**Vocabulary sheet**

Atoms, their properties and how they bond together

## Topic 2 and 12 – Atom structure

Atomic emission spectra: The characteristic line spectrum that occurs as a result of energy being supplied to individual elements. When energy is supplied to an atom, e- are **excited** from their ground state to a higher energy level. The e- **dropping** from higher energy levels to lower **emit** energy, which can be observed in a spectrum. As e- can only exist in **fixed** energy levels, the energy in the emissions are **characteristic** for each type of atom. Lines converge toward high energy end of spectrum as the energy levels themselves are convergent.

Atomic number: Number of protons in the nucleus.

Aufbau principle: The principle that states that lowest energy levels are filled first. The starting order is 1s2s2p3s3p4s3d…

Hund’s rule: Orbitals within the same sub-shell are filled singly first.

Ionization energy, first: Energy required to remove one e-­ from an atom in its gaseous state. Increases across a period (due to increasing **nuclear** **attraction**), except for slight drops (due to the commencement of **pairing** of **e-** into orbitals → greater electron repulsion). Decreases down a group (due to increased number of **energy** **levels** → increased electron **shielding**).

Isotope: An atom that contains the same number of protons but a different number of neutrons.

Mass number: Total number of nucleons.

Mass spectrometer: A device for determining relative atomic masses and their relative abundances. Composed of:

 1) **vaporizer** (where the sample is vaporized),

 2) **ionizer** (where atoms are bombarded with high energy e- → e- knocked off the atoms. Only ions with a +1 charge are formed, in practice),

 3) **accelerator** (where an electric field accelerates them),

 4) **deflector** (where a magnetic field deflects them; the smaller the mass and the higher the charge, the greater the deflection) and

 5) **detector** (which measures both the mass and relative amounts of ions present.

Orbital: A region of space around the nucleus of an atom that is occupied by a maximum of two e- at any given time. The types of orbitals are s (spherical), p (dumbbell-shaped), d and f.

Pauli’s exclusion principle: e- in single orbital must have opposite spin.

Relative abundances: The percentage of natural occurrence of an isotope of an element.

Relative atomic mass: The weighted mean of all the naturally occurring isotopes of the element relative to 12C. No units.

Valence electrons: The e- in the highest main energy level.

## Topic 3 and 13 – Periodicity

Alkali metals: Group 1 elements. Strength of metallic bond gets weaker as atoms get larger.

Anion: A – charged ion. Decrease in size across a period as they are isoelectronic but have an increased nuclear charge. Contain more e- than their parent atom, so they are larger in size.

Atomic radius: Half the distance between the nuclei of two bonded atoms of the same element. Decreases across period as nuclear charge increases. Increases down a group because of greater number of energy levels.

Catalyst: A substance that increases the rate of reaction while being recoverable in the end stage of the reaction, unchanged. Examples of catalytic transition metals: Fe is used in Haber process; V2O5 in Contact process; Ni in hydrogenation reactions; MnO2 with hydrogen peroxide.

Cation: A + charged ion. Decrease in size across a period. Contain less e- than their parent atom, so they are smaller in size.

Chloride: Chlorides of **Na** and **Mg** have an **ionic** structure, while all the **remaining** chlorides of period 3 elements have **simple** **covalent** structures. Still, Al2Cl6 can act as a poor conductor. NaCl dissolves in water to give a neutral solution, MgCl2 gives slightly acidic solution, while all other chlorides react vigorously with water to produce acidic solutions of HCl together with fumes of hydrogen chloride.

Colored complex: A complex is a compound in which molecules or ions form dative bonds to a metal atom or ion. Colors are due to e- transitions between different d orbitals.

Co-ordination number: Number of lone pairs bonded to the metal ion. Cl- often gives 4 coordinate bonds, CN- gives 6, H2O gives 6 and NH3 gives 4 or 6.

D-d transition: The excitation of an e- from the lower of two energy level groups to the upper. As the ligand approaches the metal, the five d orbitals **split**. The difference in energy between the two levels corresponds to the wavelength of visible light, as **incoming** light causes energy of a particular wavelength to be **absorbed** by e- that are **excited** from the lower level to the higher. The amount that the d orbitals are split will determine the exact color.

Degenerate: Of equal energy. In a free ion, the five d orbitals are degenerate.

Electronegativity: Relative measure of the attraction that an atom has for a shared pair of e- when it is covalently bonded to another atom. Inversely proportional to size of atom → F is most electronegative.

Group: Elements with the same number of valence e-.

Halide ions: Ions of the halogens. Their presence can be detected by the addition of silver nitrate. AgCl is white, AgBr is cream-colored, and AgI is yellow. Silver halides react with light to form silver metal.

Isoelectronic: Containing the same number of e-.

Ligand: A molecule or ion that can donate e-.

Melting point: Depends on both the **structure** of element and type of **attractive** forces holding the atoms together. Increases across the period while following elements which exhibit metallic bonding (due to increased strength following increased number of **valence** e-.) Highest m.p. of period lies with **macromolecular** **covalent** structure (with very strong bonds). Sharp decrease in m.p. with elements that exhibit **simple** **molecular** structures with only van der Waals’ forces between them. Lowest m.p. with noble gases, which are **monatomic**. Down Group 1, m.p. decreases as atoms get larger and strength of metallic bond get weaker. Down Group 7, van der Waals’ forces increase as molar mass increases → m.p. increases.

Metalloid: An element that possesses some of the properties of a metal and some of a non-metal. While metal oxides tend to be basic and non-metal oxides tend to be acidic, metalloid oxides such as aluminium oxide can be amphoteric.

Oxide: **Metal** oxides are **ionic** (explains conductivity, m.p. and b.p.) and react with water to form alkaline solutions. **SiO2** is covalent **macromolecular**, and the remaining oxides are **simple** **covalent**. Oxides of period 3 elements are **solid** from Na2O to P4O10 and P4O6, while SO3 and Cl2O7 are **liquid** and SO2 and Cl2O are **gases** at 25°C. Al2O3 and SiO2 do not react with water, but **Al2O3** can act **amphoterically** depending on whether it is reacted with a base or an acid. Remaining oxides are **acidic**.

Period: Elements with the same outer shell of valence e-.

Periodicity: The repeating pattern of physical and chemical properties as shown by the different periods.

Structure, giant covalent: Very hard but brittle. Very high m.p. and b.p. Do not conduct in any state. Insoluble.

Structure, giant ionic: Hard but brittle. High m.p. and b.p. Conduct when molten or aqueous, but not as solids.

Structure, giant metallic: Malleable, not brittle. M.p. and b.p. dependent on no. of valence e-. Good conductivity.

Structure, molecular covalent: Usually soft and malleable unless hydrogen bonded. Low m.p. and b.p. Do not conduct in any state. Often soluble in non-aqueous solvents, unless they can hydrogen bond to water.

Transition element: An element that possesses an incomplete d sub-level in one or more of its oxidation states. Often very efficient catalysts as they can exist in a variety of oxidation states (all except Ti have oxidation state of +2). Form colored complexes.

## Topic 4 and 14 – Chemical bonding and structure

Allotrope: Occur when an element can exist in different crystalline forms, such as in carbon, which can exist as graphite, buckminsterfullerene and diamond (ooh, pwetty). Diamond is exceptionally hard because there is no plane of weakness in the molecule made up of sp3 hybridized carbon atoms. In graphite, the carbon atoms are sp2 hybridized. Remaining e- after the three σ bonds, are delocalized, resulting in the fact that graphite is an excellent lubricant and a good conductor of electricity.

Bond length and strength: Depends on strength of attraction that two nuclei have for the shared e-. Generally, the stronger the bond, the shorter its length.

Bond polarity: A polarity caused by a difference in electronegativity between the elements. The greater the difference, the greater the polarity.

Bond, π: Pi bond. A radial bond caused by the sideways overlapping of p-orbitals. Has two regions of e- density. Double bonds have one π bond, while triple bonds have two which are perpendicular to each other.

Bond, σ: Sigma bond. An axial bond caused by orbitals overlapping along a line drawn through the two nuclei. Single, double and triple bonds have one σ bond.

Bond, covalent: Bonding by the sharing of e-. The e- are shared and attracted by both nuclei resulting in a directional bond between the two atoms.

Bond, dative: A bond in which both e- come from one of the atoms. Also known as coordinate bond.

Bond, ionic: A bond by which e- are transferred from one atom to another to form ions with complete outer shells. In an ionic compound the + and – ions are attracted to each other by the electrostatic force between them, and build up into a strong lattice. Have relatively high m.p. Ionic bonds occur between elements with a great difference (>1.8) in electronegativity.

Conductivity: The extent to which a substance can conduct electricity. Must possess e- or ions that are free to move.

Delocalization: The sharing of a single e- pair by more than two atoms.

Forces, dipole:dipole: Permanent electrostatic forces of attraction between polar molecules. Stronger than van der Waals’.

Forces, Hydrogen bonding: Occurs when hydrogen is bonded directly to a highly electronegative element (N, F, or O). Stronger than dipole:dipole forces.

Forces, van der Waal’s: Temporary dipole forces due to momentary unevenness in spread of e-. Weakest of intermolecular forces. Increase with increasing molar mass.

Formal charge theory: 1) If compound is **covalent**, treat the entire molecule; if **ionic**, treat each ion separately.

 2) Determine total no. of **valence** e- available.

 3) Organize atoms such that there is a **central** atom (usually the most electronegative) surrounded by **ligands**.

 4) i) One **pair** of e- between central atom and each ligand atom, ii)**three** more **pairs** on each **ligand** (except hydrogen), iii)**remaining** e- pairs around **central** atom.

 5) Calculate **formal** **charge** on central atom. Should be equal to charge on species, otherwise form a **multiple** **bond** with ligand atom that has most – formal charge. Repeat 5.

Formal charge: A hypothetical measure of the number of e- originating from an atom. F = V – (n + b/2), where V is valence e-, n is e- in lone pairs and b is e- shared as bonds.

Hybridization: The combination of orbitals to create new orbitals that are more energetically feasible for bonding. Bonds that are sp3 hybridized have a regular bond angle of 109.5°, sp2 have 120°, and sp have 180°.

Lewis structure: Diagram showing arrangement of e- in a molecule. Usually only shows valence shells.

Metallic bonding: The valence e- in metals become detached from the individual atoms so that the metals consist of a closely packed lattice of + ions in a ‘sea’ of delocalized e-. Forces of attraction are between ions and e- and not between the ions themselves, which means that metals are malleable and ductile.

Molecular polarity: Depends on both the bond polarity and the symmetry.

Resonance hybrid: Structures that arise from the possibility to draw a multiple bond in different positions equivalently. Can be better explained by delocalization.

Solubility: The extent to which one substance dissolves in another. ‘Like tends to dissolve like.’

VSEPR theory: Valence Shell Electron Pair Repulsion theory. States that pairs of e- arrange themselves around the central atom so that they are as far apart from each other as possible. Greater repulsion between lone pair of e- than bonded pairs.