WJEC (Wales) Chemistry A-level

Unit 1: The Language of Chemistry, Structure of Matter and Simple Reactions

Definitions and Concepts

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Unit 1: The Language of Chemistry, Structure of Matter and Simple Reactions

1.1 - Formulae and equations

Atom: The smallest part of an element that can exist. All substances are made up of atoms.

Compound: A compound is a substance that combines two or more different elements through the formation of chemical bonds.

Ion: An ion is formed when an atom/molecule loses or gains electrons. This gives it an overall charge - a positive charge if it has lost at least one electron and a negative charge if it has gained at least one electron.

Ionic equation: A chemical equation that involves dissociated ions.

Molecular formula: The actual number of atoms of each element in a molecule.

Oxidation number: The charge of an ion or a theoretical charge of an atom in a covalently bonded compound assuming the bond becomes ionic.

State Symbol: State symbols show the physical state of the substance during the reaction, they are usually in brackets: gas (g), liquid(l), solid(s) and aqueous(aq). Aqueous means the substance is dissolved in water.

1.2 - Basic ideas about atoms

Absorption spectra: A spectrum of frequencies of electromagnetic radiation that has been transmitted through an atom or molecule, that shows dark bands due to the absorption of the radiation at those specific wavelengths.

Alpha-decay: A type of radioactive decay, during which an atomic nucleus loses two protons and two neutrons. An alpha particle is equivalent to a helium nucleus. It reduces the atomic number by two and the mass number by four, making the element more stable.

Beta-decay: A type of radioactive decay, during which a beta particle is lost, which is equivalent to an electron and a neutron turns into a proton or a proton turns into a neutron. This changes the atomic number by one, but the mass number remains the same.

Electromagnetic spectrum: The range of frequencies of electromagnetic radiation and the respective wavelengths.

Electron transition: When an electron absorbs energy and moves from a low energy orbital to a vacant higher energy orbital.

Electronic configuration: The arrangement of electrons into orbitals and energy levels around the nucleus of an atom/ion. E.g. Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$.

Emission spectra: A spectrum of frequencies of electromagnetic radiation that has been emitted by an atom or molecule undergoing a transition from a state with higher energy to a state with lower energy.

Energy level: The shell that an electron is in.

First ionisation energy: The energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of gaseous 1^+ ions. For example, $O_{(g)} \rightarrow O^+_{(g)} + e^-$

Frequency: The number of wave oscillations per second.

$$E = hf$$

$$f = c / \lambda$$

(E = energy, h = Planck's constant, f = frequency, c = speed of light, λ = wavelength)

Gamma radiation: A type of electromagnetic wave that can be emitted when a nucleus undergoes radioactive decay.

Half-life ($t_{1/2}$): The time taken for the number of radioactive nuclei in a sample to decrease by half.

Isotope: Atoms of the same element with the same number of protons but different numbers of neutrons in the nucleus, e.g. ^{35}Cl and ^{37}Cl.

Infrared: The part of the electromagnetic spectrum that has wavelengths between 780 nm and 1 mm.

Lyman series: A group of lines present in the ultra-violet spectrum. The lines converge as frequency increases.

Orbital: An orbital is a cloud of negative charge that can hold up to two electrons. Different orbitals have different shapes.

***p* orbital:** A dumbbell shaped region in which up to two electrons can be found. There are three p orbitals at right angles to each other, so in total, the p subshell can hold up to 6 electrons.

Radioactive decay: A random process in which a radioactive nuclei loses energy by emitting radiation.

***s* orbital:** s orbitals are spherical and symmetrical regions around the nucleus, they can each hold up to two electrons.

Second ionisation energy: The energy required to remove 1 mole of electrons from each ion in 1 mole of gaseous 1+ ions to form 1 mole of gaseous 2+ ions (could be asked for any successive ionisation energy).

Shell: The energy level that an orbital is in around the nucleus of an atom. The shell closest to the nucleus is the first shell. The outermost shell that is occupied by electrons is the valence shell.

Shielding: A decrease in the nuclear attraction experienced by an outer shell electron caused by electron-electron repulsion between the outer shell electron and electrons from adjacent quantum shells.

Third ionisation energy: The energy required to remove 1 mole of electrons from each ion in 1 mole of gaseous 2+ ions to form 1 mole of gaseous 3+ ions (could be asked for any successive ionisation energy).

Ultra-Violet: The part of the electromagnetic spectrum that has wavelengths between 10 nm and 400 nm.

Visible light: The part of the electromagnetic spectrum that has wavelengths between 380 and 700 nm.

Wavelength: The distance between two successive points on a wave, such as two troughs or two crests. Inversely proportional to frequency.

1.3: Chemical calculations

Amount of substance: The quantity of a chemical species, measured in moles. Used as a way of counting atoms. The amount of substance can be calculated using:

Number of moles = Mass ÷ Mr

Number of moles = (Pressure x Volume) ÷ (Gas constant, R x Temperature)

Number of moles = Concentration x Volume

Atom economy: Measure of the proportion of reacting atoms that become part of the desired product in the balanced chemical equation.

Atom Economy = (Molar mass of desired product / Total molar mass of all products) x 100%

Avogadro constant (*L*): The number of atoms, molecules or ions in one mole of a given substance. It is the number of atoms in exactly 12 g of C (6.02×10^{23}mol).

Empirical formula: Smallest whole number ratio of atoms of each element in a compound. For example, the empirical formula of benzene (C_6H_6), cyclobutadiene (C_4H_4) and acetylene (C_2H_2) are all simply "CH".

Ideal gas equation: An equation that relates the number of moles of a gas to its volume, temperature and pressure.

PV = nRT

(P = pressure, V = volume, n = number of moles, R = gas constant, T = temperature)

Mass spectrometry: A technique used to identify compounds and determine their relative molecular mass.

Molar Mass: Mass of one mole of the substance expressed in $gmol^{-1}$.

Mole: The unit for the amount of substance. This is the amount of chemical species found in 12 g of ^{12}C. One mole is 6.02×10^{23}.

Molecular formula: The actual number of atoms of each element in a molecule.

Molecular ion peak: The peak on a mass spectrum with the highest m/z value, this is used to determine the molecular mass of a compound.

Percentage error: The degree of error in taking a measurement, this is estimated to be + or - half the smallest scale division of the apparatus.

$$\text{Percentage error} = \frac{\text{Instrument error}}{\text{Measurement}} \times 100$$

Percentage yield: The percentage ratio of the actual yield of product from a reaction compared with the theoretical yield.

$$\text{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical Yield}} \times 100$$

Relative abundance (of isotopes): The relative abundance of an isotope is the percentage of atoms found within a naturally occurring sample of an element that has a specific atomic mass.

Relative atomic mass: Average mass of an atom of an element, relative to 1/12 of the mass of an atom of carbon-12.

Relative formula mass: Average mass of a compound relative to 1/12 of the mass of an atom of carbon-12. Relative formula mass refers to compounds that have a giant structure.

Relative isotopic mass: Average mass of an atom of an isotope, relative to 1/12 of the mass of an atom of carbon-12.

Relative molecular mass: Average mass of a molecule relative to 1/12 of the mass of an atom of carbon-12.

Relative peak height: In mass spectra the peak heights show the relative abundances of the substance that made the peak.

1.4 - Bonding

Boiling temperature: The temperature at which a substance changes from a liquid state to a gaseous state.

Bond angle: The angle that is found between two bonds from the same atom in a covalently bonded compound.

Coordinate bond: A type of covalent bond in which one bonding atom provides both electrons in the bonding pair.

Covalent bond: The strong electrostatic attraction between two nuclei and the shared pair of electrons between them.

Electron pair repulsion: Pairs of electrons around a nucleus repel each other so the shape that a molecule adopts has the pairs of electrons positioned as far apart as possible.

Electronegativity: The ability of an atom to attract the bonding electrons in a covalent bond. The most electronegative elements (N,O,F) are small and have a relatively high nuclear charge.

Electrostatic attraction: The attraction between 2 species with opposite charges.

Hydrogen bonding: An interaction between a hydrogen atom and an electronegative atom, commonly nitrogen, fluorine or oxygen. The slightly positive hydrogen is attracted to the lone pair on the electronegative atom. Hydrogen bonds are stronger than van der Waals and dipole-dipole forces but weaker than ionic and covalent bonds.

Intermolecular forces: The forces which exist between molecules. The strength of the intermolecular forces impact physical properties like boiling/melting point.

Ion: An ion is formed when an atom/molecule loses or gains electrons. This gives it an overall charge - a positive charge if it has lost at least one electron and a negative charge if it has gained at least one electron.

Ionic bond: Strong electrostatic attraction between two oppositely charged ions. The strength of attraction depends on the relative sizes and charges of ions.

Ionic compound: A compound made up of anions and cations which are held together by ionic bonds, which arise due to the electrostatic attraction between oppositely charged ions. These structures are neutral overall.

Linear: The shape of a molecule when the central atom has 2 bonding pairs and no lone pairs of electrons. The bond angle is 180°.

Octahedral: The shape of a molecule in which the central atom has 6 bonding pairs. The bond angle is 90°.

Permanent dipole-dipole forces: When molecules with polar covalent bonds interact with dipoles in other molecules dipole-dipole intermolecular forces are produced between the molecules. These intermolecular forces are generally stronger than van der Waals forces but weaker than hydrogen bonding.

Polar bond: A covalent bond between two atoms in which the electrons in the bond are unevenly distributed. This causes a slight charge difference, inducing a dipole in the molecule.

Solubility: The ability of a given substance to dissolve in a solvent.

Solvent: A liquid that can dissolve other substances.

Tetrahedral: The shape of a molecule when the central atom has 4 bonding pairs. The bond angle is 109.5°.

Trigonal planar: The shape of a molecule when the central atom has 3 bonding pairs and no lone pairs of electrons. The bond angle is 120°.

Van der Waals: Also known as induced dipole–dipole, dispersion and London forces, van der Waals forces exist between all molecules. They arise due to fluctuations of electron density within a nonpolar molecule. These fluctuations may temporarily cause an uneven electron distribution, producing an instantaneous dipole. This dipole can induce a dipole in another molecule, and so on.

1.5: Solid structures

Boiling temperature: The temperature at which a substance changes from a liquid state to a gaseous state.

Delocalised electrons: The electrons that are not contained within a single atom or a covalent bond.

Diamond: A type of giant covalent structure composed of carbon atoms. Each carbon atom is bonded to four other carbon atoms in a tetrahedral structure, making diamond very hard.

Giant covalent structure: Large structures containing lots of atoms that are covalently bonded to each other, they are usually arranged in a regular lattice. E.g. diamond.

Giant ionic lattice: A regular repeating structure made up of oppositely charged ions.

Graphite: A type of giant covalent structure composed of carbon atoms. Each carbon atom is only bonded to three other carbon atoms in flat hexagonal sheets. This means there is one delocalised electron per carbon atom, so graphite can conduct electricity. The forces between the hexagonal layers in graphite are weak and can slide over each other.

Melting temperature: The melting point of a substance is the temperature at which it changes from solid state to liquid state.

Metallic bonding: Strong electrostatic attraction between positive metal ions and the sea of delocalised electrons that surround them.

Metallic structure: Layers of positive metal ions surrounded by a 'sea' of delocalised electrons.

Simple molecular substances: The structures formed by covalent molecules with weak intermolecular forces between molecules. These substances generally have low melting and boiling points.

1.6: The Periodic Table

Atomic/Proton number: The number of protons found in the nucleus of an atom of a particular element.

Cation: A positively charged ion, e.g. Na^+.

Crystallisation: A preparation technique used to form solid crystals from solution.

Displacement: A chemical reaction in which one element replaces another element in a compound. A halogen will displace a halide from solution if the halide is below it in the periodic table.

Electronegativity: The ability of an atom to attract the bonding electrons in a covalent bond. The most electronegative elements (N,O,F) are small and have a relatively high nuclear charge.

Element classification: An element is classified as s, p, d or f block according to its position in the Periodic Table.

First ionisation energy: The energy required to remove 1 mole of electrons from 1 mole of gaseous atoms to form 1 mole of gaseous 1$^+$ ions. For example, $O_{(g)} \rightarrow O^+_{(g)} + e^-$

Flame test: An analytical technique used to identify certain elements and ions based on the colour produced when a nichrome wire is dipped into a solution of the species and held in a blue bunsen flame.

Gravimetric analysis: An analytical technique used to separate ions in a solution.

Ionisation energy trend: Ionisation energy generally decreases down the group due to electron shells and shielding increasing. The nucleus, therefore, attracts the outer shell electrons less strongly.

Melting temperature: The melting point of a substance is the temperature at which it changes from solid state to liquid state.

Melting point trend: The Group 7 elements are simple covalent molecules held together with van der waals forces. The strength of these intermolecular forces increases down the group as the relative atomic mass of the molecule increases. Further down the group more energy is required to overcome the van der waals forces, resulting in higher melting points.

Oxidising ability: Oxidising ability is the ability to act as an oxidising agent. The oxidising ability of the halogens decreases down the group. This is because down the group the atoms get larger so the electrons are less strongly attracted to the nucleus so it is harder to gain an electron.

Oxidation: Process involving the loss of electrons. Results in an increase in oxidation number.

Oxidising agent: Electron acceptors. The elements/compounds which accept electrons causing itself to be reduced by oxidising another element/compound.

p-block element: Elements in groups 3-8/0 of the periodic table. p-block non-metals generally undergo reduction reactions.

Periodicity: Trends in element properties with increasing atomic number. The trends are caused by the changes in elements' atomic structure.

Precipitation: The formation of a solid from a solution.

Reducing ability: Reducing ability is the ability to act as a reducing agent. The reducing ability, or reducing power, of the halides increases down the group. This is because to act as a reducing agent the halide needs to lose an electron. As you go down the group it is easier for a halide to lose an electron because the attraction from the outer electron and nucleus decreases due to increased shielding and an increasing ionic radius.

Reactivity trend: The Group 7 elements need to gain an electron in order to react. As atomic radius increases, this becomes harder as the positive attraction of the nucleus is weakened by additional shielding. Therefore, down Group 7 it is harder to attract an electron so reactivity decreases.

Reducing agent: Electron donors. The elements/compounds which donate electrons causing itself to be oxidised by reducing another element/compound.

Reduction: Process involving the gain of electrons. Results in a decrease in oxidation number.

s-block element: Elements in Groups 1 and 2 of the periodic table. These generally undergo oxidation reactions.

Solubility: The ability of a given substance to dissolve in a solvent. Solubility of the Group 2 hydroxides increases down the group and solubility of the Group 2 sulfates decreases down the group.

Test for halide ions: When combined with acidified silver nitrate, halide ions react to form different coloured precipitates depending on the ion present. The colour of the precipitate formed can be used to identify which halide is present in a solution.

Thermal decomposition: A reaction in which a chemical substance is broken down by heating.

Thermal stability trend: As you go down the period more heat is required for the thermal decomposition of Group 2 nitrates and carbonates because the ions increase in size and therefore have greater thermal stability.

Volatility: How easily a substance evaporates in standard conditions.

Water treatment: The addition of chlorine to water to kill bacteria. The risks associated with the use of chlorine to treat water are the hazards of toxic chlorine gas and the possible risks from the formation of chlorinated hydrocarbons.

1.7 Simple equilibria and acid-base reactions

Acid: Proton donors. These species release hydrogen ions in solution.

Base: Proton acceptors. These species release hydroxide ions in solution.

Dynamic equilibrium: Dynamic equilibrium is reached when the rate of the forward reaction of a reversible reaction equals the rate of the backward reaction. The concentrations of the reactants and products remain constant.

Effect of changing concentration on equilibrium: If the concentration of a reactant increases, more products will be formed to re-establish the equilibrium.

Effect of changing pressure on equilibrium: If pressure is increased, the position of equilibrium shifts towards the side with the fewest number of molecules. If the pressure is decreased, the position of equilibrium shifts towards the side with the greatest number of molecules to oppose this change.

Effect of changing temperature on equilibrium: If the temperature of a system in equilibrium is increased, there will be an increase in the relative amount of products for an endothermic reaction and a decrease for an exothermic reaction.

K_c: A value that relates the concentrations of products and reactants present at equilibrium in a reversible reaction at a specific temperature. The equilibrium constant that is equal to the concentration of products raised to their stoichiometric coefficients divided by the concentration of reactants raised to the power of their stoichiometric coefficients.

Le Chatelier's principle: If a reaction at equilibrium is subjected to a change in concentration, temperature or pressure, the position of equilibrium will move to counteract the change.

pH: A value that represents the acidity or alkalinity of a solution. Acidic solutions have a pH of less than 7 while alkali solutions have a pH of greater than 7. Neutral solutions have a pH of 7.

$$pH = -\log[H^+]$$

$$[H^+] = 10^{-pH}$$

Reversible reaction: A reaction in which the products from the reaction can react together to form the original reactants. The direction of reversible reactions can be changed by changing the conditions.

Strong acid: An acid which dissociates/ionises almost completely in water. This means nearly all the H^+ ions will be released. E.g. HCl.

Strong base: A base which dissociates/ionises almost completely in water. E.g. NaOH.

Titration: An experimental technique used to determine the concentration of an unknown solution by using a second solution with a known concentration.

Weak acid: Acids which only dissociate/ionise very slightly in water so that only a small number of H^+ ions are released. E.g. Ethanoic acid.

Weak base: A base which only slightly dissociates/ionises in water. E.g. NH_3.