member of a homologous series. For example C_nH_{2n+2} for alkanes.

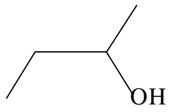
Homologous series A family of compounds with similar chemical properties, whose members differ from each other by the addition of a CH_2 group. A homologous series, like the alkanes, shows a steady increase in physical properties from one member to the next.

Hydrocarbon A compound containing carbon and hydrogen *only*.

Infrared spectroscopy A technique used to gain information about the functional groups present in an organic compound. Different functional groups absorb different frequencies of infrared energy allowing identification.

Saturated Compounds containing only single bonds. e.g. alkanes are saturated hydrocarbons.

Skeletal formula This shows just the carbon skeleton and attached functional groups, for example, the skeletal formula of butan-2-ol is:



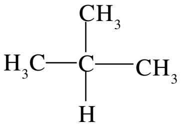
Structural formula The minimum detail needed for an unambiguous structure. e.g. structural formula of ethanol is CH_3CH_2OH rather than C_2H_6O.

Unsaturated Compounds containing double and triple bonds. Alkenes, aldehydes and ketones are unsaturated compounds.

Isomerism

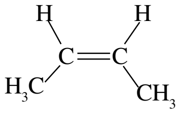
Chain isomer Isomers which belong to the same homologous series but have different carbon chains e.g. butane and methylpropane.

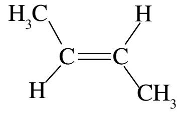
$$H_3C—CH_2—CH_2—CH_3$$



$$H_3C-\overset{\displaystyle CH_3}{\underset{\displaystyle H}{C}}-CH_3$$

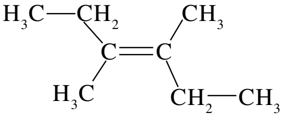
***cis-trans* (geometric) isomers** Isomers of alkenes with groups on different sides of the double bond, for example *cis*-but-2-ene and *trans*-but-2-ene.



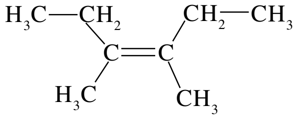


Cis and *trans* refer to alkenes with two hydrogens and two non-hydrogen groups around the double bond.

***E-Z* (geometric) isomers** Isomers of alkenes with groups on different sides of the double bond. *E* (for *entgegen*, German for against) and *Z* (for *zusammen*, German for together) refer to alkenes with two different groups attached to each carbon of the double bond.

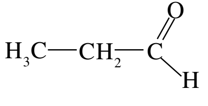


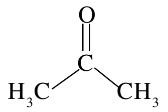
E-3,4-dimethylhex-3-ene



Z-3,4-dimethylhex-3-ene

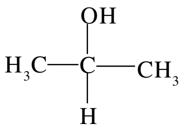
Functional group isomer Compounds with the same molecular formula but different functional groups, for example propanal and propanone.





Position isomer Isomers with the same functional group at different positions on the carbon chain. For example propan-1-ol and propan-2-ol.

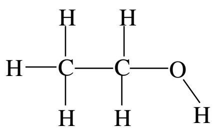
$$H_3C—CH_2—CH_2—OH$$

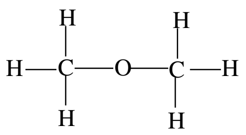


$$H_3C-\overset{\displaystyle OH}{\underset{\displaystyle H}{C}}-CH_3$$

Stereoisomers Isomers with a different spatial arrangement of functional groups. At AS, the only examples of stereoisomerism you come across are *E-Z* and *cis-trans* isomerism.

Structural isomers Compounds with the same molecular formula, but different structural formulae. For example ethanol and dimethyl ether.





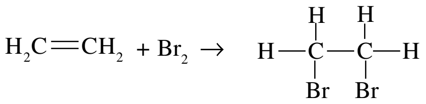
Types of Reaction

Catalytic cracking Breaking of long chain alkanes using slightly increased pressures, high temperatures and a zeolite catalyst to make high proportions of short alkanes for fuel and aromatic compounds.

Cracking Breaking of long alkane chains into smaller molecules.

Electrophile An electron pair acceptor e.g Br^+

Electrophilic addition A reaction in which an electrophile adds to a compound, to make a single product. For example the reaction of ethene with bromine.



$$H_2C{=}CH_2 + Br_2 \rightarrow H-\overset{\displaystyle H}{\underset{\displaystyle Br}{C}}-\overset{\displaystyle H}{\underset{\displaystyle Br}{C}}-H$$

Elimination A reaction in which a compound (often water) is removed from an organic compound to create an unsaturated molecule. For example, the conversion of an ethanol to ethene.

$$CH_3CH_2OH \rightarrow H_2C{=}CH_2 + H_2O$$

Free radical A species with an unpaired electron e.g $Cl^{\cdot}$

Free radical substitution A reaction in which one species swaps places with another, by a mechanism involving free radicals. For example the chlorination of methane.

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

Heterolytic fission Breaking of a covalent bond in which both electrons go to the same atom, resulting in a cation and an anion.

Homolytic fission Breaking of a covalent bond in which the electrons go to separate atoms, resulting in free radicals.

Hydration Reaction with water resulting in an addition reaction. For example, the reaction of ethene with steam to make ethanol.

$${}_2HC{=}CH_2 + H_2O \rightarrow CH_3CH_2OH$$

Hydrolysis Reaction with water resulting in a substitution reaction. For example, the reaction of a 1-bromopropane with water.

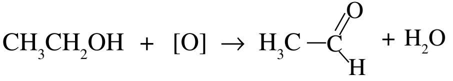
$$CH_3CH_2CH_2Br + H_2O \rightarrow CH_3CH_2CH_2OH + HBr$$

Nucleophile An electron pair donor e.g **:**NH_3.

Nucleophilic substitution A reaction in which a nucleophile swaps places with another functional group For example the reaction of bromoethane with an OH^- ion.

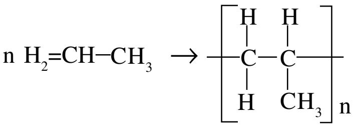
$$CH_3CH_2Br + OH^- \rightarrow CH_3CH_2OH + Br^-$$

Oxidation Addition of oxygen, or loss of hydrogen.For example the transformation of a ethanol (a primary alcohol) to ethanal (analdehyde).



$$CH_3CH_2OH + [O] \rightarrow H_3C-\overset{\displaystyle O}{\underset{\displaystyle H}{C}} + H_2O$$

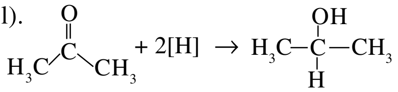
Polymerisation Reaction where many small monomer molecules join together to make polymers. e.g. the polymerisation of propene.



$$n\,H_2{=}CH{-}CH_3 \rightarrow \left[\overset{\displaystyle H}{\underset{\displaystyle H}{C}}-\overset{\displaystyle H}{\underset{\displaystyle CH_3}{C}}\right]_n$$

This is also called **addition** polymerisation since it involves repeated addition reactions.

Reduction Loss of oxygen, or addition of hydrogen. For example the transformation of a propanone (a ketone) to propan-2-ol (a secondaryalcohol).



$$H_3C-\overset{\displaystyle O}{\overset{\|}{C}}-CH_3 + 2[H] \rightarrow H_3C-\overset{\displaystyle OH}{\underset{\displaystyle H}{C}}-CH_3$$

Thermal cracking Breaking of long alkane chains using high pressure and high temperatures to make high proportions of alkenes.

Green Chemistry

Biofuel A fuel which is derived from sustainably grown plant matter. For example, ethanol derived from plant sugar by fermentation.

Carbon neutral A process which has no net annual emissions of greenhouse gases, including carbon dioxide.

Greenhouse effect The process by which energy from the Sun is trapped in the Earth's atmosphere, keeping it warm. Gases which contribute to this include carbon dioxide, water vapour and methane.

Greenhouse gas A gas which contributes to the greenhouse effect. The most common greenhouse gases are carbon dioxide, methane and water vapour.

Parts per million (ppm) A measure of the concentration of particles, for example pollutants in water or ozone in the atmosphere. One part per million means that in a sample of one million particles (of water, or air for example), there is one particle of the substance under consideration.

Practice Questions

1. Match up these key words and phrases (a-f) with their definitions (1-6):

a) Molar mass	1	Formula which shows the actual number of atoms in one molecule of a compound.
b) Atomic number	2	The number of protons in an atom.
c) Isotopes	3	The mass in grams of one mole of a substance.
d) Relative atomic mass	4	The number of protons plus the number of neutrons in an atom.
e) Mass number	5	The average mass of all the isotopes of an element, relative to carbon-12.
f) Molecular formula	6	Atoms of an element which have the same number of protons, but a different number of neutrons.

2. Define each of these enthalpy changes:
   a) $Mg\ (s) + \frac{1}{2}O_2(g) \rightarrow MgO\ (s)$
   b) $C_3H_8\ (g) + 5O_2(g) \rightarrow 3CO_2\ (g) + 4H_2O(l)$
   c) $S^-\ (g) + e^- \rightarrow S^{2-}(g)$
   d) $Mg\ (g) \rightarrow Mg^+(g) + e^-$
   e) $Al^{2+}\ (g) \rightarrow Al^{3+}(g) + e^-$
   f) $F\ (g) + e^- \rightarrow F^-(g)$

3. Classify each of these organic reaction types:
   a) $CH_3CH_2CH_2OH \rightarrow CH_3CH_2COOH$
   b) $CH_3Br + H_2O \rightarrow CH_3OH + HBr$
   c) $CH_3CH{=}CH_2 + Br_2CH_3CHBrCH_2Br$
   d) $CH_3CHO \rightarrow CH_3CH_2OH$
   e) $CH_3CH_2CH_2OH \rightarrow CH_3CH{=}CH_2$
   f) $CH_3Cl + Cl_2 \rightarrow CH_2Cl_2 + HCl$

Answers

1. a – 3 ; b – 2 ; c – 6 ; d – 5 ; e – 4 ; f – 1

2. a) Standard enthalpy of formation of MgO
   b) Standard enthalpy of combustion of propane
   c) Second electron affinity of sulphur
   d) First ionisation energy of magnesium
   e) Third ionisation energy of aluminium
   f) First electron affinity of fluorine

3. a) Oxidation
   b) Hydrolysis *and* nucleophilic substitution
   c) Electrophilic addition
   d) Reduction
   e) Elimination
   f) Free radical substitution

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