# Atomic Structure and the Periodic Table
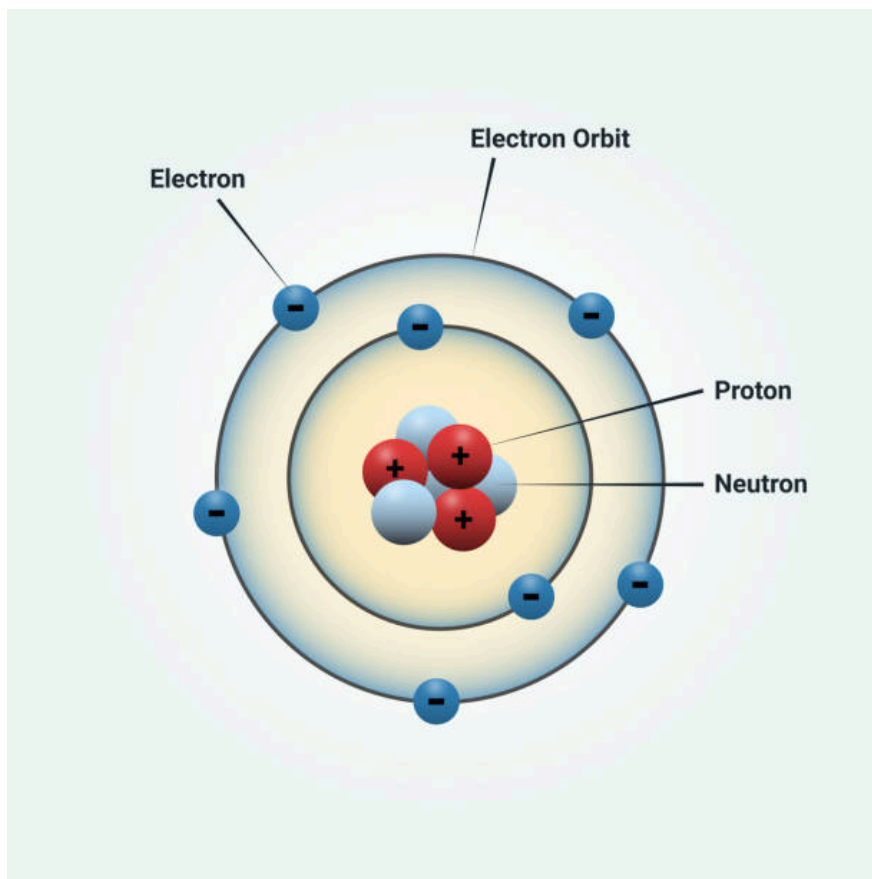


Edexcel International A-Level Chemistry

Topic 2 of Unit 1

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# Basic Atomic Structure

## ⚗ What is an Atom?

An atom is the **smallest unit of an element** that retains the chemical properties of that element. It consists of three main subatomic particles:
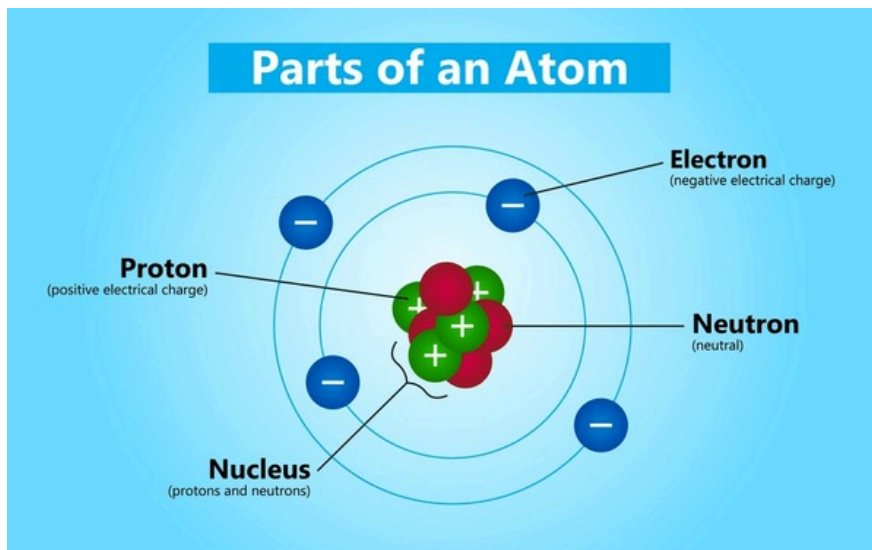


Atoms have a **central nucleus** containing protons and neutrons, with electrons orbiting around this nucleus in energy levels or shells.

## ❖ Subatomic Particles

| Particle | Symbol | Location | Relative Mass | Relative Charge |
|---|---|---|---|---|
| Proton | $p^+$ | Nucleus | 1 | +1 |
| Neutron | $n^0$ | Nucleus | 1 | 0 (neutral) |
| Electron | $e^-$ | Orbiting nucleus | 1/1840 (≈0) | -1 |

**Key Points:**

- Protons and neutrons have nearly the same mass, while electrons are much lighter
- Protons and electrons have equal but opposite charges
- Neutrons have no charge (neutral)
- The mass of an atom is concentrated in the nucleus

## 💡 Formal Analogy: The Solar System

Think of an atom like a miniature solar system. The nucleus is like the sun at the center, containing most of the mass. The electrons are like planets orbiting around the sun. Just as planets are much lighter than the sun, electrons are much lighter than the nucleus. The gravitational attraction between the sun and planets is similar to the electrostatic attraction between the positively charged nucleus and negatively charged electrons.

# Atomic Number and Mass Number

## # Atomic (Proton) Number

### 🔖 Definition

The **atomic number (Z)** is the number of protons in the nucleus of an atom. It is also called the **proton number**.

**Key Points:**

- Each element has a unique atomic number
- In a neutral atom, the number of protons equals the number of electrons
- The atomic number determines the element's identity
- Elements are arranged in the periodic table in order of increasing atomic number

## ⚗ Mass Number

### 🔖 Definition

The **mass number (A)** is the total number of protons and neutrons in the nucleus of an atom.

**Key Points:**

- Mass number = number of protons + number of neutrons
- Since electrons have negligible mass, they are not included in the mass number
- Isotopes of the same element have the same atomic number but different mass numbers
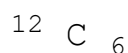
## Σ Atomic Notation

Atoms are represented using the following notation:

$$^{A}_{Z}X$$
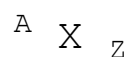
Where:

- X = chemical symbol of the element
- A = mass number (protons + neutrons)
- Z = atomic number (protons)

### ⚗ Example: Carbon-12

$$^{12}\mathrm{C}_{6}$$

- Atomic number (Z) = 6 (6 protons)
- Mass number (A) = 12 (6 protons + 6 neutrons)
- Number of neutrons = A - Z = 12 - 6 = 6
- In a neutral atom: 6 electrons

### 💡 Formal Analogy: Student ID Cards

Think of atomic number and mass number like a student ID card. The atomic number is like the student's unique ID number that identifies which school they belong to (which element they are). The mass number is like the student's total number of classes taken (protons + neutrons). Just as students in the same school have the same school ID number but may be taking different numbers of classes, atoms of the same element have the same atomic number but can have different mass numbers (isotopes).

# Determining Subatomic Particles

## ▦ Key Formulas

**For neutral atoms:**

Number of protons = Atomic number (Z)

Number of electrons = Atomic number (Z)

Number of neutrons = Mass number (A) - Atomic number (Z)

**For ions:**

Number of protons = Atomic number (Z)

Number of electrons = Atomic number (Z) - Charge

Number of neutrons = Mass number (A) - Atomic number (Z)

## ⚒ Examples for Neutral Atoms

### ✦ Example 1: Oxygen-16

$$^{16}\text{O}_{8}$$
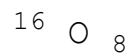
- Atomic number (Z) = 8
- Mass number (A) = 16
- Number of protons = 8
- Number of electrons = 8
- Number of neutrons = 16 - 8 = 8

### ✦ Example 2: Chlorine-35

$$^{35}\text{Cl}_{17}$$
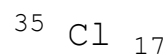
- Atomic number (Z) = 17
- Mass number (A) = 35
- Number of protons = 17
- Number of electrons = 17
- Number of neutrons = 35 - 17 = 18

## ⚡ Examples for Ions

## ✦ Example 1: Sodium ion (Na⁺)

$$^{23}\text{Na}_{11}{}^{+}$$

- Atomic number (Z) = 11
- Mass number (A) = 23
- Number of protons = 11
- Number of electrons = 11 - 1 = 10
- Number of neutrons = 23 - 11 = 12

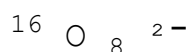## ✦ Example 2: Oxide ion (O²⁻)

$$^{16}\text{O}_{8}{}^{2-}$$

- Atomic number (Z) = 8
- Mass number (A) = 16
- Number of protons = 8
- Number of electrons = 8 - (-2) = 10
- Number of neutrons = 16 - 8 = 8

## 💡 Formal Analogy: Library Books

Think of determining subatomic particles like managing a library. The atomic number is like the unique catalog number for a book series (identifies the element). The mass number is like the total number of pages in a specific book (protons + neutrons). For neutral atoms, the number of electrons equals the number of protons, like having a complete book with all its chapters. For ions, electrons are either lost (positive charge) or gained (negative charge), like chapters being removed from or added to a book, changing its completeness but not its fundamental identity.

# Isotopes

## ⚘ What are Isotopes?

### 🔖 Definition

**Isotopes** are atoms of the same element that have the same number of protons but different numbers of neutrons.

**Key Properties:**

- Same atomic number (Z) - same number of protons
- Different mass numbers (A) - different numbers of neutrons
- Same chemical properties - determined by electron arrangement
- Different physical properties - due to different masses

## ⚗ Examples of Isotopes

### ✨ Hydrogen Isotopes

- $^1H$ (Protium): 1 proton, 0 neutrons
- $^2H$ (Deuterium): 1 proton, 1 neutron
- $^3H$ (Tritium): 1 proton, 2 neutrons

### ✨ Carbon Isotopes

- $^{12}C$ : 6 protons, 6 neutrons (98.9% abundance)
- $^{13}C$ : 6 protons, 7 neutrons (1.1% abundance)
- $^{14}C$ : 6 protons, 8 neutrons (radioactive)

### ✨ Chlorine Isotopes

- $^{35}Cl$ : 17 protons, 18 neutrons (75.8% abundance)
- $^{37}Cl$ : 17 protons, 20 neutrons (24.2% abundance)

## ⇋ Comparison of Isotopes

| Property | Same for Isotopes | Different for Isotopes |
|---|:---:|:---:|
| Atomic number (Z) | ✓ | ✗ |
| Number of protons | ✓ | ✗ |
| Number of electrons (in neutral atoms) | ✓ | ✗ |
| Chemical properties | ✓ | ✗ |
| Mass number (A) | ✗ | ✓ |
| Number of neutrons | ✗ | ✓ |
| Physical properties (density, melting point) | ✗ | ✓ |

## ✪ Importance of Isotopes

**Applications:**

- **Radioactive dating:** Carbon-14 used to determine the age of organic materials
- **Medical applications:** Iodine-131 for thyroid treatment, Technetium-99m for diagnostics
- **Tracer studies:** Deuterium and Oxygen-18 used to track chemical reactions
- **Nuclear power:** Uranium-235 used as fuel in nuclear reactors
- **Stable isotope analysis:** Used in archaeology, geology, and environmental science

## 💡 Formal Analogy: Siblings in a Family

Think of isotopes like siblings in the same family. They share the same "parents" (element) and have the same fundamental characteristics (same number of protons). However, they have different "weights" or "sizes" (different numbers of neutrons). Just as siblings might have different heights or weights but share the same family name and genetic traits, isotopes have different masses but share the same chemical properties and belong to the same element.

# Introduction to Mass Spectrometry

## ⚗ What is Mass Spectrometry?

### 🔖 Definition

**Mass spectrometry** is an analytical technique that measures the mass-to-charge ratio (m/z) of ions to identify and quantify molecules in a sample.

**Key Applications:**

- Determining isotopic composition of elements
- Calculating relative atomic masses
- Identifying unknown compounds
- Quantifying substances in complex mixtures
- Studying molecular structures

## 🔬 The Mass Spectrometer
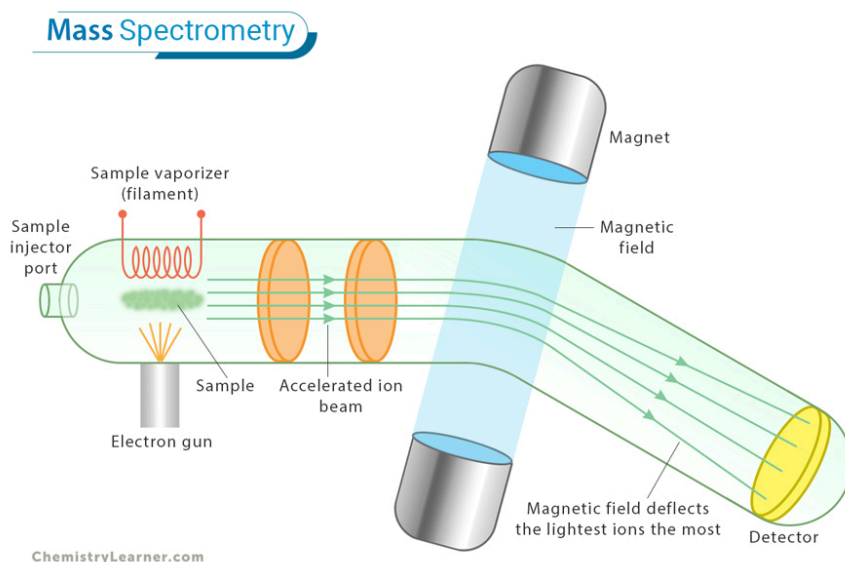


A mass spectrometer consists of four main components that work together to separate and detect ions based on their mass-to-charge ratios.

## ☰ How It Works

1. **Ionization**
   Sample atoms are converted into positive ions by bombarding them with high-energy electrons. This removes electrons from the atoms,

creating positively charged ions.

**2** **Acceleration**

The positive ions are accelerated by an electric field, giving them the same kinetic energy. Lighter ions achieve higher velocities than heavier ions.

**3** **Deflection**

The ions pass through a magnetic field that deflects their paths. The degree of deflection depends on the mass-to-charge ratio (m/z). Lighter ions are deflected more than heavier ions.

**4** **Detection**

Ions with specific m/z values reach the detector, which measures their abundance. The detector produces an electrical signal proportional to the number of ions hitting it.

## ⚡ Mass Spectrum

The output of a mass spectrometer is a **mass spectrum**, which is a graph showing:

- X-axis: Mass-to-charge ratio (m/z)
- Y-axis: Relative abundance (percentage of total ions)

Each peak in the spectrum represents ions with a specific m/z value. The height of the peak indicates the relative abundance of those ions.

## 💡 Formal Analogy: Olympic High Jump

Think of mass spectrometry like an Olympic high jump competition. All athletes (ions) start with the same energy (acceleration phase). When they approach the bar (magnetic field), lighter athletes can jump higher (are deflected more) than heavier athletes. The judges (detector) record how high each athlete jumped and how many athletes cleared each height. This creates a record (mass spectrum) showing the distribution of athletes by their jumping ability (mass-to-charge ratio).

# Analyzing Mass Spectra

## 📊 Understanding Mass Spectra

A **mass spectrum** is a graph that shows the relative abundance of ions with different mass-to-charge ratios (m/z). Each peak represents ions with a specific m/z value.

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**Example: Chlorine Mass Spectrum**

75%

25%

35                    37

X-axis: Mass-to-charge ratio (m/z)

Y-axis: Relative abundance (%)

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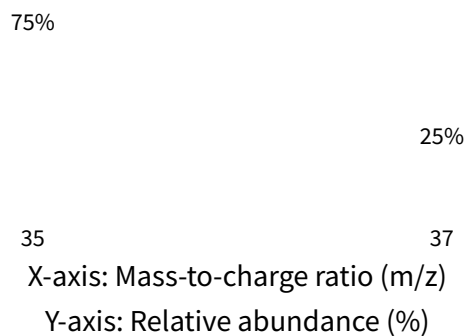
**Key Information from Mass Spectra:**

- Position of peaks indicates m/z values
- Height of peaks shows relative abundance
- Pattern of peaks reveals isotopic composition

## ⚌ Calculating Relative Atomic Mass

To calculate the relative atomic mass (Ar) from a mass spectrum:

1 **Identify Isotopes**

Determine the m/z values of all peaks, which correspond to the isotopes present

2 **Determine Abundance**

Find the relative abundance of each isotope from the peak heights

3 **Apply Formula**

Use the formula: `Ar = Σ (isotope mass × fractional abundance)`

> ⚗ **Example: Chlorine**
>
> From the spectrum above:
> - Isotope 35Cl: 75% abundance

- Isotope 37Cl: 25% abundance

  Ar(Cl) = (35 × 0.75) + (37 × 0.25) = 26.25 + 9.25 = 35.5

## 🔬 Analyzing Molecular Mass Spectra

For molecules, mass spectra can reveal:

- **Molecular ion peak (M⁺):** The peak with the highest m/z value, representing the intact molecule
- **Fragment ions:** Peaks with lower m/z values, representing fragments of the molecule
- **Isotopic patterns:** Characteristic patterns that help identify elements like Cl, Br, S

**Special Cases:**

- **2+ ions:** Some molecules form ions with 2+ charge, appearing at m/z = $M/2$
- **Isotope clusters:** Elements with multiple isotopes create characteristic patterns

## 💡 Formal Analogy: Orchestra Performance

Think of analyzing a mass spectrum like listening to an orchestra performance. Each instrument section (isotope) produces a distinct sound (peak) at a specific pitch (m/z value). The volume of each section (peak height) tells you how many musicians are playing (relative abundance). By listening to all the sounds together, you can identify the composition of the entire orchestra (element or molecule) and even determine the average "weight" of the performance (relative atomic mass).

# Calculating Relative Atomic Mass

## Σ  The Formula

**Relative Atomic Mass (Ar)** is the weighted average of the masses of all naturally occurring isotopes of an element.

$$Ar = Σ \text{ (isotope mass × fractional abundance)}$$

Where Σ means "sum of all isotopes"

**Key Points:**

- Use fractional abundance (decimal), not percentage
- Convert percentages to fractions by dividing by 100
- The result is a weighted average, not a simple average
- More abundant isotopes have greater influence on the final value

## ▦  Step-by-Step Method

1  **Identify Isotopes**
List all naturally occurring isotopes of the element with their masses

2  **Convert Abundances**
Convert percentage abundances to fractional abundances (divide by 100)

3  **Multiply**
Multiply each isotope mass by its fractional abundance

4  **Sum**
Add all the products together to get the relative atomic mass

## ⚗  Worked Examples

### ✦ Example 1: Chlorine

Chlorine has two naturally occurring isotopes:
- $^{35}$Cl: 75.77% abundance
- $^{37}$Cl: 24.23% abundance

| Isotope | Mass | Abundance | Product |
|---|---|---|---|
| $^{35}Cl$ | 34.969 | 0.7577 | 26.49 |
| $^{37}Cl$ | 36.966 | 0.2423 | 8.96 |
| | | Sum: | 35.45 |

Ar(Cl) = 35.45

## ✦ Example 2: Magnesium

Magnesium has three naturally occurring isotopes:
- $^{24}Mg$: 78.99% abundance
- $^{25}Mg$: 10.00% abundance
- $^{26}Mg$: 11.01% abundance

Ar(Mg) = (23.985 × 0.7899) + (24.986 × 0.1000) + (25.983 × 0.1101)
Ar(Mg) = 18.95 + 2.50 + 2.87 = 24.32

## ⚠ Common Mistakes to Avoid

- Using percentages instead of fractional abundances
- Forgetting to convert abundances to decimals (divide by 100)
- Using simple average instead of weighted average
- Not including all naturally occurring isotopes
- Using incorrect isotope masses

## 💡 Formal Analogy: Class Grades

Think of calculating relative atomic mass like calculating a weighted average grade for a class. Different assignments (isotopes) have different weights (abundances) in the final grade. A test worth 75% of the grade (more abundant isotope) will have a much greater impact on the final grade than a quiz worth only 25% (less abundant isotope). Just as you'd multiply each assignment grade by its weight before adding them together, you multiply each isotope mass by its fractional abundance before summing to get the relative atomic mass.

# Determining Relative Molecular Mass

## ⚗ Relative Molecular Mass

### 🔖 Definition

**Relative Molecular Mass (Mr)** is the sum of the relative atomic masses of all atoms in a molecule.

**Key Points:**

- Used for discrete molecules (not ionic compounds with giant structures)
- Calculated by adding the relative atomic masses of all atoms in the molecular formula
- Has no units (it's a ratio compared to 1/12 of the mass of a carbon-12 atom)

## 📊 From Mass Spectra

Mass spectrometry can be used to determine the relative molecular mass of a compound by:

1. **Identify Molecular Ion Peak**
   Find the peak with the highest m/z value, which represents the intact molecule ($M^+$)

2. **Check for 2+ Ions**
   Some molecules form ions with 2+ charge, appearing at m/z = M/2

3. **Consider Isotopic Patterns**
   Elements with multiple isotopes create characteristic patterns that help identify the molecule

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**Example: Mass Spectrum of Ethanol ($C_2H_6O$)**

$M^+$

M-1

$CH_3O^+$

$C_2H_5^+$

$C_2H_3^+$

| 46 | 45 | 31 | 29 | 27 |
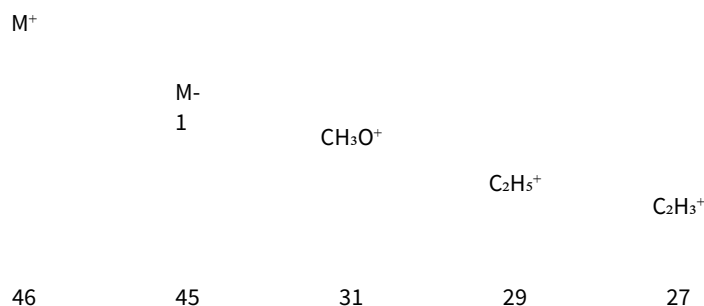
> The molecular ion peak at m/z = 46 indicates Mr = 46

## ⊞ Calculating Mr from Formula

### ✦ Example 1: Carbon Dioxide ($CO_2$)

$Mr(CO_2) = Ar(C) + 2 \times Ar(O)$
$Mr(CO_2) = 12.0 + 2 \times 16.0 = 44.0$

### ✦ Example 2: Glucose ($C_6H_{12}O_6$)

$Mr(C_6H_{12}O_6) = 6 \times Ar(C) + 12 \times Ar(H) + 6 \times Ar(O)$
$Mr(C_6H_{12}O_6) = 6 \times 12.0 + 12 \times 1.0 + 6 \times 16.0 = 72.0 + 12.0 + 96.0 = 180.0$

## 🔬 Identifying Molecules

Mass spectrometry can help identify unknown compounds by:

- Determining the molecular mass from the molecular ion peak
- Analyzing fragmentation patterns to deduce structural information
- Comparing with reference spectra in databases
- Using isotopic patterns to identify specific elements

**Special Cases:**

- **Chlorine-containing compounds:** Show characteristic 3:1 ratio of M and M+2 peaks
- **Bromine-containing compounds:** Show characteristic 1:1 ratio of M and M+2 peaks
- **Sulfur-containing compounds:** Show characteristic 95:5 ratio of M and M+2 peaks

### 💡 Formal Analogy: Fingerprint Identification

Think of determining relative molecular mass from mass spectra like identifying a person from their fingerprint. The molecular ion peak is like the primary fingerprint pattern that tells you the person's identity (the molecule's mass). The fragment ions are like the smaller details in the fingerprint that provide additional information about the person's characteristics (the molecule's structure). Just as experts can identify people by comparing fingerprints to a database, chemists can identify compounds by comparing their mass spectra to reference spectra.

# Predicting Mass Spectra

## ⚗ Diatomic Molecules in Mass Spectrometry

### 🔖 Key Concept

When diatomic molecules containing isotopes are analyzed in a mass spectrometer, they produce characteristic patterns of peaks that can be predicted based on the isotopic abundances.

**Important Points:**

- Diatomic molecules can form from different combinations of isotopes
- The relative peak heights are proportional to the probability of each combination
- Chlorine and bromine are classic examples with distinctive patterns

## 🔢 Predicting Chlorine Mass Spectrum

Chlorine has two naturally occurring isotopes:

- $^{35}Cl$: 75.8% abundance
- $^{37}Cl$: 24.2% abundance

1   **Identify Possible Combinations**
    $^{35}Cl$-$^{35}Cl$, $^{35}Cl$-$^{37}Cl$, and $^{37}Cl$-$^{37}Cl$

2   **Calculate Masses**
    35+35=70, 35+37=72, 37+37=74

3   **Calculate Probabilities**
    $^{35}Cl$-$^{35}Cl$: $0.758 \times 0.758 = 0.575$ (57.5%)
    $^{35}Cl$-$^{37}Cl$: $2 \times (0.758 \times 0.242) = 0.367$ (36.7%)
    $^{37}Cl$-$^{37}Cl$: $0.242 \times 0.242 = 0.059$ (5.9%)

---

### Predicted Mass Spectrum of Chlorine ($Cl_2$)

57.5%

36.7%

5.9%

70                72                74

## ⚓ Other Diatomic Examples

### ✦ Bromine (Br₂)

Bromine has two isotopes with nearly equal abundance:
- $^{79}Br$: 50.7% abundance
- $^{81}Br$: 49.3% abundance

This produces peaks at m/z = 158, 160, and 162 with approximately 1:2:1 ratio

### ✦ Nitrogen (N₂)

Nitrogen has two isotopes:
- $^{14}N$: 99.63% abundance
- $^{15}N$: 0.37% abundance

This produces a very small peak at m/z = 30 ($^{14}N$-$^{15}N$) and an almost negligible peak at m/z = 32 ($^{15}N$-$^{15}N$)

## ⚠ Important Considerations

- For molecules with two different isotopes, the middle peak (containing one of each isotope) is twice as likely as the outer peaks
- Peak heights are proportional to the probability of each combination
- For isotopes with very different abundances, the peaks containing the rarer isotope will be much smaller
- These patterns help identify elements in unknown compounds

### 💡 Formal Analogy: Coin Flipping

Think of predicting mass spectra of diatomic molecules like predicting the outcomes of flipping two coins. If you have a fair coin (50% heads, 50% tails), the possible outcomes are HH, HT, TH, and TT. The probability of getting two heads (HH) is $0.5 \times 0.5 = 0.25$, the same as getting two tails (TT). However, getting one head and one tail (HT or TH) has a probability of $2 \times (0.5 \times 0.5) = 0.5$, which is twice as likely as the other outcomes. Similarly, in diatomic molecules with two isotopes of similar abundance, the combination with one of each isotope is twice as likely as the combinations with two of the same isotope.

# Ionisation Energies

## ⚡ What is Ionisation Energy?

### 🔖 Definition

**Ionisation energy** is the energy required to remove an electron from a gaseous atom or ion.
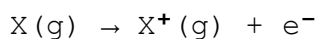
**Key Points:**

- Measured in kilojoules per mole ($kJ\ mol^{-1}$)
- All ionisation energies are **endothermic** (require energy input)
- Higher ionisation energies require more energy to remove electrons

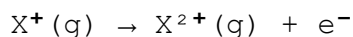## 🔢 First, Second and Third Ionisation Energies

### 1️⃣ First Ionisation Energy

The energy required to remove the **first electron** from a neutral gaseous atom.

$$X(g)\ \rightarrow\ X^+(g)\ +\ e^-$$

### 2️⃣ Second Ionisation Energy

The energy required to remove the **second electron** from a singly charged gaseous ion.

$$X^+(g)\ \rightarrow\ X^{2+}(g)\ +\ e^-$$

### 3️⃣ Third Ionisation Energy

The energy required to remove the **third electron** from a doubly charged gaseous ion.

$$X^{2+}(g)\ \rightarrow\ X^{3+}(g)\ +\ e^-$$
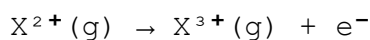
## 📈 Why Ionisation Ion Energies are Endothermic

**Electrostatic Attraction:**

- Electrons are negatively charged and attracted to the positively charged nucleus
- Energy must be supplied to overcome this electrostatic attraction
- The stronger the attraction, the more energy required

> Energy must be supplied to move an electron from a lower energy state (bound to atom) to a higher energy state (free electron)

## ↱ Successive Ionisation Energies

**Pattern:**

- Second ionisation energy is always greater than the first
- Third ionisation energy is always greater than the second
- Large jumps occur when removing electrons from a more stable electron configuration

> ### ⚗ Example: Sodium (Na)
>
> - First IE: 496 kJ mol$^{-1}$ (removing $3s^1$ electron)
> - Second IE: 4562 kJ mol$^{-1}$ (removing $2p^6$ electron)
> - Large jump occurs when removing electrons from a noble gas configuration

### 💡 Formal Analogy: Climbing a Mountain

Think of ionisation energy like the effort required to climb a mountain. The first ionisation energy is like climbing from the base camp to the first camp - it requires significant effort. The second ionisation energy is like climbing from the first camp to the second camp - it requires even more effort because you're already tired and the path might be steeper. The third ionisation energy requires even more effort. Each successive climb requires more energy than the previous one, just as each successive ionisation requires more energy than the previous one.

# Atomic Orbitals

## ❖ What is an Orbital?

### 🔖 Definition

An **orbital** is a region within an atom that can hold up to two electrons with opposite spins. It represents the most probable location of finding an electron.
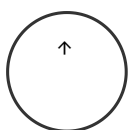
**Key Concepts:**

- Orbitals are three-dimensional regions around the nucleus
- Each orbital can hold a maximum of two electrons
- Electrons in the same orbital must have opposite spins
- Orbitals have different shapes and energy levels
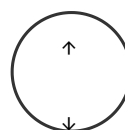
## ↻ Electron Spin

Electrons have a property called **spin**, which can be:

- **Spin up (↑)** - often represented as +½
- **Spin down (↓)** - often represented as -½

**Pauli Exclusion Principle:** No two electrons in an atom can have the same set of quantum numbers. This means that if two electrons occupy the same orbital, they must have opposite spins.
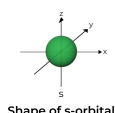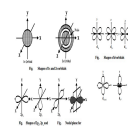


**One electron (spin up)**



**Two electrons (opposite spins)**

## ▲ Types of Orbitals

There are four main types of orbitals, each with a distinct shape:



Shape of s-orbital

**s-orbital**
**(spherical)**

**p-orbital**
**(dumbbell-shaped)**

**s-orbitals:** Spherical shape, can hold 2 electrons

**p-orbitals:** Dumbbell shape, come in sets of 3 (px, py, pz), each can hold 2 electrons

**d-orbitals:** More complex shapes, come in sets of 5, each can hold 2 electrons

**f-orbitals:** Even more complex shapes, come in sets of 7, each can hold 2 electrons

## ⭐ Key Points to Remember

- Each orbital can hold a maximum of two electrons
- Electrons in the same orbital must have opposite spins
- Orbitals are grouped into shells (energy levels) and sub-shells
- The first shell has only one s-orbital (1s)
- The second shell has one s-orbital (2s) and three p-orbitals (2px, 2py, 2pz)
- The third shell has one s-orbital (3s), three p-orbitals (3px, 3py, 3pz), and five d-orbitals

## 💡 Formal Analogy: Hotel Rooms

Think of orbitals like hotel rooms in a building (atom). Each room (orbital) can accommodate at most two guests (electrons). If two guests share a room, they must have opposite preferences (spins) - one likes the window side, the other prefers the door side. Just as hotels have different types of rooms (single, double, suite) with different capacities and layouts, atoms have different types of orbitals (s, p, d, f) with different shapes and capacities. The arrangement of guests in rooms follows specific rules, just as electrons fill orbitals according to specific principles.

# Factors Affecting Ionisation Energies

## 📈 Key Factors

Three main factors influence ionisation energy:

- **Nuclear charge** - Number of protons in the nucleus
- **Electron shielding** - Repulsion from inner electrons
- **Sub-shell** - Type of orbital the electron occupies

## ⊕ Nuclear Charge

### 🔖 Effect

Higher nuclear charge → Stronger attraction to electrons → Higher ionisation energy

+1          +2
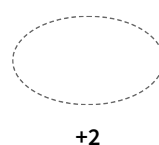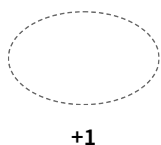
**Higher nuclear charge → Higher ionisation energy**

## 🛡 Electron Shielding

### 🔖 Effect

More shielding → Weaker attraction to nucleus → Lower ionisation energy

Less shielding          More shielding

**More shielding → Lower ionisation energy**

## ◆ Sub-shell

## 🔖 Effect

Electrons in different sub-shells have different ionisation energies:

- s < p < d < f (in the same shell)
- Lower energy sub-shells are closer to the nucleus
- Electrons in lower energy sub-shells are held more tightly

**s < p < d < f → Increasing ionisation energy**

## ↱ Combined Effects

**Across a period (left to right):**

- Nuclear charge increases
- Shielding remains similar
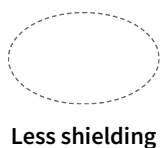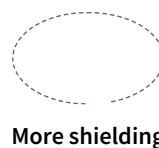- Sub-shell may change
- Overall: Ionisation energy generally increases

**Down a group:**

- Nuclear charge increases
- Shielding increases significantly
- Sub-shell remains the same
- Overall: Ionisation energy generally decreases

## 💡 Formal Analogy: Magnet and Paper Clips

Think of ionisation energy like the force needed to pull paper clips away from a magnet. The strength of the magnet (nuclear charge) determines how strongly it attracts the paper clips (electrons). If you add more paper clips between the magnet and the one you're trying to remove (electron shielding), it becomes easier to pull away. The distance from the magnet (sub-shell) also affects the force needed - paper clips closer to the magnet are held more tightly than those farther away. The actual force needed depends on all these factors working together.

# Evidence for Electronic Structure

## ⚗ Ionisation Energies as Evidence

Ionisation energy data provides crucial evidence for the existence of:

- **Quantum shells** - Energy levels where electrons reside
- **Electron sub-shells** - Divisions within quantum shells

By analyzing patterns in successive ionisation energies, we can deduce the electronic structure of atoms.

## ◈ Evidence for Quantum Shells

### ▮ Successive Ionisation Energies

Large jumps in successive ionisation energies indicate when electrons are being removed from a new, inner shell.

| Ionisation | Na (kJ mol$^{-1}$) | Mg (kJ mol$^{-1}$) | Al (kJ mol$^{-1}$) |
|------------|--------------------|--------------------|--------------------|
| 1st | 496 | 738 | 578 |
| 2nd | 4562 | 1451 | 1817 |
| 3rd | 6910 | 7733 | 2745 |
| 4th | 9543 | 10542 | 11577 |

**Key Observations:**

- Na: Large jump between 1st and 2nd IE → 1 electron in outer shell
- Mg: Large jump between 2nd and 3rd IE → 2 electrons in outer shell
- Al: Large jump between 3rd and 4th IE → 3 electrons in outer shell

## ♣ Evidence for Electron Sub-shells

### ▮ First Ionisation Energies Across a Period

The pattern of first ionisation energies across a period provides evidence for the existence of sub-shells.

## First Ionisation Energies of Period 3 Elements
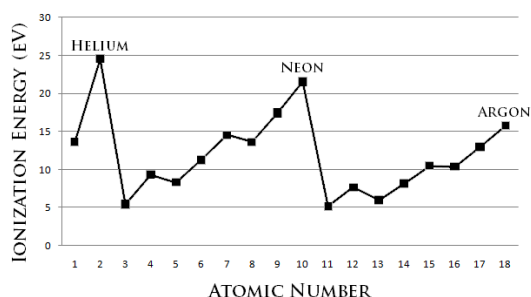


**Key Observations:**

- General increase across the period (due to increasing nuclear charge)
- Dips at Al and S (electrons entering new sub-shells)
- Al: electron enters 3p sub-shell (higher energy than 3s)
- S: electron pairs in 3p sub-shell (electron-electron repulsion)

## 💡 Formal Analogy: Apartment Building

Think of electronic structure like an apartment building. The floors represent quantum shells, and the apartments on each floor represent sub-shells. Just as it takes more energy to move someone from a lower floor to outside the building (higher ionisation energy), it takes more energy to remove electrons from inner shells. The pattern of energy required to move people out reveals the building's structure - some floors have more residents (more electrons), and some apartments are easier to leave than others (different sub-shells). By analyzing how much energy is needed to remove each person, we can deduce the building's layout, just as ionisation energies reveal electronic structure.

# Shapes of s and p Orbitals

## ❖ What are Orbitals?

**Orbitals** are three-dimensional regions around the nucleus where electrons are most likely to be found. They have specific shapes that determine the probability distribution of electrons.

**Key Points:**

- Orbitals are not fixed paths but probability distributions
- Different types of orbitals have different shapes
- The shape of an orbital affects the chemical properties of elements

## ◯ s Orbitals
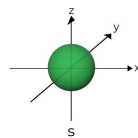


Shape of s-orbital

**s Orbital**
Spherical shape

**Characteristics of s Orbitals:**

- **Spherical shape** - Uniform probability distribution in all directions
- **One per energy level** - Only one s orbital in each principal quantum shell
- **No angular nodes** - No regions where electron probability is zero
- **Lowest energy** - s orbitals have the lowest energy in each shell

**Examples:** 1s, 2s, 3s, 4s...

## △ p Orbitals
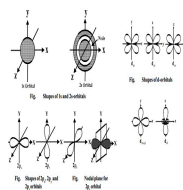


**p$_z$ Orbital**
Dumbbell shape
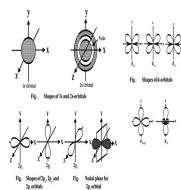


**p$_x$ Orbital**
Dumbbell shape
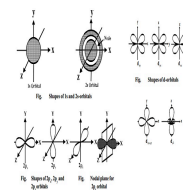


**p$_y$ Orbital**
Dumbbell shape

**Characteristics of p Orbitals:**

- **Dumbbell shape** - Two lobes with a node at the nucleus
- **Three per energy level** - px, py, and pz orbitals (from n=2 onwards)
- **One angular node** - A region where electron probability is zero
- **Higher energy than s** - p orbitals have higher energy than s orbitals in the same shell

**Examples:** 2px, 2py, 2pz, 3px, 3py, 3pz...

## ✪ Key Points to Remember

- s orbitals are spherical, p orbitals are dumbbell-shaped
- Each orbital can hold a maximum of two electrons with opposite spins
- p orbitals come in sets of three (px, py, pz) oriented perpendicular to each other
- The shape of orbitals affects how atoms bond and form molecules
- Higher energy shells have more complex orbital shapes (d and f orbitals)

## 💡 Formal Analogy: Cloud Shapes

Think of orbitals like different shapes of clouds in the sky. An s orbital is like a perfectly spherical cloud, with the same density in all directions. A p orbital is like a dumbbell-shaped cloud, with two distinct lobes and a clear region of no cloud in the middle. Just as different cloud shapes affect how light passes through them, different orbital shapes affect how atoms interact with each other and form chemical bonds. The "cloud" represents where you're most likely to find an electron, not a fixed path.

# Electron Filling Rules

## ≔ Principles of Electron Filling

Electrons fill orbitals according to specific rules that determine the electronic configuration of atoms.

### 1 Hund's Rule

**Each orbital in a sub-shell is singly occupied before any orbital is doubly occupied.** All electrons in singly occupied orbitals have the same spin.

### 2 Pauli Exclusion Principle

**No two electrons in an atom can have the same set of quantum numbers.** This means that if two electrons occupy the same orbital, they must have opposite spins.

### 3 Aufbau Principle

**Electrons fill the lowest energy orbitals first.** The order of filling is determined by the (n+l) rule, where n is the principal quantum number and l is the azimuthal quantum number.

## ⊞ Orbital Filling Diagram

2p ☐ ☐ ☐

2s ◯

1s ◯
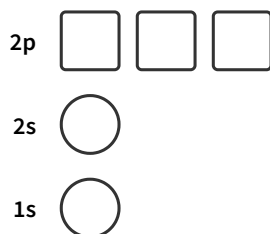
**Key Observations:**

- Each 2p orbital gets one electron before any gets a second electron
- All singly occupied orbitals have electrons with the same spin (up in this case)
- When pairing occurs, electrons have opposite spins

## ⚗ Examples

### ✦ Carbon (6 electrons)

Electronic configuration: $1s^2\ 2s^2\ 2p^2$

**2p** ☐ ☐ ☐

### ✦ Nitrogen (7 electrons)

Electronic configuration: $1s^2\ 2s^2\ 2p^3$

**2p** ☐ ☐ ☐

### ✦ Oxygen (8 electrons)

Electronic configuration: $1s^2\ 2s^2\ 2p^4$

**2p** ☐ ☐ ☐

### 💡 Formal Analogy: Bus Seats

Think of electron filling like passengers boarding a bus with empty seats. According to Hund's rule, passengers prefer to sit in empty rows before sitting next to someone (each orbital gets one electron before pairing). When they do sit next to someone, they prefer to sit on opposite sides of the aisle (opposite spins). Just as passengers fill the bus in an orderly way, electrons fill orbitals according to specific rules that minimize energy and maximize stability.

# Electronic Configuration

## ⚙ What is Electronic Configuration?

### 🔖 Definition

**Electronic configuration** is the arrangement of electrons in the orbitals of an atom, showing how electrons are distributed among the available atomic orbitals.

**Notation Systems:**

- **s, p, d notation:** Shows the principal quantum number (n) and sub-shell type
- **Electron-in-boxes notation:** Visual representation of electrons in orbitals

### ☰ Order of Filling

Electrons fill orbitals in order of increasing energy, following the (n+l) rule:

$$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p \rightarrow 6s \rightarrow 4f \rightarrow 5d \rightarrow 6p$$
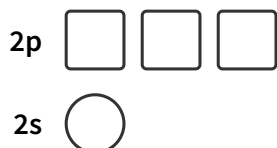
**Key Points:**

- Lower (n+l) values fill first
- If (n+l) values are equal, lower n fills first
- Each orbital can hold a maximum of 2 electrons with opposite spins
- Hund's rule applies when filling degenerate orbitals (same energy)

### ⚗ Examples of Electronic Configuration

#### ✦ Carbon (C, 6 electrons)

**s, p notation:** $1s^2\, 2s^2\, 2p^2$

2p ☐ ☐ ☐

2s ◯

1s ◯

### ✦ Sodium (Na, 11 electrons)

**s, p notation:** $1s^2\, 2s^2\, 2p^6\, 3s^1$
**Noble gas shorthand:** $[Ne]\, 3s^1$

### ✦ Chlorine (Cl, 17 electrons)

**s, p notation:** $1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^5$
**Noble gas shorthand:** $[Ne]\, 3s^2\, 3p^5$

## ⚡ Electronic Configuration of Ions

**Cations (positive ions):**

- Formed by losing electrons
- Electrons are removed from the highest energy level first
- Example: $Na^+$ has configuration $[Ne]$ (lost one 3s electron)

**Anions (negative ions):**

- Formed by gaining electrons
- Electrons are added to the lowest available energy orbital
- Example: $Cl^-$ has configuration $[Ar]$ (gained one 3p electron)

| Species | Electrons | Configuration |
|---------|-----------|---------------|
| Na | 11 | $[Ne]\, 3s^1$ |
| $Na^+$ | 10 | $[Ne]$ |
| Cl | 17 | $[Ne]\, 3s^2\, 3p^5$ |
| $Cl^-$ | 18 | $[Ar]$ |

## 💡 Formal Analogy: Apartment Building

Think of electronic configuration like residents in an apartment building. The building has different floors (energy levels) with different types of apartments (orbitals). Residents (electrons) fill the lowest available apartments first, following specific rules. When someone moves out (cation), they leave from the highest floor. When someone

new moves in (anion), they take the lowest available apartment. The arrangement of residents determines the building's character, just as electronic configuration determines an element's chemical properties.

# Electronic Configuration and Chemical Properties

## 🔗 Connection Between Configuration and Properties

### 🔖 Key Principle

The **electronic configuration** of an element determines its chemical properties by influencing:
- Valence electrons (outermost electrons)
- Ionisation energy
- Electronegativity
- Reactivity patterns
- Types of bonds formed

Elements with similar electronic configurations in their outermost shells exhibit similar chemical properties. This is the basis for the periodic arrangement of elements.

## 🔺 Groups and Electronic Configuration

Elements in the same group have the same number of valence electrons, leading to similar chemical properties:

### 1️⃣ Group 1 (Alkali Metals)

**Configuration:** $ns^1$
**Properties:** Highly reactive, lose 1 electron to form +1 ions, soft metals, low melting points

### 2️⃣ Group 2 (Alkaline Earth)

**Configuration:** $ns^2$
**Properties:** Reactive, lose 2 electrons to form +2 ions, harder metals, higher melting points

### 7️⃣ Group 7 (Halogens)

**Configuration:** $ns^2\,np^5$
**Properties:** Highly reactive non-metals, gain 1 electron to form -1 ions, diatomic molecules

### 8️⃣ Group 8 (Noble Gases)

**Configuration:** $ns^2\,np^6$ (except He: $1s^2$)
**Properties:** Inert, full valence shells, very low reactivity, monatomic gases

## ⚗ Examples of Chemical Behavior

### ✨ Sodium (Na) vs. Magnesium (Mg)

**Na:** [Ne] $3s^1$

**Mg:** [Ne] $3s^2$

Both are in Period 3, but Na has 1 valence electron while Mg has 2. This difference explains why:

- Na forms $Na^+$ ions (loses 1 electron)
- Mg forms $Mg^{2+}$ ions (loses 2 electrons)
- Na is more reactive than Mg
- Na has lower ionisation energy than Mg

## ✦ Fluorine (F) vs. Chlorine (Cl)

**F:** [He] $2s^2\ 2p^5$

**Cl:** [Ne] $3s^2\ 3p^5$

Both are halogens with 7 valence electrons, but F is in Period 2 while Cl is in Period 3. This difference explains why:

- F is more reactive than Cl
- F has higher electronegativity than Cl
- F has higher ionisation energy than Cl
- Both gain 1 electron to form -1 ions

## ⟁ Periodic Trends

**Across a Period (left to right):**

- Valence electrons increase from 1 to 8
- Ionisation energy generally increases
- Electronegativity generally increases
- Metallic character decreases

**Down a Group:**

- Valence electrons remain the same
- Ionisation energy generally decreases
- Electronegativity generally decreases
- Metallic character increases

## 💡 Formal Analogy: Social Behavior

Think of electronic configuration like a person's social behavior. Just as a person's personality and interactions are influenced by their inner characteristics and desires, an element's chemical behavior is determined by its electronic configuration. People with similar personalities (same group) tend to behave in similar ways, just as elements with similar valence electron configurations exhibit similar

chemical properties. The desire to achieve stability (full valence shell) drives chemical reactions, just as the pursuit of happiness or fulfillment drives human behavior.

# Periodic Table Blocks

## ▦ What are Periodic Table Blocks?

The Periodic Table is divided into **blocks** based on the sub-shell (s, p, d, f) that the valence electrons of elements occupy.

**Key Points:**

- Each block corresponds to the type of orbital being filled with electrons
- Elements in the same block have similar properties
- The block structure helps predict electronic configurations

## ▦ The Four Blocks

Periodic table of elements



☐ s-block   ☐ p-block   ☐ d-block   ☐ f-block

## ◈ Electron Capacity of Sub-shells

| Sub-shell | Number of Orbitals | Electrons per Orbital | Total Electrons | Elements in Block |
|-----------|--------------------|-----------------------|-----------------|-------------------|
| s | 1 | 2 | 2 | 2 per period |
| p | 3 | 2 | 6 | 6 per period (from period 2) |
| d | 5 | 2 | 10 | 10 per period (from period 4) |
| f | 7 | 2 | 14 | 14 per period (from period 6) |

## ♠ Characteristics of Each Block

### ○ s-Block Elements

- Groups 1 and 2 (alkali metals and alkaline earth metals)
- Valence electrons in ns orbital
- Highly reactive metals
- Form +1 or +2 ions

### △ p-Block Elements

- Groups 13 to 18
- Valence electrons in $ns^2 np^{1-6}$ configuration
- Include metals, metalloids, and non-metals
- Wide range of chemical properties

### ※ d-Block Elements

- Transition metals (Groups 3 to 12)
- Valence electrons in $(n-1)d$ orbitals
- Form colored compounds
- Multiple oxidation states

### 💡 Formal Analogy: City Districts

Think of the Periodic Table blocks like different districts in a city. Each district (block) has its own character and types of residents (elements). The s-block is like the financial district with its distinctive high-rise buildings (alkali and alkaline earth metals). The p-block is like the diverse downtown area with a mix of businesses and residences (metals, metalloids, and non-metals). The d-block is like the industrial district with specialized factories (transition metals). Just as city districts are organized by function and character, the Periodic Table is organized by electronic structure and chemical properties.

# Periodic Properties
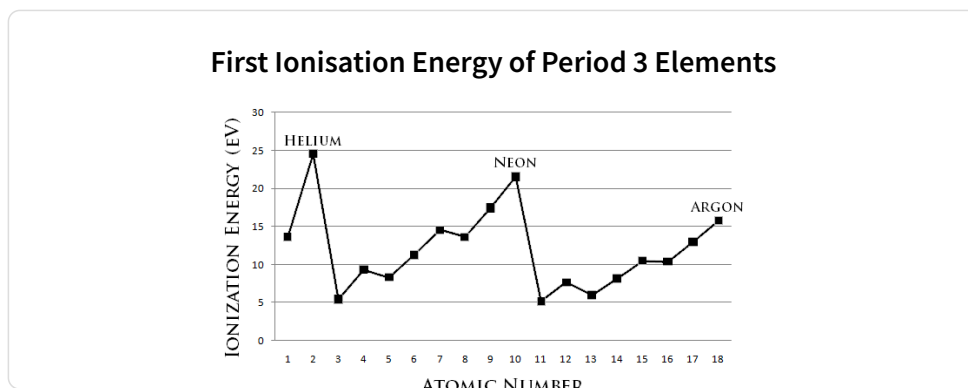
## 〰️ What are Periodic Properties?

### 🔖 Definition

**Periodic properties** are physical or chemical properties of elements that show a regular pattern or trend when moving across periods or down groups in the Periodic Table.

**Key Points:**

- Properties repeat periodically with increasing atomic number
- Trends can be visualized using graphs
- Help predict properties of unknown elements
- Reveal underlying electronic structure

## 〰️ Graphical Representation

Periodic properties are often represented graphically to visualize trends:

**First Ionisation Energy of Period 3 Elements**



**Logarithmic Scale:** For properties with large variations (like ionisation energies), a logarithmic scale is often used to better visualize the data.

## 〰️ Common Periodic Properties

### ⚡ Ionisation Energy

Energy required to remove an electron from an atom. Generally increases across a period and decreases down a group.

### ⚞ Atomic Radius

Distance from the nucleus to the outermost electron. Generally decreases across a period and increases down a group.

### 💧 Electronegativity

Tendency of an atom to attract electrons in a chemical bond. Generally increases across a period and decreases down a group.

### 🌡 Melting and Boiling Points

Temperatures at which elements change state. Varies based on bonding type and structure.

### 🧠 Understanding Periodic Trends

**Across a Period (left to right):**

- Nuclear charge increases
- Atomic radius decreases
- Ionisation energy generally increases
- Electronegativity generally increases

**Down a Group:**

- Nuclear charge increases
- Atomic radius increases
- Ionisation energy generally decreases
- Electronegativity generally decreases

### 💡 Formal Analogy: Seasonal Weather Patterns

Think of periodic properties like seasonal weather patterns. Just as temperature, rainfall, and daylight hours change predictably throughout the year, chemical properties change predictably across the Periodic Table. We can graph these patterns to visualize trends and make predictions about future conditions. The underlying causes (Earth's tilt and orbit for weather; electronic structure for chemical properties) explain why these patterns occur and repeat periodically.

# Trends in Melting and Boiling Points

## 🌡 Melting and Boiling Points

Melting and boiling points are physical properties that indicate the temperatures at which elements change state. These properties vary across periods due to differences in structure and bonding.

**Key Factors:**

- Type of bonding between atoms/molecules
- Structure of the element (atomic, molecular, or giant)
- Strength of intermolecular forces
- Atomic/molecular size

## 〽 Period 2 Trends

| Element | Structure | Bonding | MP (°C) | BP (°C) |
|---------|-----------|---------|---------|---------|
| Li | Metallic | Metallic | 180 | 1342 |
| Be | Metallic | Metallic | 1287 | 2470 |
| B | Giant covalent | Covalent | 2075 | 4000 |
| C | Giant covalent | Covalent | 3550 | 4827 |
| N | Molecular | Covalent | -210 | -196 |
| O | Molecular | Covalent | -218 | -183 |
| F | Molecular | Covalent | -220 | -188 |
| Ne | Molecular | None | -249 | -246 |

### 〽 Key Observations

- Metals (Li, Be) have high melting/boiling points due to strong metallic bonding
- Non-metals with giant structures (B, C) have very high melting/boiling points
- Molecular elements ($N_2$, $O_2$, $F_2$, Ne) have very low melting/boiling points
- Large jump between C and N due to change from giant covalent to molecular structure

# 〽 Period 3 Trends

## Melting Points of Period 3 Elements
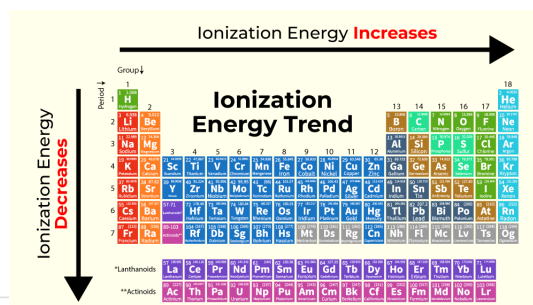


## 〽 Key Observations

- Metals (Na, Mg, Al) have relatively high melting points due to metallic bonding
- Si has a very high melting point due to its giant covalent structure
- Non-metals with molecular structures ($P_4$, $S_8$, $Cl_2$, Ar) have low melting points
- Large drop between Si and P due to change from giant covalent to molecular structure

## 💡 Explaining the Trends

**Structure and Bonding Effects:**

- **Metals:** High melting/boiling points due to strong metallic bonds
- **Giant covalent structures:** Very high melting/boiling points due to strong covalent bonds throughout the structure
- **Molecular elements:** Low melting/boiling points due to weak intermolecular forces between molecules
- **Noble gases:** Lowest melting/boiling points due to very weak van der Waals forces

**Across a Period:**

- Melting/boiling points generally increase for metals
- Sharp decrease when changing from giant covalent to molecular structures
- Gradual decrease across molecular elements

## 💡 Formal Analogy: Building Materials

Think of melting and boiling points like the strength of different building materials. Giant covalent structures are like steel frameworks - very strong and require high temperatures to break down. Metallic structures are like reinforced concrete - strong but not as strong as steel. Molecular structures are like individual bricks held together with weak mortar - easy to separate with little heat. Just as different materials require different amounts of energy to break apart, different elements require different amounts of energy to change state.

# Trends in Ionisation Energies
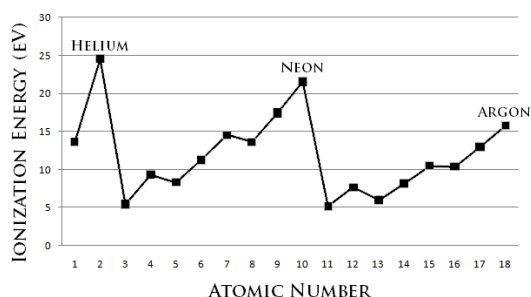
## ⚡ Ionisation Energy Trends

Ionisation energy shows distinct patterns across periods due to changes in electronic structure. Understanding these trends helps explain chemical reactivity.

**General Trend Across a Period:**

- Generally increases from left to right
- Due to increasing nuclear charge
- With similar shielding across the period
- Resulting in stronger attraction to valence electrons

## ∿ Period 2 Ionisation Energies

### First Ionisation Energies of Period 2 Elements



### ↘ Specific Trends

- **Dip at B:** Electron enters 2p sub-shell (higher energy than 2s)
- **Dip at O:** Electron pairs in 2p sub-shell (electron-electron repulsion)
- **Peak at Ne:** Full valence shell (stable configuration)

## ∿ Period 3 Ionisation Energies

| Element | Electronic Configuration | First IE (kJ mol$^{-1}$) | Trend Explanation |
|---------|--------------------------|--------------------------|-------------------|
| Na | [Ne] 3s$^1$ | 496 | Low - one valence electron |
| Mg | [Ne] 3s$^2$ | 738 | Higher - full 3s sub-shell |

| Element | Electronic Configuration | First IE (kJ mol$^{-1}$) | Trend Explanation |
| --- | --- | --- | --- |
| Al | [Ne] 3s$^2$ 3p$^1$ | 578 | Dip - electron enters 3p sub-shell |
| Si | [Ne] 3s$^2$ 3p$^2$ | 789 | Higher - increasing nuclear charge |
| P | [Ne] 3s$^2$ 3p$^3$ | 1012 | Higher - half-filled 3p sub-shell |
| S | [Ne] 3s$^2$ 3p$^4$ | 1000 | Dip - electron pairing in 3p |
| Cl | [Ne] 3s$^2$ 3p$^5$ | 1251 | Higher - nearly full 3p sub-shell |
| Ar | [Ne] 3s$^2$ 3p$^6$ | 1521 | Highest - full valence shell |

## 🧠 Explaining the Trends

**General Increase:**

- Nuclear charge increases across the period
- Electrons added to same principal energy level
- Similar shielding effect across the period
- Stronger attraction to valence electrons

**Specific Dips:**

- **Group 13 elements (B, Al):** Electron enters p sub-shell (higher energy)
- **Group 16 elements (O, S):** Electron pairing in p sub-shell (repulsion)

**Peaks:**

- **Group 15 elements (N, P):** Half-filled p sub-shell (stable)
- **Group 18 elements (Ne, Ar):** Full valence shell (very stable)

## 💡 Formal Analogy: Stacking Books

Think of ionisation energy like the effort needed to remove books from a stack. As you add more books to the stack (increase nuclear charge), it becomes harder to remove books from the top (higher ionisation energy). However, when you start a new layer (enter a new sub-shell), it's easier to remove the first book in that layer (dip in ionisation energy). When you have to place books side by side in the same layer

(electron pairing), it becomes slightly easier to remove one (another dip). A completely filled layer (full valence shell) is very stable and hard to disrupt (highest ionisation energy).

# Trends Down Groups

## ↘ Ionisation Energy Down Groups

First ionisation energy **decreases down a group** in the Periodic Table. This trend is consistent across all groups and can be explained by several factors.

**Key Factors:**

- Increasing atomic radius
- Increasing electron shielding
- Decreasing effective nuclear charge
- Greater distance between nucleus and valence electrons

## ⚗ Atomic Structure Down Groups
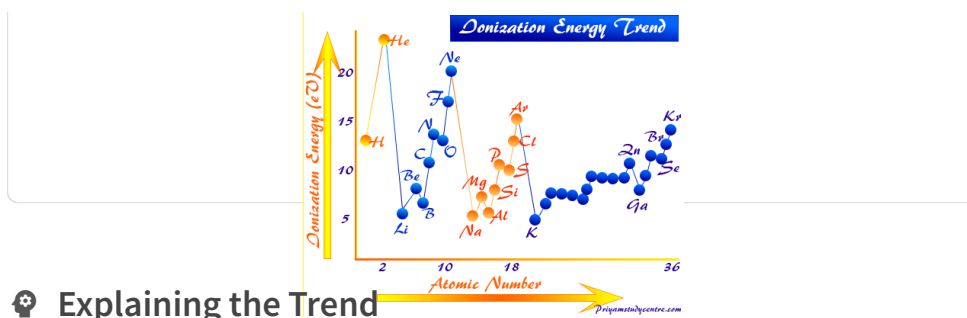


Li          Na          K

As we move down a group:

- More electron shells are added
- Atomic radius increases
- Valence electrons are farther from the nucleus
- Inner electrons shield the valence electrons

## 〽 Ionisation Energy Data

| Group | Period 2 | Period 3 | Period 4 | Period 5 |
|-------|----------|----------|----------|----------|
| Group 1 | Li: 520 kJ/mol | Na: 496 kJ/mol | K: 419 kJ/mol | Rb: 403 kJ/mol |
| Group 2 | Be: 900 kJ/mol | Mg: 738 kJ/mol | Ca: 590 kJ/mol | Sr: 550 kJ/mol |
| Group 17 | F: 1681 kJ/mol | Cl: 1251 kJ/mol | Br: 1140 kJ/mol | I: 1008 kJ/mol |

**First Ionisation Energies of Group 1 Elements**

## Explaining the Trend

### ⤵ Why Ionisation Energy Decreases Down Groups

- **Atomic Radius:** Increases down the group, so valence electrons are farther from the nucleus
- **Electron Shielding:** Inner electrons shield the valence electrons from the nuclear charge
- **Effective Nuclear Charge:** Decreases for valence electrons due to increased shielding
- **Distance:** Greater distance between nucleus and valence electrons reduces attraction

**Result:** Less energy is required to remove an electron from an atom lower in a group because the valence electrons are held less tightly by the nucleus.

### 💡 Formal Analogy: Magnet and Paper Clips

Think of ionisation energy down a group like the force needed to pull paper clips away from a magnet at different distances. When the paper clips are close to the magnet (top of the group), the magnetic force is strong, making it difficult to pull them away (high ionisation energy). As you move the paper clips farther from the magnet (down the group), the magnetic force weakens, making it easier to pull them away (lower ionisation energy). Even if you add more paper clips between the magnet and the one you're trying to remove (electron shielding), it becomes even easier to pull away the outermost paper clip.