Cambridge (CIE) A Level Chemistry

Your notes

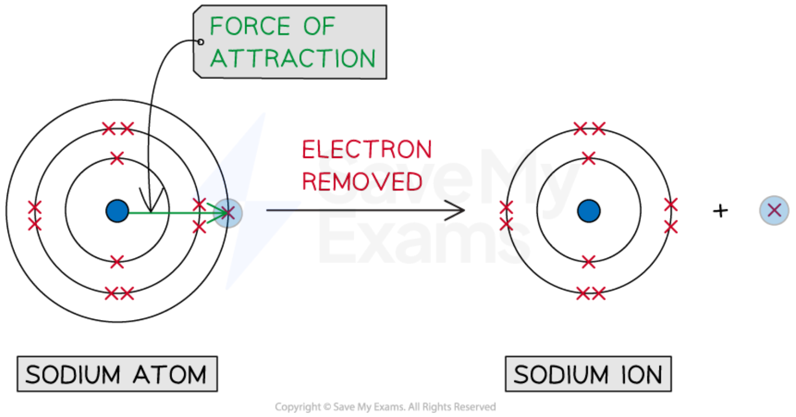
Ionisation Energy

Contents

Ionisation Energy Definition & Equations

- **Ionisation** is the process by which an electron is removed from an atom or a molecule

Process of ionisation



In the first ionisation of sodium, the single outer electron is lost

- The **ionisation energy** (**IE**) of an element is the amount of energy required to remove **one mole** of electrons from **one mole** of gaseous atoms of an element to form **one mole** of gaseous ions
- Ionisation energies are measured under **standard conditions** which are 298 K and 101 kPa
- The units of IE are **kilojoules per mole** (kJ mol^{-1})
- The values for ionisation energies are always positive as this is an endothermic process
  - This is because **energy is required** to break the force of attraction between the electron and the central positive nucleus

First ionisation energy

- The **first ionisation energy** (IE_1) is the energy required to remove **one mole of electrons** from one mole of gaseous atoms of an element to form one mole of gaseous 1+ ions
  - E.g. the first ionisation energy of gaseous calcium:

$$\mathbf{Ca\,(g) \rightarrow Ca^{+}\,(g) + e^{-}} \qquad IE_1 = +590\ \text{kJ mol}^{-1}$$

Second ionisation energy

- The **second ionisation energy** (IE_2) of an element is the amount of energy required to remove **one mole** of electrons from **one mole** of gaseous 1+ ions of an element to form

one mole of gaseous **2+ ions**

- E.g. the second ionisation energy of gaseous calcium:

$$\mathbf{Ca^{+}\,(g) \rightarrow Ca^{2+}\,(g) + e^{-}} \qquad IE_2 = +1145.4\ \text{kJ mol}^{-1}$$

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Examiner Tips and Tricks

- It is a common mistake to write ionisation energies in more than one step
  - For example; $Ca\,(g) \rightarrow Ca^{2+}\,(g) + 2e^{-}$
- Ionisation energies should take place one step at a time
- Students often miss out the gaseous state symbol in the equation and also in the definition, so make sure you include this

Trends in Ionisation Energy

- Ionisation energies show **periodicity**, a recurring trend across each **period of the Periodic Table**
  - As you move **left to right across a period**, **first ionisation energy generally increases**
  - This trend is linked to the **electronic configuration** of elements:
    - **Group 1 metals** (e.g. Na) have **low first ionisation energies** because they have a single outer electron that is easy to remove
    - **Noble gases** (e.g. Ne, Ar) have **very high ionisation energies** due to their full outer shells and strong nuclear attraction
- There are four factors that affect ionisation energy
  - Nuclear charge
  - Distance from the nucleus
  - Electron shielding
  - Spin-pair repulsion

Nuclear charge (number of protons)

- As atomic number increases, so does the **positive charge of the nucleus**
- A higher nuclear charge creates a **stronger electrostatic attraction** between the nucleus and outer electrons
- Therefore more energy is required to remove the electron means a **higher ionisation energy**

Distance from the nucleus (atomic radius)

- Outer electrons that are **further from the nucleus** experience **weaker attraction**
- The **greater the distance**, the **easier it is to remove the electron**
- Ionisation energy decreases with increased distance from the nucleus

Electron shielding (inner shell repulsion)

- **Electrons in inner shells** repel electrons in outer shells and **partially block** the attraction of the nucleus
- This is known as the **shielding effect**
- More inner electron shells means greater shielding and therefore a lower ionisation energy

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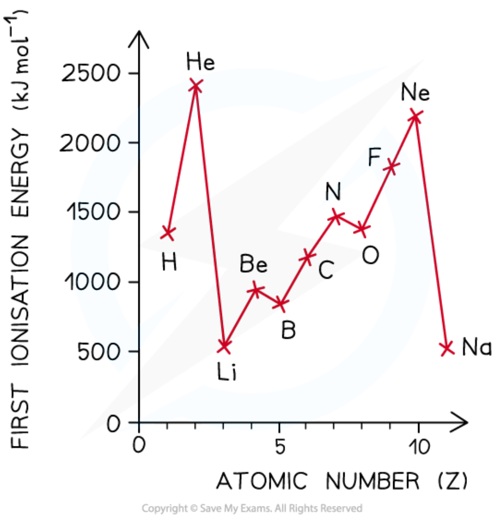
Spin-pair repulsion (electron–electron repulsion)

- In **orbitals that contain two electrons**, the electrons experience **repulsion** due to their like charges
- This repulsion slightly reduces the energy needed to remove one of the electrons
- This explains **small drops in ionisation energy across a period**, such as:
  - **Oxygen vs nitrogen** (O's paired 2p electrons are easier to remove)
  - **Aluminium vs magnesium** (Al's $3p^1$ electron is higher in energy than Mg's $3s^2$)

Ionisation Energy & the Periodic Table

- **Across a period**:
  - first ionisation energy **increases** due to higher nuclear charge and same shielding
- **Down a group**:
  - first ionisation energy **decreases** due to greater atomic radius and more shielding

Graph of first ionisation energies from H to Na



There are ionisation energy trends within periods and groups

Ionisation energy across a period

- **Across a period**, ionisation energy **increases** because:
  - **Nuclear charge increases**, pulling electrons closer
  - **Atomic radius decreases**, reducing the distance to outer electrons
  - **Shielding stays roughly constant**, since electrons are added to the same shell
  - Outer electrons are held more tightly, so **more energy is needed to remove them**

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Dips in the trend

- There is a slight **decrease** in IE_1 between **beryllium** and **boron** as the fifth electron in boron is in the 2p subshell, which is further away from the nucleus than the 2s subshell of beryllium
  - **Beryllium** has a first ionisation energy of **900 kJ mol^{-1}** as its electron configuration is **$1s^2\ 2s^2$**
  - **Boron** has a first ionisation energy of **800 kJ mol^{-1}** as its electron configuration is **$1s^2\ 2s^2\ 2p_x^1$**
- There is a slight **decrease** in IE_1 between **nitrogen** and **oxygen** due to **spin-pair repulsion** in the $2p_x$ orbital of oxygen
  - **Nitrogen** has a first ionisation energy of **1400 kJ mol^{-1}** as its electron configuration is $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$
  - **Oxygen** has a first ionisation energy of **1310 kJ mol^{-1}** as its electron configuration is $1s^2\ 2s^2\ \mathbf{2p_x^2}\ 2p_y^1\ 2p_z^1$
  - In oxygen, there are 2 electrons in the $2p_x$ orbital, so the repulsion between those electrons makes it slightly easier for one of those electrons to be removed

Ionisation energy down a group

- The ionisation energy down a group **decreases** due to the following factors:
  - The number of protons in the atom is increased, so the **nuclear charge** increases
  - But, the atomic radius of the atoms increases as you add more shells of electrons, making the atoms bigger
  - So, the **distance** between the nucleus and outer electron **increases** as you descend the group
  - The **shielding** by inner shell electrons **increases** as there are more shells of electrons
  - These factors outweigh the increased nuclear charge, meaning it becomes **easier to remove the outer electron** as you descend a group
  - So, the ionisation energy decreases

Table summarising ionisation energy trends across a period & down a group

Across a period: Ionisation energy increases	Down a group: Ionisation energy decreases
Increase in nuclear charge	Increase in nuclear charge
The same number of shells	Increased number of shells

Distance from the outer electron to the nucleus decreases	Distance from the outer electron to the nucleus increases
Shielding remains relatively constant	Shielding increases
Decreased atomic / ionic radius	Increased atomic / ionic radius
The attraction between the outer electron and the nucleus gets stronger so the outer electron is **harder** to remove	The attraction between the outer electron and the nucleus gets weaker so the outer electron is **easier** to remove

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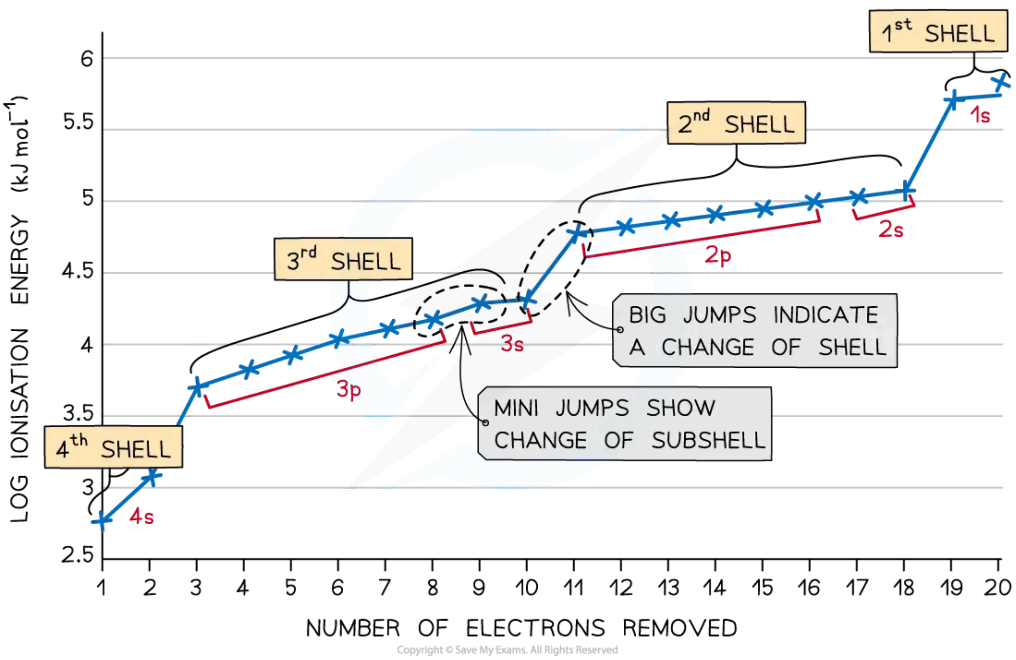
Successive Ionisation Energies of an Element

- **Successive ionisation energies** of an element **increase**
- This is because once the outer electron is removed, the atom becomes a **positive ion**, making further electron removal **more difficult**
- As more electrons are removed:
  - **Shielding decreases**
  - The **proton-to-electron ratio increases**
  - **Attraction between nucleus and remaining electrons increases**
- The increase is **not constant**, it depends on the **electronic configuration**
- Taking calcium as an example:

Table showing the successive ionisation energies of calcium table

Electronic Configuration	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$**	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$**	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$**	**$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$**
IE	First	Second	Third	Fourth
IE ($kJ\ mol^{-1}$)	590	1150	4940	6480

Successive ionisation energies of calcium



Graph to show the successive ionisation energies for the element calcium

- The **first ionisation energy** is relatively low due to **spin-pair repulsion** in the **4s orbital**
- The **second electron** is harder to remove because this **repulsion is no longer present**
- The **third ionisation energy** increases sharply as the electron is removed from the **3p subshell**, which is **closer to the nucleus**
- The **fourth electron** is also more difficult to remove due to **reduced spin-pair repulsion** within the **3p orbital**
- **Successive ionisation energies always increase** because electrons are being removed from an **increasingly positive ion**
- A **large jump** in ionisation energy indicates a **change in shell**, while **smaller jumps** reflect changes **within the same shell or subshell**

Using successive ionisation data

- Helps **predict or confirm electronic configuration**.
- Identifies the **number of outer-shell electrons**.
- Indicates the **group number** by locating the **position of a significant jump** in ionisation energy.
- Commonly applied to elements like **sodium (Na)**, **magnesium (Mg)**, and **aluminium (Al)** to deduce their place in the **Periodic Table**.

Successive ionisation energies table

Element	Atomic Number	Ionisation Energy (kJ mol^{-1})

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		First	Second	Third	Fourth
Na	11	494	4560	6940	9540
Mg	12	736	1450	7740	10500
Al	13	577	1820	2740	11600

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Sodium

- A large increase between the **first and second ionisation energies** shows that the first electron is much easier to remove than the second
- This first electron is in the **3s subshell**
  - This is the outermost shell
  - This electron experiences **less nuclear attraction** due to shielding and distance from the nucleus
- This suggests Na has **one valence electron**, so it belongs to **Group 1**
- The jump corresponds to removing the next electron from the **2p subshell**, which is closer to the nucleus and more strongly held
  - **Na:** $1s^2 2s^2 2p^6 \mathbf{3s^1}$

Magnesium

- A large increase between the **second and third ionisation energies** shows that the first two electrons are easier to remove than the third
- These two electrons are in the **3s subshell**
  - This is the outermost shell
  - These two electrons experience **less nuclear attraction** due to shielding and distance from the nucleus
- This suggests Mg has **two valence electrons**, so it belongs to **Group 2**
- The jump corresponds to removing the next electron from the **2p subshell**, which is held more tightly
  - **Mg:** $1s^2 2s^2 2p^6 \mathbf{3s^2}$

Aluminium

- A large increase between the third and fourth ionisation energies shows that the first three electrons are easier to remove
- These three electrons are in the **3s and 3p subshells**
  - These are part of the outermost shell

- These three electrons experience **less nuclear attraction** due to shielding and distance from the nucleus

- This suggests Al has **three valence electrons**, so it belongs to **Group 13 (Group III)**

- The jump corresponds to removing the next electron from the **2p subshell**, which is closer to the nucleus and more strongly held
  - **Al:** $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2\ 3p^1}$

Examiner Tips and Tricks

- It is easy to remove electrons from a full subshell as they undergo **spin-pair repulsion.**
- It gets more difficult to remove electrons from **principal quantum shells** that get closer to the nucleus as there is less **shielding** and an increase in **attractive forces** between the electrons and nuclear charge.

Your notes

Ionisation Energies: Electronic Configuration

Using Successive Ionisation Energy Data

- Each ionisation energy refers to the energy required to remove one one mole of electrons from one mole of a specific species in the gaseous state
  - **First ionisation energy** is the energy required to remove one mole of electrons from one mole of **gaseous atoms**
  - **Successive ionisation energies** are the energy required to remove one mole of electrons from one mole of **gaseous positive ions**
- As electrons are removed, the ion becomes more positively charged
  - This increases the **electrostatic attraction** between the nucleus and remaining electrons
  - Therefore, **ionisation energies increase** progressively
- A **large jump** in successive ionisation energies indicates that:
  - An electron is being removed from a new, **inner shell** closer to the nucleus
  - These electrons experience a stronger attraction to the nucleus and require more energy to remove

How to deduce electronic configuration

- **Look for the largest increase** between two successive ionisation energies
- The number of electrons removed **before** this jump = **number of outer-shell (valence) electrons**
- This corresponds to the element's **group number** in the Periodic Table (for s- and p-block elements)

Worked Example

Deduce the electronic configuration of element Z, a Period 3 element, using successive ionisation energy data.

IE	5th	6th	7th	8th
IE (kJ mol^{-1})	6542	9362	11,018	33,606

Answer

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- You are told that the element is in **Period 3**
- Period 3 elements have electrons up to the 3rd shell:
  - 3s and 3p subshells are included
- There is a large jump between the 7th and 8th IE, so element Z must be in Group 17
- So element Z is in Period 3 and Group 17:
  - This corresponds to chlorine, Cl
- The electronic configuration of Cl is:
  - **$1s^2 2s^2 2p^6 3s^2 3p^5$**