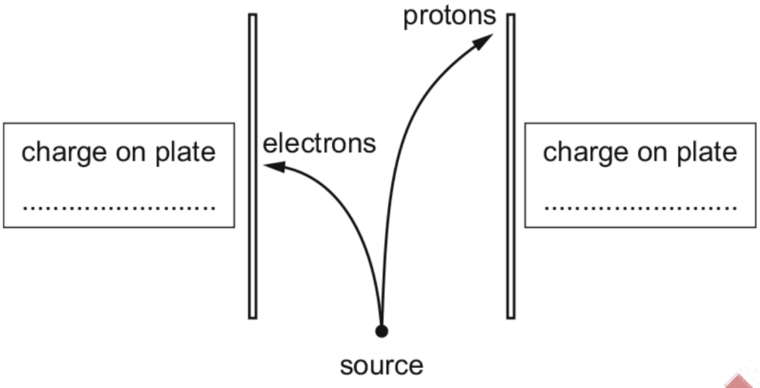
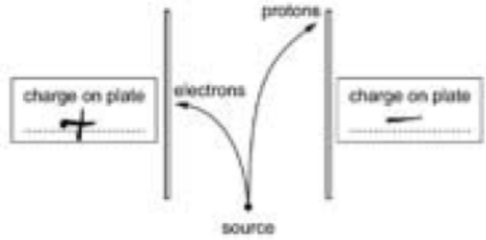
1.

The diagram shows the behaviour of separate beams of electrons and protons in an electric field.



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

positive / + on left **AND** negative / – on right



2.

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

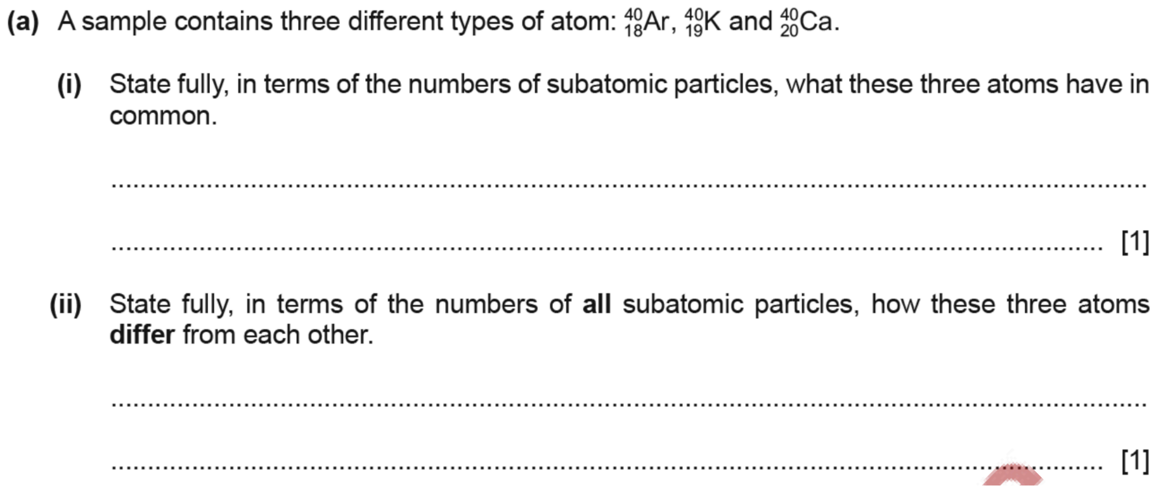
State, in terms of subatomic particles in the nucleus, what is the same about each of these ions and what is different.

same ....

different ....

NOTE: the question specifies subatomic particles <u>in the nucleus</u>. So don't write 'same number of electrons'!

3.



(a) A sample contains three different types of atom: $^{40}_{18}Ar$, $^{40}_{19}K$ and $^{40}_{20}Ca$.

(i) State fully, in terms of the numbers of subatomic particles, what these three atoms have in common.

....

.... [1]

(ii) State fully, in terms of the numbers of **all** subatomic particles, how these three atoms **differ** from each other.

....

.... [1]

- They have the same nucleon number OR the same sum/ total number of protons + neutrons
- They have different numbers of protons, neutrons **and** electrons.

4. State the trend in ionic radius shown by P^{3-}, S^{2-} and Cl^{-}, and explain your answer.
   - Ionic radius decreases from P^{3-} to Cl^{-}.
   - Explanation:
     - increased attraction (of outer e−) for nucleus
     - increased nuclear charge
     - same shielding

5. Define *relative isotopic mass*
   - mass of an atom of an isotope
   - compared to the mass of the unified atomic mass unit (amu) OR on a scale on which a carbon-12 atom has a mass of exactly 12 units OR divided by 1/12 mass of a carbon-12 atom

6. Define *relative atomic mass*
   - mean/average mass of the isotopes of an atom
   - relative to 1/12 of the mass of an atom of ^{12}C.

7. Define the *unified atomic mass unit*.
   - (the mass of) one twelfth of the mass of a C-12 / ^{12}C atom (or isotope)

8. Define relative atomic mass, Ar , in terms of the unified atomic mass unit.
   - average mass of the isotopes of an element compared to the unified atomic mass unit.

9. Calculate the relative isotopic mass of ^{11}B, giving answer to 6s.f.

isotope	relative isotopic mass	abundance / %
^{10}B	10.0129	19.78
^{11}B	to be calculated	80.22

Use the relative atomic mass given in the periodic table = 10.8

$$\frac{(10.0129 \times 19.78) + (80.22x)}{100} = 10.8$$

$x = 10.9941$

10.

A sample of copper contains two isotopes, ^{63}Cu and ^{65}Cu. The relative atomic mass of the copper in this sample is 63.55.

Calculate the percentage abundance of each of these isotopes. Show your working.

% abundance of ^{63}Cu = 72.5%
% abundance of ^{65}Cu = 27.5%
M1 *correct algebraic expression AND correct calculation of x for one isotope*
% ab of $^{63}Cu = x$ $(x/100 \times 63) + ((1- x)/100 \times 65) = 63.55$ so $x = 72.5$
OR
% ab of $^{65}Cu = x$ $(1- x)/100 \times 63) + x/100 \times 65) = 63.55$ so $x = 27.5$

M2 *calculation of abundance of other isotope by 100- x*

11.

A sample of strontium exists as a mixture of four isotopes. Information about three of these isotopes is given in the table.

mass number	86	87	88
abundance	9.86%	7.00%	82.58%

The relative atomic mass of this sample of strontium is 87.71.

Calculate the mass number of the fourth isotope.

$$\frac{(A \times 0.56) + (86 \times 9.86) + (87 \times 7.00) + (88 \times 82.58)}{100} = 87.71$$

$A = 84$

NOTE: the mass number will be given as a whole number.

12.

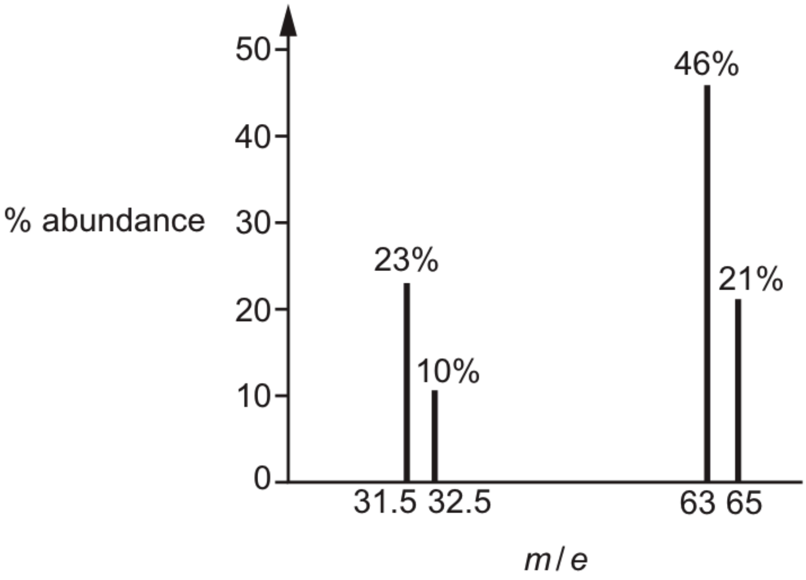
isotope	relative isotopic mass	% abundance
^{63}Cu	62.930	69.15
^{65}Cu	64.928	30.85

Calculate the Ar of copper

$(62.93 \times 69.15) \div 100 + (64.928 \times 30.85) \div 100 = 63.55 / 63.546$

13.

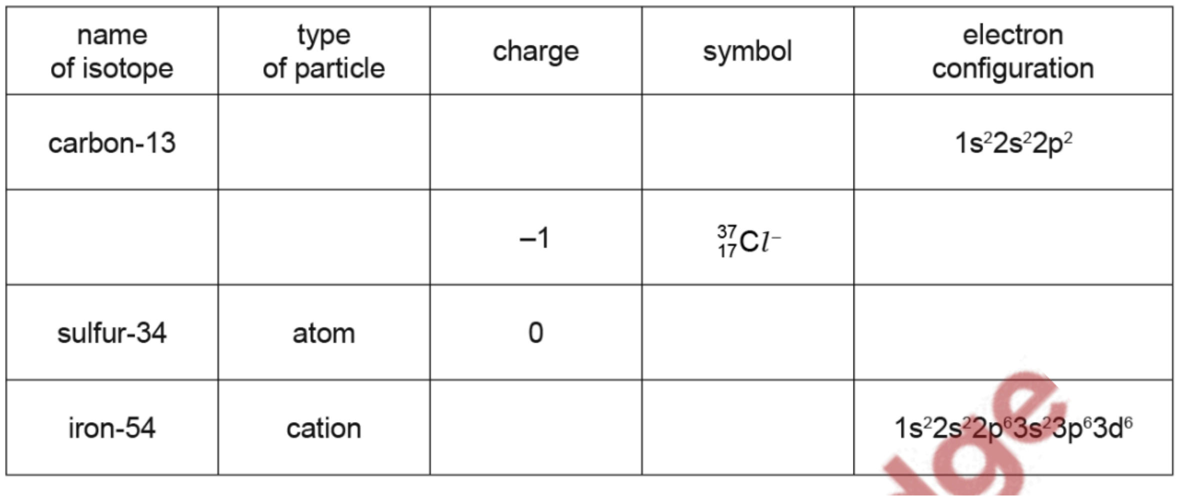
The mass spectrum of a sample of pure copper is shown in Fig. 1.1.



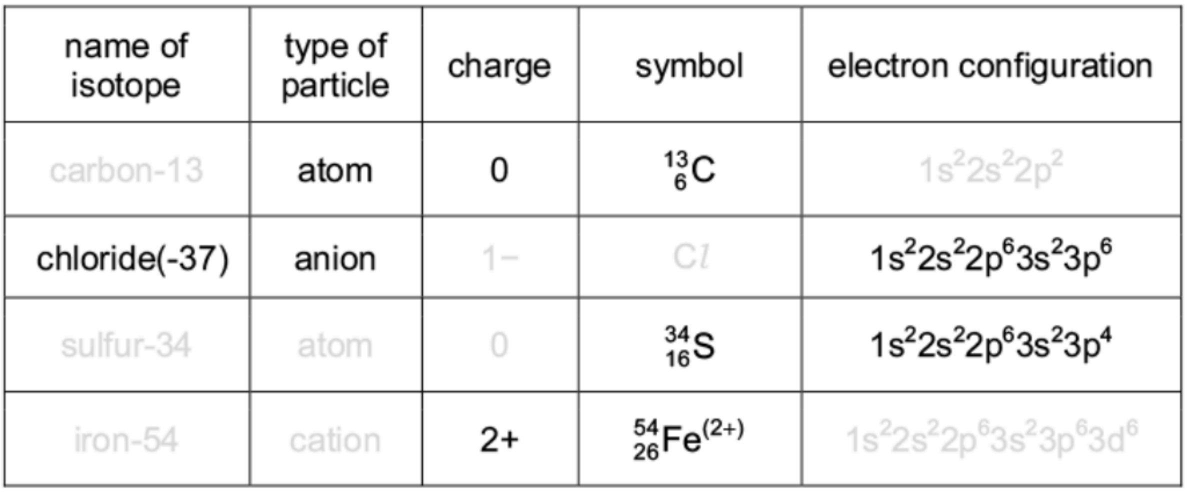
Identify the ion with an abundance of 23% in the sample.

$^{63}Cu^{2+}$

14. Fill in the table:



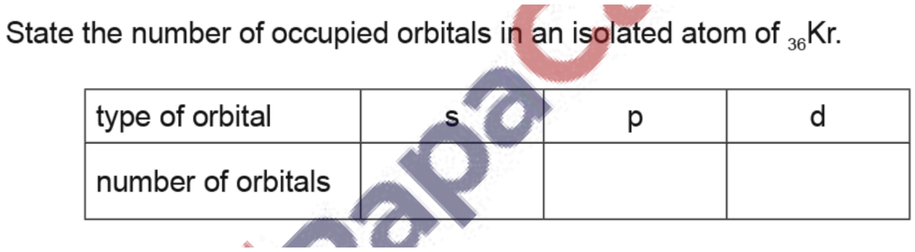
name of isotope	type of particle	charge	symbol	electron configuration
carbon-13				$1s^2 2s^2 2p^2$
		–1	$^{37}_{17}Cl^-$	
sulfur-34	atom	0		
iron-54	cation			$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$



name of isotope	type of particle	charge	symbol	electron configuration
carbon-13	atom	0	$^{13}_{6}C$	$1s^2 2s^2 2p^2$
chloride(-37)	anion	1–	Cl	$1s^2 2s^2 2p^6 3s^2 3p^6$
sulfur-34	atom	0	$^{34}_{16}S$	$1s^2 2s^2 2p^6 3s^2 3p^4$
iron-54	cation	2+	$^{54}_{26}Fe^{(2+)}$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$

NOTE: chloride–37 NOT chlorine–37!!

15.



State the number of occupied orbitals in an isolated atom of $_{36}Kr$.

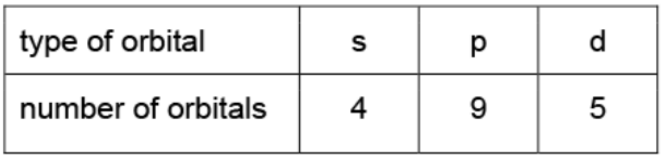
type of orbital	s	p	d
number of orbitals			

Electronic configuration of Kr:
1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6

s subshell = 1 orbital
p subshell = 3 orbitals
d subshell = 5 orbitals

4 s subshells = 4 s orbitals

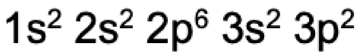
3 p subshells = 3 x 3 = 9 orbitals
1 d subshell = 1 x 5 = 5 orbitals



type of orbital	s	p	d
number of orbitals	4	9	5

16.

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



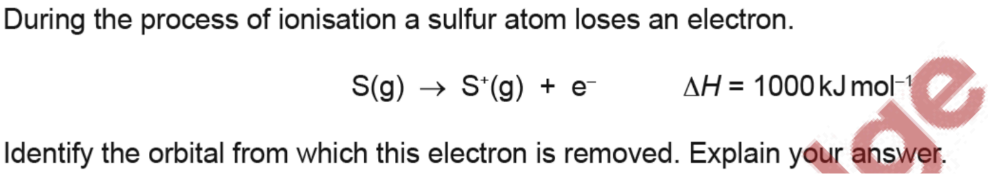
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

CAREFUL!! S2+ NOT S2–. 2 electrons are lost!

17. Explain what is meant by the term *first ionisation energy*.
    - The amount of energy required to remove one electron / energy change when one electron is removed
    - From each atom in one mole
    - Of gaseous atoms

18. Suggest why first ionisation energies of elements in first row of d-block are similar. For all these elements, a 4s electron is lost during first ionisation.
    - Nuclear charge increases
    - Extra electron(s) in inner shell / n=3 / d-subshell / d-orbital
    - Increased shielding (of 4s electrons by electrons in n=3/ 3rd shell/ 3d)
    - (overall) similar nuclear attraction (for outer electron)

19.



During the process of ionisation a sulfur atom loses an electron.

$$S(g) \rightarrow S^{+}(g) + e^{-} \qquad \Delta H = 1000\,kJ\,mol^{-1}$$

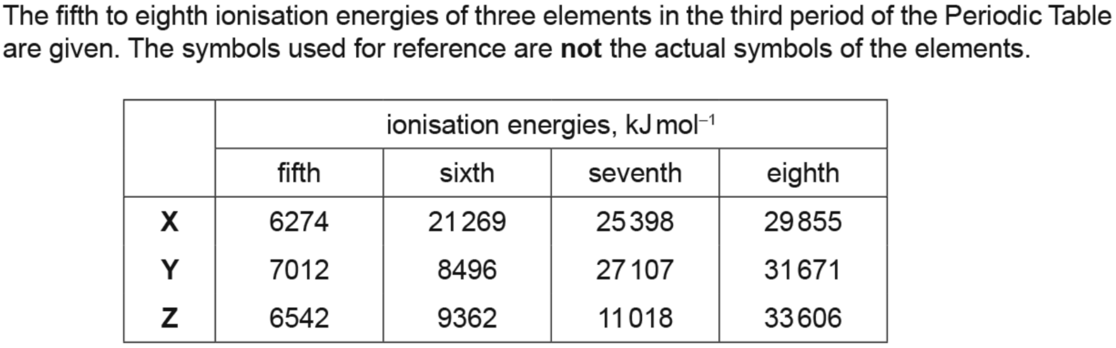
Identify the orbital from which this electron is removed. Explain your answer.

- Orbital: 3p
- Explanation: it is less attracted to the nucleus (so takes less energy to lose) OR it is the highest energy orbital (which is occupied) / it is in the highest energy orbital.

20. Explain why first ionisation energy of sulphur is less than that of phosphorus
    - In S, the electron is removed from a 3p orbital/ subshell which has a pair of electrons.

- Paired electrons repel.

21.



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

	ionisation energies, kJ mol^{-1}			
	fifth	sixth	seventh	eighth
X	6274	21269	25398	29855
Y	7012	8496	27107	31671
Z	6542	9362	11018	33606

a. State & explain group number of element Y

- Group number: 16 / 6
- Explanation: big increase/ big difference/ big jump/ big gap after 6th IE

b. Explain why first ionisation energy of element Y is less than of element X.

- Electron pair repulsion
- Y has a pair of electrons in a 3p orbital / a 3p orbital is full

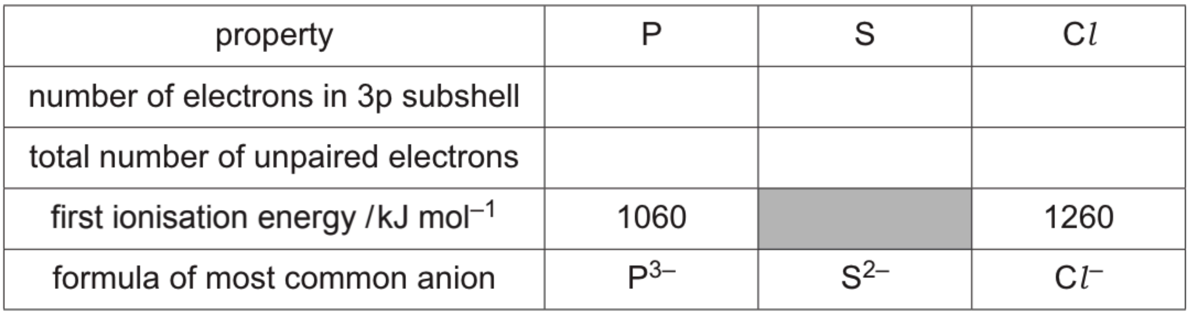
22. State and explain the general trend in first ionisation energies across the third period.
    - Increases across period due to increasing nuclear attraction of nucleus for electrons.
    - Due to increasing nuclear charge / number of protons in the nucleus
    - Similar shielding of outer electrons

23. Equation for the first ionisation energy of element X:

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

DO NOT FORGET to include state symbols!!

24.



property	P	S	Cl
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-}	Cl^{-}

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 Cl.

- Correct value = 1000 kJ mol^{-1}
- Explanation:
  - Nuclear attraction for Cl is stronger compared to S and P.
  - S has less nuclear charge than Cl
  - S has a greater nuclear charge than P BUT S has two electrons in 3p orbital AND resulting in spin-pair repulsion

25.

property	Si	P	S
total number of electrons in s subshells			
total number of electrons in p subshells			
first ionisation energy / kJ mol^{-1}	786	▨	1000
formula of most common chloride	$SiCl_4$	PCl_5	SCl_2

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

619 kJ mol^{-1} **893 kJ mol^{-1}** **1060 kJ mol^{-1}**

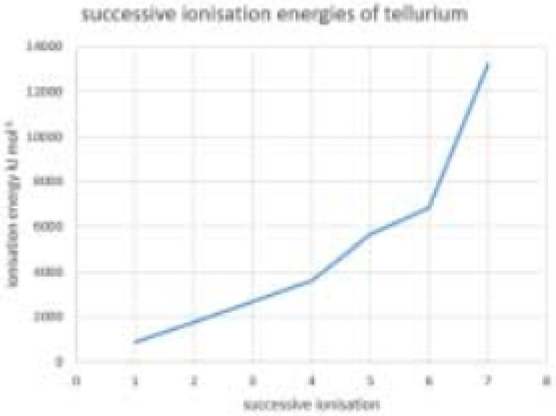
Circle the correct value.

Explain your choice, including a comparison of your chosen value to those of Si and S.

- 1060
- Explanation:
  - P has greater attraction of nucleus for outer electrons compared to Si
  - P has greater nuclear charge than Si
  - S has two electrons in a (3)p orbital AND resulting in spin-pair repulsion

26. The radius of Te ions decreases after each successive ionisation.
    Factors responsible for the increase in the first six ionisation energies of Te.
    - increased nuclear attraction / attraction for nucleus
    - decreased shielding by sub-shells (IE 4 to 5)
    - spin-pair repulsion (IE 1 to 2)

27. Sketch a graph to show the trend in the first seven ionisation energies of Te.

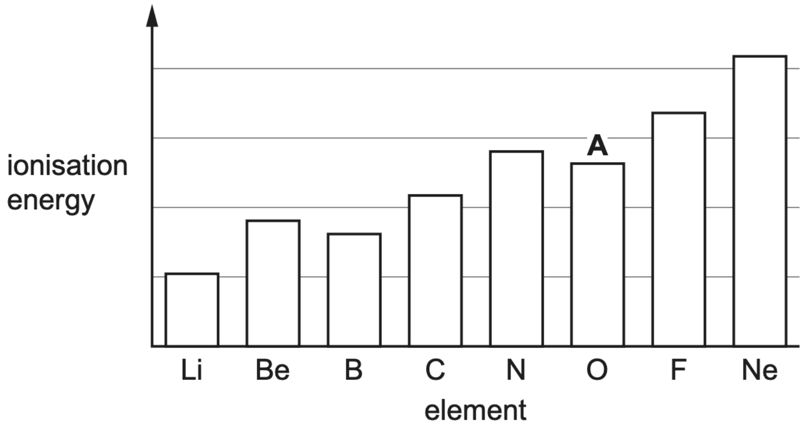


M1 general increase
M2 jumps from 4 to 5 and (largest) 6 to 7

28. Explain why the first ionisation energy of neon is greater than that of fluorine.
    - Greater nuclear charge/ number of protons
    - Same shielding/ number of shells/ energy level

29. Explain why H does not have 2nd ionisation energy
    - H cannot undergo 2nd ionisation because it has only 1 electron / H+ has no electrons

30.



Explain why ionisation energy A in Fig. 1.1 does not follow the general trend in first ionisation energies across Period 2.
- Spin pair repulsion of electrons in 2p orbital
- Outweighs increased nuclear charge

31. Explain what is meant by the term relative molecular mass.
- The mass of a molecule
- Relative to 1/12th the mass of an atom of carbon-12

1. State 1 similarity and 1 difference in properties of isotopes of an element. Explain your answer.
    - Identical chemical properties AND same electronic configuration
    - Different physical properties AND different number of neutrons

2. Define first ionisation energy
    - energy required when 1 electron is removed
    - from each atom in 1 mole of gaseous atoms

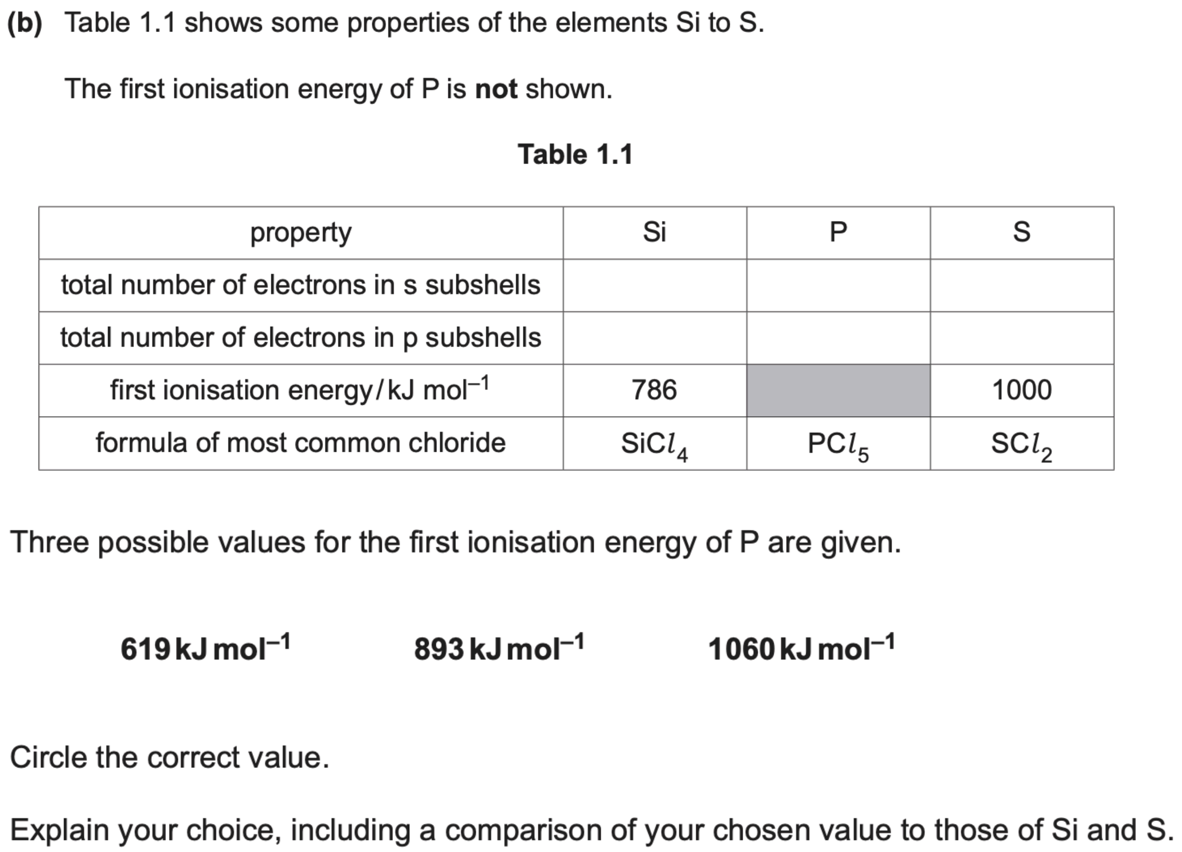
3. Explain why first ionisation energy of sulfur is less than first ionisation energy of phosphorus
    - spin pair repulsion in 3p orbital in S
    - outweighs increased nuclear charge (in S)

4. 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.
    - similarity in electronic structure: both remove electron from 3s OR Al^{2+} and Na have same electronic configuration OR shielding (of outer electron) is same
    - Al^{2+} has greater nuclear charge / greater number of protons (13p compared to 11p)

5. Define relative atomic mass
    - The average mass of the isotopes of an element
    - Compared to the mass of the unified atomic mass unit

6. The first, second and third ionisation energies of potassium are 418, 3070 and 4600kJmol–1, respectively. Use this information to explain why potassium is in Group 1.
    - big increase in IE between first and second
    - second (and third) electron(s) is removed from inner shell OR
      second (and third) electron(s) is removed from a shell closer to the nucleus OR second (and third) electron(s) has a stronger nuclear attraction

7.



(b) Table 1.1 shows some properties of the elements Si to S.

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

Table 1.1

property	Si	P	S
total number of electrons in s subshells			
total number of electrons in p subshells			
first ionisation energy / kJ mol^{-1}	786		1000
formula of most common chloride	$SiCl_4$	PCl_5	SCl_2

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

619 kJ mol^{-1} **893 kJ mol^{-1}** **1060 kJ mol^{-1}**

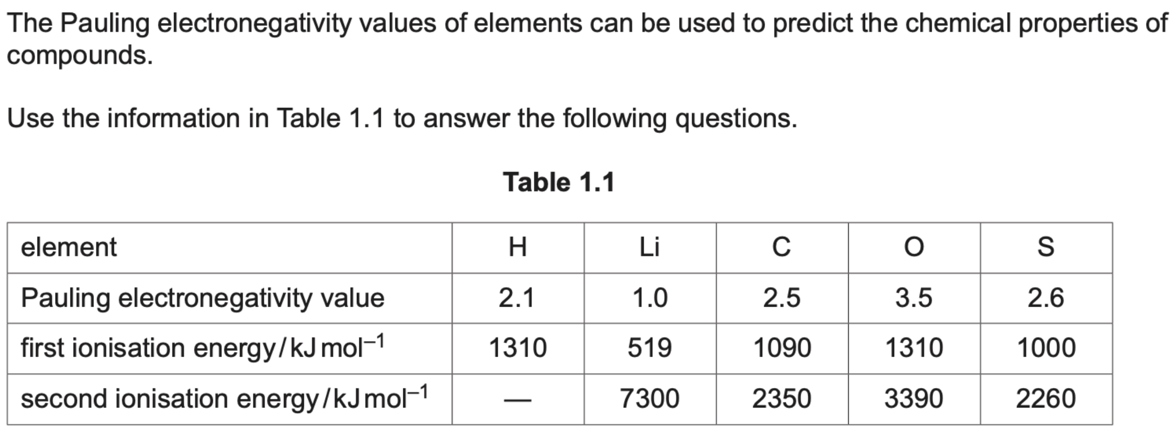
Circle the correct value.

Explain your choice, including a comparison of your chosen value to those of Si and S.

- 1060 (is circled)
- P / it has greater attraction of nucleus for outer electrons compared to Si
- P / it has greater nuclear charge than Si
- S has two electrons in a (3)p orbital AND resulting in spin-pair repulsion

8. Define electronegativity
   - Power of an atom to attract electrons to itself

9.



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 / kJ mol^{-1}	1310	519	1090	1310	1000
second ionisation energy / kJ mol^{-1}	—	7300	2350	3390	2260

O and S are in Group 16. Explain the difference in the Pauling electronegativity values of O and S.

- O lower nuclear charge / lower proton number
- O has one fewer shell than S / less shielding
- Greater attraction for nucleus in O

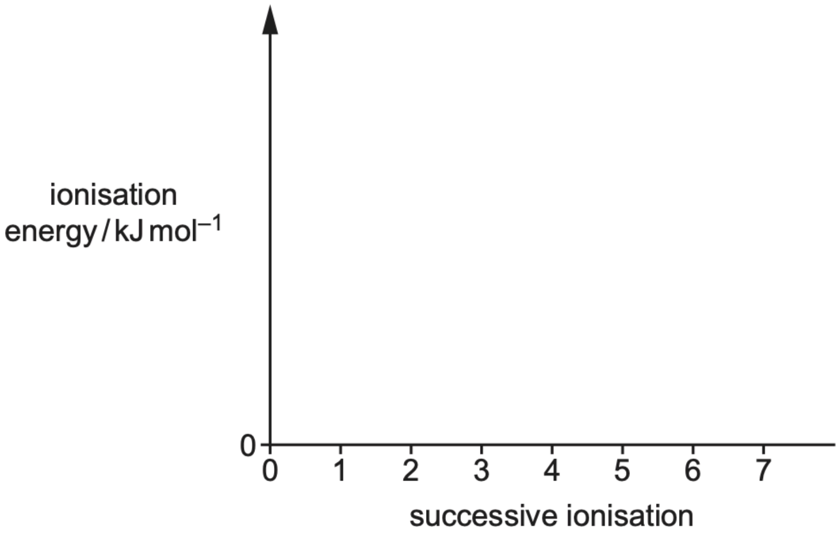
10. Explain why there is a general increase in first ionisation energies of the elements across Period 3.
    - similar shielding AND increase in proton no / atomic no / nuclear charge
    - increased nuclear attraction

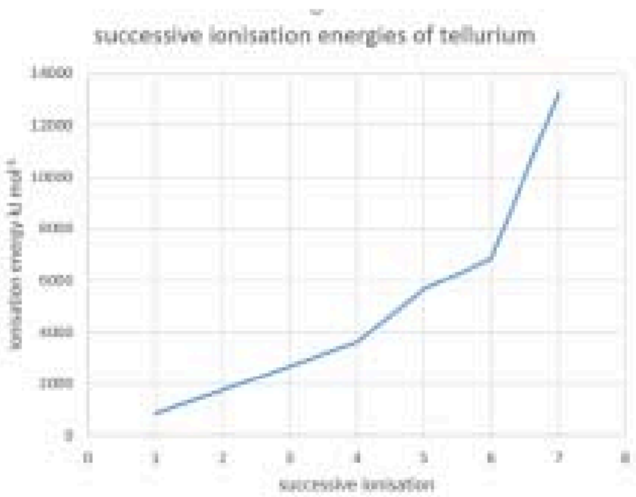
11. The radius of Te ions decreases after each successive ionisation.
    State 2 factors responsible for increase in the first six ionisation energies of Te.
    - Increased nuclear attraction
    - Decreased shielding by sub-shells (between IE 4 and 5)
    - Reduced spin pair repulsion (between IE 1 and 2)

12.

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

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





M1 general increase
M2 jumps from 4 to 5 and (largest) 6 to 7

- NOTE: Jump from 4 to 5 is because p subshell has only 4 electrons. After 4 electrons removed, electrons will then be removed from the s subshell. This can be derived from the electronic configuration given.

13. State the full electronic configuration of Cu2+.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9(4s^0)$