Topic 1: Atomic Structure and the Periodic Table

Edexcel A Level Chemistry • Paper 1 (9CH0/01)

1.1	Know the structure of an atom in terms of electrons, protons and neutrons. Atoms have a small, dense, positively charged nucleus containing protons and neutrons, surrounded by electrons in quantum shells. Most of the atom is empty space.
1.2	Know the relative mass and relative charge of protons, neutrons and electrons. Proton: relative mass = 1, relative charge = +1 Neutron: relative mass = 1, relative charge = 0 Electron: relative mass = 1/1836 (≈0), relative charge = −1
1.3	Know what is meant by the terms 'atomic (proton) number' and 'mass number'. Atomic (proton) number (Z): the number of protons in the nucleus of an atom. This defines the element. Mass number (A): the total number of protons and neutrons in the nucleus. Be able to determine the number of each type of sub-atomic particle in an atom, molecule or ion
1.4	from the atomic (proton) number and mass number. Protons = Z (atomic number) Neutrons = A – Z (mass number minus atomic number) Electrons = Z for a neutral atom. For ions: subtract the charge for cations, add the charge for anions.
1.5	Understand the term 'isotopes'. Isotopes are atoms of the same element (same number of protons) with different numbers of neutrons, and therefore different mass numbers. They have identical chemical properties but different physical properties (e.g. density, rate of diffusion). Be able to define the terms 'relative isotopic mass' and 'relative atomic mass', based on the ^{12}C

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1.6 scale.

Relative isotopic mass: the mass of one atom of an isotope compared to 1/12 of the mass of one atom of carbon-12.

Relative atomic mass (Ar): the weighted mean mass of one atom of an element compared to 1/12 of the mass of one atom of carbon-12.

Understand the terms 'relative molecular mass' and 'relative formula mass', including

1.7 calculating these values from relative atomic masses.

Relative molecular mass (Mr): the sum of the relative atomic masses of all atoms in a molecular formula. Used for simple covalent molecules.

Relative formula mass: the sum of the relative atomic masses of all atoms in a formula unit. Used for compounds with giant structures (e.g. ionic compounds like NaCl).

Be able to analyse and interpret data from mass spectrometry to calculate relative atomic mass

1.8 from relative abundance of isotopes and vice versa.

Ar = Σ(isotopic mass × relative abundance) / Σ(relative abundances)

In a mass spectrum, the x-axis is m/z and the y-axis is relative abundance (%). Each peak represents an isotope. Multiply each mass by its abundance, sum, and divide by total abundance.

Be able to predict the mass spectra, including relative peak heights, for diatomic molecules,

1.9 including chlorine.

For Cl_2: chlorine has two isotopes, ^{35}Cl (75%) and ^{37}Cl (25%).

Possible combinations: ^{35}Cl–^{35}Cl (m/z = 70), ^{35}Cl–^{37}Cl (m/z = 72), ^{37}Cl–^{37}Cl (m/z = 74).

Peak height ratio = 9 : 6 : 1 (using probability: $0.75^2 : 2\times0.75\times0.25 : 0.25^2$).

Understand how mass spectrometry can be used to determine the relative molecular mass of a

1.10 molecule.

The molecular ion peak (M^+) is the peak with the highest m/z value (ignoring the M+1 peak). This gives the relative molecular mass of the molecule directly. The M^+ peak is formed when the molecule loses one electron in the mass spectrometer.

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1.11 Be able to define the terms 'first ionisation energy' and 'successive ionisation energies'.

First ionisation energy: the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous 1+ ions.

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

Successive ionisation energies: the energies required to remove each subsequent electron from the increasingly positive ion.

Understand how ionisation energies are influenced by the number of protons, the electron

1.12 shielding and the electron sub-shell from which the electron is removed.

Ionisation energy increases with:

- More protons in the nucleus (greater nuclear attraction)
- Less shielding from inner electrons (outer electron feels greater effective nuclear charge)
- Electron removed from a sub-shell closer to the nucleus (e.g. s < p in the same shell for removal difficulty)

Electrons in the same sub-shell experience similar shielding.

1.13 Understand reasons for the general increase in first ionisation energy across a period.

Across a period, nuclear charge increases (more protons) but shielding remains approximately constant (electrons are added to the same shell). This means the effective nuclear charge increases, so outer electrons are held more tightly, and more energy is required to remove them.

Exceptions: drop from Group 2 to 3 (new sub-shell, p electron easier to remove) and Group 5 to 6 (paired electron in p-orbital experiences repulsion, easier to remove).

1.14 Understand reasons for the decrease in first ionisation energy down a group.

Down a group: more electron shells means the outer electron is further from the nucleus and there is increased shielding from inner electrons. The effective nuclear charge decreases, so less energy is needed to remove the outer electron, despite the increase in nuclear charge.

Understand how ideas about electronic configuration developed from: (i) atomic emission spectra provide evidence for quantum shells, (ii) successive ionisation energies provide

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1.15 Understand how ideas about electronic configuration developed from: (i) atomic emission spectra provide evidence for quantum shells, (ii) successive ionisation energies provide evidence for quantum shells and the group, (iii) first ionisation energy of successive elements provides evidence for electron sub-shells.

provides evidence for electron sub-shells.

(i) Atomic emission spectra show discrete lines (not continuous), proving electrons can only exist at specific, quantised energy levels. Each line corresponds to a transition between shells.

(ii) A plot of successive IEs for an element shows large jumps where electrons are removed from a shell closer to the nucleus. The number of electrons removed before the first big jump = the group number.

(iii) Plotting first IE across a period shows sub-shell structure: drops between Groups 2→3 (s→p) and 5→6 (half-full→paired p) indicate different sub-shell energies.

1.16 Know the number of electrons that can fill the first four quantum shells.

Shell 1: 2 electrons ($1s^2$)

Shell 2: 8 electrons ($2s^2\ 2p^6$)

Shell 3: 18 electrons ($3s^2\ 3p^6\ 3d^{10}$)

Shell 4: 32 electrons ($4s^2\ 3d^{10}\ 4p^6\ 4d^{10}\ 4f^{14}$)

Maximum electrons per shell = $2n^2$.

Know that an orbital is a region within an atom that can hold up to two electrons with opposite

1.17 spins.

An orbital is a region of space around the nucleus where there is a high probability of finding an electron. Each orbital holds a maximum of 2 electrons, which must have opposite (antiparallel) spins (↑↓)

1.18 Know the shape of an s-orbital and a p-orbital.

s-orbital: spherical shape, centred on the nucleus.

p-orbital: dumbbell (figure-of-eight) shape, with a node at the nucleus. There are three p-orbitals per sub-shell (px, py, pz), oriented at 90° to each other.

1.19 Know the number of electrons that occupy s, p and d-subshells.

s-subshell: 1 orbital → 2 electrons

p-subshell: 3 orbitals → 6 electrons

d-subshell: 5 orbitals → 10 electrons

Know that electrons fill subshells singly, before pairing up, and that two electrons in the same

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1.20 orbital must have opposite spins.

Electrons occupy orbitals singly first with parallel spins before pairing up. When electrons do pair, they must have opposite spins. This minimises electron-electron repulsion.

Be able to predict the electronic configurations, using 1s notation and electrons-in-boxes

1.21 notation, of: (i) atoms up to Z = 36, (ii) ions (s and p block only) up to Z = 36.

Fill in order: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Key anomalies: Cr = [Ar] $3d^5\ 4s^1$ and Cu = [Ar] $3d^{10}\ 4s^1$ (half-full/full d-subshell stability).

For ions: remove electrons from the outermost shell first (i.e. 4s before 3d for d-block cations).

Electrons-in-boxes: draw boxes for each orbital, fill with arrows (↑↓) singly first with parallel spins before pairing up.

1.22 Know that elements can be classified as s, p and d-block elements.

Elements are classified by the sub-shell into which the last electron is added:

- s-block: Groups 1 and 2 (including H and He)
- p-block: Groups 3, 4, 5, 6, 7, 0
- d-block: transition metals

1.23 Understand that electronic configuration determines the chemical properties of an element.

Elements with the same outer shell configuration show similar chemical properties. This is why elements in the same group of the Periodic Table react in similar ways. E.g. all Group 1 metals have one electron in their outer s-subshell and form 1+ ions.

1.24 Understand periodicity in terms of a repeating pattern across different periods.

Periodicity is the regular recurring variation in properties of elements with increasing atomic number. As each new period begins, the pattern of properties repeats because the outer shell electronic configuration repeats.

Understand reasons for the trends in: (i) ionisation energy across Periods 2 and 3, (ii) melting

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1.25 and boiling temperatures across Periods 2 and 3.

(i) Ionisation energy: general increase across the period due to increasing nuclear charge with constant shielding. Drops at Group 3 (p-electron easier to remove than s) and Group 6 (paired p-electron repulsion).

(ii) Melting/boiling points: metals (Na, Mg, Al) increase due to increasing charge density and number of delocalised electrons → stronger metallic bonding. Si has very high m.p. (giant covalent). P_4, S_8, Cl_2, Ar are simple molecular – weak London forces, so low m.p. $S_8 > P_4 > Cl_2 > Ar$ (more electrons → stronger London forces).

Be able to illustrate periodicity using data, including electronic configurations, atomic radii,

1.26 melting and boiling temperatures and first ionisation energies.

You must be able to interpret and sketch graphs showing periodic trends. Look for repeating patterns across Periods 2 and 3. Atomic radii decrease across a period (increased effective nuclear charge pulls electrons closer). All properties show a repeating pattern that resets at the start of each new period.

ABOUT THESE NOTES

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