

# Edexcel Chemistry A-level

## Topic 1 - Atomic Structure and Periodic Table

### Flashcards

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# What was stated in Dalton's atomic theory? (4)



What was stated in Dalton's atomic theory?

- Atoms are tiny particles made of elements
- Atoms cannot be divided
- All the atoms in a element are the same
- Atoms of one element are different to those of other elements



# What did Thompson discover about electrons? (3)



What did Thompson discover about electrons?

- They have a negative charge
- They can be deflected by electromagnetic fields
- They have very small mass



# Explain the current model of the atom.



## Explain the current model of the atom

- Protons and neutrons are found in the nucleus
- Electrons orbit the nucleus in shells
- The nucleus is tiny compared to the total volume of atom
- Most of atom's mass is in the nucleus
- Most of the atom is empty space between the nucleus and the electrons



What is the charge of a proton  
and an electron?





What is the charge of a proton and an electron?

Proton = +1

Electron = -1



Which particle has the same mass as proton?



Which particle has the same mass as proton?

Neutron



Which two particles make up most of an atom's mass?



Which two particles make up most of atom's mass?

Protons and neutrons



# What does the atomic number show about an element?



What does the atomic number tell about an element?

Atomic number = number of protons in  
an atom



# How is mass number calculated?





How is mass number calculated?

Mass number = number of protons +  
number of neutrons



# How to calculate the number of neutrons?



How to calculate the number of neutrons?

Number of neutrons = mass number -  
atomic number



Define the term, isotope.



Define the term, isotope.

Atoms of the same element with different number of neutrons and therefore different mass number



Why do different isotopes of the same element react in the same way? (2)



Why do different isotopes of the same element react in the same way?

- Neutrons have no impact on the chemical reactivity
- Reactions involve electrons, isotopes have the same number of electrons in the same arrangement



Define relative atomic mass.





Define relative atomic mass.

The weighted mean mass of an atom of an element compared with one twelfth of the mass of an atom of carbon -12



# Define relative isotopic mass.



Define relative isotopic mass

The mass of an atom of an isotope compared with one twelfth of the mass of an atom of carbon-12



The relative isotopic mass is  
same as which number?



The relative isotopic mass is same as which number?

Mass number



What two assumptions are made when calculating mass number?



What two assumptions are made when calculating mass number?

1. Contribution of the electron is neglected
2. Mass of both proton and neutron is taken as  $1.0 \text{ u}$



# How to calculate the relative molecular mass and relative formula mass?





How to calculate the relative molecular mass and relative formula mass?

Both can be calculated by adding the relative atomic masses of each of the atom making up the molecule or the formula



# What are the uses of mass spectrometry? (3)

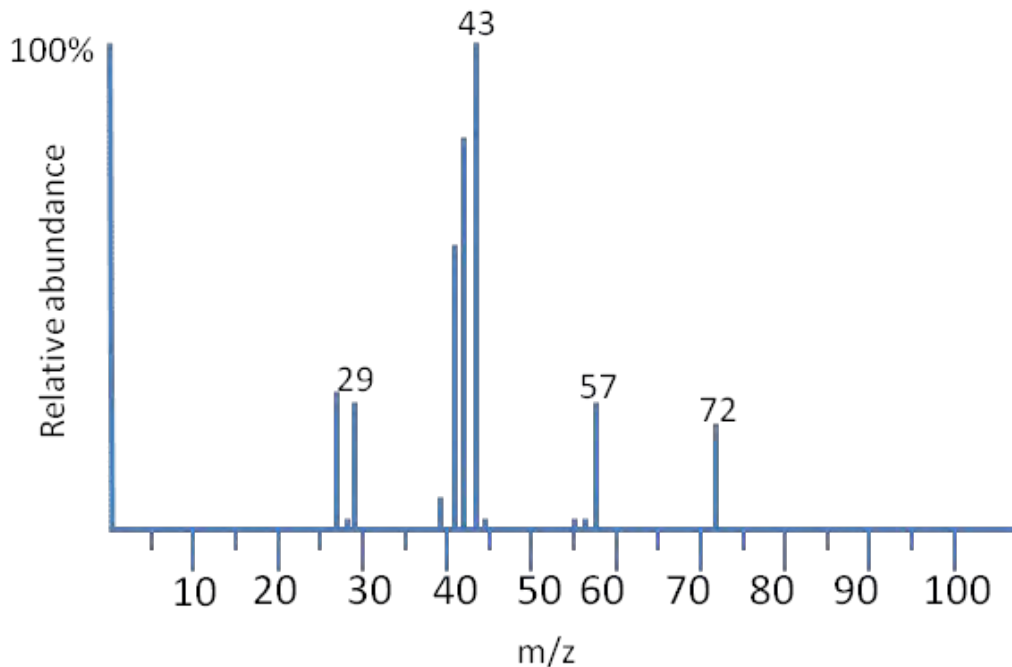


# What are the uses of mass spectrometry?

- Identify unknown compounds
- Find relative abundance of each isotope of an element
- Determine structural information



# What is the $m/z$ value of the $M^+$ ion



What is the  $m/z$  value of the  $M^+$  ion

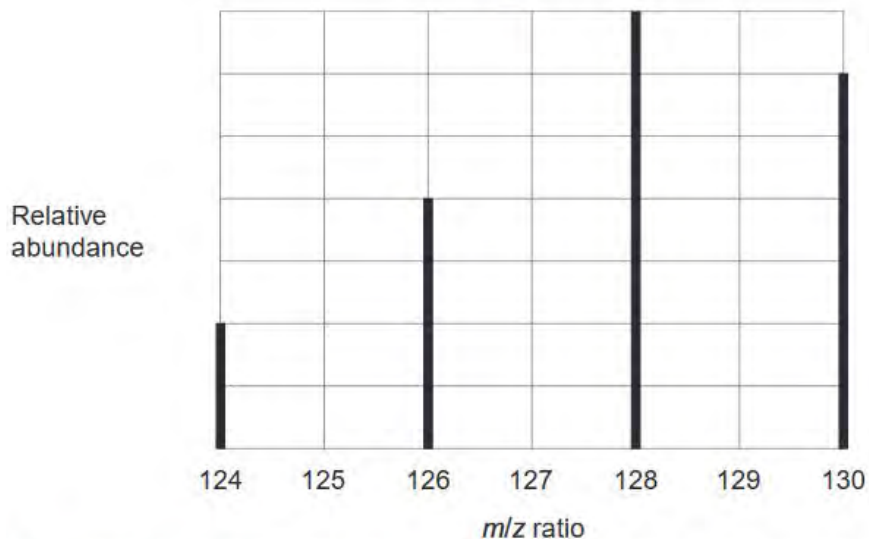
The  $m/z$  value of the  $M^+$  ion is the value of the last peak - **72**



# Complete this question...

- ) The mass spectrum of a sample of tellurium is shown in **Figure 1**.

**Figure 1**



- ) (i) Use **Figure 1** to calculate the relative atomic mass of this sample of tellurium.

Give your answer to one decimal place.

[3 marks]

$\frac{(124 \times 2) + (126 \times 4) + (128 \times 7) + (130 \times 6)}{19} \text{ or } \frac{2428}{19}$ <p><u>127.8</u></p>	<p>1</p> <p>1</p> <p>1</p>	<p>M1 for top line</p> <p>M2 for correct denominator</p> <p>127.8 with no working shown scores 3 marks</p>
<p>Or</p> $\frac{(124 \times 10.5) + (126 \times 21.1) + (128 \times 36.8) + (130 \times 31.6)}{100}$ <p><u>127.8</u></p>	<p>Or</p> <p>1</p> <p>1</p> <p>1</p>	<p>Mark for 100 dependent on top line correct</p>



# What does the principal quantum number indicate?





What does the principal quantum number indicate?

The shell occupied by the electrons



# What is a shell?



What is a shell?

A group of orbitals with the same principal quantum number



# How many electrons can the 1st shell hold?



How many electrons can the 1st shell hold?

2



# How many electrons can the 2nd shell hold?



How many electrons can the 2nd shell hold?

8



How many electrons can the  
3rd shell hold?





How many electrons can the 3rd shell hold?

18



How many electrons can the  
4th shell hold?



How many electrons can the 4th shell hold?

32



# What is an orbital?



What is an orbital?

A region around the nucleus that can hold up to two electrons with opposite spins



# How many electrons can an orbital hold?



How many electrons can an orbital hold?

2



# What are the 4 types of orbitals?





What are the 4 types of orbitals?

- s orbital
- p orbital
- d orbital
- f orbital



# What is the shape of a s-orbital?



What is the shape of a s-orbital?

Spherical



# What is the shape of a p-orbital?



What is the shape of a p-orbital?

Dumb-bell shape



How many orbitals are found  
in a  $S$  subshell?



How many orbitals are found in a S subshell?

1



How many electrons can be held in a S subshell?





How many electrons can be held in a S subshell?

2



How many orbitals does P  
subshell have?



How many orbitals does P subshell have?

3



How many electrons can be held in a P subshell?



How many electrons can be held in a P subshell?

6



How many orbitals are present  
in a D subshell?



How many orbitals are present in a D subshell?

5



How many electrons can be held in a D subshell?





How many electrons can be held in a d-sub shell?

10



How many orbitals are found  
in a F subshell?



How many orbitals are found in a F subshell?

7



# How many electrons can fill F subshell?



# How many electrons can fill F subshell?

14



When using 'electrons in box' representation, what shape is used to represent the electrons?

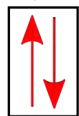


When using 'electrons in box' representation, what shape is used to represent the electrons?

# Arrows

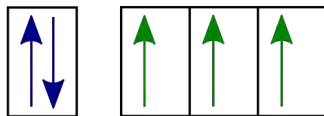
Image courtesy of wikibooks

The first two electrons pair up in the 1s orbital



1s

The last three electrons singly occupy the three 2p orbitals. They all have the same spin!



2s

2p

The second two electrons pair up in the 2s orbital

What letter used to represent  
shell number?





What letter is used to represent the shell number?

$n$



From which shell onwards is S orbital present?



From which shell onwards is S orbital present?

$$n = 1$$



From which shell onwards is  
P-orbital present?



From which shell onwards is P orbital present?

$$n = 2$$



From which shell onwards is  
D-orbital present?



From which shell onwards is D orbital present?

$$n = 3$$



From which shell onwards is  
F-orbital present?





From which shell onwards is F orbital present?

$$n = 4$$



What are the rules by which electrons are arranged in the shell? (5)



What are the rules by which electrons are arranged in a shell?

- Electrons are added one at a time
- Lowest available energy level is filled first
- Each energy level must be filled before the next one can fill
- Each orbital is filled singly before pairing
- 4s is filled before 3d



# Why does 4s orbital fill before 3d orbital?



Why does 4s orbital fill before 3d orbital?

4s orbital has a lower energy than 3d  
before it is filled



# What is the electron configuration of krypton?



What is the electron configuration of krypton?



How can the electron configuration be written in short?





How can the electron configuration be written in short?

The noble gas before the element is used to abbreviate

E.g  $\text{Li} \rightarrow 1s^2 2s^1$  ;  $\text{Li} \rightarrow [\text{He}] 2s^1$



# How are the elements arranged in a periodic table?



How is the group number related to the number of electrons?

Group number = number of electrons in  
the outer shell



# What is a period on a periodic table?



What is a period on a periodic table?

The horizontal rows



# What is a group on a periodic table?



What is a group on a periodic table?

The vertical columns



How is the group number  
related to the number of  
electrons?





How are the elements arranged in a periodic table?

They are arranged in the order of increasing atomic numbers



Does the group number  
indicate horizontal or vertical  
columns in the periodic table?



Does the group number indicate horizontal or vertical columns in the periodic table?

Vertical column



# What is meant by periodicity?



# What is meant by periodicity?

The repeating trends in chemical and physical properties



# What change happens across each period?



What change happens across each period?

Elements change from metals to  
non-metals



# Define first ionisation energy.





Define first ionisation energy.

The energy required to remove a mole of electrons from a mole of gaseous atoms to form one mole of gaseous  $1+$  ions under standard conditions



Write an equation for the first ionisation energy of magnesium.



Write an equation for the first ionisation energy of magnesium



# What are the factors that affect ionisation energy?

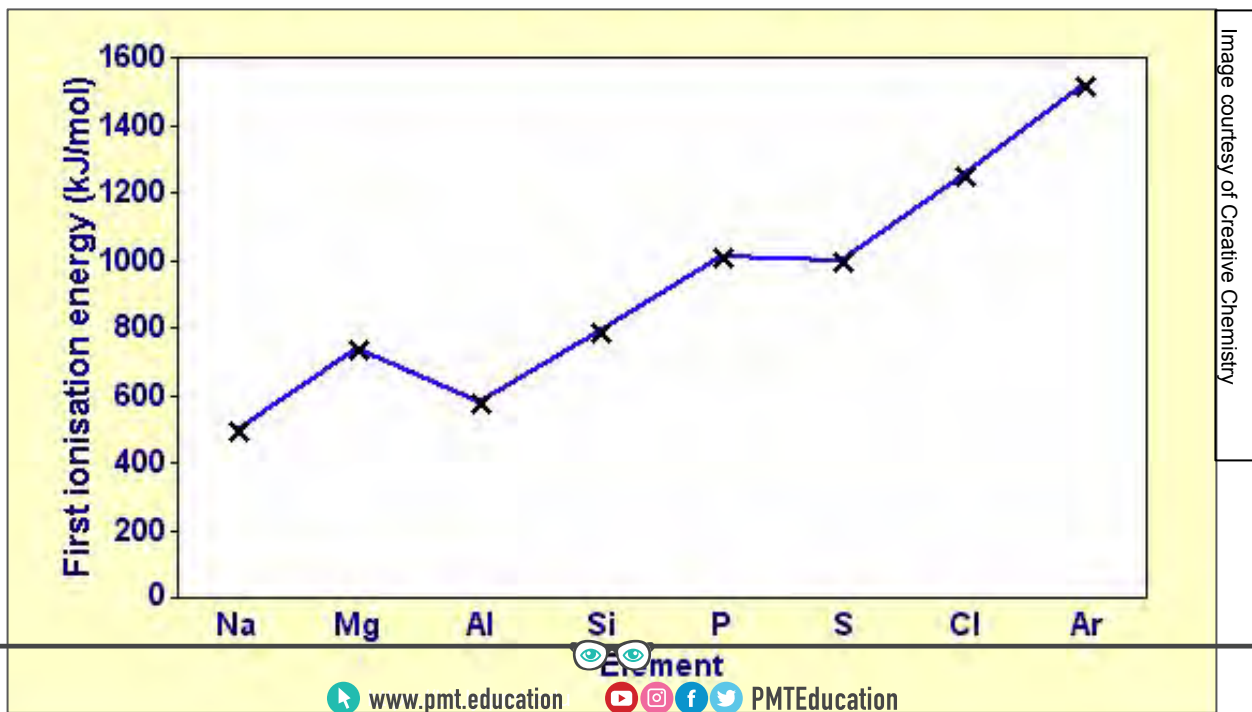


What are the factors that affect ionisation energy?

- Atomic radius
- Nuclear charge
- Electron shielding or screening



# Explain the trend on this graph.



## Explain the trend on this graph.

- *First Ionisation energy increases across period 3* because of:
  - Increased nuclear charge
  - Decreased atomic radius
  - Same electron shielding
- This means more energy is needed to remove the first electron.
- *Dips at Al* because: outer electron is in a 3p orbital, higher energy than 3s orbital → less energy needed to remove electron
- *Dips at S* because one 3p orbital contains two electrons → repulsion between paired electrons → less energy needed to remove one



Why does first ionisation energy decrease between group 2 to 3?





Why does first ionisation energy decrease between group 2 to 3 ?

- Decreases between 2 to 3 because in group 3 the outermost electrons are in p orbitals.
- Whereas in group 2 they are in s orbital, so the electrons are easier to be removed.



Why does first ionisation energy decrease between group 5 to 6?



Why does first ionisation energy decrease between group 5 to 6?

- The decrease between 5 to 6 is due to the group 5 electrons in p orbital which are single electrons.
- In group 6 the outermost electrons are spin paired, with some repulsion.
- Therefore the electrons are slightly easier to remove.



Does first ionisation increase or decrease between the end of one period and the start of next? Why?



Does first ionisation increase or decrease between the end of one period and the start of next? Why?

- Decrease
- There is increase in atomic radius
- Increase in electron shielding



Does first ionisation increase  
or decrease down a group?  
Why?



Does first ionisation increase or decrease down a group? Why?

- Decrease
- Shielding increases → weaker attraction
- Atomic radius increases → distance between the outer electrons and nucleus increases → weaker attraction
- Increase in number of protons is outweighed by increase in distance and shielding



Describe the structure, forces and bonding in every element across period 2.





Describe the structure, forces and bonding in every element across period 2

- Li & Be → giant metallic ; strong attraction between positive ions and delocalised electrons ; metallic bonding
- B & C → giant covalent ; strong forces between atoms ; covalent
- $N_2$  ,  $O_2$  ,  $F_2$  , Ne → simple molecular ; weak intermolecular forces between molecules ; covalent bonding within molecules and intermolecular forces between molecules



Describe the structure, forces and bonding in every element across period 3.



Describe the structure, forces and bonding in every element across period 3

- Na, Mg, Al → giant metallic ; strong attraction between positive ions and delocalised electrons ; metallic bonding
- Si → giant covalent ; strong forces between atoms ; covalent
- P<sub>4</sub> , S<sub>8</sub> , Cl<sub>2</sub> , Ar → simple molecular ; weak intermolecular forces between molecules ; covalent bonding within molecules and intermolecular forces between molecules

