Structure 1.3 HL Answers

IB CHEMISTRY HL

Structure 1.3.6

Understandings:

- In an emission spectrum, the limit of convergence at higher frequency corresponds to ionisation.

Learning outcome(s):

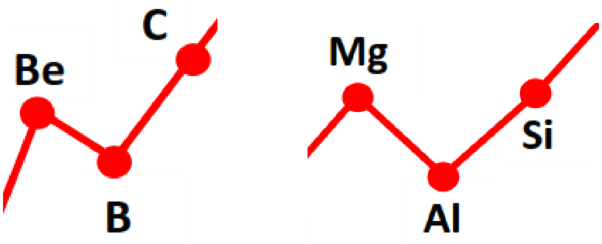
- Explain the trends and discontinuities in first ionisation energy (IE) across a period and down a group.
- Calculate the value of the first IE from spectral data that gives the wavelength or frequency of the convergence limit.

Additional notes:

- The value of the Planck constant h and the equations $E = hf$ and $c = \lambda f$ are given in the data booklet.

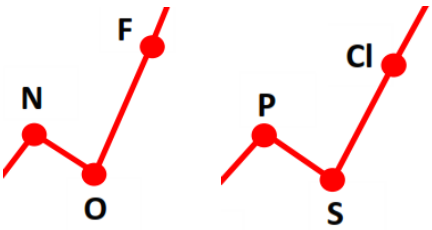
Discontinuities in first ionisation energy across a period

Be to B and Mg to Al

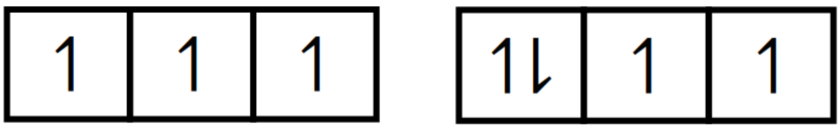


- Be has the electronic configuration $1s^2\ 2s^2$
- B has the electronic configuration $1s^2\ 2s^2\ 2p^1$
- Electrons in p orbitals are of higher energy and further from the nucleus than electrons in s orbitals, therefore they require less energy to remove.
- The same explanation can be applied for the drop in ionisation energy from Mg to Al, except that the electron configurations are $1s^2\ 2s^2\ 2p^6\ 3s^2$ and $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

N to O and P to S



- N has the electronic configuration $1s^2\ 2s^2\ 2p^3$
- has the electronic configuration $1s^2\ 2s^2\ 2p^4$
- For oxygen, the electron is removed from a doubly occupied p orbital. An electron in a doubly occupied orbital is repelled by the other electron and requires less energy to remove than an electron in a half-filled orbital.



Exercises:

1. Explain the reason for the decrease in first ionisation energy between Mg and Al.
   Mg has the electron configuration $1s^2 2s^2 2p^6 3s^2$, Al has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^1$. Electrons in p orbitals are of higher energy and further from the nucleus than electrons in s orbitals, therefore, they require less energy to remove.

2. Explain the reason for the decrease in first ionisation energy between P and S.
   P has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$. S has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^4$. For sulfur, the electron is removed from a doubly occupied p orbital. An electron in a doubly occupied orbital is repelled by the other electron and requires less energy to remove than an electron in a half-filled orbital.

Calculations involving $E = hv$

$$E = hv \qquad c = \lambda v$$

E is energy in Joules (J)
h is Planck's constant (6.63×10^{-34} J·s)
v is the frequency in s^{-1} (or Hertz, Hz)
c is the speed of light (3.00×10^{8} m s^{-1})
λ is the wavelength in m or nm (1 m = 1×10^{9} nm)

Exercises: use the formulas and constants above to answer the following questions.

1. What is the frequency, in s^{-1}, of a photon of light with an energy of 2.24×10^{-19} J?

   $E = hv$
   $2.24 \times 10^{-19} = 6.63 \times 10^{-34}\ v$
   $v = 2.24 \times 10^{-19} / 6.63 \times 10^{-34}$
   $v = 3.38 \times 10^{14}\ s^{-1}$

2. What is the wavelength, in m, of light with a frequency of $7.11 \times 10^{14}\ s^{-1}$?

   $c = \lambda v$
   $3.00 \times 10^{8} = \lambda\ 7.11 \times 10^{14}$
   $\lambda = 3.00 \times 10^{8} / 7.11 \times 10^{14}$
   $\lambda = 4.22 \times 10^{-7}$ m

3. A photon of light has a wavelength of 6.98×10^{-7} m. How much energy does it have in J?

   $c = \lambda v$
   $1.00 \times 10^{8} = 6.98 \times 10^{-7}\ v$
   $v = 3.00 \times 10^{8} / 6.98 \times 10^{-7}$
   $v = 4.30 \times 10^{14}\ s^{-1}$
   $E = hv$
   $E = 6.63 \times 10^{-34} \times 4.30 \times 10^{14}$
   $E = 2.85 \times 10^{-19}$ J

4. How much energy, in J, does a photon of light have if it has a wavelength of 5.26 $\times 10^{-7}$ m?

   $c = \lambda v$
   $3.00 \times 10^{8} = 5.26 \times 10^{-7}\ v$
   $v = 3.00 \times 10^{8} / 5.26 \times 10^{-7}$
   $v = 5.70 \times 10^{14}\ s^{-1}$
   $E = hv$
   $E = 6.63 \times 10^{-34} \times 4.30 \times 10^{14}$
   $E = 3.78 \times 10^{-19}$ J

5. What is the wavelength, in m, of a photon of light if it has an energy of 4.01×10^{-19} J?

$E = hv$
$4.01 \times 10^{-19} = 6.63 \times 10^{-34}\, v$
$v = 4.01 \times 10^{-19} / 6.63 \times 10^{-34}$
$v = 6.05 \times 10^{14}\ s^{-1}$
$c = \lambda v$
$3.00 \times 10^{8} = \lambda\ 6.05 \times 10^{14}\ s^{-1}$
$\lambda = 3.00 \times 10^{8} / 6.05 \times 10^{14}\ s^{-1}$
$\lambda = 4.96 \times 10^{-7}$ m

6. What is the wavelength, in m, of a photon of light with an energy of 1.66×10^{-19} J?

$E = hv$
$1.66 \times 10^{-19} = 6.63 \times 10^{-34}\, v$
$v = 1.66 \times 10^{-19} / 6.63 \times 10^{-34}$
$v = 2.50 \times 10^{14}\ s^{-1}$
$c = \lambda v$
$3.00 \times 10^{8} = \lambda\ 2.50 \times 10^{14}$
$\lambda = 3.00 \times 10^{8} / 2.50 \times 10^{14}$
$\lambda = 1.20 \times 10^{-6}$ m

Ionisation energy and the convergence limit

- In an atom, the highest possible energy level corresponds to the frequency at which the spectral lines converge (n=∞).
- If enough energy is supplied, the electron in the hydrogen atom can be promoted from n=1 to the infinity energy level, n=∞.
- At this point, the electron has been removed from the attraction of the nucleus and the atom has been ionised.

Calculating ionisation energy

Example 1: In the hydrogen emission spectrum, the transition from n=∞ to n=1 produces a line in the UV spectrum with a wavelength of 9.12×10^{-8} m. Calculate the ionisation energy of a hydrogen atom.

- First, convert from wavelength to frequency:

$$c = \lambda\nu$$

$$3.00 \times 10^{8} = 9.12 \times 10^{-8}\ \nu$$

$$\nu = \frac{3.00\times10^{8}}{9.12\times10^{-8}} = 3.29 \times 10^{15}\ \text{s}^{-1}$$

- Next, calculate the energy to remove one electron from one hydrogen atom:

$$E=h\nu$$

$$E = 6.63 \times 10^{-34} \times 3.29 \times 10^{15}$$

$$E = 2.18 \times 10^{-18}\ \text{J}$$

- Finally, calculate the energy to remove one mole of electrons from one mole of hydrogen atom, in kJ mol^{-1}

$$(6.02 \times 10^{23} \times 2.18 \times 10^{-18}) / 1000 = 1310\ \text{kJ mol}^{-1}$$

Example 2: Determine the wavelength of a photon that will cause the first ionisation of helium. The ionisation energy of helium is 2372 kJ mol^{-1}.

- First, calculate the energy to remove one electron from one helium atom:

  $E = \frac{2372000}{6.02 \times 10^{23}}$ = 3.94 × 10^{-18} J

- Next, calculate the frequency of the photon:

  $E = h\nu$

  3.94 × 10^{-18} = 6.63 × 10^{-34} ν

  ν = 5.94 × 10^{15} s^{-1}

- Finally convert from frequency to wavelength:

  $c = \lambda\nu$

  3.00 × 10^{8} = 5.94 × 10^{15} λ

  $\lambda = \frac{3.00 \times 10^{8}}{5.94 \times 10^{15}}$ = 5.05 × 10^{-8} m

Exercises:

1. Which transition corresponds to the ionisation of hydrogen in the ground state?
   The transition is from energy level n = 1 to n = ∞.

2. What has occurred when the electron is in the n=∞ energy level?
   When the electron is in the n = ∞ energy level, the atom has been ionised. At this point, the electron has been removed from the attraction of the nucleus.

3. The convergence limit for the sodium atom has a wavelength of 2.42 × 10^{-7} m. Calculate the first ionization energy of sodium from this data.
   The two equations you need are $c = \lambda\nu$ and E=hv
   Planck's constant = 6.63 × 10^{-34} J s
   c = 3.00 × 10^{8} m s^{-1}

   You are given the wavelength, so use the equation $c = \lambda\nu$ to find the frequency.

   $c = \lambda\nu$

   3.00 × 10^{8} m s^{-1} = 2.42 × 10^{-7} ν

   $\nu = \frac{3.00 \times 10^{8}}{2.42 \times 10^{-7}}$ = 1.24 × 10^{15} s^{-1}

Next use the equation below to find the energy to remove one electron.

$E = h\nu$

$E = (6.63 \times 10^{-34})\ (1.24 \times 10^{15})$

$E = 8.22 \times 10^{-19}$ J

Multiply by Avogadro's constant to get the energy to remove one mole of electrons (multiply by 6.02×10^{23} and divide by 1000)

$(6.02 \times 10^{23} \times 8.22 \times 10^{-19})\ /1000 = 495$ kJ mol^{-1}

4. Calculate the frequency of a photon that will cause the first ionisation of copper. The first ionisation energy of copper is 745 kJ mol^{-1}.

First, calculate the energy to remove one electron from one helium atom:

$E = \frac{745000}{6.02 \times 10^{23}} = 1.24 \times 10^{-18}$ J

Next, calculate the frequency of the photon:

$E = h\nu$

$1.24 \times 10^{-18} = 6.63 \times 10^{-34}\ \nu$

$\nu = 1.87 \times 10^{15}$ s^{-1}

Structure 1.3.7

Understandings:

- Successive ionization energy (IE) data for an element give information about its electron configuration.

Learning outcomes:

- Deduce the group of an element from its successive ionization data.

Additional notes:

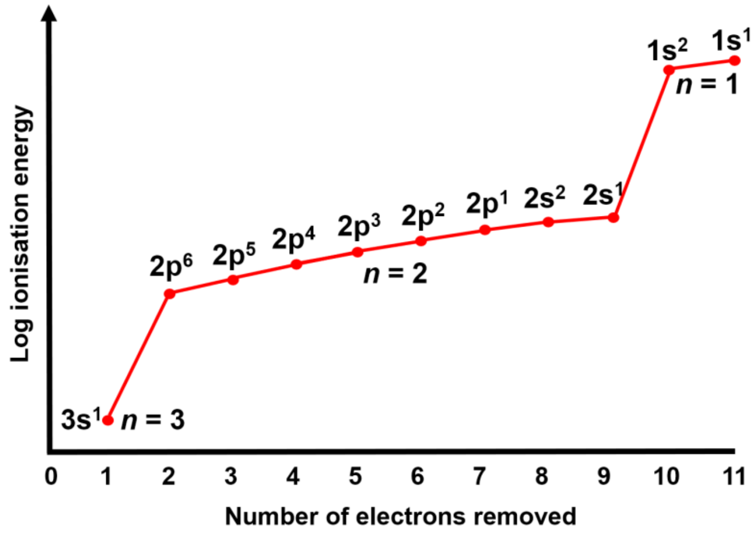
- Databases are useful for compiling graphs of trends in IEs.

Linking questions:

- AHL Structure 3.1—How do patterns of successive IEs of transition elements help to explain the variable oxidation states of these elements?

Successive ionisation energies

- We can determine to which group an element belongs by looking at a graph of successive ionisation energies.
- The graph below shows the successive ionisation energies of sodium.



- The large increases in ionisation energy show the existence of main energy levels in the atom.
- Successive ionisation energies show an increase because as more electrons are removed, the nucleus attracts the remaining electrons more strongly.
- The is a large increase in ionisation energy after the first electron is removed. This tells us that the element is located in group one of the periodic table.

Table of successive ionisation energies for elements Na to S

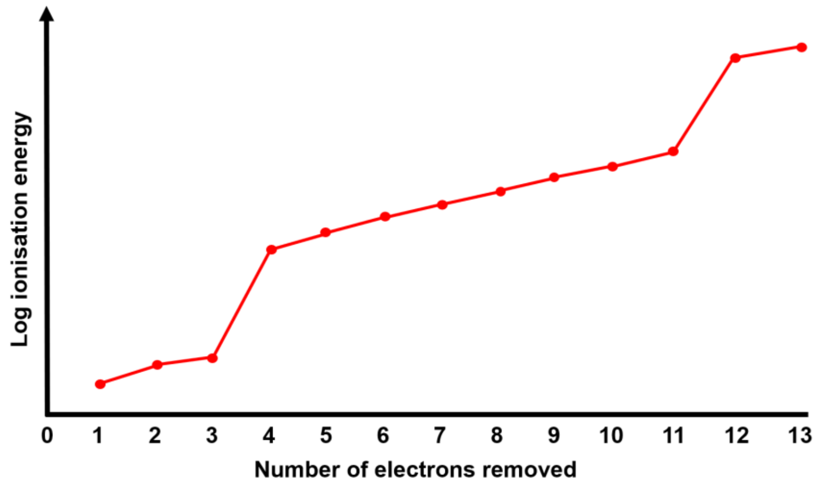
Element	Group number	IE_1 kJ mol^{-1}	IE_2 kJ mol^{-1}	IE_3 kJ mol^{-1}	IE_4 kJ mol^{-1}1	IE_5 kJ mol^{-1}	IE_6 kJ mol^{-1}	IE_7 kJ mol^{-1}
Na	1	496	**4562**	**6912**	**9543**	**13353**	**16610**	**20114**
Mg	2	738	1451	**7733**	**10540**	**13630**	**17995**	**21703**
Al	13	578	1817	2745	**11575**	**14380**	**18376**	**23293**
Si	14	787	1577	3231	4356	**16091**	**19784**	**23783**
P	15	1012	1903	2912	4956	6273	**22233**	**25397**
S	16	1000	2251	3361	4564	7013	8495	**27106**

- From the table, we can see that the large increase in ionisation energy corresponds to the group number and the number of valence electrons.
- For group 1 elements with one valence electron, the large increase occurs for the second ionisation.
- For group 2 elements with two valence electrons, the large increase occurs for the third ionisation and so on.

Exercises:

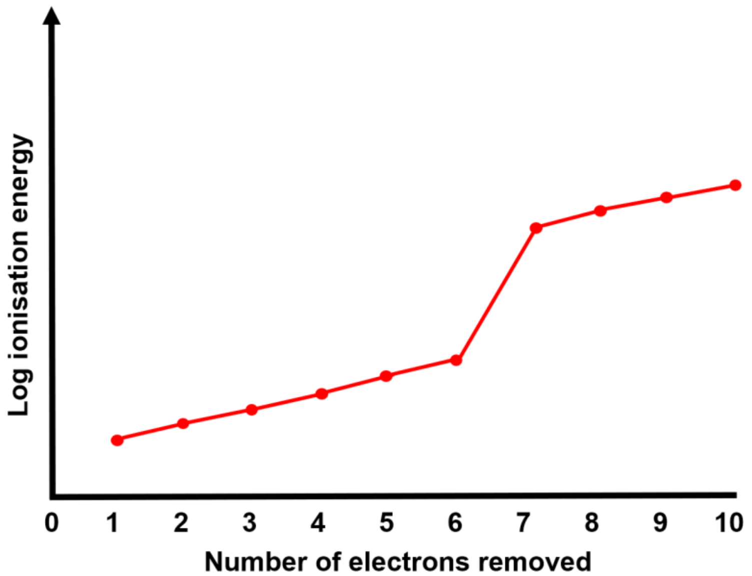
1. Explain why the successive ionisation energies of an element show an increase.
Successive ionisation energies increase because as more electrons are removed, the nucleus attracts the remaining electrons more strongly (the electrons are being removed from an increasingly positive ion).

2. From the graph of successive ionisation energies below, explain to which group of the periodic table the elements belong.



The element is located in group 13; there is a large increase in ionisation energy between the 3rd and 4th ionisation energies. The 4th electron is being removed from an energy level closer to the nucleus, which experiences a stronger electrostatic attraction from the nucleus and requires more energy to remove.

3. The graph below shows the first ten ionisation energies of an element. To which group does the element belong?



The element is located in group 16; there is a large increase in ionisation energy between the 6th and 7th ionisation energies. The 7th electron is being removed from an energy level closer to the nucleus, which experiences a stronger electrostatic attraction from the nucleus and requires more energy to remove.

4. From the table of data shown below, explain to which group the element belongs.

	1st	2nd	3rd	4th	5th
Ionisation energy ($kJ\ mol^{-1}$)	1087	2353	4621	6223	37831

The element is located in group 14; there is a large increase in ionisation energy between the 4th and 5th ionisation energies. The 5th electron is being removed from an energy level closer to the nucleus, which experiences a stronger electrostatic attraction from the nucleus and requires more energy to remove.

5. Sketch a graph to show the relative values of the successive ionisation energies of boron.