1.1 Atomic Structure

YOUR NOTES

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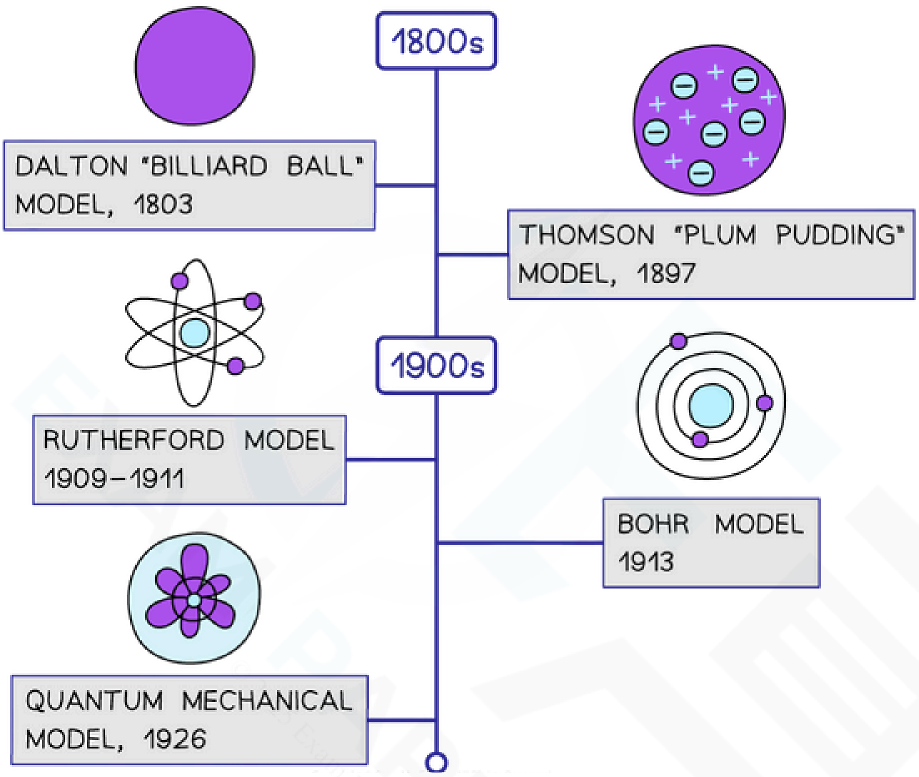
1.1.1 CHANGING MODELS OF THE ATOM

Changing Models of the Atom

- The atomic model developed by scientists has changed over time as experimental evidence has improved our understanding of the structure of atoms
- In 1803 John Dalton presented his atomic theory based on three key ideas:
    - Matter is made of atoms which are tiny particles that cannot be created, destroyed, or **divided**
    - Atoms of the same element are **identical,** and atoms of different elements are **different**
    - Different atoms combine together to form new substances
- At the time, the theory was correct but as science developed some parts of Dalton's theory were disproved
- This is a fundamental feature of science: new experimental evidence may lead to a scientific model being changed or replaced

1.1 Atomic Structure



The evolution of models of atomic structure

Discovery of the electron

- In 1897 physicist J.J. Thomson discovered the **electron**
- Using a **cathode-ray tube** he conducted an experiment which identified the electron as a **negatively charged subatomic** particle, hence proving that atoms are **divisible**

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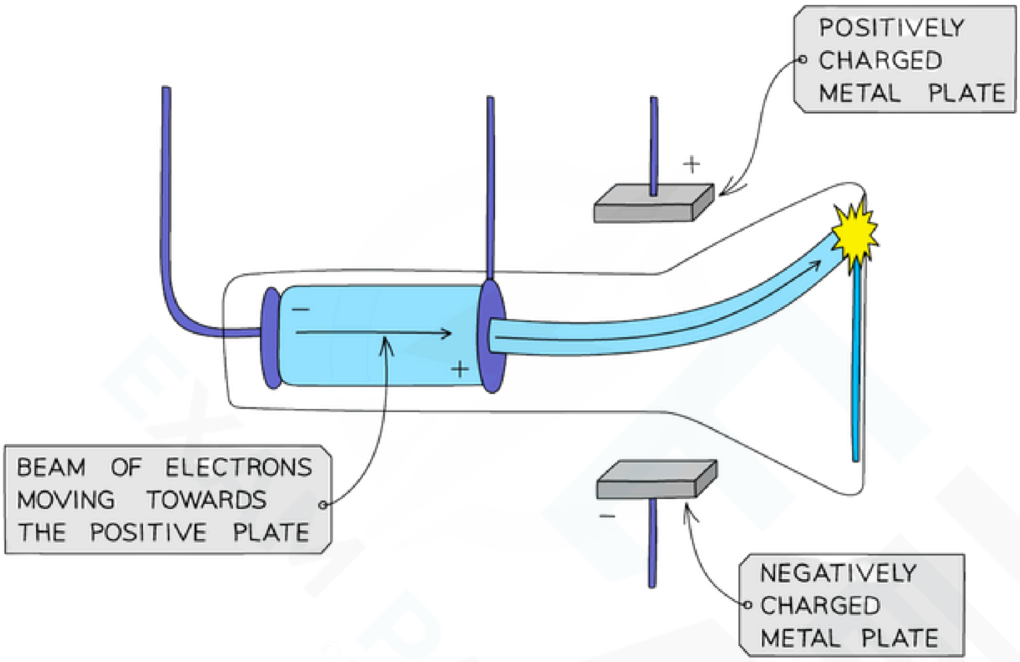

Diagram showing an electron beam deflecting towards the positive plate, proving electrons are negatively charged

- Based on his investigations Thomson proposed a model of the atom known as the **plum pudding** model which depicted negative electrons spread throughout soft globules of positively charged material

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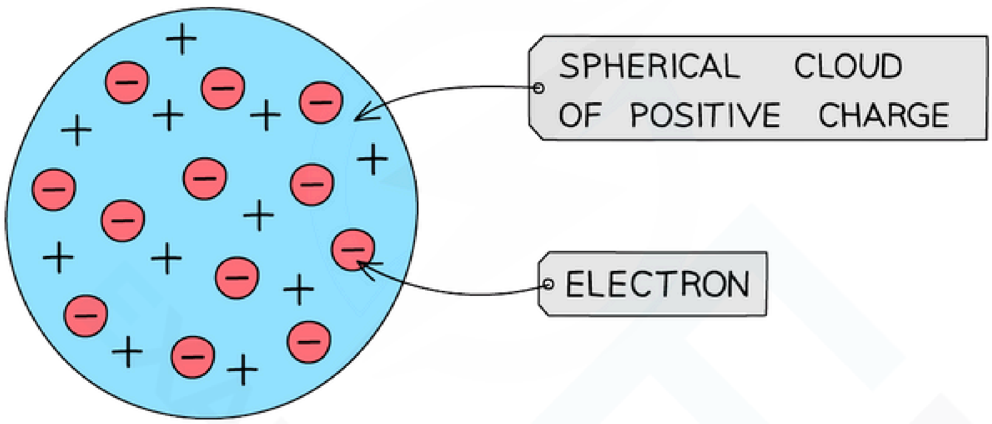

Diagram showing the plum pudding model of the atom

The discovery of the nucleus

- In 1909 Ernest Rutherford presented his model of the atom based on the famous gold foil experiment
- Rutherford shot a beam of positively charged particles at a thin sheet of gold foil and based on the plum pudding model, expected the particles to pass through the foil because the positive charge of the nucleus was thought to be evenly spread out
- Some particles were **scattered**, however, and a few were deflected directly back, which led him to postulate that most of an atom's mass is concentrated in a region of space at the centre of the atom called the **nucleus**
- The results of Rutherford's scattering experiments did not support the idea that atoms were as described in the plum pudding model, so the model had to be changed
- In Rutherford's model the atom consists mainly of **empty space** with the **nucleus** at the centre and the **electrons orbiting** in paths around the nucleus
- This model was known as the **nuclear model** of the atom

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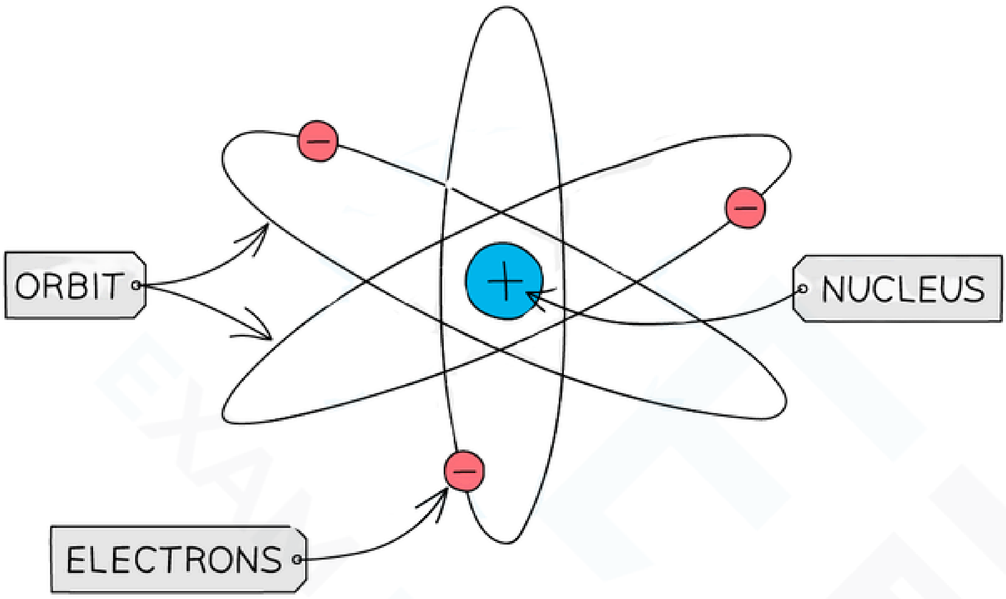

Diagram showing Ernest Rutherford's nuclear model of the atom

- In the plum pudding model, atoms were described as being made from electrons embedded within a positive sphere, whereas in the nuclear model the nucleus is a positive structure at the centre of the atom, with negative (and much smaller) electrons 'orbiting' around the outside of it

The Bohr model

- In 1913 Niels Bohr further developed the nuclear model by proposing that electrons orbit the nucleus in **fixed** shells or orbitals located at **set distances** from the nucleus
- Each orbital has a **different energy** associated with it, with the higher energy orbitals being located **further** away from the nucleus
- This model solved the question of why the atom does not collapse inwards due to the attraction between the positive nucleus and negative electrons circling the nucleus
- Bohr's theory and calculations agreed with experimental results
- Further investigation and experimentation revealed that the nucleus could be divided into smaller particles, each one having the same mass and charge
    - This work led to the discovery of the **proton**

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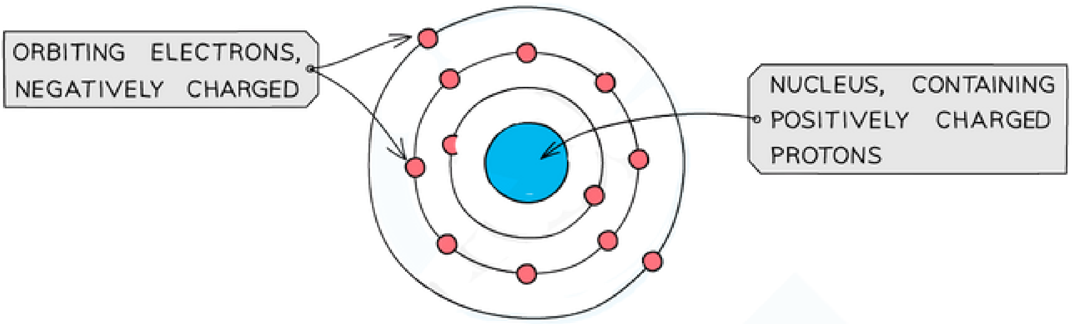

Diagram showing Niels Bohr's model of the atom

Discovery of the neutron

- In 1920 Rutherford put forward the idea of the existence of large, neutral particles within the nucleus
- His idea was based on the differences between the atomic mass and the atomic number of atoms
- In 1932 James Chadwick published a paper based on an experiment carried out by Frédéric and Irène Joliot-Curie which provided evidence for the existence of these neutral particles which were called **neutrons**

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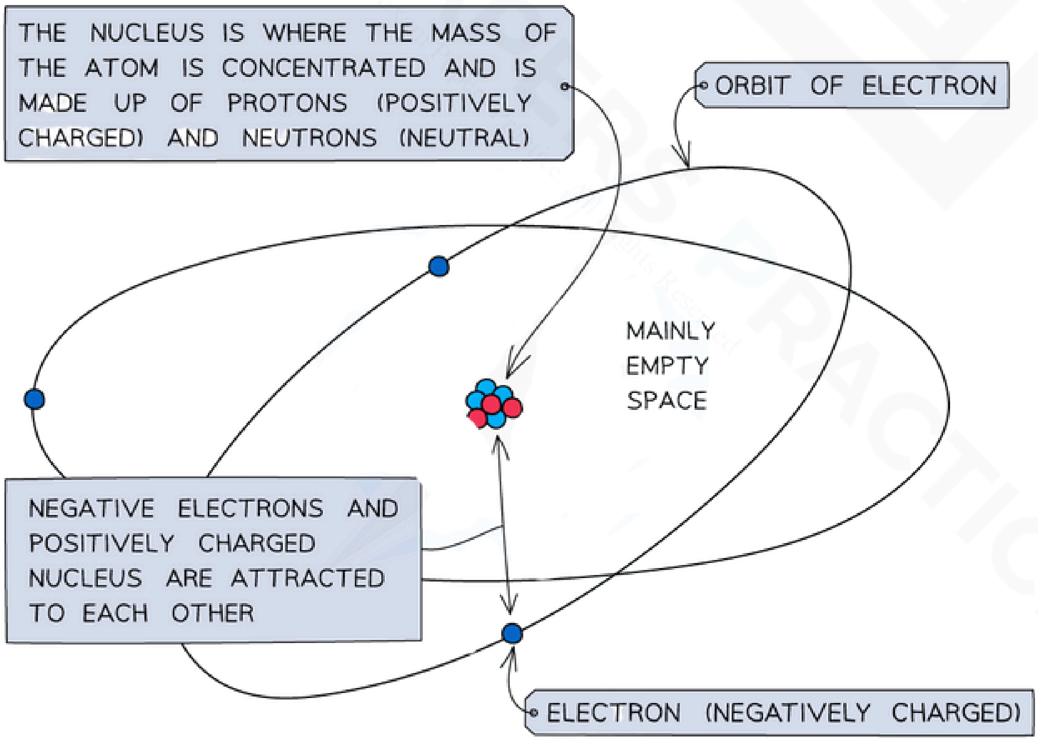
1.1.2 Atomic Structure

The Structure of an Atom

- Elements are made of tiny particles of matter called **atoms**
- Each atom is made of subatomic particles called **protons**, **neutrons** and **electrons**
- Their size is so tiny that we can't really compare their masses in conventional units such as kilograms or grams, so a unit called the relative atomic mass is used
- The mass of an atom is **concentrated** in the nucleus, because the nucleus contains the heaviest subatomic particles (the neutrons and protons)
- The nucleus is also **positively charged** due to the protons
- Electrons orbit the nucleus of the atom, contributing very little to its overall mass, but creating a **'cloud'** of negative charge
- The **electrostatic attraction** between the **positive nucleus** and **negatively charged electrons** orbiting around it is what holds an atom together



The mass of the atom is concentrated in the positively charged nucleus which is attracted to the negatively charged electrons orbiting around it

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The Sub-atomic Particles

- The **protons**, **neutrons** and **electrons** that an atom is made up of are called **subatomic particles**
- These subatomic particles are so small that it is not practical to measure their masses and charges using **conventional units** (such as grams or coulombs)
- Instead, their masses and charges are compared to each other, and so are called **'relative atomic masses'** and '**relative atomic charges**'
- These are not actual charges and masses, but rather charges and masses of particles relative to each other
    - Protons and neutrons have a very similar mass, so each is assigned a relative mass of **1**
    - Electrons are 1840 times smaller than a proton and neutron, and so their mass is often described as being negligible
- The relative mass and charge of the subatomic particles are:

The Mass & Charge of Subatomic Particles Table

PARTICLE	RELATIVE MASS	CHARGE
PROTON	1	+1
NEUTRON	1	0 (NEUTRAL)
ELECTRON	$\frac{1}{1840}$	−1

- Atoms are electrically **neutral**
- This is achieved by having the **same number** of electrons as protons
- The **negative charge** of an electron exactly cancels out the **positive charge** of a proton

Exam Tip

The mass of an electron can just be stated as 'negligible' or 'very small' in an exam. You do not need to learn the value.

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1.1.3 THE NUCLEUS

The Nucleus

- Atoms are extremely small with a radius of about 1 x 10-10 metres
- The central nucleus contains **protons** and **neutrons** only which are packed close together in a small region of space
- The radius of the nucleus is about 10 000 times smaller than that of the atom, so it is an **extremely small** region of space compared to the overall size of the atom
- This means that rather than being evenly spread out throughout the atom, virtually all of the atom's mass is concentrated inside the nucleus
- Electrons have a **much smaller mass** than protons and neutrons (1 proton has the same mass of around 1840 electrons) and move in the space outside the nucleus in **orbits**

Exam Tip

Most of the atom is actually empty space, with the mass being concentrated in the nucleus and the electrons orbiting in shells around it.

Mass Number & Atomic Number

- You need to know the following terms to describe the properties and characteristics of atoms

Atomic Structure Key Terms Table

Term	Definition
Atomic number	The number of protons in the nucleus of an atom
Mass number	The sum of the number of protons andd neutrons in the nucleus of an aatom

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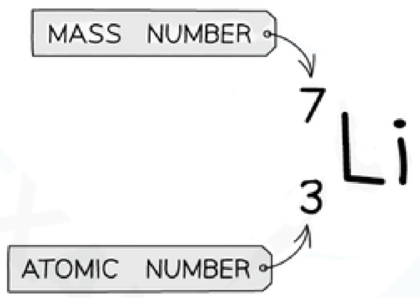
- The atomic number and mass number are represented by writing them next to the symbol of the element
- By convention the mass number is usually written as a superscript and the atomic number as a subscript



Conventions for showing atomic number and mass number

Exam Tip

The term nucleon number is an alternative to mass number and means the same thing. A nucleon is a collective name for protons and neutrons.

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1.1.4 ISOTOPES

Isotopes

- Isotopes are atoms of the **same element** that contain the same number of **protons** and electrons but a different number of **neutrons**
- The symbol for an isotope is the chemical symbol (or word) followed by a dash and then the mass number
- So, C-14 is the isotope of carbon which contains 6 protons and 6 electrons, but the 14 signifies that it has 8 neutrons (14 – 6 = 8)
    - It can also be written as 14C
- Isotopes display the **same chemical characteristics**
- This is because they have the same number of electrons in their outer shells, and this is what determines their chemistry
- The difference between isotopes is the neutrons which are neutral particles within the nucleus and add mass only

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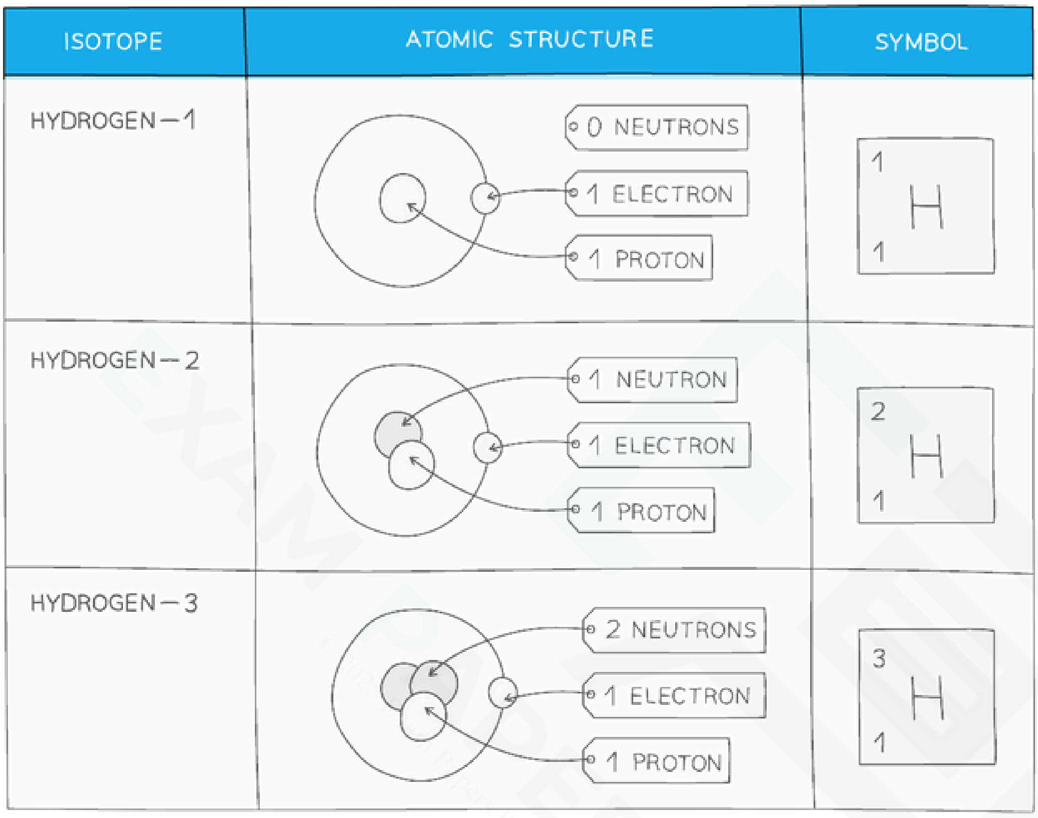

The atomic structure and symbols of the three isotopes of hydrogen

Deducing protons, neutrons & electrons

Finding the protons

- The atomic number of an atom and ion determines which element it is
- Therefore, all atoms and ions of the **same element** have the same number of protons (atomic number) in the nucleus
  - E.g. lithium has an atomic number of 3 (three protons) whereas beryllium has atomic number of 4 (4 protons)
- The number of protons equals the **atomic (proton) number**
- The number of protons of an **unknown** element can be calculated by using its mass number and number of neutrons:

Mass number = number of protons + number of neutrons

Number of protons = mass number – number of neutrons

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Finding the electrons

- An atom is **neutral** and therefore has the **same** number of **protons** and **electrons**

Finding the neutrons

- The **mass** and **atomic numbers** can be used to find the number of **neutrons** in **ions** and **atoms:**

Number of neutrons = mass number – number of protons

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Worked Example

Determine the number of protons, electrons and neutrons in an atom of element X with atomic number 29 and mass number 63

Answer:

- The number of protons of element **X** is the same as the atomic number

Number of protons = 29

- The **neutral atom** of element **X** therefore also has **29 electrons**
- The atomic number of an element **X** atom is 29 and its mass number is 63

Number of neutrons = mass number – number of protons

Number of neutrons = 63 – 29

Number of neutrons = 34

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Is mass number and relative atomic mass the same thing?

- On a GCSE periodic table you will see that lithium has a relative atomic mass of **7**
- Although it seems that this is the same as the mass number, they are not the same thing because the relative atomic mass is a **rounded** number
- Relative atomic mass takes into account the existence of isotopes when calculating the mass
- Relative atomic mass is an **average mass** of all the isotopes of that element
- For simplicity **relative atomic masses** are often shown to the nearest whole number

The relative atomic mass of lithium to two decimal places is 6.94 when rounded to the nearest whole number, the RAM is 7, which is the same as the mass number shown on this isotope of lithium

Exam Tip

For atoms to be isotopes of each other, they must both be from the same element, hence they must have the same atomic number. E.g., C-13 and C-14 are isotopes whereas C-13 and H-2 are not

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1.1.5 CALCULATING RELATIVE ATOMIC MASS

Higher Tier Only

Calculating Relative Atomic Mass

- The relative atomic mass of each element is calculated from the **mass number** and **relative abundances** of all the isotopes of a particular element
- The equation below is used where the top line of the equation can be extended to include the number of different isotopes of a particular element present
- So, if there were 3 isotopes present then the top line of the equation would read:

(% of isotope A x mass of isotope A) + (% of isotope B x mass of isotope B) + (% of isotope C x mass of isotope C)

RAM Calculation Formula

$$A_r = \frac{(\%\ \boldsymbol{of\ isotope\ a} \times \boldsymbol{mass\ of\ isotope\ a}) + (\%\ \boldsymbol{of\ isotope\ b} \times \boldsymbol{mass\ of\ isotope\ b})}{100}$$

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Worked Example

The table shows information about the Isotopes in a sample of rubidium with 72% 85Rb and 28% ^{87}Rb

Isotope	Number of Protons	Number of Neutrons	Percentage of Isotope in Sample
1	37	48	72
2	37	50	28

Use information from the table to calculate the relative atomic mass of this sample of Rubidium. Give your answer to one decimal place:

Answer

$$\frac{(72 \times 85) + (28 \times 87)}{100} = 85.6$$

Relative Atomic Mass = 85.6

Exam Tip

Isotopes are easy to recognize from their notation as they have the same symbol but different mass numbers. For example, the two stable isotopes of copper are 63Cu and 65Cu

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