Cambridge Advanced Subsidiary Level Notes
9701 Chemistry

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Contents

1 Atomic structure

1.1 Particles in the atom and atomic radius

Understand that atoms are mostly empty space surrounding a very small, dense nucleus that contains protons and neutrons; electrons are found in shells in the empty space around the nucleus

Self explanatory.

Identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses

Considering that the proton is a particle present in the nucleus of an atom with charge +1 and mass 1, we see that neutrons are particles with no charge and a mass equal to that of the proton – and electrons are particles with no (neglible) mass but a charge opposite to that of the proton (-1).

Understand the terms atomic and proton number; mass and nucleon number

For an atom:

- *Atomic/Proton number* is the number of protons in the atom.

- *Mass/Nucleon number* is the combined number of neutrons and protons in the atom.

Describe the distribution of mass and charge within an atom

Observe that the massive particles in an atom (protons and neutrons) are present in the nucleus, hence, the nucleus is the massive part of an atom. The charges however, are present in both the nucleus and the electron shells, as the charges of the protons and electrons, respectively.

Describe the behaviour of beams of protons, neutrons and electrons moving at the same velocity in an electric field

Knowing that unlike charges attract and like charges repel, we can see that protons move to the negative end of an electric field whereas electrons deflect to the positive end. Knowing also that a proton is more massive than an electron and hence harder to move, we can deduce that a proton deflects to a lesser extent than an electron in the same electric field while travelling at the same velocity.

Determine the numbers of protons, neutrons and electrons present in both atoms and ions given atomic or proton number, mass or nucleon number and charge

For an atom,

$$
\begin{aligned}
q, & \quad \text{is the } \textit{net charge} \\
N_e, & \quad \text{is the } \textit{number of electrons} \\
Z, & \quad \text{is the } \textit{atomic number} \\
N, & \quad \text{is the } \textit{number of neutrons} \\
A, & \quad \text{is the } \textit{mass number.}
\end{aligned}
$$

From the discussions in sections above, we can deduce that:
$$A = Z + N$$

Note that,
$$Z - N_e = q$$

and hence for $q = 0$,

$$Z = N_e$$

State and explain qualitatively the variations in atomic radius and ionic radius across a period and down a group

Across a period, the number of electron shells in the atom is the same. As such, the atomic and ionic radii remain constant across a period. However, down the group the number of electron shells increases by one, as does atomic and ionic radii.

1.2 Isotopes

Define the term isotope in terms of protons and neutrons

Atoms with the same atomic number but a different number of neutrons are isotopes of each other.

Understand the notation $_y^x A$ for isotopes, where x is the mass or nucleon number and y is the atomic or proton number

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Self explanatory. Also note that $x = A$ and $y = Z$ from the above sections.

State that and explain why isotopes of the same element have the same chemical properties

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These isotopes have the same number of electrons and the same electronic configurations, thus they show same chemical properties.

State that and explain why isotopes of the same element have different physical properties, limited to mass and density

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A change in the number of neutrons per nucleus causes a change in mass of the same volume of a given samples of two isotopes of the same element. As such, density is also affected. In general, the higher the mass number the higher the mass and density.

1.3 Electrons, energy levels and atomic orbitals

Understand the terms:

- *shells, sub-shells and orbitals*
- *principal quantum number, (n)*
- *ground state, limited to electronic configuration*

In atoms, electrons are arranged in orbits around the nucleus. These orbits are called shells. In those shells, the electrons are in subshells. In those subshells, the electrons are arranged in orbitals.

The *principle quantum number* identifies the electron shell. The shell closest to the nucleus has principal quantum number, $n = 1$, and the further the nucleus is, the greater the n.

The *ground state* of an atom is where the electrons are arranged at the lowest possible energy levels.

Describe the number of orbitals making up s, p and d sub-shells, and the number of electrons that can fill s, p and d sub-shells

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The s orbital has one subshell, p has three, and d has five. Each orbital can hold a pair of electrons, so 2 electrons fill an s subshell. 6 electrons fill a p subshells and 10 electrons fill a d subshell.

Describe the order of increasing energy of the subshells within the first three shells and the 4s and 4p sub-shells

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The first electron shell has one s subshell. The second electron shell has an s subshell and a p subshell. The third electron shell has s, p and d. The fourth electron shell has shells s, p, d and f (the f subshell is not included in the specification). Electrons fill subshells in the following order:

$$1s, \ 2s, \ 2p, \ 2d, \ 3s, \ 3p, \ 4s, \ 3d, \ 4p$$

Note that the 4s subshell fills before 3d. This is because the 4s subshell is at a lower energy level than 3d.

Describe the electronic configurations to include the number of electrons in each shell, sub-shell and orbital

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Electronic configurations are now represented as follows

$$n\lambda^e$$

for each subshell.

Here, n stands for the principle quantum number of the shell, λ is the symbol of the subshell, and e is the number of electrons in that subshell.

Explain the electronic configurations in terms of energy of the electrons and inter-electron repulsion
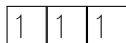

Electrons fill subshells as per their energy. They tend to fill the subshell with the lowest possible energy first.

Consider an example at the orbital level. Below is a p orbital without any electrons.

If we are to populate it with 3 electrons, its state would be:

1	1	1

This is as such as two electrons will not fill the same orbital if there is the option to fill another, since the electrons in the same orbital repel each other.

Determine the electronic configuration of atoms and ions given the atomic or proton number and charge, using either of the following conventions:
e.g. for Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ (full electronic configuration)
or [Ar] $3d^6 4s^2$ (shorthand electronic configuration)

Described above.

Understand and use the electrons in boxes notation
e.g. for Fe: [Ar] $\boxed{1\downarrow}$ $\boxed{1\ \ 1\ \ 1\ \ 1}$ $\boxed{1\downarrow}$

Each box represents an orbital. Each arrow represents an electron in that orbital. The arrowhead shows the spin of the electron, and arrows in the same orbital are in opposite directions to show that electrons in the same orbital must have opposite spin, as otherwise, the repulsion is too great.

Describe and sketch the shapes of s and p orbitals

Look it up.

Describe a free radical as a species with one or more unpaired electrons

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Self explanatory.

1.4 Ionisation energy

*Define and use the term first ionisation energy, IE*

First ionisation energy, IE is the energy needed to remove 1 mole of electrons from 1 mole of atoms of an element in the gaseous state to form 1 mole of gaseous ions.

*Construct equations to represent first, second and subsequent ionisation energies*

An equation that represents first ionisation energy of an element Λ is as follows

$$\Lambda(g) \longrightarrow \Lambda^+(g) + e^-$$
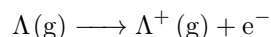
the enthalpy change (ΔH) of this reaction is the first IE.

The subsequent ionisation energy, would be the energy required to remove one mole of electrons from one mole of gaseous ions of an element with charge $+1$ to form one mole of gaseous ions with charge $+2$. The equation is as follows

$$\Lambda^+(g) \longrightarrow \Lambda^{2+}(g) + e^-$$
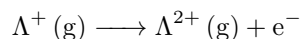
As such, in general, the nth ionisation energy would be the enthalphy change of a reaction

$$\Lambda^{n-1}(g) \longrightarrow \Lambda^n(g) + e^-$$
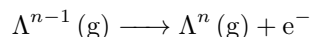
Identify and explain the trends in ionisation energies across a period and down a group of the Periodic Table

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Across a period from left to right, ionisation energy increases. This is because nuclear charge increases, the distance between the nucleus and the outer electron remains reasonably constant. The shielding by inner shells remains reasonably constant.

2 Atoms, molecules and stoichiometry

2.1 Relative masses of atoms and molecules

Define the unified atomic mass unit as one twelfth of the mass of a carbon-12 atom

Self explanatory.

Define relative atomic mass, A_r, relative isotopic mass, relative molecular mass, M_r, and relative formula mass in terms of the unified atomic mass unit

<u>Relative atomic mass</u>, A_r, is the weighted average mass of atoms in a given sample of an element compared to the value of the unified atomic mass unit.

<u>Relative isotopic mass</u> is the mass of a particular atom of an isotope compared to the value of the unified atomic mass unit.

<u>Relative molecular mass</u>, M_r, is the weighted average mass of a molecule in a given sample of the molecule compared to the value of the unified atomic mass unit.

<u>Relative formula mass</u>, is the weighted average mass of one formula unit compared to the value of the unified atomic mass unit.

2.2 The mole and the Avogadro constant

Define and use the term mole in terms of the Avogadro constant

A <u>mole</u> is an amount of substance which contains 6.02×10^{23} specified particles (atoms, molecules, ions or electrons).

The <u>Avogadro constant</u>, L, is the number of specified particles, atoms, ions, molecules or electrons, in a mole of those particles. The numerical value is $L = 6.02 \times 10^{23}$.

2.3 Formulas

Write formulas of ionic compounds from ionic charges and oxidation numbers (shown by a Roman numeral), including:

(a) *the prediction of ionic charge from the position of an element in the Periodic Table*

(b) *recall of the names and formulas for the following ions: NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+}, Ag^+, HCO_3^-, PO_4^{3-}*

The number of valence electrons of an element is the same as the group number for that element. Loss or gain of electrons in the valence shell is what forms ions. Usually, for elements with less than 4 valence electrons, they form cations with the magnitude of the charge equal to the group number. For elements with more than 4 valence electrons, they gain electrons, forming anions with magnitude equal to $(8 - X)$, where X is the group number[1].

(a) *Write and construct equations (which should be balanced), including ionic equations (which should not include spectator ions)*

(b) *Use appropriate state symbols in equations*

I mean.

Define and use the terms empirical and molecular formula

<u>Empirical formula</u> is the simplest whole number ratio of the elements present in one molecule or formula unit of the compound. <u>Molecular formula</u> is the formula that shows the number and type of each atom present in a molecule.

Understand and use the terms anhydrous, hydrated and water of crystallisation

Ionic compounds may form crystals when they chemically bind with water. The water bound with is called the <u>water of crystallisation</u>. In presence and

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[1]Consider group numbers 1 to 8, in these cases.

absence of this water, the molecule is said to be hydrated and anhydrous, respectively.

Calculate empirical and molecular formulas, using given data

Given that there are n elements in a compound, with percentage by mass p_1, p_2, ... p_n and A_r of A_1, A_2, ... A_n. The simplest ratio of atoms present here is

$$\frac{p_1}{A_1} : \frac{p_2}{A_2} : \cdots : \frac{p_n}{A_n}$$

To find the molecular formula from the above information, we may find a "multiplier" k.

$$k = \frac{M_r(\text{actual molecule})}{M_r(\text{empirical molecule})}$$

2.4 Reacting masses and volumes (of solutions and gases)

Perform calculations including use of the mole concept, involving:

(a) reacting masses (from formulas and equations) including percentage yield calculations

(b) volumes of gases (e.g. in the burning of hydrocarbons)

(c) volumes and concentrations of solutions

(d) limiting reagent and excess reagent (When performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question. When rounding up or down, candidates should ensure that significant figures are neither lost unnecessarily nor used beyond what is justified (see also Mathematical requirements section).)

(e) deduce stoichiometric relationships from calculations such as those in 2.4.1(a)–(d)

The moles present of a compound, can be calculated from its mass from the following formula

$$\text{moles, } n = \frac{(\text{mass})}{M_r(\text{compound})}$$

One mole of a gas has, at room temperature and pressure, volume of $24\,\text{dm}^3 = 24\,000\,\text{cm}^3$. This is called the molar gas constant and thus follows the formula

$$\text{volume of gas} = (n)(24\,\text{dm}^3)$$
$$= (n)(24 \times 10^3\,\text{cm}^3)$$

For a solution with concentration c and volume V, the moles of compound present in it is

$$n = cV$$

In a reaction, the compounds reacting are in specific ratios. Often, it is the case that one compound is present in more an amount than necessary. This compound is called the *excess reagent* whereas the others are called *limiting reagent*. To find which is which, we find moles of all compounds present in one part of the ratios in which they are to be present. Whichever has least moles is limiting and whichever has more, is excess.

3 Chemical bonding

4   States of matter

5 Chemical energetics

5.1 Enthalpy change, ΔH

Understand that chemical reactions are accompanied by enthalpy changes and these changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive)

When chemical reactions occur, there is some change in the surroundings, usually seen in the form of temperature decrease or increase. This shows that the system of the reaction has either absorbed or released certain amount of energy, respectively. The amount of energy absorbed or released by a chemical reaction is said to be its *enthalpy change*. This can be quantified, and its mathematical symbol is ΔH, with unit kJ mol^{-1}.

Energy is absorbed to break bonds, $\Delta H > 0$, in reactions that break bonds, which are called *endothermic*. Energy is released when bonds are formed, so $\Delta H < 0$ in such cases, which are called *exothermic*. You may consider ΔH to be the energy added to the system, the logic for the positive and negative natures of the above described changes in energy is now apparent.

Construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy

I refuse.

Define and use the terms:

(a) standard conditions (this syllabus assumes that these are 298 K and 101 kPa) shown by $^{\ominus}$.

(b) enthalpy change with particular reference to: reaction, ΔH_r, formation, ΔH_f, combustion, ΔH_c, neutralisation, ΔH_{neut}

For a reaction occuring under standard conditions (stated above), the enthalpy change is represented by the symbol $\Delta H^{\ominus}$.

Standard enthalpy change of reaction, $\Delta H_r^{\ominus}$, is the enthalpy change when the amounts of reactants shown in the stoichiometric equation react to give products under standard conditions.

Standard enthalpy change of formation, $\Delta H_f^{\ominus}$, is the enthalpy change when one mole of a compound is formed from its elements under standard conditions.

Standard enthalpy change of combustion, $\Delta H_c^{\ominus}$, is the enthalpy change when one mole of a substance is burnt in excess oxygen under standard conditions.

Standard enthalpy change of neutralisation, ΔH_{neut} is the enthalpy change when one mole of water is formed by the reaction of an acid with an alkali under standard conditions.

Understand that energy transfers occur during chemical reactions because of the breaking and making of chemical bonds

Breaking bonds absorbs energy, (breaking anything does, really). Making chemical bonds releases energy.

Use bond energies (ΔH positive, i.e. bond breaking) to calculate enthalpy change of reaction, ΔH_r

Each bond takes a certain amount of energy to break. We say that the bond stores a certain amount of energy. The amount of energy stored by a bond is its *bond energy*. We represent the energy of a certain bond mathematically as

$$E(R)$$

where R is a certain bond such as $\text{Br}\!-\!\text{Br}$, $\text{O}\!=\!\text{O}$, etc.

In a reaction, the bonds of the reactants are first broken, then the bonds of the products are formed. The energy required to break the bonds is added to the enthalpy change, before the energy released to form the bonds is subtracted. Thus, in terms of bond energies,

$$\Delta H = \sum(\text{bond energies of reactants})$$
$$- \sum(\text{bond energies of products})$$

The *exact* bond energy of a bond depends on the bonds surrounding the bond itself. That is, $E(\text{C}\!-\!\!-\!\text{H})$ is not exactly equal in the molecules CH_4 and HCOOH.

Thus, we often use *average bond energies*, which are the average energy needed to break a specific covalent bond averaged from a variety of molecules in the gaseous state.

Calculate enthalpy changes from appropriate experimental results, including the use of the relationships $q = mc\Delta T$ and $\Delta H = -mc\Delta T/n$

Calorimetry is used to find enthalpy changes of reactions practically. We use the fact that it takes $4.18\,\text{J}$ per gram of water to change its temperature by 1 Kelvin. This is known as the *specific heat capacity, c* of water.

$$c = 4.18$$

with units $\text{kJ}\,\text{g}^{-1}\,\text{K}^{-1}$.

Each cubic centimetre of water weighs one gram. Thus, if we know the volume of a solution in which a reaction occurs, and measure the change in its temperature, we can find the energy change caused by the reaction using

$$q = mc\Delta T$$

where q is the energy change, m is the mass of water, c is as defined above and ΔT is the change in temperature.

Thus, the enthalpy change can now be

$$E = -mc\Delta T/n$$

where n is the moles of defined reactant or product. The negative sign comes from the fact that enthalpy change is the energy change from the perspective of the chemical system and the energy change (calculated as q) is the energy change from the perspective of any outside observer.

5.2 Hess's law

Apply Hess's law to construct simple energy cycles

Hess's law states that the enthalpy change in a chemical reaction is the same independent of the route by which the chemical reaction takes place as long as the initial and final conditions and states of reactants are products are the same.

Carry out calculations using cycles and relevant energy terms, including:

(a) determining enthalpy changes that cannot be found by direct experiment

(b) use of bond energy data

In summary, given enthalpy changes of combustion,

$$\Delta H_r = \sum \Delta H_c(\text{reactants}) - \sum \Delta H_c(\text{products})$$

and given enthalpy changes of formation,

$$\Delta H_r = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$$

In the case where bond energies are that which are given,

$$\Delta H = \sum(\text{bond energies of reactants})$$
$$- \sum(\text{bond energies of products})$$

6 Electrochemistry

6.1 Redox processes: electron transfer and changes in oxidation number (oxidation state)

Calculate oxidation numbers of elements in compounds and ions

The sum of the oxidation numbers of every element in a compound or a molecular ion is equal to the charge on the ion or compound. We can figure oxidation states out by arranging equations according to the above.

Use changes in oxidation numbers to help balance chemical equations

Figure that out.

Explain and use the terms redox, oxidation, reduction and disproportionation in terms of electron transfer and changes in oxidation number

Reduction is loss of electrons, *oxidation* is gain of electrons. A reaction where both reduction and oxidation occur simultaneously is called a *redox* reaction. A *disproportionation* reaction is where both reduction and oxidation happens to the same species.

Explain and use the terms oxidising agent and reducing agent

In a redox reaction, the species that is oxidised is the *reducing agent* and the species which is reduced is the *oxidising agent*.

Use a Roman numeral to indicate the magnitude of the oxidation number of an element

An element with oxidation state +1 can be represented by Roman numeral (I), and so on and so forth.

7 Equilibria

7.1 Chemical equilibria: reversible reactions, dynamic equilibrium

(a) *Understand what is meant by a reversible reaction*

(b) *Understand what is meant by dynamic equilibrium in terms of the rate of forward and reverse reactions being equal and the concentration of reactants and products remaining constant*

(c) *Understand the need for a closed system in order to establish dynamic equilibrium*

A <u>reversible reaction</u> is one where products can be changed back to reactants by reversing the conditions.

<u>Dynamic equilibrium</u> is the state of a reversible reaction where the rate of the forward reaction equals that of the backward reaction. In this state, the concentrations of products and reactants are unchanging but the reaction is ongoing. Such a state requires a <u>closed system</u>, where no material can enter the reaction mixture, and none can exit it either.

Define Le Chatelier's principle as: if a change is made to a system at dynamic equilibrium, the position of equilibrium moves to minimise this change

Self explanatory :).

Use Le Chatelier's principle to deduce qualitatively (from appropriate information) the effects of changes in temperature, concentration, pressure or presence of a catalyst on a system at equilibrium

For an exothermic reaction, increasing temperature makes equilibrium shift to left and vice versa. It is the opposite case for an endothermic reaction.

Increase in pressure causes equilibrium to shift to the side which has lower gaseous moles, and vice versa.

Increase in concentration or amount of a substance in the equilibrium mixture causes equilibrium to shift to the side opposite to the substance and vice versa.

Deduce expressions for equilibrium constants in terms of concentrations, K_c

Consider the reaction

$$m\,A + n\,B \rightleftharpoons p\,C + q\,D$$
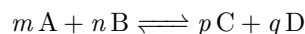
it will have K_c

$$K_c = \frac{[C]^p[D]^q}{[A]^m[B]^n}$$

where [X] is concentration of X.

Use the terms mole fraction and partial pressure

The <u>mole fraction</u> is the number of moles of a particular gas divided by the total number of moles of all gases in the mixture.

$$(\text{mole fraction of X}) = \frac{n(X)}{\Sigma n}$$

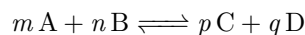<u>Partial pressure</u> is the pressure exerted by a particular gas in a mixture of gases.

$$(\text{partial pressure of X}),\, pX = (\text{mole fraction of X})$$
$$\times (\text{total pressure})$$

Deduce expressions for equilibrium constants in terms of partial pressures, K_p (use of the relationship between K_p and K_c is not required)

Consider again, the reaction

$$m\,A + n\,B \rightleftharpoons p\,C + q\,D$$

where every reagent is gaseous.

$$K_p = \frac{pA^m pB^n}{pC^p pD^q}$$

Use the K_c and K_p expressions to carry out calculations (such calculations will not require the solving of quadratic equations)

Just. do. it.

Calculate the quantities present at equilibrium, given appropriate data

js do it fr cuh.

State whether changes in temperature, concentration or pressure or the presence of a catalyst affect the value of the equilibrium constant for a reaction
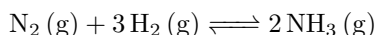
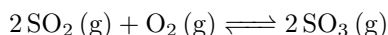Judge changes in concentrations of reagents and answer accordingly.

Describe and explain the conditions used in the Haber process and the Contact process, as examples of the importance of an understanding of dynamic equilibrium in the chemical industry and the application of Le Chatelier's principle

Haber process

$$N_2 \text{ (g)} + 3\,H_2 \text{ (g)} \rightleftharpoons 2\,NH_3 \text{ (g)}$$

Contact process

$$2\,SO_2 \text{ (g)} + O_2 \text{ (g)} \rightleftharpoons 2\,SO_3 \text{ (g)}$$

7.2 Brønsted–Lowry theory of acids and bases

State the names and formulas of the common acids, limited to hydrochloric acid, HCl, sulfuric acid, H_2SO_4, nitric acid, HNO_3, ethanoic acid, CH_3COOH

Self explanatory.

State the names and formulas of the common alkalis, limited to sodium hydroxide, NaOH, potassium hydroxide, KOH, ammonia, NH_3

Self explanatory.

Describe the Brønsted–Lowry theory of acids and bases

This theory defines <u>acids</u> as proton donors and <u>bases</u> as proton acceptors. Note that a proton is a hydrogen cation H^+.

Describe strong acids and strong bases as fully dissociated in aqueous solution and weak acids and weak bases as partially dissociated in aqueous solution

Self explanatory.

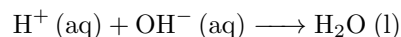*Appreciate that water has pH of 7, acid solutions pH of below 7 and alkaline solutions pH of above 7*

Self explanatory.

Understand that neutralisation reactions occur when H^+ (aq) and OH^- (aq) form H_2O (l)

All neutralisation reactions have the following ionic equation

$$H^+ \text{ (aq)} + OH^- \text{ (aq)} \longrightarrow H_2O \text{ (l)}$$

Understand that salts are formed in neutralisation reactions

Understand that.

Sketch the pH titration curves of titrations using combinations of strong and weak acids with strong and weak alkalis

Figure that out.

Select suitable indicators for acid-alkali titrations, given appropriate data (pKa values will not be used)

All indicators have a range of pH, above which it is a certain colour and below which it is another colour. We first determine the range across which the pH of a neutralisation reaction changes. We select the indicator which has range within the range of neutralisation.

8 Reaction kinetics

9 The Periodic Table: chemical periodicity

9.1 Periodicity of physical properties of the elements in Period 3

Describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements

Across the period, atomic radius decreases as positive nuclear charge increases, shielding remains the same so the valence electrons are attracted more strongly.

When these atoms form anions, the ionic radius formed is greater than the atomic radius, and the opposite occurs when cations are formed. The trends across the period in case of ionic radius (decrease across the group) remains the same.

For the metallic elements, melting point increases across the period. Melting point reaches a maximum in group 4, where the elements form giant covalent molecules. After this, melting point decreases.

Explain the variation in melting point and electrical conductivity in terms of the structure and bonding of the elements

Metals are conductive because they have delocalised electrons which are mobile. They have high melting points because of the electrostatic attraction between cations and the sea of delocalised electrons. Giant molecules have covalent bonds which take significant amount of energy to break.

9.2 Periodicity of chemical properties of the elements in Period 3

Describe, and write equations for, the reactions of the elements with oxygen (to give Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2), chlorine (to give $NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5) and water (Na and Mg only)

Reactions with oxygen

Sodium reacts vigorously, giving a bright yellow flame

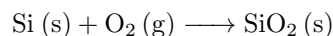$$4\,Na\,(s) + O_2\,(g) \longrightarrow 2\,Na_2O\,(s)$$
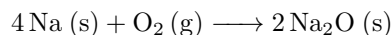
Magnesium also reacts vigorously, giving a bright white flame
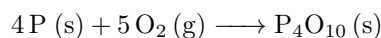
$$2\,Mg\,(s) + 2\,(g) \longrightarrow 2\,MgO\,(s)$$
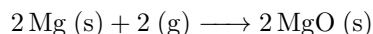
Aluminium metal is protected by a layer of aluminium oxide, but powdered aluminium reacts well with oxygen. This too, gives a bright white flame
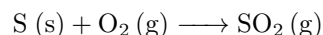
$$4\,Al\,(s) + 3\,O_2\,(g) \longrightarrow 2\,Al_2O_3\,(s)$$
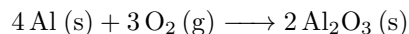
Silicon reacts slowly with oxygen

$$Si\,(s) + O_2\,(g) \longrightarrow SiO_2\,(s)$$

Phosphorus reacts vigorously, giving a yellow/white flame, and clouds of phosphorus (V) oxide are produced

$$4\,P\,(s) + 5\,O_2\,(g) \longrightarrow P_4O_{10}\,(s)$$

Sulfur, in powder form, once ignited, burns gently with a blue flame in a gas jar of oxygen gas, producing toxic fumes of sulfur gas
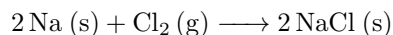
$$S\,(s) + O_2\,(g) \longrightarrow SO_2\,(g)$$

further oxidation of sulfur dioxide gives sulfur trioxide

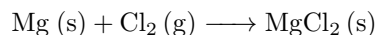$$2\,SO_2\,(g) + O_2\,(g) \underset{}{\overset{V_2O_5 \text{ catalyst}}{\rightleftharpoons}} 2\,SO_3\,(g)$$
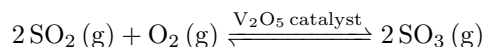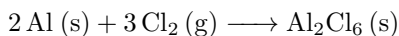
Chlorine and argon do not react with oxygen.

Reactions with chlorine

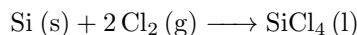Sodium metal is heated then plunged into a gas jar of chlorine there is a vigorous reaction, forming sodium chloride

$$2\,Na\,(s) + Cl_2\,(g) \longrightarrow 2\,NaCl\,(s)$$

Magnesium and aluminium also react vigorously with chlorine gas, giving chlorides

$$Mg\,(s) + Cl_2\,(g) \longrightarrow MgCl_2\,(s)$$

$$2\,\mathrm{Al\ (s)} + 3\,\mathrm{Cl_2\ (g)} \longrightarrow \mathrm{Al_2Cl_6\ (s)}$$

<u>Silicon</u> reacts slowly with chlorine

$$\mathrm{Si\ (s)} + 2\,\mathrm{Cl_2\ (g)} \longrightarrow \mathrm{SiCl_4\ (l)}$$

<u>Phosphorus</u> also reacts slowly with excess chlorine gas

$$2\,\mathrm{P\ (s)} + 5\,\mathrm{Cl_2\ (g)} \longrightarrow 2\,\mathrm{PCl_5\ (L)}$$
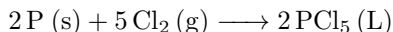
The rest of Period 3 does not react with chlorine.

Reactions with water

<u>Sodium</u> reacts vigorously with cold water, melting into a ball of molten metal, moving across the surface of water, giving off hydrogen gas as it gets smaller and smaller until it disappears. It leaves a strongly alkaline (pH 14) solution of sodium hydroxide behind

$$2\,\mathrm{Na\ (s)} + 2\,\mathrm{H_2O\ (l)} \longrightarrow 2\,\mathrm{NaOH\ (aq)} + \mathrm{H_2\ (g)}$$
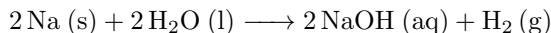
<u>Magnesium</u> metal reacts extremely slowly with cold water, a weakly alkaline of slightly soluble magnesium hydroxide solution is formed of pH 11.

$$\mathrm{Mg\ (s)} + 2\,\mathrm{H_2O\ (l)} \longrightarrow \mathrm{Mg(OH)_2\ (aq)} + \mathrm{H_2\ (g)}$$
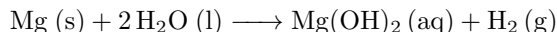
Under heat, however, magnesium reacts vigorously with steam, giving magnesium oxide and hydrogen gas

$$\mathrm{Mg\ (s)} + \mathrm{H_2O\ (g)} \longrightarrow \mathrm{MgO\ (s)} + \mathrm{H_2\ (g)}$$
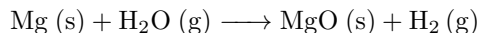
State and explain the variation in the oxidation number of the oxides (Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 only) and chlorides ($NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 only) in terms of their outer shell (valence shell) electrons

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All of the compounds mentioned above show increase in oxidation states across the group. This is because across the group, the number of valence electrons increases, as does the number of electrons that can be lost.
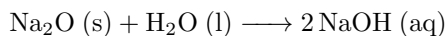

Describe, and write equations for, the reactions, if any, of the oxides Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10},
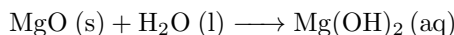
SO_2 and SO_3 with water including the likely pHs of the solutions obtained

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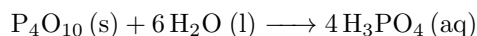<u>Sodium</u> and <u>magnesium</u> oxides react with water, giving hydroxides

$$\mathrm{Na_2O\ (s)} + \mathrm{H_2O\ (l)} \longrightarrow 2\,\mathrm{NaOH\ (aq)}$$

where NaOH has pH 14.

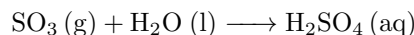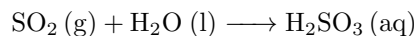$$\mathrm{MgO\ (s)} + \mathrm{H_2O\ (l)} \longrightarrow \mathrm{Mg(OH)_2\ (aq)}$$

where $\mathrm{Mg(OH)_2}$ has pH 10.

<u>Phosphorus</u> (V) oxide reacts vigorously and dissolves in water to form an acidic solution of phosphoric (V) acid

$$\mathrm{P_4O_{10}\ (s)} + 6\,\mathrm{H_2O\ (l)} \longrightarrow 4\,\mathrm{H_3PO_4\ (aq)}$$

where $\mathrm{H_3PO_4}$ has pH 2.

Both <u>sulfur</u> oxides react and dissolve in water, giving acidic solutions

$$\mathrm{SO_2\ (g)} + \mathrm{H_2O\ (l)} \longrightarrow \mathrm{H_2SO_3\ (aq)}$$

$$\mathrm{SO_3\ (g)} + \mathrm{H_2O\ (l)} \longrightarrow \mathrm{H_2SO_4\ (aq)}$$

where the solutions containing products of the above reactions have pH 1.

Describe, explain, and write equations for, the acid/base behaviour of the oxides Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 and the hydroxides $NaOH$, $Mg(OH)_2$ and $Al(OH)_3$ including, where relevant, amphoteric behaviour in reactions with acids and bases (sodium hydroxide only)


The oxides of the metals <u>sodium</u> and <u>magnesium</u>, with purely ionic bonding, produce alkaline solutions with water as their oxide ions $\mathrm{O^{2-}}$ (aq), become hydroxide ions, $\mathrm{OH^-}$ (aq). *The oxide ions behave as bases by accepting protons from the water molecules.*

$$\mathrm{O^{2-}\ (aq)} + \mathrm{H_2O\ (l)} \longrightarrow 2\,\mathrm{OH^-\ (aq)}$$
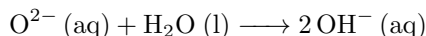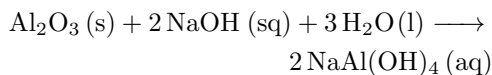
<u>Aluminium</u> oxide is amphoteric, meaning it can display both acidic and alkaline properties. It does

so because it shows characteristics of both ionic and covalent bonding. In reaction with acids
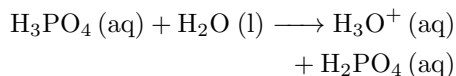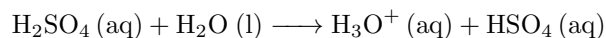
$$Al_2O_3 \, (s) + 3\,H_2SO_4 \, (sq) \longrightarrow Al_2(SO_4)_3 \, (aq)$$
$$+ 3\,H_2O \, (l)$$
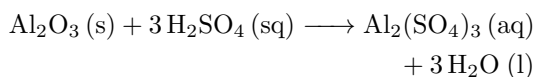
and in reaction with *hot, concentrated alkali*

$$Al_2O_3 \, (s) + 2\,NaOH \, (sq) + 3\,H_2O \, (l) \longrightarrow$$
$$2\,NaAl(OH)_4 \, (aq)$$

Silicon dioxide is insoluble, but it reacts with *hot, concentrated alkali*

$$SiO_2 \, (s) + 2\,NaOH \, (sq) \longrightarrow Na_2SiO_3 \, (aq)$$
$$+ H_2O \, (l)$$
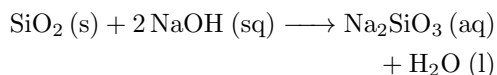
The covalently bonded non metal oxides of phosphorus and sulfur dissolve and react in water to form acidic solutions. The acid molecules formed donate H^+ ions to water molecules, behaving as typical acids

$$H_2SO_4 \, (aq) + H_2O \, (l) \longrightarrow H_3O^+ \, (aq) + HSO_4 \, (aq)$$

$$H_3PO_4 \, (aq) + H_2O \, (l) \longrightarrow H_3O^+ \, (aq)$$
$$+ H_2PO_4 \, (aq)$$

The hydroxides mentioned are just plain ol' hydroxides.

Describe, explain, and write equations for, the reactions of the chlorides NaCl, MgCl₂, AlCl₃, SiCl₄, PCl₅ with water including the likely pHs of the solutions obtained

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Sodium and magnesium chlorides do not react with water.

In aluminium chloride crystals, the compound has formula $AlCl_3$. When anhydrous, the compoound exists as Al_2Cl_6. This can be thought of as a dimer of $AlCl_3$. Al_2Cl_6 is covalent molecular, whereas $AlCl_3$ is ionic. Addition of water causes aluminium and chloride ions enter the solution, where each relatively small and highly charged aluminium ion is hydra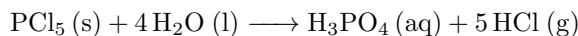ted and causes a water molecule bonded to it to lose a proton. This turns the solution acidic, which is shown as follows

$$[Al(H_2O)_6]^{3+} \, (aq) \longrightarrow [Al(H_2O)_5OH]^{2+} \, (aq)$$
$$+ H^+ \, (aq)$$
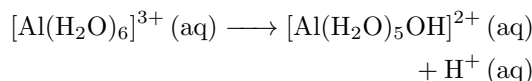
Phoshphorus and silicon chlorides are hydrolysed in water, releasing white fumes of hydrogen chloride gas in a rapid reaction.

$$SiCl_4 \, (l) + 2\,H_2O \, (l) \longrightarrow SiO_2 \, (s) + 4\,HCl \, (g)$$
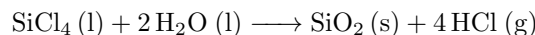
where SiO_2 is seen as an offwhite precipitate.

$$PCl_5 \, (s) + 4\,H_2O \, (l) \longrightarrow H_3PO_4 \, (aq) + 5\,HCl \, (g)$$

Explain the variations and trends in 9.2.2, 9.2.3, 9.2.4 and 9.2.5 in terms of bonding and electronegativity

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Some shit happens 'cuz shit's ionic, otherwise 'cuz shit's covalent. Figure it out.

Suggest the types of chemical bonding present in the chlorides and oxides from observations of their chemical and physical properties

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The difference in electronegativity of the compounds is what determines the nature of the bonding in this compound. Significant difference implies ionic bond, otherwise covalent.

9.3 Chemical periodicity of other elements

Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity

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Intuition.

Deduce the nature, possible position in the Periodic Table and identity of unknown elements from given information about physical and chemical properties

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Intuition.