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 ^x_yA 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*

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2 Atoms, molecules and stoichiometry

3 Chemical bonding

3.1 Electronegativity and bonding

Define electronegativity as the power of an atom to attract electrons to itself

Self explanatory.

Explain the factors infuencing the electronegativities of the elements in terms of nuclear charge, atomic radius and shielding by inner shells and sub-shells

Electronegativity is influenced by the following:

- Nuclear charge: Atoms with a greater positive nuclear charge have a greater electronegativity, as they can pull attract electrons to themselves.
- Atomic radius: The larger the atomic radius, the smaller the electronegativity, since it is harder to cause attraction over a larger distance.
- Shielding: The greater the number of inner electron shells and sub shells, the lower the effective nuclear charge on bonding electrons as these inner electrons are repelling the outer ones.

State and explain the trends in electronegativity across a period and down a group of the Periodic Table

Electronegativity increases across a period from Group 1 to Group 17, and it decreases down each group.

Across a period, nuclear charge increases whilst shielding remains the same, as does atomic radius, increasing electronegativity.

Down a group, atomic radius and shielding increase, as down nucleaer charge, however, the increase in shielding is the most influential of these factors and this is what causes the decrease in electronegativity.

Use the differences in Pauling electronegativity values to predict the formation of ionic and covalent bonds (the presence of covalent character in some ionic compounds will not be assessed) (Pauling electronegativity values will be given where necessary)

Every atom has its own electronegativity, which is measured in the Pauling electronegativity scale. Note that carbon and hydrogen have low electronegativities. The value of electronegativity is denoted N_{p}

For elements chemically bonded, the differences in their electronegativities can be used to predict the ionic or covalent character of their bonds. A high difference (> 2.0) signifies an ionic bond, whereas a lower difference (< 1.0) signifies a covalent bond. A zero value shows absence of any ionic character and any value between 1 and 2 shows that the compound is not entirely covalent and has some ionic character.

3.2 Ionic bonding

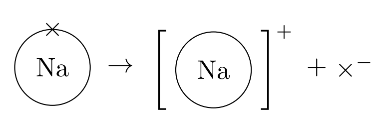
Define ionic bonding as the electrostatic attraction between oppositely charged ions (positively charged cations and negatively charged anions)

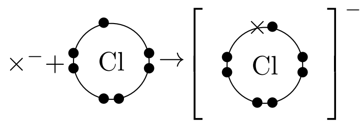
Self explanatory.

Describe ionic bonding including the examples of sodium chloride, magnesium oxide and calcium fluoride

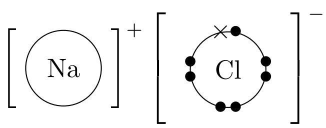
Ionic bonds form between metals and non-metals, where the metallic element loses its valence electrons, which are gained by the non-metal, producing a metal cation and a non-metal anion which bond together via electrostatic attraction.

Consider sodium chloride, NaCl, which consists of the ions Na^{+} and Cl^{-}, formed when the Na atom loses its valence electron which is gained by the Cl atom. We can represent this in a dot and cross diagram:

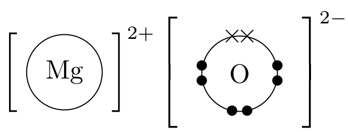


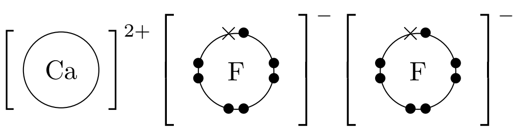


Thus a molecule of NaCl is shown as follows in a dot and cross diagram



Given below are the cases for MgO and CaF_2





3.3 Metallic bonding

Define metallic bonding as the electrostatic attraction between postiive metal ions and delocalised electrons

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Self explanatory.

3.4 Covalent bonding and coordinate (dative covalent) bonding

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Define covalent bonding as electrostatic attraction between the nuclei of two atoms and a shared pair of electrons

(a) Describe covalent bonding in molecules including:

- *hydrogen,* H_2
- *oxygen,* O_2
- *nitrogen,* N_2
- *chlorine,* Cl_2
- *hydrogen chloride, HCl*
- *carbon dioxide,* CO_2
- *ammonia,* NH_3
- *methane,* CH_4
- *ethane,* C_2H_6
- *ethene,* C_2H_4

(b) Understand that elements in period 3 can expand their octet including in the compounds sulfur dioxide, SO_2*, phosphorus pentachloride,* PCl_5 *, and sulfur hexafluoride,* SF_6

(c) Describe coordinate (dative covalent) bonding, including in the reaction between ammonia and hydrogen chloride gases to form the ammonium ion, NH_4^+*, and in the* Al_2Cl_6 *molecule*
