Cambridge International AS & A Level Chemistry

Candidate Name		College Name	

CHAPTER 1 **9701/1 & 9701/2**

Atomic Structure **2025 – 2027 syllabus**

Prerequisites: States of matter (IGCSE)
Atoms, elements & compounds (IGCSE)

BOOKLET SUMMARY

- Particles in the atom and atomic radius
- Isotopes
- Electrons, energy levels and atomic orbitals
- Ionisation energy
- Tutorial questions

INFORMATION

- This booklet (version 01) is designed by Vince Hoor © 2024.
- The PDF copy of this booklet is available in the Google Classroom.

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Learning Outcome

1.1	**Particles in the atom and atomic radius**
1	understand that atoms are mostly empty space surrounding a very small, dense nucleus that contains protons and neutrons; electrons are found in shells in the empty space around the nucleus
2	identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses
3	understand the terms atomic and proton number; mass and nucleon number
4	describe the distribution of mass and charge within an atom
5	describe the behaviour of beams of protons, neutrons and electrons moving at the same velocity in an electric field
6	determine the number of protons, neutrons and electrons present in both atoms and ions given atomic or proton number, mass or nucleon number and charge
7	state and explain qualitatively the variations in atomic radius and ionic radius across a period and down a group

1.2	**Isotopes**
1	define the term isotope in terms of numbers of protons and neutrons
2	understand the notation ${}^{x}_{y}A$ for isotopes, where x is the mass or nucleon number and y is the atomic or proton number
3	state that and explain why isotopes of the same element have the same chemical properties
4	state that and explain why isotopes of the same element have different physical properties, limited to mass and density

1.3	**Electrons, energy levels and atomic orbitals**
	• In 1.3 each atom or ion described will be in the ground state. • Only the elements hydrogen to krypton will be assessed.
1	understand the terms: • shells, sub-shells and orbitals • principal quantum number (n) • ground state, limited to electronic configuration
2	describe the number of orbitals making up s, p, and d sub-shells, and the number of electrons that can fill s, p, and d-subshells
3	describe the order of increasing energy of the sub-shells within the first three shells and the 4s and 4p subshells
4	describe the electronic configurations to include the number of electrons in each shell, sub-shell and orbital.
5	explain the electronic configurations in terms of energy of the electrons and inter-electron repulsion
6	determine the electronic configuration of atoms and ions given the atomic or proton number and charge, using either of the following conventions: e.g. for Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (full electronic configuration) or $[Ar]3d^6 4s^2$ (shorthand electronic configuration)
7	understand and use the electrons in the boxes notation e.g. for Fe: [Ar] 
8	describe and sketch the shapes of s and p orbitals
9	describe a free radical as a species with one or more unpaired electrons

1.4	**Ionisation energy**
	• In 1.4 each atom or ion described will be in the ground state. • Only the elements hydrogen to krypton will be assessed.
1	define and use the term first ionisation energy, IE
2	construct equations to represent first, second and subsequent ionisation energies
3	identify the explain the trends in ionisation energies across a period and down a group of the Periodic Table.
4	identify and explain the variation in successive ionisation energies of an element
5	understand that ionisation energies are due to the attraction between nucleus and the outer electron.
6	explain the factors influencing the ionisation energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and sub-shells and spin-pair repulsion
7	deduce the electronic configurations of elements using successive ionisation energy data
8	deduce the position of an element in the Periodic Table using successive ionisation energy data

Quick Revision Notes

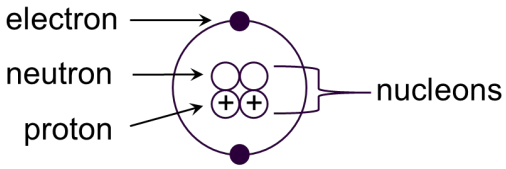
terms	definition
proton number	number of protons present in the nucleus of an atom
nucleon number	total number of protons and neutrons present in the nucleus of an atom
remark	proton number : atomic number nucleon number : mass number

Comprehensive Notes

1 Particles in the atom and atomic radius

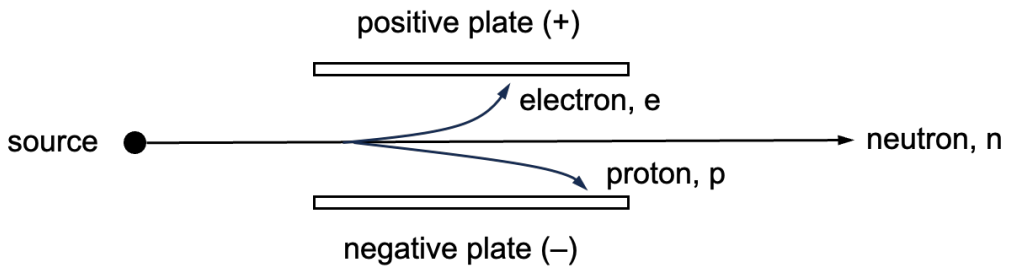
Particles in the atom

idea	Atoms are mostly empty space surrounding a very small, dense nucleus that contains protons and neutrons
	electrons are negatively charged subatomic particles found in shells in the empty space around the nucleus.



subatomic particle	symbol	relative mass	charge
proton	p	1	+1
neutron	n	1	0
electron	e	$\frac{1}{1836}$	–1

note	evidence of charges present in subatomic particles



info	neutron is uncharged and it is not deflected by the electrically charged plate
	electron is deflected towards positively charged plate
	proton is deflected towards negatively charged plate

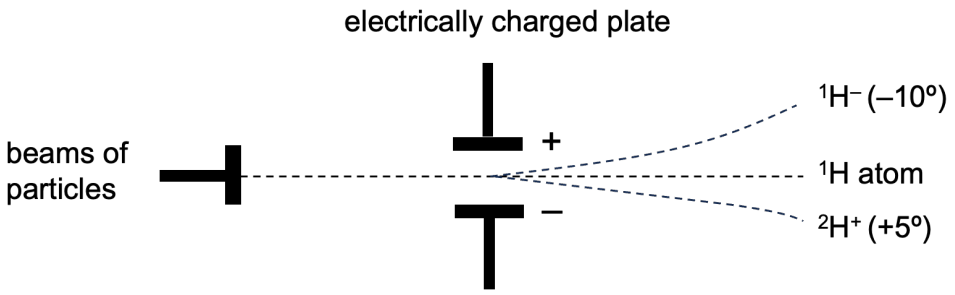
Case Study	deflection angle of H atom and H ions

The angle of deflection is **directly** proportional to *charge/mass* ratio.

The $^{1}H^{-}$ deflects with an angle of –10º.
^{1}H atom does not reflect in this experiment.

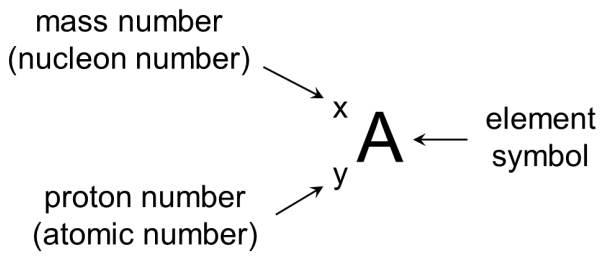
Determine the angle of deflection of $^{2}H^{+}$.

particle	charge	mass	charge / mass ratio	angle of deflection
^{1}H	0	1	0	0
$^{1}H^{-}$	1	1	(–)1 : 1	–10º
$^{2}H^{+}$	1	2	(+)1 : 2	+5º



2 Isotopes

	definition and properties of isotopes
definition	atoms of the same elements with the same proton number but with different nucleon number / number of neutrons

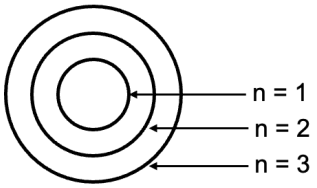


properties	isotopes have same chemical properties because they have the same number of electrons
	isotopes have different physical properties because they have different mass numbers

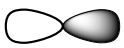
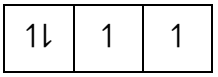
3 Electrons, energy levels and atomic orbitals

(a) Principal quantum number (shell), subshell and atomic orbitals

order →	principal quantum number	subshells (s, p, d)	atomic orbitals

Level 1	principal quantum number, n (number of electron shell)	
description & example		C atom has 6 electrons 2 electrons are in *n* = 1 4 electrons are in *n* = 2

Level 2	subshells (s, p, d)		
description & example	• principal quantum number is assigned to subshells • e.g. the first electron shell contains s subshell, this subshell is known as 1s • each subshell can occupy a specific number of electrons		
	subshell	maximum number of electrons	n + subshell (up to Kr atom)
	s	2	1s, 2s, 3s, 4s
	p	6	2p, 3p, 4p
	d	10	3d

Level 3	atomic orbitals		
description & example	• an atomic orbital has its own distinct shape and position in the shell. • each atomic orbital can be filled with two electrons with opposite spin. • an atomic orbital is represented by box diagram.		
	subshell	atomic orbital(s)	shape of atomic orbital
	s	s	
	p	p_x, p_y, p_z	 on *x*, *y*, *z* axis
	d	not required for AS Level	
box diagram	electronic configuration of O atom	$1s^2$ 	$2s^2$  $2p^4$ 

(b) Energy level of subshells and ground state electronic configuration

order of energy level of subshells

lowest	→					highest
1s	2s	2p	3s	3p	4s	3d	4p

guideline and rules of ground state electronic configuration

guideline	1s subshell is the closest to the nucleus, so it has the lowest energy level energy level increases with increasing distance from the nucleus
rules of filling electrons	• electrons always fill the lowest energy orbital first before the higher ones. • In an atomic orbital, the spin of electrons (1↓) must be in the opposite direction. • In a subshell, fill each orbital with an electron first before pairing them up.
remark	ground state refers to the lowest energy state
	excited state will **not** be tested in AS Level examination

types of electronic configuration (e.g. H → Mg)

element	full electronic configuration	shorthand notation	box notation								
			1s		2s		2p				3s
He	$1s^2$	[He]	1↓								
Li	$1s^2\ 2s^1$	[He] $2s^1$	1↓		1						
Be	$1s^2\ 2s^2$	[He] $2s^2$	1↓		1↓						
B	$1s^2\ 2s^2\ 2p^1$	[He] $2s^2\ 2p^1$	1↓		1↓		1				
C	$1s^2\ 2s^2\ 2p^2$	[He] $2s^2\ 2p^2$	1↓		1↓		1	1			
N	$1s^2\ 2s^2\ 2p^3$	[He] $2s^2\ 2p^3$	1↓		1↓		1	1	1		
O	$1s^2\ 2s^2\ 2p^4$	[He] $2s^2\ 2p^4$	1↓		1↓		1↓	1	1		
F	$1s^2\ 2s^2\ 2p^5$	[He] $2s^2\ 2p^5$	1↓		1↓		1↓	1↓	1		
Ne	$1s^2\ 2s^2\ 2p^6$	[Ne]	1↓		1↓		1↓	1↓	1↓		
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	[Ne] $3s^1$	1↓		1↓		1↓	1↓	1↓		1
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$	[Ne] $3s^2$	1↓		1↓		1↓	1↓	1↓		1↓

electronic configuration of Cr and Cu

element	chromium, Cr																								
full	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$																								
shorthand	[Ar] $3d^5\ 4s^1$																								
box notation	1s		2s		2p				3s		3p				3d										4s
	1↓		1↓		1↓	1↓	1↓		1↓		1↓	1↓	1↓		1	1	1	1	1						1

element	copper, Cu																								
full	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$																								
shorthand	[Ar] $3d^{10}\ 4s^1$																								
box notation	1s		2s		2p				3s		3p				3d										4s
	1↓		1↓		1↓	1↓	1↓		1↓		1↓	1↓	1↓		1↓	1↓	1↓	1↓	1↓						1

Electronic configuration of free radicals and ions

terms	definition
free radical	a species that contains an unpaired electron
cation	a positively charged ion
anion	a negatively charged ion

species	electronic configuration
C*l* radical	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Mg^{2+}	$1s^2\ 2s^2\ 2p^6$
P^{3-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Cr^{3+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$
Cu^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9$
guideline	for cations, remove electrons from the orbital of the highest energy
	for anions, add electrons to the orbital of the highest energy
	for **all** transition elements (Ti → Cu), always remove electrons from 4s orbital before 3d orbital

4 Ionisation energy

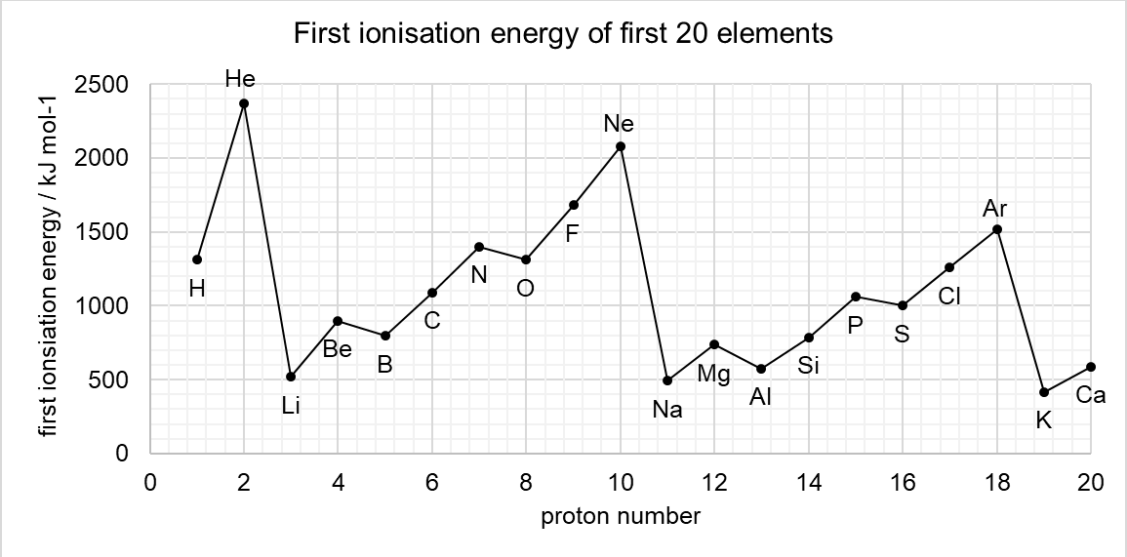
first ionisation energy

definition	energy needed to remove an electron from each atom in one mole of **gaseous** atoms		
equation	$\mathbf{M}(g) \rightarrow \mathbf{M}^{+}(g) + e^{-}$		
e.g.	metal	$Mg(g) \rightarrow Mg^{+}(g) + e^{-}$	1^{st} IE = +736 kJ mol^{-1}
	non-metal	$S(g) \rightarrow S^{+}(g) + e^{-}$	1^{st} IE = +1000 kJ mol^{-1}
concept	• an electron is removed from the outer shell during first ionisation process • first ionisation energy is an **endothermic** process as the energy is absorbed to overcome electrostatic attraction between the nucleus and outer electron		
factors affecting first ionisation energy	nuclear charge	shielding effect	
	nuclear charge increases with increasing proton number in the nucleus.	shielding effect increases with increasing number of filled electron shells.	
	electrostatic attraction between nucleus and outer electron increases with increasing nuclear charge.	electrostatic attraction between nucleus and outer electrons decreases with increasing shielding effect.	
remark	other factor affecting first ionisation energy: spin-pair repulsion		

Case Study 1	Compare the first ionisation energy of Na and Mg (via nuclear charge)
$Na(g) \rightarrow Na^{+}(g) + e^{-}$ 1^{st} IE = +496 kJ mol^{-1} $Mg(g) \rightarrow Mg^{+}(g) + e^{-}$ 1^{st} IE = +736 kJ mol^{-1} Explain why the first ionisation energy of Mg is higher than Na.	
Mg has greater nuclear charge than Na. Mg has stronger electrostatic attraction between the nucleus and outer electron.	

Case Study 2	Compare the first ionisation energy of Mg and Ca (via shielding effect)
$Mg(g) \rightarrow Mg^{+}(g) + e^{-}$ 1^{st} IE = +736 kJ mol^{-1} $Ca(g) \rightarrow Ca^{+}(g) + e^{-}$ 1^{st} IE = +590 kJ mol^{-1} Explain why the first ionisation energy of Mg is higher than Ca.	
Ca has greater number of filled electron shell than Mg, so the outer electron experiences greater shielding effect, the electrostatic attraction between the nucleus and outer electron is weaker in Ca.	

Trend of first ionisation energies (H → Ca)



Explaining the trend of first ionisation energies

trend 1	first ionisation energy across Period 3							
element	Na	Mg	Al	Si	P	S	Cl	Ar
proton number	11	12	13	14	15	16	17	18
IE / kJ mol^{-1}	494	736	577	786	1060	1000	1260	1520
trend explanation	• nuclear charge increases across Period 3 due to increasing proton number. • the shielding effect experienced by outer electron remains almost constant as electrons are added to the same shell (n = 3). • overall, the electrostatic attraction between the nucleus and outer electron becomes stronger.							

trend 2	first ionisation energy across down the group				
element	Be	Mg	Ca	Sr	Ba
proton number	4	12	20	38	56
1st IE / kJ mol^{-1}	900	736	590	548	502
trend explanation	• number of filled electron shells increases down the group • this increases the shielding effect experienced by outer electron • electrostatic attraction between the nucleus and outer electron becomes weaker				
remark	nuclear charge and shielding effect both increase down the group; the increase in shielding effect outweighs the increasing nuclear charge.				

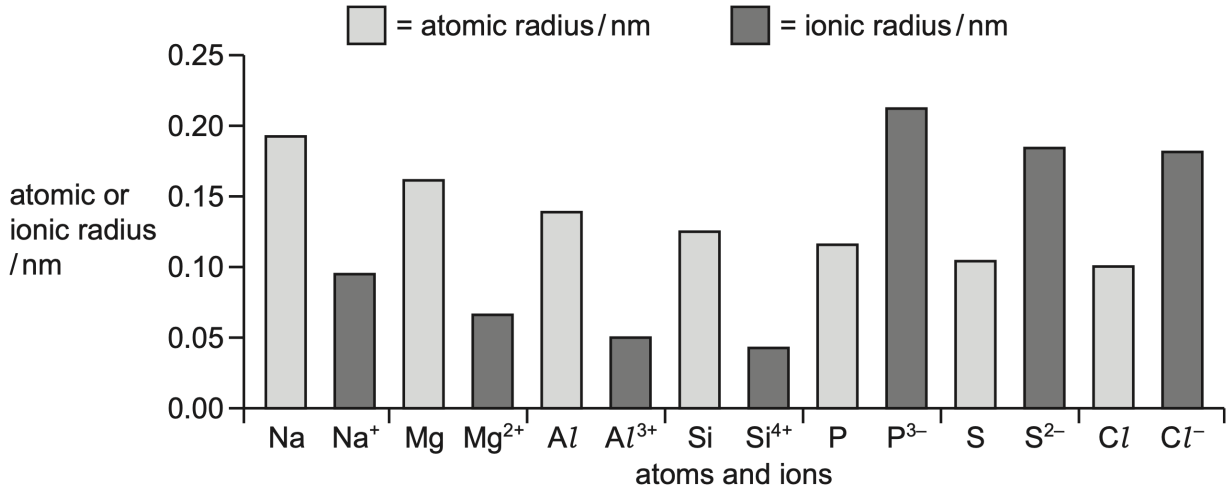
trend 3	anomalous trend across the Period			
	Group 2 → Group 13		Group 15 → Group 16	
element	Mg	Al	P	S
configuration	$1s^2\ 2s^2\ 2p^6\ 3s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
shorthand notation (n = 3)	3s | 3p ↿⇂ | ↿⇂ ⬜ ⬜	3s | 3p ↿⇂ | ↿⇂ ↑ ⬜ ⬜	3p ↑ ↑ ↑	3p ↿⇂ ↑ ↑
1st IE / kJ mol^{-1}	736	577	1060	1000
explanation	first electron to be removed from Al is from 3p orbital		there is spin-pair repulsion between paired electrons in one 3p orbital	
	3p orbital experiences greater shielding effect than 3s orbital		this effect outweighs the increasing nuclear charge in S atom	
remark	similar trend can be seen between: Be and B (Group 2 → Group 13), N and O (Group 15 → Group 16)			

Successive ionisation energies

equation	2nd IE	$\mathbf{M}^+(g) \rightarrow \mathbf{M}^{2+}(g) + e^-$	e.g.	$Ca^+(g) \rightarrow Ca^{2+}(g) + e^-$
	3rd IE	$\mathbf{M}^{2+}(g) \rightarrow \mathbf{M}^{3+}(g) + e^-$	e.g.	$Al^{2+}(g) \rightarrow Al^{3+}(g) + e^-$
trend	• IE value generally increases with stages • When electron is removed from an atom, the resulting ion becomes smaller as there are more protons attracting fewer number of electrons.			
application	identify the possible group of the element based on successive ionisation energies			

Case Study	Four successive ionisation energies of Mg			
stages of IE	1st IE	2nd IE	3rd IE	4th IE
value / kJ mol^{-1}	736	1450	7740	10500
trend	2nd IE is higher than 1st IE because electron is removed from a smaller Mg^+ ion, the second electron to be removed is subject to stronger nuclear attraction			
	3rd IE is much higher than 2nd IE because third electron is removed from the inner shell of Mg, where it is subject to stronger nuclear attraction			
conclusion	Mg has two outer electrons as there is a big jump between 2nd and 3rd IE values			

5 Atomic and ionic radii of Period 3 elements



trend	explanation
atomic radius decreases across the Period (Na → Cℓ)	• nuclear charge increases across Period 3 • the shielding effect remains almost constant as electrons are added to the same principal quantum number (n = 3) • overall, the electrostatic attraction between the nucleus and electron becomes stronger across the Period
cation is always smaller than its corresponding atom	• cation has one less electron shell than its corresponding atom
anion is always larger than its corresponding atom	• electrons are added to the outer shell, increasing repulsion between electrons in the outer shell
$Na^+ > Mg^{2+} > Aℓ^{3+} > Si^{4+}$	• nuclear charge increases from $Na^+ \rightarrow Mg^{2+} \rightarrow Aℓ^{3+} \rightarrow Si^{4+}$ • shielding effect experienced by outer electrons remains constant • overall, there is an increasing electrostatic attraction between nucleus and outer electrons.
$P^{3-} > S^{2-} > Cℓ^-$	

case study	compare the radius of Ar, Ca^{2+}, K^+		
species	Ar	K^+	Ca^{2+}
number of protons	18	19	20
number of electrons	18	18	18
trend of radius	Ar > K^+ > Ca^{2+}		
explanation	• Ar, K^+ and Ca^{2+} are isoelectronic species • nuclear charge increases from Ar → K^+ → Ca^{2+} • shielding effect experienced by outer electrons remains constant • there is an increasing electrostatic attraction between the nucleus and outer electrons from Ar → K → Ca^{2+}.		

Appendix 1 Electron configuration of first 36 elements (H → Kr) in the Periodic Table

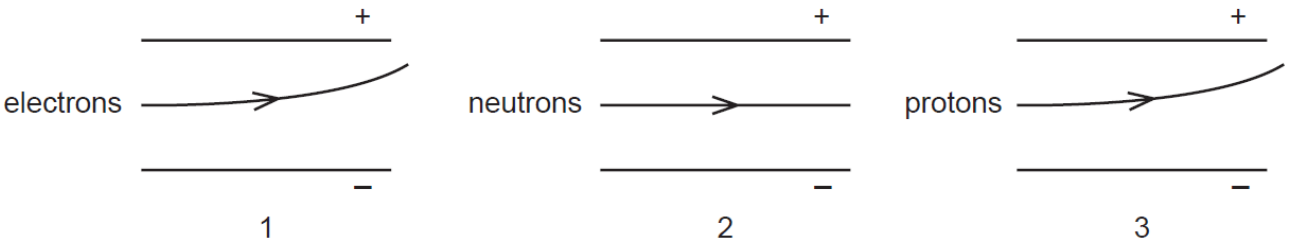
element	symbol	atomic number	full electronic configuration	shorthand notation
hydrogen	H	1	$1s^1$	
helium	He	2	$1s^2$	[He]
lithium	Li	3	$1s^2\ 2s^1$	[He] $2s^1$
beryllium	Be	4	$1s^2\ 2s^2$	[He] $2s^2$
boron	B	5	$1s^2\ 2s^2\ 2p^1$	[He] $2s^2\ 2p^1$
carbon	C	6	$1s^2\ 2s^2\ 2p^2$	[He] $2s^2\ 2p^2$
nitrogen	N	7	$1s^2\ 2s^2\ 2p^3$	[He] $2s^2\ 2p^3$
oxygen	O	8	$1s^2\ 2s^2\ 2p^4$	[He] $2s^2\ 2p^4$
fluorine	F	9	$1s^2\ 2s^2\ 2p^5$	[He] $2s^2\ 2p^5$
neon	Ne	10	$1s^2\ 2s^2\ 2p^6$	[Ne]
sodium	Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$	[Ne] $3s^1$
magnesium	Mg	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$	[Ne] $3s^2$
aluminium	Al	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	[Ne] $3s^2\ 3p^1$
silicon	Si	14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	[Ne] $3s^2\ 3p^2$
phosphorus	P	15	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	[Ne] $3s^2\ 3p^3$
sulfur	S	16	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	[Ne] $3s^2\ 3p^4$
chlorine	Cl	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	[Ne] $3s^2\ 3p^5$
argon	Ar	18	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	[Ar]
potassium	K	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	[Ar] $4s^1$
calcium	Ca	20	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	[Ar] $4s^2$
scandium	Sc	21	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$	[Ar] $3d^1\ 4s^2$
titanium	Ti	22	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$	[Ar] $3d^2\ 4s^2$
vanadium	V	23	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$	[Ar] $3d^3\ 4s^2$
chromium	Cr	24	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \mathbf{3d^5\ 4s^1}$	[Ar] $\mathbf{3d^5\ 4s^1}$
manganese	Mn	25	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$	[Ar] $3d^5\ 4s^2$
iron	Fe	26	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$	[Ar] $3d^6\ 4s^2$
cobalt	Co	27	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$	[Ar] $3d^7\ 4s^2$
nickel	Ni	28	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$	[Ar] $3d^8\ 4s^2$
copper	Cu	29	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \mathbf{3d^{10}\ 4s^1}$	[Ar] $\mathbf{3d^{10}\ 4s^1}$
zinc	Zn	30	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$	[Ar] $3d^{10}\ 4s^2$
gallium	Ga	31	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$	[Ar] $3d^{10}\ 4s^2\ 4p^1$
germanium	Ge	32	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^2$	[Ar] $3d^{10}\ 4s^2\ 4p^2$
arsenic	As	33	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$	[Ar] $3d^{10}\ 4s^2\ 4p^3$
selenium	Se	34	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$	[Ar] $3d^{10}\ 4s^2\ 4p^4$
bromine	Br	35	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$	[Ar] $3d^{10}\ 4s^2\ 4p^5$
krypton	Kr	36	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$	[Kr]

Tutorial Questions

Section A

For each question there are four possible answers, **A**, **B**, **C** and **D**. Choose the **one** you consider to be correct.

1 The diagrams show the possible path of subatomic particles moving in an electric field in a vacuum.



Which diagrams are correct?

A 1 and 2 only **B** 1 and 3 only **C** 2 and 3 only **D** 1, 2 and 3

2 Neutrons are passed through an electric field. The mass of one neutron relative to unified atomic mass and any deflection in the electric field is recorded.

Which row is correct?

	mass of neutron	behaviour of beam of neutrons in an electric field
A	0	deflected
B	1	deflected
C	0	not deflected
D	1	not deflected

3 Beams of charged particles are deflected by an electrical field. The angle of deflection of a particle is proportional to its charge/mass ratio.

In an experiment protons are deflected by an angle of +15°.

In another experiment under identical conditions $^{2}H^{-}$ ions are deflected by an angle of Y°.

What is the value of Yº?

A –30.0 **B** –7.5 **C** +7.5 **D** +30.0

4 The table shows the numbers of protons, neutrons and electrons in four different particles, W, X, Y and Z.

	number of protons	number of neutrons	number of electrons
W	32	40	32
X	32	40	34
Y	32	42	32
Z	34	40	34

Which pair represents the atoms of two isotopes of the same element?

A W and Y **B** W and Z **C** X and Y **D** X and Z

5 Helium, He is the second element in the Periodic Table.

Tritium is the isotope of hydrogen ^{3}H.

What is the same in an atom of ^{4}He and an atom of ^{3}H?

A the number of electrons

B the number of neutrons

C the number of protons

D the relative atomic mass

6 A radioactive isotope of thallium $^{201}_{81}Tl$, ,is used to assess damage in heart muscles after a heart attack.

Which statement about $^{201}_{81}Tl$ is correct?

A This isotope has a nucleon number of 120.

B The number of electrons in one atom of this isotope is 81.

C The number of neutrons in one atom of this isotope is 201.

D $^{201}_{82}X$ is an isotope of $^{201}_{81}Tl$.

7 It is now thought that where an element exists as several isotopes, the stable ones usually contain a ‘magic number’ of neutrons. One of these magic number is 126.

Which isotope is unstable?

A ^{209}Bi **B** ^{208}Pb **C** ^{210}Po **D** ^{208}Tl

8 Unnilpentium is an artificial element. One of its isotopes is $^{262}_{105}$Unp.

Which of the following statements is correct?

A $^{262}_{105}$Unp has a nucleon number of 105.

B The atom $^{260}_{105}$X is an isotope of $^{262}_{105}$Unp.

C There are 262 neutrons in $^{262}_{105}$Unp.

D The proton number of $^{262}_{105}$Unp is 262.

9 A single ^{32}P nucleus can be produced when a single ^{32}S nucleus joins with particle X.
In the process a proton is emitted.

What is particle X?

A a deuteron, $^{2}_{1}H^{+}$

B an electron

C a neutron

D a proton

10 Skin cancer can be treated using a radioactive isotope of phosphorus, $^{32}_{15}$P. A compound containing the phosphide ion $^{32}_{15}P^{3-}$, wrapped in a plastic sheet, is strapped to the affected area.

What is the composition of the phosphide ion, $^{32}_{15}P^{3-}$?

	protons	neutrons	electrons
A	15	17	18
B	15	17	32
C	17	15	17
D	32	17	15

11 What number of protons, neutrons and electrons are present in the ion $^{54}Fe^{3+}$?

	protons	neutrons	electrons
A	26	28	23
B	26	28	29
C	29	25	23
D	29	25	26

12 Which atom contains four times as many neutrons as the $^{7}_{3}Li$ atom?

A $^{40}_{20}Ca$ **B** $^{24}_{12}Mg$ **C** $^{31}_{15}P$ **D** $^{26}_{14}Si$

13 In which species are the numbers of electrons and neutrons equal?

A $^{9}_{4}Be$ **B** $^{19}_{9}F$ **C** $^{23}_{11}Na^{+}$ **D** $^{18}_{8}O^{2-}$

14 Which isotope of an element in the third period of the Periodic table contains the same number of neutrons as $^{32}_{16}S$?

A $^{24}_{12}Mg$ **B** $^{28}_{14}Si$ **C** $^{23}_{11}Na$ **D** $^{31}_{15}P$

15 In which species are the number of protons, neutrons and electrons all different?

A $^{11}_{5}B$ **B** $^{19}_{9}F^{-}$ **C** $^{23}_{11}Na^{+}$ **D** $^{24}_{12}Mg^{2+}$

16 In which species are the number of protons, neutrons and electrons all different?

A $^{27}_{13}Al$ **B** $^{35}_{17}Cl^{-}$ **C** $^{32}_{16}S^{2-}$ **D** $^{39}_{19}K^{+}$

17 Which atom has the same number of electrons as the hydroxide ion, OH^{-}?

A F **B** Ne **C** Na **D** Mg

18 Which ion has more electrons than protons and more protons than neutrons?

[H = $^{1}_{1}H$; D = $^{2}_{1}H$; O = $^{16}_{8}O$]

A D^{-} **B** H_3O^{+} **C** OD^{-} **D** OH^{-}

19 In which set do all species contain the same number of electrons?

A Co^{2+}, Co^{3+}, Co^{4+}

B F^{-}, Br^{-}, Cl^{-}

C Na^{+}, Mg^{2+}, Al^{3+}

D K_2SO_4, K_2SeO_4, K_2TeO_4

20 Sodium azide, NaN_3 is an explosive used to inflate airbags in cars when they crash.

It consists of positive sodium ions and negative azide ions.

What are the number of electrons in the sodium ion and the azide ion?

	sodium ion	azide ion
A	10	20
B	10	22
C	12	20
D	12	22

21 Which statement about the electrons in a ground state carbon atom is correct?

A Electrons are present in four different energy levels.

B There are more electrons in p orbitals than there are in s orbitals.

C The occupied orbital of highest energy is spherical.

D The occupied orbital of lowest energy is spherical.

22 Which statement about a 3p orbital is correct?

A It can hold a maximum of 6 electrons.

B It has the highest energy of the orbitals with principal quantum number 3.

C It is at a higher energy level than a 3s orbital but has the same shape.

D It is occupied by one electron in an isolated phosphorus atom.

23 An atom has eight protons.

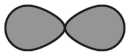
Which diagram shows the electronic configuration of this atom in its lowest energy state?

	1s	2s	2p			3s
A	↿⇂	↿⇂	↿⇂	↿⇂		
B	↿⇂	↿⇂	↿⇂	↑	↑	
C	↿⇂	↿⇂	↿⇂	↑		↑
D	↿⇂	↿⇂	↑	↑	↑	↑

24 Which atomic orbitals are occupied in an atom of phosphorus?

A 1p 2s 2p **B** 2s 2p 2d **C** 2s 2p 3s **D** 2p 3s 3d

25 Which atom has its outermost electron in an orbital of the shape shown, with principal quantum number 3?



A sodium **B** chlorine **C** calcium **D** bromine

26 What is the order of increasing energy of the listed orbitals in the atom of titanium?

A 3s 3p 3d 4s

B 3s 3p 4s 3d

C 3s 4s 3p 3d

D 4s 3s 3p 3d

27 Which of the elements below has an **equal** number of electron pairs and of unpaired electrons with orbitals of principal quantum number 2?

A beryllium **B** carbon **C** nitrogen **D** oxygen

28 What could be the proton number of an element that has three unpaired electrons in each of its atoms?

A 5 **B** 13 **C** 15 **D** 21

29 Which atom has exactly three unpaired electrons in the ground state?

A an isolated gaseous aluminium atom

B an isolated gaseous carbon atom

C an isolated gaseous chromium atom

D an isolated gaseous phosphorus atom

30 This question refers to isolated gaseous atoms.

In which atom are all electrons paired?

A Ba **B** Br **C** S **D** Si

31 In which pair of species do both species have only one unpaired p electron?

A Ar^+ and C^- **B** B and Ti^+ **C** F and Ga **D** Se^- and Si^-

32 The table refers to the electron distribution in the second shell of an atom with eight protons.

Which row is correct for this atom?

	orbital shape 		orbital shape 	
	orbital type	number of electrons	orbital type	number of electrons
A	p	2	s	4
B	p	4	s	2
C	s	2	p	4
D	s	4	p	2

33 The electronic configuration of an atom of sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

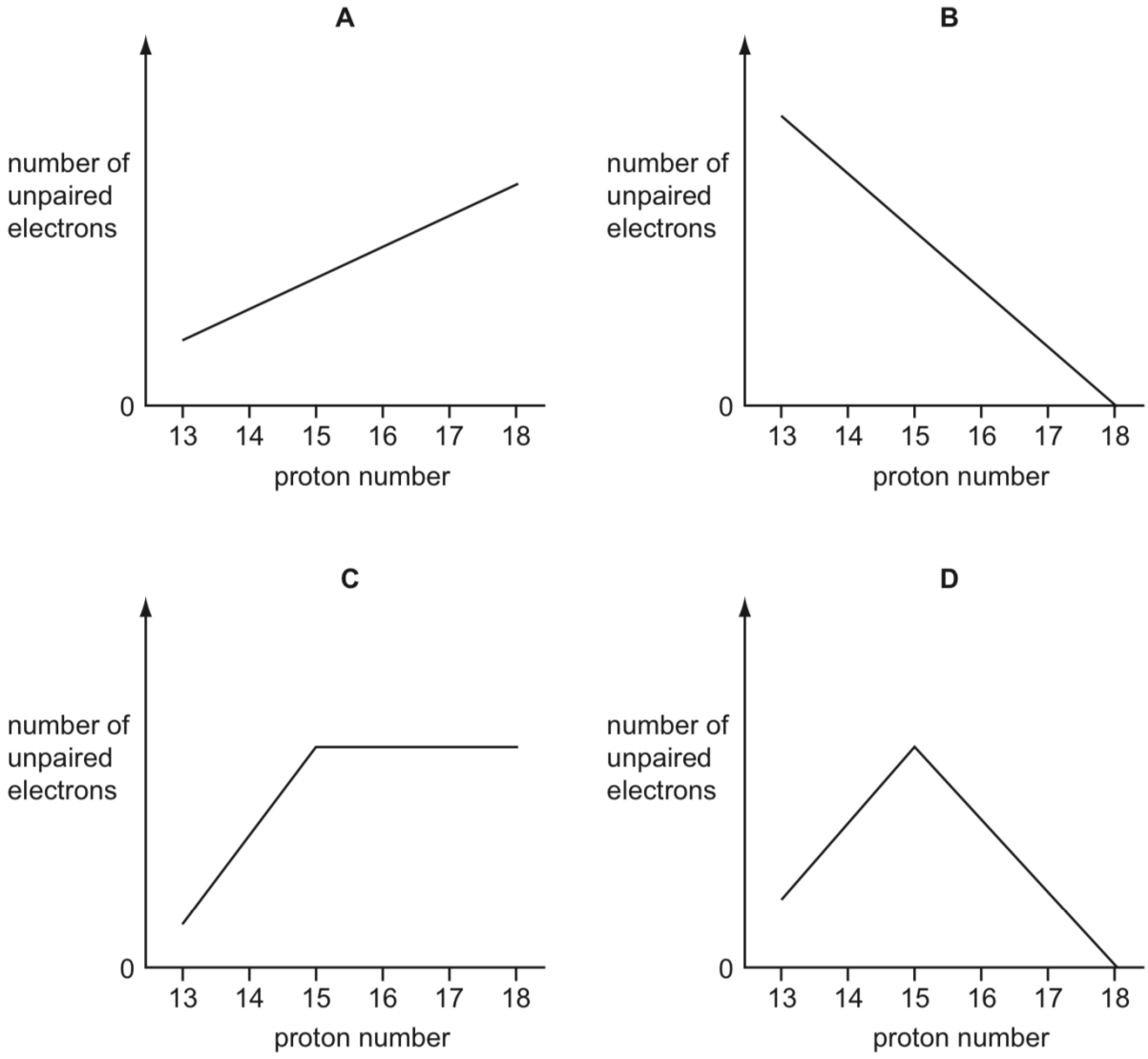
How many valence shell and unpaired electrons are present in one sulfur atom?

	valence shell electrons	unpaired electrons
A	2	1
B	4	2
C	6	0
D	6	2

34 Where in the Periodic Table is the element that has an outer electron shell arrangement of $4s^2\ 4p^3$?

	Group	Period
A	13	3
B	13	4
C	15	3
D	15	4

35 Which graph represents the number of unpaired p orbital electrons for atoms with proton numbers 13 to 18?



36 The electronic structures of calcium, krypton, phosphorus and an element **Y** are shown.

Which electronic structure is that of element **Y**?

A $1s^2 2s^2 2p^6 3s^2 3p^3$

B $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

C $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

D $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

37 What is the electronic configuration of Mg^{2+}?

A $1s^2\ 2s^2\ 2p^6$

B $1s^2\ 2s^2\ 2p^6\ 3s^2$

C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

38 Gallium nitride, GaN, is an ionic compound containing the Ga^{3+} ion.

What is the electron arrangement of the nitrogen ion in gallium nitride?

A $1s^2\ 2s^2$

B $1s^2\ 2s^2\ 2p^3$

C $1s^2\ 2s^2\ 2p^4$

D $1s^2\ 2s^2\ 2p^6$

39 The ion X^{2+} has the same electronic configuration as the atom Kr. What is the electronic configuration of an atom of X?

A [Ar] $4s^2 3d^{10} 4p^6$

B [Ar] $4s^2 3d^{10} 4p^6 5s^2$

C [Ar] $4s^2 4d^{10} 4p^6$

D [Ar] $4s^2 4d^{10} 4p^6 5s^2$

40 When chlorine gas is analysed in a mass spectrometer, $^{35}C\ell^+$ ions are detected.

Which row is correct?

	number of neutrons in $^{35}C\ell^+$	electronic configuration of $^{35}C\ell^+$
A	17	$1s^2 2s^2 2p^6 3s^2 3p^4$
B	17	$1s^2 2s^2 2p^6 3s^2 3p^6$
C	18	$1s^2 2s^2 2p^6 3s^2 3p^4$
D	18	$1s^2 2s^2 2p^6 3s^2 3p^6$

41 Elements generally form an ion with the electronic structure of a noble gas.

Which ion does **not** have a noble gas electronic structure?

A I^- **B** Rb^+ **C** Sn^{2+} **D** Sr^{2+}

42 Elements generally form an ion with the electronic structure of a noble gas.

Which ion does not have a noble gas electronic structure?

A K^+ **B** Pb^{2+} **C** Ba^{2+} **D** P^{3-}

43 Which of the following particles would, on losing an electron, have a half-filled set of p orbitals?

A C^- **B** N **C** N^- **D** O^+

44 Which species is a free radical?

A He **B** Be^- **C** O^{2-} **D** Zn

45 One of the most important chemical species responsible for the removal of ozone from the stratosphere is a free radical of chlorine, $^{35}Cl^{\bullet}$. Free radical of chlorine contains an unpaired electron.

What does $^{35}Cl^{\bullet}$ contain?

	protons	neutrons	electrons
A	17	18	16
B	17	18	17
C	18	17	17
D	18	17	18

46 Which equation represents the first ionisation energy of iodine?

A $\frac{1}{2}I_2(g) + e^- \rightarrow I^-(g)$

B $I(g) + e^- \rightarrow I^-(g)$

C $\frac{1}{2}I_2(g) \rightarrow I^+(g) + e^-$

D $I(g) \rightarrow I^+(g) + e^-$

47 The value of the second ionisation energy of magnesium is 1450 kJ mol^{-1}.

Which equation correctly represents this statement?

A $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$; $\Delta H^{\ominus} = +1450$ kJ mol^{-1}

B $Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$; $\Delta H^{\ominus} = +1450$ kJ mol^{-1}

C $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$; $\Delta H^{\ominus} = -1450$ kJ mol^{-1}

D $Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$; $\Delta H^{\ominus} = -1450$ kJ mol^{-1}

48 The value of the third ionisation energy of aluminium is 2740 kJ mol^{-1}.

Which equation correctly represents this statement?

A $Al(g) \rightarrow Al^{3+}(g) + 3e^-$; $\Delta H^{\ominus} = +2740$ kJ mol^{-1}

B $Al^{2+}(g) \rightarrow Al^{3+}(g) + e^-$; $\Delta H^{\ominus} = +2740$ kJ mol^{-1}

C $Al(g) \rightarrow Al^{3+}(g) + 3e^-$; $\Delta H^{\ominus} = -2740$ kJ mol^{-1}

D $Al^{2+}(g) \rightarrow Al^{3+}(g) + e^-$; $\Delta H^{\ominus} = -2740$ kJ mol^{-1}

49 Astatine is a radioactive element which has the same group as chlorine.

What is the same in an atom of astatine and an atom of chlorine?

A the total number of electrons

B the number of unpaired electrons in p orbital.

C the size of the nuclear charge

D the magnitude of the shielding effect from inner electrons.

50 Which statement about the first ionisation energies of magnesium and neon is correct?

A Magnesium has the greater numerical value and both are endothermic.

B Magnesium has the greater numerical value and both are exothermic.

C Neon has the greater numerical value and both are endothermic.

D Neon has the greater numerical value and both are exothermic.

51 Which factor helps to explain why the first ionisation energies of the Group 1 elements decrease from lithium to sodium to potassium to rubidium?

A The nuclear charge of the element increases.

B The outer electron is in an ‘s’ subshell.

C The repulsion between spin-paired electrons increases.

D The shielding effect of the inner shells increases.

52 Why is the first ionisation energy of phosphorus greater than the first ionisation energy of silicon?

A A phosphorus atom has one more proton in its nucleus.

B The atomic radius of a phosphorus atom is greater.

C The outer electron in a phosphorus atom is more shielded.

D The outer electron in a phosphorus atom is paired.

53 The first ionisation energy of potassium, K, is 418 kJ mol^{-1}. The first ionisation energy of strontium, Sr, is 548 kJ mol^{-1}.

Which statement helps to explain why Sr has a greater first ionisation energy than K?

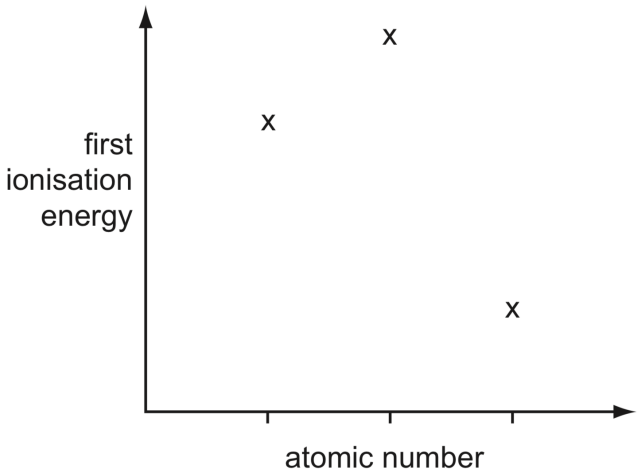
A The charge on a Sr nucleus is greater than the charge on a K nucleus.

B The outer electron in a Sr atom experiences greater shielding than the outer electron in a K atom.

C The outer electron in a Sr atom experiences spin-pair repulsion.

D The outer electron in a Sr atom is further from the nucleus than the outer electron in a K atom.

54 Three successive elements in the Periodic Table have first ionisation energies which have the pattern shown in the diagram.



What could be the first element of this sequence?

A C **B** N **C** F **D** Na

55 Equations involving four enthalpy changes are shown.

$Na(g) \rightarrow Na^{+}(g) + e^{-}$ $\Delta H = W$

$Na(g) \rightarrow Na^{2+}(g) + 2e^{-}$ $\Delta H = X$

$Na(s) \rightarrow Na(g)$ $\Delta H = Y$

$Na(s) \rightarrow Na^{2+}(g) + 2e^{-}$ $\Delta H = Z$

Which equation represents the second ionisation energy of sodium?

A X **B** X + Y – W **C** X – W **D** Z – W

56 Which of these elements has the highest fifth ionisation energy?

A C **B** N **C** P **D** Si

57 For the element sulfur, which pair of ionisation energies has the largest difference between them?

A third and fourth ionisation energies

B fourth and fifth ionisation energies

C fifth and sixth ionisation energies

D sixth and seventh ionisation energies

58 The successive ionisation energies, in kJ mol^{-1}, of an element **X** are given below.

870 1800 3000 3600 5800 7000 13200

What is **X**?

A ^{33}As **B** ^{40}Zr **C** ^{52}Te **D** ^{53}I

59 The first seven ionisation energies of an element between lithium and neon in the Periodic Table are as follows.

1310 3390 5320 7450 11 000 13 300 71 000

What is the outer electronic configuration of the element?

A $2s^2$ **B** $2s^2\ 2p^1$ **C** $2s^2\ 2p^4$ **D** $2s^2\ 2p^6$

60 Which statement is correct?

A The first ionisation energy of chlorine is more than the first ionisation energy of argon.

B The second ionisation energy of calcium is more than the second ionisation energy of magnesium.

C The second ionisation energy of sulfur is equal to the first ionisation energy of phosphorus.

D The eighth ionisation energy of chlorine is more than the first ionisation energy of neon.

61 The second ionisation energy of oxygen is greater than the second ionisation energy of fluorine. Which factor explains this difference?

A The atomic radius of an oxygen atom is smaller than that of fluorine.

B The covalent bond in a fluorine molecule is weaker than the bond in an oxygen molecule.

C A spin-paired electron is removed from fluorine but not from oxygen.

D Fluorine has more electrons in total than oxygen. This causes a greater shielding of the nuclear attraction in fluorine.

62 Element X has six more protons than element Y. Which statement must be correct?

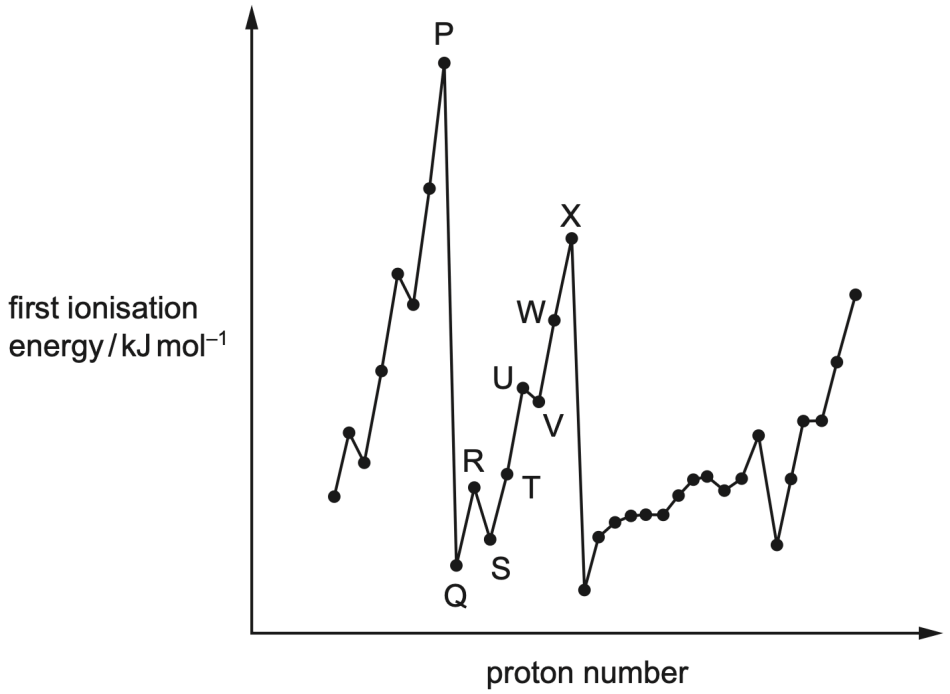
A Atoms of element Y are smaller than atoms of element X.

B Element X has a full shell of electrons.

C Element X and element Y are in the same group.

D Element X and element Y are in the same period.

63 The graph shows the variation of the first ionisation energy with proton number for some elements. The letters used are not the actual symbols for the elements.



Which statement about the elements is correct?

A P and X are in the same period in the Periodic Table.

B The general increase from Q to X is due to increasing atomic radius.

C The small decrease from R to S is due to decreased shielding.

D The small decrease from U to V is due to repulsion between paired electrons.

64 Three statements about potassium and chlorine and their ions are listed.

1 The atomic radius of a potassium atom is greater than the atomic radius of a chlorine atom.

2 The first ionisation energy of potassium is greater than the first ionisation energy of chlorine.

3 The ionic radius of a potassium ion is greater than the ionic radius of a chloride ion.

Which statements are correct?

A 1 only **B** 2 only **C** 1 and 3 **D** 2 and 3

Section B

Answer all the questions in the spaces provided.

1 Atoms contain the subatomic particles electrons, protons and neutrons. Protons and electrons were discovered by observations of their behaviours in electric fields.

(a) Fig. 1.1 shows the behaviour of separate beams of electrons and protons in an electric field.

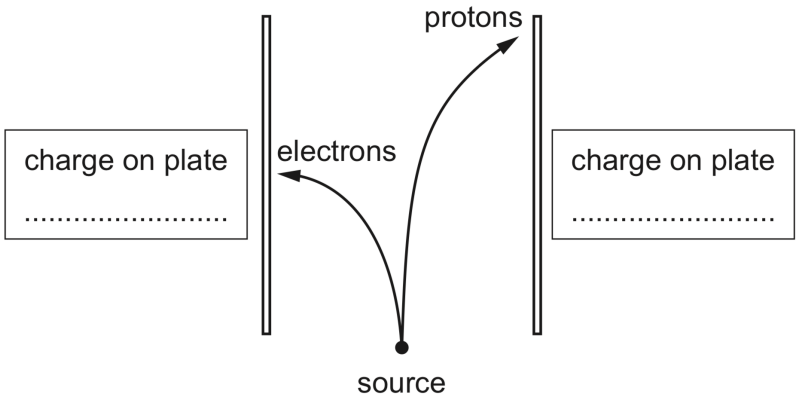


Fig. 1.1

(i) Complete the diagram with the relative charge of each of the electrically charged plates. [1]

(ii) On the diagram, draw the line to show how a separate beam of neutrons from the same source behaves in the same electric field. [1]

(b) *Hydron* is a general term used to represent the ions, $^{1}_{1}H^{+}$, $^{2}_{1}H^{+}$ and $^{3}_{1}H^{+}$.

State, in terms of subatomic particles in the nucleus, what is the same about each of these ions and what is different.

same .....

different .....

[1]

(c) Gallium has two stable isotopes, ^{69}Ga and ^{71}Ga.

Complete Table 1.1 to show the number of protons, neutrons and electrons in the two stable isotopes of gallium.

Table 1.1

isotope	number of protons	number of neutrons	number of electrons
^{69}Ga			
^{71}Ga			

[2]

(d) Complete Table 1.1 to show the composition and identity of some atoms and ions.

Table 1.2

name of element	nucleon number	atomic number	number of protons	number of neutrons	number of electrons	overall charge
lithium	6	3				+1
oxygen				9	10	
	54	26	26		24	
			17	18		0

[4]

(e) Complete Table 1.2 to show the composition and identity of some atoms and ions.

Table 1.3

name of element	nucleon number	atomic number	number of protons	number of neutrons	number of electrons	overall charge
boron	10	5				0
nitrogen				8	10	
	208	82	82		80	
			3	3		+1

[4]

(f) Complete Table 1.4 to show the total numbers of protons and electrons in the molecular ions NH_4^+, CO_3^{2-} and PO_4^{3-}.

Table 1.4

molecular ion	total number of protons	total number of electrons
NH_4^+		
CO_3^{2-}		
PO_4^{3-}		

[3]

2 **(a)** $^{25}_{12}Mg$ is an isotope of magnesium.

(i) Define isotope.

.....

..... [1]

(ii) Determine the number of protons and neutrons in an atom of $^{25}_{12}Mg$.

number of protons .....

number of neutrons .....

[1]

(iii) State the full electronic configuration of an atom of $^{25}_{12}Mg$.

..... [1]

(b) $^{25}_{12}Mg$ and $^{24}_{12}Mg$ are isotopes of magnesium.

State **one** similarity and **one** difference in the properties of these isotopes of magnesium. Explain your answer.

.....

.....

..... [2]

(c) An isotope of copper has a relative isotopic mass of 65.

Complete Table 2.1 for an atom of copper-65.

Table 2.1

	atomic number	nucleon number	number of neutrons	electronic arrangement
copper-65				

[3]

3 **(a)** Period 2 elements (B → Ne) occupies orbitals in both 2s and 2p subshells.

Sketch the shapes of a 2s and $2p_x$ orbital on the axes in Fig. 3.1.

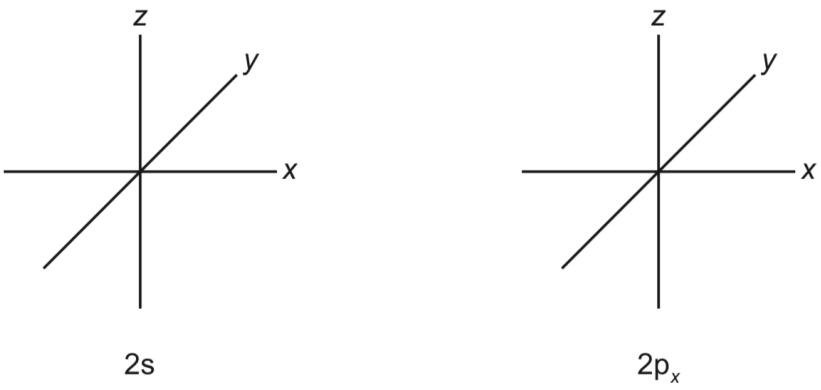


Fig. 3.1

[2]

(b) Identify the highest energy orbital which contains electrons in a calcium atom. Sketch the shape of this orbital.

identity of highest energy orbital in Ca .....

shape

[1]

(c) Fluorine is an atom with 9 protons in its nucleus.

(i) Complete the electrons in boxes diagram in Fig. 3.2 to show the electronic configuration of a fluorine atom.

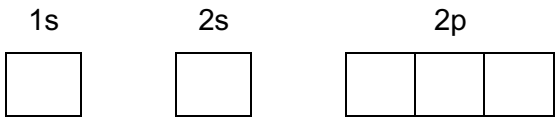


Fig. 3.2

[1]

(ii) Identify the lowest energy orbital which contains electrons in a fluorine atom.
State the number of unpaired electron in a fluorine atom.

identify of lowest energy orbital in F .....

number of unpaired electron .....

[1]

(d) Electrons in atoms up to $_{36}Kr$ are distributed in s, p and d orbitals.

(i) Use Table 3.1 to state the number of occupied orbitals in an isolated atom of $_{36}Kr$.

Table 3.1

type of orbital	s	p	d
number of orbitals			

[3]

(ii) Complete Fig. 3.3 to show the number and relative energies of the electrons in an isolated atom of $_{14}Si$.

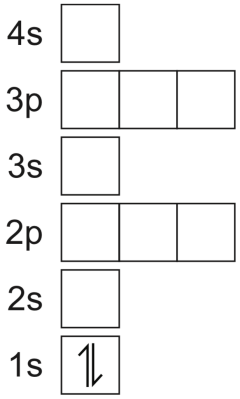


Fig. 3.3

[2]

(iii) The diagram shows a type of orbital.

State the total number of electrons that exists in all orbitals of this type in an atom of $_{9}F$.

..... [1]

(e) Use Table 3.2 to state full and shorthand electronic configuration of $_{29}Cu$.

Table 3.2

full electronic configuration	shorthand electronic configuration

[2]

(f) Complete the following electronic configuration of Fe and Fe^{3+} in Fig. 3.4.

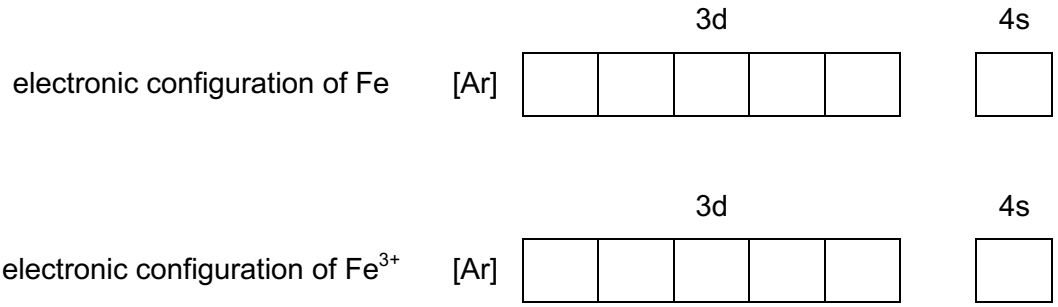


Fig. 3.4

[2]

(g) A stable ion of Cr atom has three unpaired electrons.

Complete the electronic configuration of Cr and stable ion of Cr in Fig. 3.5.

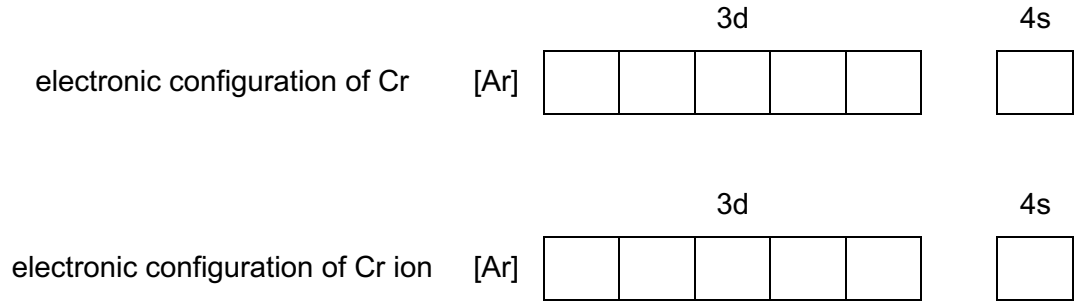


Fig. 3.5

[2]

(h) Potassium is the third element in Group 1.

(i) State the full electronic configuration of an atom of potassium.

..... [1]

(ii) State the number of fully filled orbitals present in an atom of potassium.

..... [1]

Rubidium is the fourth element in Group 1.

(iii) Identify the highest energy orbital which contains electron in rubidium.

..... [1]

4 **(a)** Define first ionisation energy.

.....

.....

..... [2]

(b) **(i)** Write an equation that represents the first ionisation energy of Na.

..... [1]

(ii) Write an equation that represents the second ionisation energy of Be.

..... [1]

(iii) Write an equation that represents the third ionisation energy of A*l*.

..... [1]

(c) State **two** dominant factors that affects the magnitude of first ionisation energy

factor 1 .....

factor 2 .....

[1]

(d) **(i)** Li and Na are in Group 1 of the Periodic Table.
Explain why the first ionisation energy of Li is higher than Na.

.....

.....

..... [2]

(ii) Na and Mg are Period 3 elements.
Explain why the first ionisation energy of Mg is higher than Na.

.....

.....

..... [2]

(e) The first ionisation energies of Period 3 elements are shown in Table 4.1.

Table 4.1

element	Na	Mg	Al	Si	P	S	Cl
IE1 kJ mol^{-1}	494	736	577	786	1060	1000	1260

State and explain the general trend in **first** ionisation energies across the period.

.....

.....

.....

.....

..... [3]

(f) The first ionisation energies of Group 2 elements are shown in Table 4.2.

Table 4.2

element	Be	Mg	Ca	Sr	Ba
IE1 kJ mol^{-1}	900	736	590	548	502

State and explain the general trend in **first** ionisation energies across down the group.

.....

.....

.....

.....

..... [3]

5 Use the information in Table 5.1 to answer the following questions.

Table 5.1

element	H	Li	C	O	S
first ionisation energy / kJ mol^{-1}	1310	519	1090	1310	1000
second ionisation energy / kJ mol^{-1}	–	7300	2350	3390	2260

(a) (i) Write an equation that represents the first ionisation energy of H.

..... [1]

(ii) Explain why there is no information give in Table 5.1 for the second ionisation energy of H.

..... [1]

(iii) Give the full electronic configuration of $S^{2+}(g)$.

..... [1]

(b) (i) Write an equation that represents the second ionisation energy of Li.

..... [1]

(ii) Explain why the second ionisation energy of Li is much higher than the first ionisation energy of Li.

.....

.....

..... [2]

(c) Explain why the second ionisation energy of C is higher than the first ionisation energy of C.

.....

.....

.....

..... [2]

(d) O and S are elements of Group 16 of the Periodic Table.

(i) State the identity of orbital where the electron is removed from S atom during the first ionisation energy.

..... [1]

(ii) Suggest why the first ionisation energy of O is higher than S.

.....

.....

..... [2]

(e) There is missing information about first ionisation energy of N in Table 4.1.

Three possible values for the first ionisation energy of N are given.

799 kJ mol^{-1} **900 kJ mol^{-1}** **1400 kJ mol^{-1}**

(i) Circle the correct value. [1]

(ii) Explain your choice by comparing your chosen value to those of C and O.

.....

.....

.....

.....

.....

..... [4]

6 **(a)** Element E is in Period 3 of the Periodic Table.
The first eight ionisation energy values of E are shown in Table 6.1.

Table 6.1

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy / kJ mol^{-1}	577	1820	2740	11 600	14 800	18 400	23 400	27 500

Deduce the full electronic configuration of **E**.
Explain your answer.

full electronic configuration of **E** = ...

explanation .....

.....

..... [3]

(b) Successive ionisation energies for element A are shown in Table 6.2.

Table 6.2

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy / kJ mol^{-1}	1310	3390	5320	7450	11 000	13 300	71 000	84 100

Use Table 6.2 to deduce the group of the Periodic Table that A belongs to.
Explain your answer.

Group .....

..... [1]

(c) The first, second and third ionisation energies of potassium are 418, 3070 and 4600 kJ mol^{-1}, respectively.

Use this information to explain why potassium is in Group 1.

.....

.....

.....

..... [2]

(d) The fifth to eight ionisation energies of three elements in the third period of the Periodic Table are given in Table 6.3. The symbols used for reference are **not** the actual symbols of the elements.

Table 6.3

element	ionisation energies, kJ mol^{-1}			
	fifth	sixth	seventh	eighth
X	6274	21 269	25 398	29 855
Y	7012	8496	27 107	31 671
Z	6542	9362	11 018	33 606

(i) State and explain the group number of element **Y**.

group number .....

explanation .....

..... [2]

(ii) Explain why the first ionisation energy of element **Y** is less than that of element **X**.

.....

.....

.....

..... [2]

(iii) Suggest the trend of atomic radius of **X**, **Y** and **Z**. Explain your answer.

.....

.....

.....

.....

..... [3]

(e) Across Period 3 there is a general trend for first ionisation energies to increase due to the increase in attraction between the nucleus and the outer electron.

(i) Explain why the first ionisation energy of aluminium is less than the first ionisation energy of magnesium.

.....

.....

..... [2]

(ii) Explain why the first ionisation energy of sulfur is less than the first ionisation energy of phosphorus.

.....

.....

..... [2]

(f) In an Al^{2+} ion the nuclear attraction for the outer electron is stronger than in an atom of Na.

Compare the electronic structures of Al^{2+} and an atom of Na and explain why the third ionisation energy of aluminium is greater than the first ionisation energy of sodium.

.....

.....

.....

..... [2]

7 Tellurium is an element in Group 16. The most common isotope of tellurium is ^{130}Te. Its electronic configuration is [Kr] $4d^{10}\ 5s^2\ 5p^4$.

(a) Complete Table 7.1.

Table 7.1

	nucleon number	number of neutrons	number of electrons
^{130}Te			

[3]

(b) Identify the sub-shell in an atom of Te that contains electrons with the lowest energy.

..... [1]

(c) Construct an equation to represent the first ionisation energy of Te.

..... [1]

(d) The radius of Te ions decreases after each successive ionisation.

(i) State two factors that are responsible for the increase in the first six ionisation energies of Te.

.....

.....

..... [2]

(ii) Sketch a graph in Fig. 7.1 to show the trend in the first **seven** ionisation energies of Te.

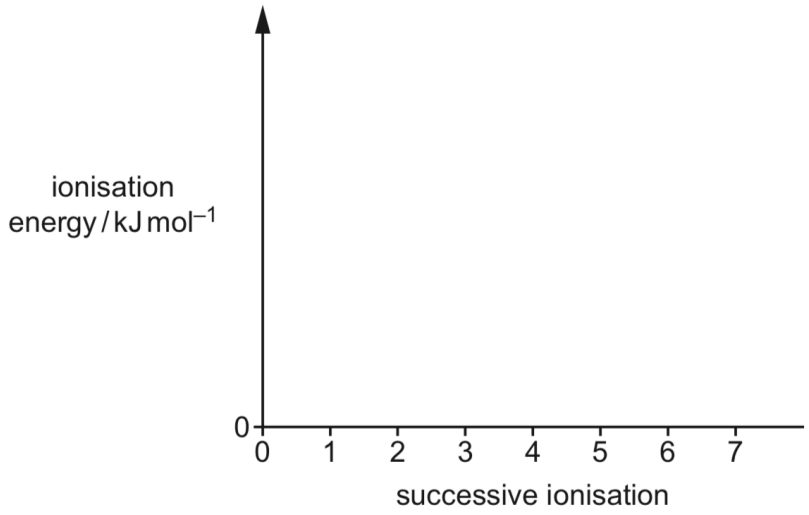


Fig. 7.1

[2]

8 The elements phosphorus, sulfur and chlorine are in Period 3 of the Periodic Table.
Table 8.1 shows some properties of the elements P to C*l*.

The first ionisation energy of S is **not** shown.

Table 8.1

property	P	S	C*l*
number of electrons in 3p subshell			
total number of unpaired electrons			
first ionisation energy / kJ mol^{-1}	1060		1260
formula of most common anion	P^{3-}	S^{2-}	C*l*$^-$

(a) (i) Complete Table 8.1 to show the number of electrons in the 3p subshell and the total number of unpaired electrons in an atom of P, S, C*l*. [2]

(ii) Construct an equation to represent the first ionisation energy of P.

..... [1]

(iii) Three possible values for the first ionisation energy of S are given.

1000 kJ mol^{-1} **1160 kJ mol^{-1}** **1320 kJ mol^{-1}**

Circle the correct value.

Explain your choice by comparing your chosen value to those of P and C*l*.

.....

.....

.....

.....

.....

..... [4]

(b) P^{3-}, S^{2-} and C*l*$^-$ have the same number of electrons.

(i) Give the full electronic configuration of P^{3-}.

..... [1]

(ii) State the trend in ionic radius shown by P^{3-}, S^{2-} and C*l*$^-$.

Explain your answer.

.....

.....

.....

..... [2]