1.4 IONISATION ENERGY

Learning outcomes
Candidates should be able to:
1. Define and use the term first ionisation energy, IE
2. Understand that ionisation energies are due to the attraction between the nucleus and the outer electron
3. Identify and explain the trends in ionisation energies across a period and down a group of the Periodic Table
4. Construct equations to represent first, second and subsequent ionisation energies
5. Identify and explain the variation in successive ionisation energies of an element
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

First ionisation energy

'The first ionisation energy is the **energy required to remove one** mole of **electrons from one** mole of the **gaseous atom** to form 1 mole of gaseous ions.'

e.g., Na (g) → Na^+(g) + e^-

Exothermic or endothermic?

What is this energy required for?

Unit ? = kJ/mol or $KJmol^{-1}$

1. Which equation represents the first ionisation energy of iodine?

   A 1/2 I_2(g) + e^- → I^-(g)
   B I(g) + e^- → I^-(g)
   C 1/2 I_2(g) → I^+(g) + e^-
   D I(g) → I^+(g) + e^-

2. Which equation represents the first ionisation energy of chlorine?

   A 1/2 Cl_2(g) + e^- → Cl^-(g)

   B Cl(g) + e^- → Cl^-(g)

   C 1/2 Cl_2(g) → Cl^+(g) + e^-

   D Cl(g) → C^+(g) + e^-

3. Which equation has an energy change that is equal to the first ionisation energy of bromine?

   A Br(g) → Br^+(g) + e^-
   B Br(g) → Br^-(g) – e^-
   C ½ Br_2(g) → Br^+(g) + e^-
   D 1/2 Br_2(g) → Br^-(g) – e^-

1. Define the first ionisation energy. [2]
2. Explain what is meant by the term first ionization energy? [3]

3. Chlorine has the highest first ionisation energy of the Period 3 elements Na to Cl.
Construct an equation for the first ionisation energy of chlorine.
Include state symbols. [1]

4. Write an equation to show the first ionisation energy of silicon. [1]

Successive ionisation energies

It is possible to remove not only the last (outermost) electron but practically all the electrons from elements. But this requires a successively higher amount of energy.

Can you construct equations to show three successive IE of Mg atom?

$Mg(g) \rightarrow Mg^{+}(g) + e^{-}$ first ionization energy

$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-}$ second ionization energy

$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-}$ third ionization energy

What is the trend of the successive ionization of an element?

Element	I IE	II IE	III IE	IV IE	All units are kJ mol^{-1}
Na	494	4560	6940	9540	
Mg	736	1450	7740	10500	
Al	577	1820	2740	11600	

What could be the reason for increasing successive ionization energies?

- Takes more energy to release electron from positively charged ion.

What could be the reason for huge difference between first ionization energy and second ionization energy of Na?

- Electrons being released from the inner shell/nearer to the nucleus
- The stronger force of attraction between oppositely charged species at shorter distance.
- Less **shielding effect**

5. Construct an equation for the second ionisation energy of argon. [1]

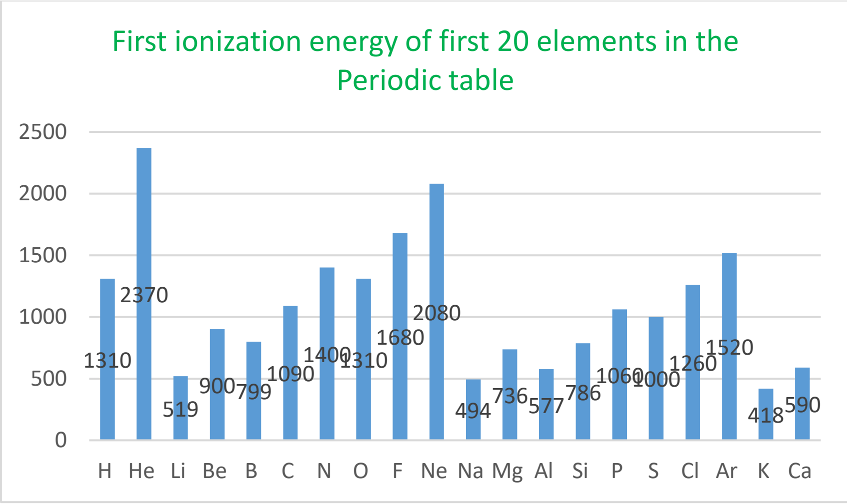
6. The Pauling electronegativity values of elements can be used to predict the chemical properties of compounds. Use the information in Table 1.1 to answer the following questions.

Table 1.1

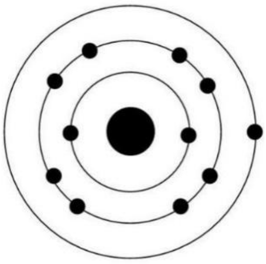
element	H	Li	C	O	S
Pauling electronegativity value	2.1	1.0	2.5	3.5	2.6
first ionisation energy/kJmol^{-1}	1310	519	1090	1310	1000
second ionisation energy/kJmol^{-1}	—	7300	2350	3390	2260

(i) Write an equation that represents the first ionisation energy of H. **[1]**

(ii) Explain why there is no information given in Table 1.1 for the second ionisation energy of H. [1]

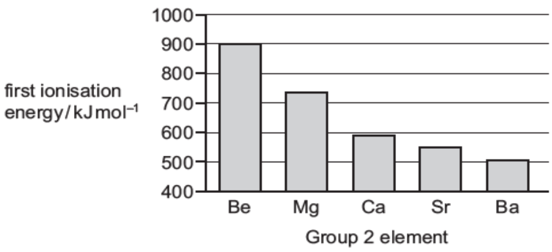


Activity: Consider the electronic structure of an element given below, and state if the following statements are true or false. Discuss in pairs and explain to the class.



1. Energy is required to remove an electron from the atom.
2. After the atom is ionised, it then requires more energy to remove a second electron because the second electron is nearer the nucleus.
3. The atom will spontaneously lose one electron to become stable.
4. Only one electron can be removed from the atom, as it then has a stable electronic configuration.
5. The nucleus is not attracted to the electrons.
6. After the atom is ionised, it then requires more energy to remove a second electron because the second electron experiences less shielding from the nucleus.
7. The nucleus is attracted towards the outermost electron less than it is attracted towards the other electrons.
8. After the atom is ionised, it then requires more energy to remove a second electron because it experiences a greater core charge than the first.

7. The graph shows the first ionisation energies of some of the elements in Group 2.



(i) Explain the observed trend in first ionisation energies down Group 2. [3]

(ii) The second ionisation energy of Be is 1757 kJ mol^{-1}.
Explain why the second ionisation energy of Be is higher than the first ionisation energy of Be. [1]

Nuclear charge: The bigger the nuclear charge, the bigger the attraction between the nucleus and the electrons, hence the bigger ionisation energy.
Distance of outer electrons from the nucleus: the further the electrons from the nucleus, the smaller the attraction of the nucleus and electrons, hence smaller ionisation energy.
Spin pair repulsion: Electrons in the same atomic orbital in a sub-shell repel each other more than electrons in different atomic orbitals.
This increased repulsion makes it easier to remove an electron. So, the first ionisation energy is decreased.
Shielding: The ability of inner shell electrons to reduce the effect of the nuclear charge on outer shell electrons. The electrons, being negatively charged, repel each other. Electrons in the full inner shells repel outer electrons and prevent the attraction of the full nuclear charge. This is called shielding.

The variation in the values of first ionisation energy is a **periodic property** because when plotted against atomic number, it shows a repeating pattern. The same goes for atomic and ionic radii.

8. Explain the general increase in the first ionisation energies of the Period 3 elements. [2]

9. Group 2 elements share common chemical properties.
Calcium reacts in cold water more quickly than magnesium because more energy is required to remove the outer electrons in magnesium. This occurs even though calcium atoms have a greater nuclear charge.
Explain why more energy is required to remove the outer electrons in magnesium than in calcium. [2]

10. Suggest how the value for the first ionisation energy of ^{54}Fe compares to the first ionisation energy of ^{56}Fe. Explain your answer in terms of the factors that affect ionisation energy.

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.... [4]

4. Which factor causes helium to have a higher first ionisation energy than hydrogen?

A In the 1s orbital in helium, electrons are paired.
B The lowest energy level in helium is filled.
C The nuclear charge in helium is higher than in hydrogen.
D There is less shielding of the outer shell in helium.

5. 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.

6. This question is about the first ionisation energies of magnesium and neon.
Which row is correct?

	first ionisation energy	type of electron removed	
		From Mg	from Ne
A	Mg > Ne	P	s
B	Mg > Ne	s	p
C	Ne > Mg	p	s
D	Ne > Mg	s	p

8. 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.

9. 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

11. 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]

12. The elements phosphorus, sulfur and chlorine are in Period 3 of the Periodic Table.
Table 1.1 shows some properties of the elements P to Cl.
The first ionisation energy of S is not shown.

Table 1.1

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

(i) Complete Table 1.1 to show the number of electrons in the 3p subshell and the total number of unpaired electrons in an atom of P, S and Cl. [2]

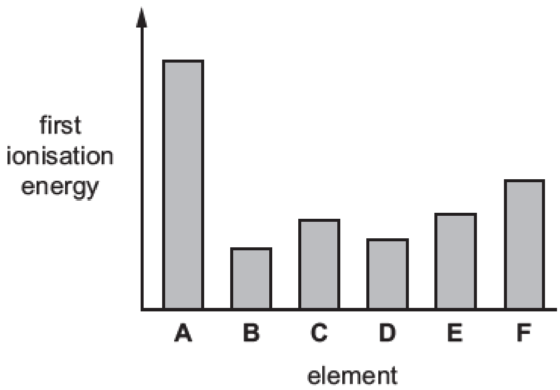
(ii) Construct an equation to represent the first ionisation energy of P. [1]

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

1000kJmol^{-1}	1160kJmol^{-1}	1320kJmol^{-1}

Circle the correct value.
Explain your choice by comparing your chosen value to those of P and Cl. [4]

13. The graph shows a sketch of the first ionisation energies of six successive elements in the Periodic Table.
The letters are **not** the symbols of the elements.



Suggest why the first ionisation energy of B is much less than that of A. [3]

14. Fig. 1.1 shows how **first** ionisation energies vary across Period 2.

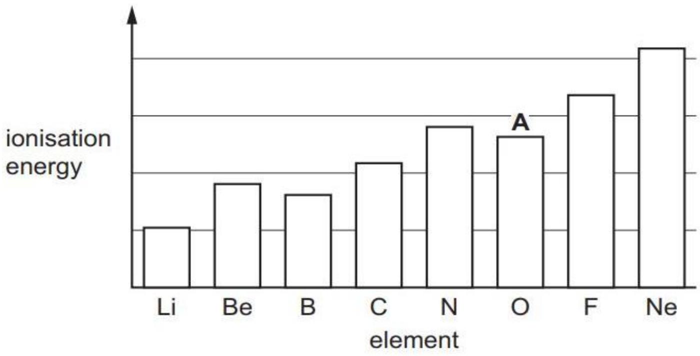


Fig. 1.1

(i) Construct an equation to represent the first ionisation energy of oxygen.
Include state symbols. [1]

(ii) State and explain the general trend in first ionisation energies across Period 2. [3]

(iii) Explain why ionisation energy A in Fig. 1.1 does not follow the general trend in first ionisation energies across Period 2. [2]

15. 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.
Explain why the first ionisation energy of sulfur is less than the first ionisation energy of phosphorus. [2]

10. Nitrogen has a higher first ionisation energy than oxygen.
Which statement explains this observation?

A The radius of an oxygen atom is smaller.
B An oxygen atom has more electron shells occupied.
C Oxygen has paired electrons in the 2p sub-shell.
D An oxygen atom has more protons in the nucleus.

11. Why is the first ionisation energy of oxygen less than that of nitrogen?

A The nitrogen atom has its outer electron in a different subshell.
B The nuclear charge on the oxygen atom is greater than that on the nitrogen atom.
C The oxygen atom has a pair of electrons in one p orbital that repel one another.
D There is more shielding in an oxygen atom.

12. 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.

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

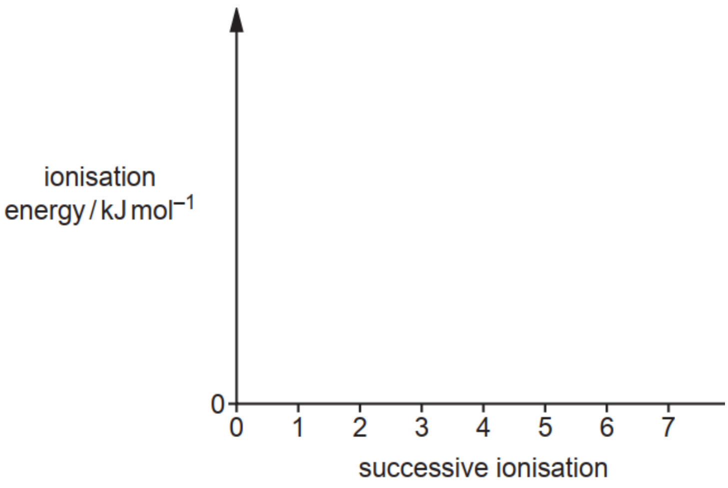
14. The second ionisation energy of oxygen is greater than the second ionisation energy of fluorine.
Which factor explains this difference?

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.

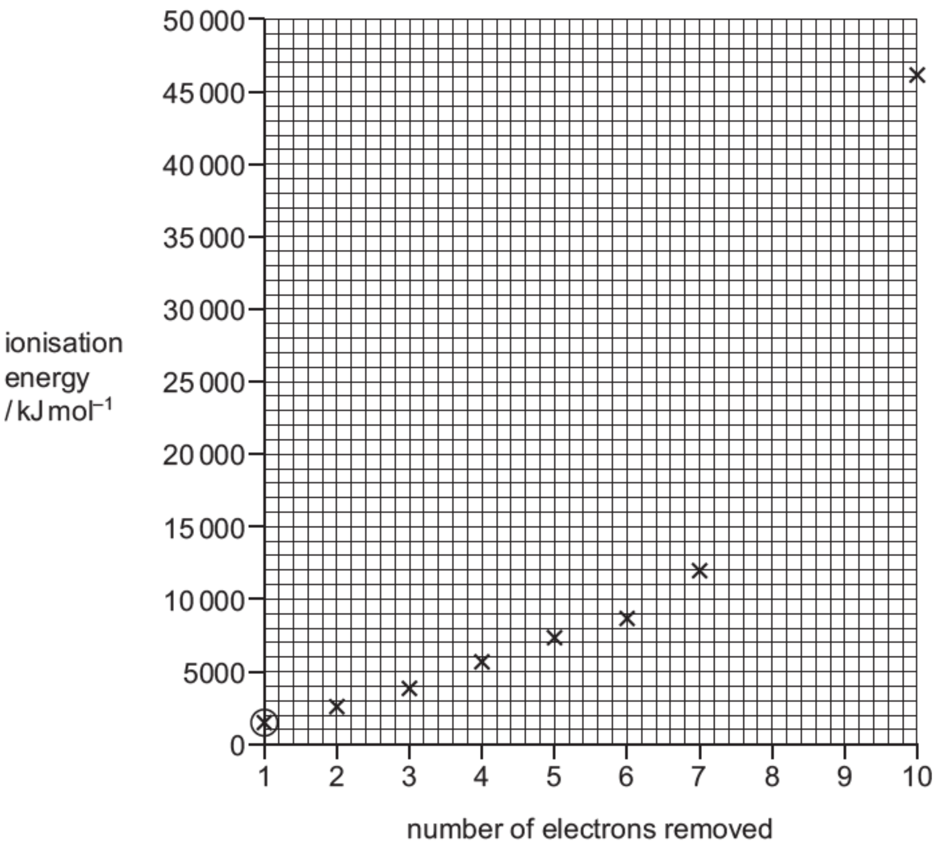
16. Tellurium is an element in Group 16. The most common isotope of tellurium is ^{130}Te. Its electronic configuration is [Kr] $4d^{10}\ 5s^2\ 5p^4$.
(i) Construct an equation to represent the first ionisation energy of Te. [1]

(ii) The radius of Te ions decreases after each successive ionisation.
State **two** factors that are responsible for the increase in the first six ionisation energies of Te. [2]

(iii) Sketch a graph in Fig. 1.1 to show the trend in the first **seven** ionisation energies of Te.



17. The graph shows successive ionisation energies for the element argon.
    (i) Complete the graph with predictions for the eighth and ninth ionisation energies of argon. Use a cross (x) for each data point. [2]



(ii) The energy value required to remove the first electron from an atom of argon is circled on the graph. Sketch the shape of the orbital that contains this electron. [1]

15. Which process has the largest enthalpy change per mole?

A $Al^{3+}(g) \rightarrow Al^{4+}(g) + e^-$
B $P^{5+}(g) \rightarrow P^{6+}(g) + e^-$
C $S^{6+}(g) \rightarrow S^{7+}(g) + e^-$
D $Si^{4+}(g) \rightarrow Si^{5+}(g) + e^-$

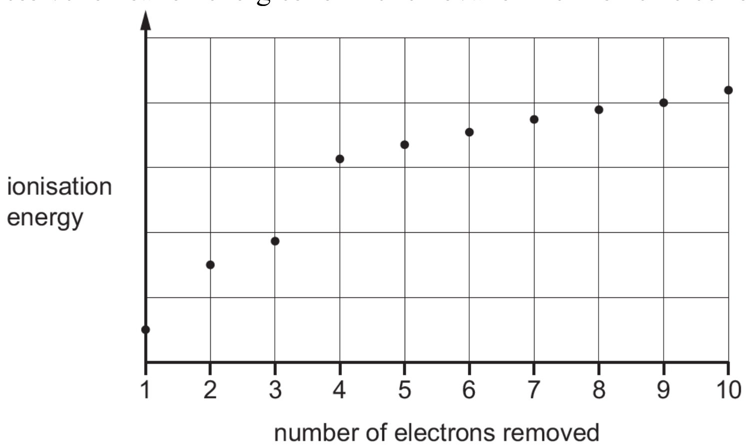
Deducing the position of an element in the Periodic Table using successive ionisation energy data

16. 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

18. The first, second and third ionisation energies of potassium are 418, 3070 and 4600 $kJmol^{-1}$, respectively. Use this information to explain why potassium is in Group 1. [2]

19. Element **A** is in the p block.
The graph shows the successive ionisation energies for the removal of the first ten electrons of **A**.



State and explain the group of the Periodic Table that element A belongs to. [2]

The diagram shows the logarithm of the first 13 ionisation energies of an element.
Which statement is correct?

A A proton is lost for each successive ionisation energy.
B The element is aluminium.
C The element is silicon.
D The element has only one outer electron.

20. Successive ionisation energies for element **A** are shown in the following table.

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy/kJmol^{-1}	1310	3390	5320	7450	11000	13300	71000	84100

Use the above table to deduce the group of the Periodic Table that **A** belongs to. Explain your answer. [1]

21. Element E is in Period 3 of the Periodic Table.
The first eight ionisation energy values of E are shown in Table 1.1.

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy/kJmol^{-1}	577	1820	2740	11600	14800	18400	23400	27500

Deduce the full electronic configuration of **E**.
Explain your answer. [3]

17. Which of these elements has the highest fifth ionisation energy?

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

18. The table shows the first five ionisation energies of element X.

Element	ionisation energy / kJ mol^{-1}				
	1st	2nd	3rd	4th	5th
X	736	1450	7740	10 500	13 600

Element X is in Period 3 of the Periodic Table.
What is element X?

A sodium B magnesium C silicon D argon

19. The first seven ionisation energies of an element between lithium and neon in the Periodic Table are shown.

1310 3390 5320 7450 11 000 13 300 71 000 kJ mol^{-1}

What is the outer electronic configuration of the element?

A $2s^2$ **B** $2s^22p^1$ **C** $2s^22p^4$ **D** $2s^22p^6$

20. X and Y are elements from the same group of the Periodic Table.
The 5th to 9th ionisation energies for X and Y are shown.

	ionisation energy / kJ mol^{-1}				
	5th	6th	7th	8th	9th
element X	11 020	15 160	17 870	92 040	106 437
element Y	6 540	9 360	11 020	33 360	38 600

Which row identifies elements X and Y?

	element X	element Y
A	argon	neon
B	chlorine	fluorine
C	fluorine	chlorine
D	neon	argon

21. Elements X and Y are in the same group of the Periodic Table.
The table shows the first six ionisation energies of X and Y in kJ mol–1.

	1st	2nd	3rd	4th	5th	6th
X	800	1600	2400	4300	5400	10 400
Y	1000	1800	2700	4800	6000	12 300

What could be the identities of X and Y?

	X	Y
A	antimony, Sb	arsenic, As
B	arsenic, As	antimony, Sb
C	selenium, Se	tellurium, Te
D	tellurium, Te	selenium, Se

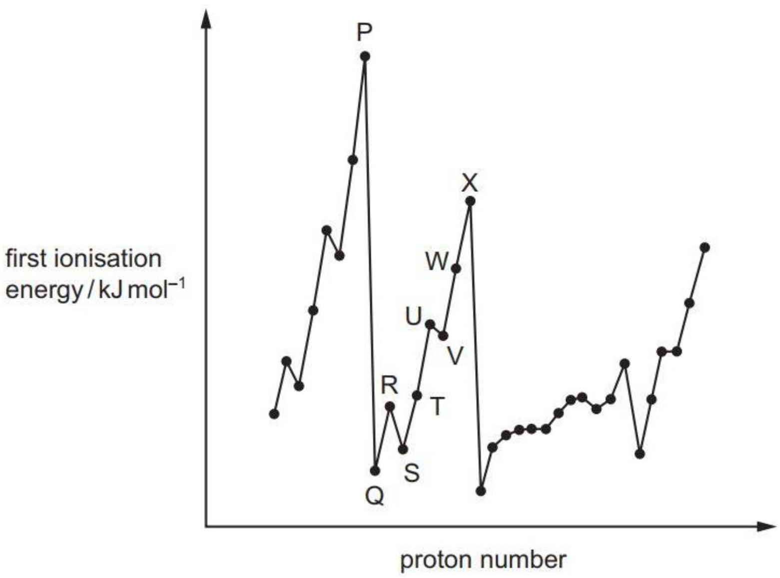
22. The data in the table gives the 5th to the 10th ionisation energies of three elements from Period 3 of the Periodic Table.

Element	ionisation energy / kJ mol^{-1}					
	5th	6th	7th	8th	9th	10th
X	6274	21 269	25 398	29 855	35 868	40 960
Y	7012	8 496	27 107	31 671	36 579	43 140
Z	6542	9 362	11 018	33 606	38 601	43 963

What are the correct identities of these three elements?

	element X	element Y	element Z
A	Na	Mg	A*l*
B	Mg	A	Si
C	P	S	C*l*
D	S	C*l*	Ar

23. 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.

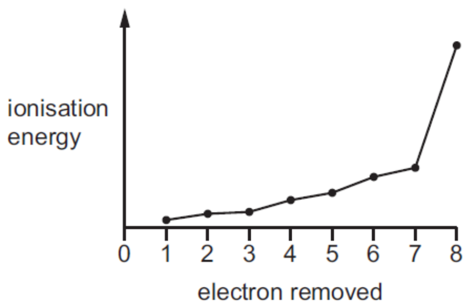
24. Four successive ionisation energies (IE) of element E are shown.
Element E is in Period 3 of the Periodic Table.

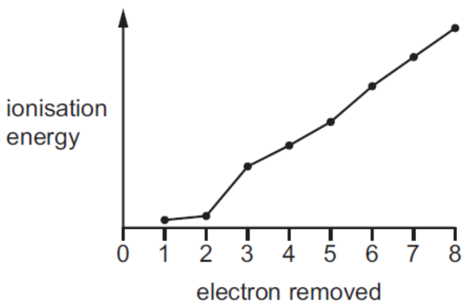
fifth IE / kJ mol^{-1}	sixth IE / kJ mol^{-1}	seventh IE / kJ mol^{-1}	eighth IE / kJ mol^{-1}
16000	20000	24000	29000

In which group of the Periodic Table is E?

A 14 **B** 15 **C** 16 **D** 17

25. The first eight successive ionisation energies for two elements of Period 3 of the Periodic Table are shown in the graphs.

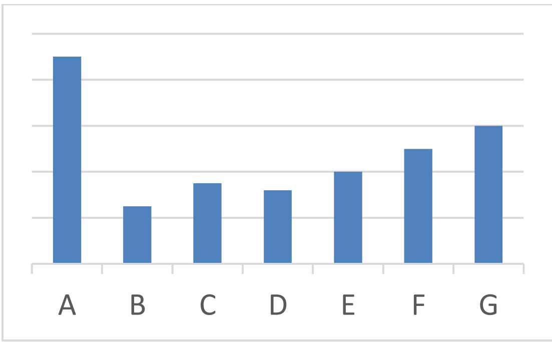




What is the formula of the ionic compound formed from these elements?

A $MgCl_2$ B $CaBr_2$ C Na_2S D K_2Se

22. The bar chart below shows the first ionisation energies of 7 successive elements in the periodic table.
The letters are not the symbols of the elements
▶ Suggest why the first IE of B is much less than that of A.



*****This is not a complete note. It is to guide you. It is recommended to study the prescribed textbooks along with this material. *****