

Electrons in atoms: shells, sub-shells, orbitals and ionisation energy (AS Chemistry notes)

These notes cover the electrons-in-atoms part of the atomic structure topic in Cambridge International AS & A Level Chemistry 9701: energy levels, sub-shells and orbitals, how electrons fill them, electronic configurations of atoms and ions, and ionisation energies. Ionisation energy is the main experimental evidence for the arrangement of electrons, so the two halves of the topic belong together. Learn the boxed definitions word for word, work through the examples, then try the Quick check at the end.

1. Energy levels: shells and the principal quantum number

Electrons in an atom cannot have any energy they like. They occupy fixed energy levels, called **shells**, around the nucleus. Each shell is labelled by its **principal quantum number**, $n = 1, 2, 3, 4$ and so on. The first shell ($n = 1$) is closest to the nucleus and lowest in energy. As n increases, the shells are on average further from the nucleus, the electrons in them are higher in energy, and the energy gap between successive shells gets smaller.

Learn.

A **shell** (main energy level) is a group of orbitals with the same principal quantum number, n .

The **principal quantum number**, n , is the number of the shell; it indicates the main energy level of an electron and its average distance from the nucleus.

The maximum number of electrons in shell n is $2n^2$: 2 in $n = 1$, 8 in $n = 2$, 18 in $n = 3$, 32 in $n = 4$.

The **ground state** of an atom is its lowest energy arrangement, with every electron in the lowest available energy level. If an atom absorbs energy, an electron can be promoted to a higher level; the atom is then in an **excited state**. Unless a question says otherwise, all electronic configurations you write are ground-state configurations.

2. Sub-shells and atomic orbitals

2.1 Sub-shells

Each shell is divided into **sub-shells** of slightly different energy, labelled s, p, d and f. At AS level you need s, p and d. The first shell has only an s sub-shell; the second has s and p; the third has s, p and d; the fourth has s, p, d and f.

Shell (n)	Sub-shells	Orbitals	Maximum electrons
1	1s	1	2
2	2s, 2p	$1 + 3 = 4$	8
3	3s, 3p, 3d	$1 + 3 + 5 = 9$	18
4	4s, 4p, 4d, 4f	$1 + 3 + 5 + 7 = 16$	32



2.2 Orbitals

Definition to learn. An **atomic orbital** is a region of space around the nucleus in which there is a high probability of finding an electron. Each orbital can hold a maximum of two electrons, and these must have opposite spins.

Learn: orbitals per sub-shell.

An s sub-shell has 1 orbital and holds up to 2 electrons.

A p sub-shell has 3 orbitals and holds up to 6 electrons.

A d sub-shell has 5 orbitals and holds up to 10 electrons.

(An f sub-shell has 7 orbitals and holds up to 14 electrons; f is not needed in detail at AS.)

All the orbitals within one sub-shell have the same energy.

2.3 Shapes of s and p orbitals

An **s orbital** is spherical, centred on the nucleus. The 2s orbital has the same shape as 1s but is larger, and 3s is larger again: the electron is, on average, further from the nucleus.

A **p orbital** is dumbbell shaped (two lobes on opposite sides of the nucleus, with zero probability of finding the electron at the nucleus itself). The three p orbitals in a sub-shell point along three mutually perpendicular axes and are labelled p_x , p_y and p_z . As with s orbitals, 3p orbitals have the same shape as 2p but are larger and higher in energy.

When you sketch them, draw an s orbital as a circle (representing a sphere) with the nucleus at the centre, and draw each p orbital as two equal lobes touching at the nucleus along one axis. The d orbitals have more complex shapes; you will need them later in the A Level transition elements topic.

Exam tip. "Describe the shape" questions want a word and a direction: "an s orbital is spherical", "a p orbital is dumbbell shaped, and the three p orbitals lie along the x, y and z axes at 90° to each other". A neat sketch with labelled axes earns the same marks.

3. Filling orbitals: the order of sub-shell energies

3.1 The order of filling

In the ground state, electrons enter the lowest-energy sub-shell available. The sub-shell energies do not simply follow the shell number: the 4s sub-shell is slightly lower in energy than 3d for the atoms where it fills, so 4s fills before 3d.

Learn: order of filling for the first 36 elements.

$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$

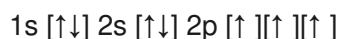
3.2 Three rules for placing electrons

1. **Lowest energy first.** Electrons occupy the lowest-energy sub-shell available (the order above).
2. **Two electrons per orbital, with opposite spins.** No orbital holds more than two electrons, and two electrons sharing an orbital have opposite spins (shown as $\uparrow\downarrow$).
3. **Single occupation first within a sub-shell.** In a set of orbitals of equal energy (for example the

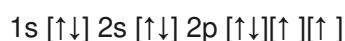
three 2p orbitals), electrons occupy separate orbitals singly, with parallel spins, before any pairing occurs. This is because two electrons in the same orbital repel each other; this extra repulsion is called spin-pair repulsion.

3.3 Electrons-in-boxes notation

Each orbital is drawn as a box and each electron as a half-arrow pointing up or down to show its spin. Writing the boxes in a row, the configuration of nitrogen (7 electrons) is:



and that of oxygen (8 electrons) is:



Nitrogen has three unpaired electrons, one in each 2p orbital. In oxygen, the eighth electron has to pair up in one of the 2p orbitals, so oxygen has two unpaired electrons and one pair in a 2p orbital. This difference explains a key feature of the ionisation energy pattern in section 6.

Definition to learn. A **free radical** is a species with one or more unpaired electrons.

Single atoms such as Cl and H have an unpaired electron, so they are free radicals; they are very reactive and appear in the mechanism of free-radical substitution in organic chemistry.

4. Electronic configurations of atoms

4.1 Writing full and shorthand configurations

The full configuration lists every occupied sub-shell with the number of electrons as a superscript: sodium (11 electrons) is $1s^2 2s^2 2p^6 3s^1$. The shorthand form replaces the inner electrons with the symbol of the previous noble gas in square brackets: sodium is $[\text{Ne}] 3s^1$. Check your answer by adding the superscripts: they must total the atomic number for an atom.

Worked example 1. Write the full electronic configurations of sulfur ($Z = 16$), potassium ($Z = 19$) and iron ($Z = 26$).

S: $1s^2 2s^2 2p^6 3s^2 3p^4$ ($2 + 2 + 6 + 2 + 4 = 16$).

K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. The 19th electron goes into 4s, not 3d, because 4s is lower in energy at this point.

Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$, or $[\text{Ar}] 3d^6 4s^2$. After $4s^2$ is full (calcium), the next ten electrons from scandium to zinc enter 3d.

Cambridge accepts 3d written before or after 4s, as long as the numbers of electrons are right. Writing 3d before 4s groups the sub-shells by shell, which makes ion configurations easier to get right.

4.2 Configurations and the Periodic Table

The Periodic Table is arranged in blocks that match the sub-shell being filled by the outermost (highest-energy) electron: Groups 1 and 2 form the s-block, Groups 13 to 18 form the p-block, and the transition elements form the d-block. The period number equals the number of the outermost shell. So an element in Period 3, Group 15 has outer configuration $3s^2 3p^3$ (this is phosphorus). Reading the

configuration from the table position, and the position from the configuration, are both common short questions.

4.3 The exceptions: chromium and copper

Two elements in the first row of the d-block do not follow the simple filling order. Their actual ground-state configurations are:

Learn.

Chromium ($Z = 24$): $[\text{Ar}] 3d^5 4s^1$ (not $[\text{Ar}] 3d^4 4s^2$)

Copper ($Z = 29$): $[\text{Ar}] 3d^{10} 4s^1$ (not $[\text{Ar}] 3d^9 4s^2$)

The usual explanation at this level is that the 3d and 4s sub-shells are very close in energy, and a half-filled ($3d^5$) or completely filled ($3d^{10}$) d sub-shell with one electron in 4s gives a more stable, lower-energy arrangement overall. In chromium, the arrangement $3d^5 4s^1$ has six unpaired electrons, each in its own orbital, which avoids the spin-pair repulsion in 4s. You will be expected to recall both configurations; the explanation is rarely worth more than a mark.

5. Electronic configurations of ions

5.1 Ions of main-group elements

For a positive ion, remove electrons from the highest-energy occupied sub-shell of the atom. For a negative ion, add electrons to the next available orbitals. Most simple ions of s- and p-block elements have the same configuration as a noble gas.

Worked example 2. Write the full electronic configurations of Ca^{2+} , Cl^- and Al^{3+} .

Ca is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$; remove the two 4s electrons: Ca^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6$ (18 electrons, the same as argon).

Cl is $1s^2 2s^2 2p^6 3s^2 3p^5$; add one electron to 3p: Cl^- is $1s^2 2s^2 2p^6 3s^2 3p^6$.

Al is $1s^2 2s^2 2p^6 3s^2 3p^1$; remove the 3p electron and both 3s electrons: Al^{3+} is $1s^2 2s^2 2p^6$ (the same as neon).

Species with the same number of electrons, and so the same configuration, are called **isoelectronic**. For example O^{2-} , F^- , Ne, Na^+ , Mg^{2+} and Al^{3+} all have the configuration $1s^2 2s^2 2p^6$. Across this series the nuclear charge increases from 8 to 13 while the number of electrons stays at 10, so the species get smaller.

5.2 Ions of transition elements: 4s out first

This is the point students most often get wrong. Although 4s fills before 3d, once the 3d orbitals are occupied the 4s electrons are the outermost electrons and are lost first when a transition element forms a positive ion.

Learn. When a transition element forms a positive ion, the 4s electrons are removed before any 3d electrons.

Worked example 3. Write the electronic configurations of Fe^{2+} , Fe^{3+} and Cu^{2+} in shorthand form.

Fe is $[\text{Ar}] 3d^6 4s^2$.

Fe^{2+} : remove both 4s electrons: $[\text{Ar}] 3d^6$.

Fe^{3+} : remove both 4s electrons and then one 3d electron: $[\text{Ar}] 3d^5$.

Cu is $[\text{Ar}] 3d^{10} 4s^1$.

Cu^{2+} : remove the 4s electron and then one 3d electron: $[\text{Ar}] 3d^9$.

A quick check: the number of electrons in the ion must equal the atomic number minus the charge. Fe^{3+} has $26 - 3 = 23$ electrons, and $[\text{Ar}] 3d^5$ has $18 + 5 = 23$.

Exam tip. Never write Fe^{2+} as $[\text{Ar}] 3d^4 4s^2$. That is the most common error in this topic, and it costs the mark every time. Remove 4s first, then 3d.

6. Ionisation energy

6.1 Definitions and equations

Definition to learn. The **first ionisation energy** of an element is the energy needed to remove one electron from each atom in one mole of gaseous atoms of the element to form one mole of gaseous $1+$ ions.

Equation for sodium: $\text{Na}(\text{g}) \rightarrow \text{Na}^+(\text{g}) + \text{e}^-$

Definition to learn. The **second ionisation energy** is the energy needed to remove one electron from each ion in one mole of gaseous $1+$ ions to form one mole of gaseous $2+$ ions.

Equation for magnesium: $\text{Mg}^+(\text{g}) \rightarrow \text{Mg}^{2+}(\text{g}) + \text{e}^-$

In general, the n th ionisation energy is $\text{X}^{(n-1)+}(\text{g}) \rightarrow \text{X}^{n+}(\text{g}) + \text{e}^-$.

Ionisation energies are measured in kJ mol^{-1} and are always positive (endothermic), because energy must be supplied to pull a negatively charged electron away from the attraction of the positive nucleus. State symbols matter: the atoms and ions must be gaseous, and an equation without (g) usually loses the mark.

Worked example 4. Write an equation, with state symbols, for the third ionisation energy of aluminium.

The third electron is removed from a $2+$ ion: $\text{Al}^{2+}(\text{g}) \rightarrow \text{Al}^{3+}(\text{g}) + \text{e}^-$.

A common error is to write $\text{Al}(\text{g}) \rightarrow \text{Al}^{3+}(\text{g}) + 3\text{e}^-$, which is the sum of the first three ionisation energies, not the third.

6.2 The factors that affect ionisation energy

Every explanation in this topic is built from four factors. Use them by name.

Learn: four factors.

1. **Nuclear charge:** more protons give a stronger attraction for the electrons, so a higher ionisation energy.
2. **Distance from the nucleus** (atomic radius, or which shell the electron is in): the further the electron is from the nucleus, the weaker the attraction, so a lower ionisation energy.
3. **Shielding:** electrons in inner shells (and, to a smaller extent, inner sub-shells) repel the outer electron and reduce the attraction it feels from the nucleus, so more shielding means a lower ionisation energy.
4. **Spin-pair repulsion:** an electron that shares an orbital with another electron is repelled by it, so it is easier to remove than an unpaired electron in a similar orbital.

Factors 2 and 3 often act together. In a good answer, state which factor is the deciding one and say clearly that it is the attraction between the nucleus and the outer electron that changes.

7. Trends in first ionisation energy

7.1 Across a period

Across Period 3 from sodium to argon, first ionisation energy shows a general increase. The approximate values below show the pattern; your data booklet gives the values to use in the exam.

Element	Na	Mg	Al	Si
First IE / kJ mol ⁻¹	496	738	578	787

Element	P	S	Cl	Ar
First IE / kJ mol ⁻¹	1012	1000	1251	1521

General increase. Across the period the nuclear charge increases by one each time, while the outer electrons are added to the same shell ($n = 3$). Shielding by the inner electrons stays almost the same, so the outer electrons are attracted more strongly and the atomic radius decreases. More energy is needed to remove an outer electron.

There are two small decreases, and each has its own explanation.

Worked example 5. Explain why the first ionisation energy of aluminium is lower than that of magnesium.

Mg is [Ne] 3s²; Al is [Ne] 3s² 3p¹.

The electron removed from aluminium is in a 3p orbital, which is higher in energy than the 3s orbital from which the magnesium electron is removed.

The 3p electron is also on average further from the nucleus and is partly shielded by the 3s electrons.

So, despite the extra proton in aluminium, less energy is needed to remove the 3p electron.

Worked example 6. Explain why the first ionisation energy of sulfur is lower than that of phosphorus.

P is [Ne] 3s² 3p³: the three 3p electrons are each in a separate orbital (3p [↑][↑][↑]).

S is $[\text{Ne}] 3s^2 3p^4$: two of the 3p electrons share an orbital ($3p \uparrow\downarrow \uparrow \uparrow$).

In sulfur, the electron removed is one of the paired electrons. It is repelled by the other electron in the same orbital (spin-pair repulsion), so it is easier to remove, even though sulfur has one more proton.

The same two features appear in Period 2: boron is lower than beryllium (2p electron removed instead of 2s), and oxygen is lower than nitrogen (paired 2p electron removed). The noble gas at the end of each period has the highest first ionisation energy in that period, and the Group 1 metal at the start of the next period has the lowest, because its outer electron is in a new shell further from the nucleus.

7.2 Down a group

First ionisation energy generally decreases down a group. In Group 2, for example, the values fall from beryllium to barium.

Worked example 7. Explain why the first ionisation energy decreases down Group 2 from magnesium to calcium.

Calcium has one more occupied shell than magnesium, so its outer electrons (4s) are further from the nucleus than magnesium's (3s).

Calcium also has more inner shells of electrons, so its outer electrons are more shielded.

Although the nuclear charge increases (12 to 20), the greater distance and greater shielding outweigh it.

The attraction between the nucleus and the outer electrons is weaker in calcium, so less energy is needed to remove one.

Exam tip. "Increased nuclear charge" must always be mentioned in a down-the-group answer, and then outweighed. Saying only "the atom is bigger" is not enough: give the reason (more shells), the effect (more shielding, greater distance) and the result (weaker attraction for the outer electron).

8. Successive ionisation energies

8.1 Why successive ionisation energies always increase

For any element, each successive ionisation energy is larger than the one before. After each electron is removed, the remaining electrons are held by the same number of protons, so the attraction per electron is greater. The ion is also smaller, so the next electron is closer to the nucleus, and there is less repulsion between the remaining electrons. More energy is needed to remove each further electron.

8.2 What the big jumps reveal

The increase is not smooth. There is a very large jump whenever the next electron has to come from a shell closer to the nucleus. That electron is much nearer the nucleus and much less shielded, so far more energy is needed to remove it.

For sodium, the first ionisation energy is about 496 kJ mol^{-1} and the second is about 4560 kJ mol^{-1} , almost ten times larger. The first electron comes from the 3s orbital in the third shell; the second must

come from the 2p sub-shell in the second shell. If you plot the logarithm of all eleven successive ionisation energies of sodium against the number of electrons removed, the points fall into three groups: 1 electron (shell 3), then 8 electrons (shell 2), then 2 electrons (shell 1), with a large jump between groups. This pattern is direct evidence for shells.

Smaller jumps within a shell can also be seen in accurate data. For sodium, electrons 2 to 7 come from the 2p sub-shell and electrons 8 and 9 from 2s. The step from the last 2p electron to the first 2s electron is larger than the steps between the 2p electrons, which is evidence for sub-shells within a shell.

Method: deducing the group from successive ionisation energies.

1. Look for the first very large jump (work out the ratio of each value to the one before if it is not obvious).
2. The number of electrons removed before that jump is the number of electrons in the outer shell.
3. For main-group elements this gives the group: 1 electron before the jump means Group 1, 2 means Group 2, 3 means Group 13, 4 means Group 14, and so on up to 8 (Group 18).

Worked example 8. The first four ionisation energies of element X, in kJ mol^{-1} , are 590, 1145, 4912 and 6491. Deduce the group of X in the Periodic Table and explain your answer.

Ratios: $1145/590 \approx 1.9$; $4912/1145 \approx 4.3$; $6491/4912 \approx 1.3$.

The large jump comes between the second and third ionisation energies.

So the first two electrons are removed from the outer shell, and the third must be removed from a shell closer to the nucleus, which is less shielded and more strongly attracted.

X has two outer electrons, so it is in Group 2.

If you know the total number of ionisation energies (all of the electrons), you can also find the period. For example, an element with 12 electrons whose data show groups of 2, 8 and 2 has three occupied shells, so it is in Period 3 (this is magnesium).

Worked example 9. Element Q is in Period 3. Its successive ionisation energies show a large jump between the fifth and sixth values. Identify Q and write its full electronic configuration.

Five outer electrons means Group 15. The Group 15 element in Period 3 is phosphorus.

P: $1s^2 2s^2 2p^6 3s^2 3p^3$. The five outer electrons are $3s^2 3p^3$; the sixth electron would be removed from the 2p sub-shell.

Quick check

Answer all the questions, then check your work against the full answers below. Data for real elements are approximate; in an exam, use the values given in the question or in the data booklet.

1. (a) Define the term first ionisation energy. (b) Write an equation, with state symbols, for the first ionisation energy of chlorine.
2. (a) State the number of orbitals in a 3d sub-shell. (b) State the maximum number of electrons in the third shell ($n = 3$), and show how this number is made up from its sub-shells.
3. Write the full electronic configuration of phosphorus ($Z = 15$) and draw its outer electrons using

electrons-in-boxes notation. State the number of unpaired electrons.

- Write the shorthand electronic configurations of chromium ($Z = 24$) and copper ($Z = 29$). Explain briefly why they do not follow the usual filling order.
- Write the shorthand electronic configurations of (a) Mn^{2+} ($Z = 25$), (b) Ni^{2+} ($Z = 28$) and (c) V^{3+} ($Z = 23$).
- Identify each species. (a) An atom with configuration $[\text{Ne}] 3s^2 3p^5$. (b) An ion with a 3+ charge and configuration $[\text{Ar}] 3d^5$. (c) An ion with a 2- charge and configuration $1s^2 2s^2 2p^6$.
- Describe the shapes of an s orbital and a p orbital. State two ways in which a 3p orbital differs from a 2p orbital.
- Explain why the first ionisation energy of oxygen is lower than that of nitrogen.
- Explain why the first ionisation energy of potassium is lower than that of sodium.
- The first eight successive ionisation energies of element Y, in kJ mol^{-1} , are: 787, 1577, 3232, 4356, 16091, 19805, 23780, 29287. Deduce the group to which Y belongs and explain your reasoning.
- Explain why the second ionisation energy of sodium is much larger than its first ionisation energy.
- Na^+ and Ne are isoelectronic. (a) State what is meant by isoelectronic. (b) Suggest, with a reason, which of Na^+ and Ne needs more energy to remove one electron.

1. (a) The energy needed to remove one electron from each atom in one mole of gaseous atoms (of the element) to form one mole of gaseous 1+ ions.

(b) $\text{Cl(g)} \rightarrow \text{Cl}^+\text{(g)} + \text{e}^-$. Note that it is chlorine atoms, Cl(g) , not $\text{Cl}_2\text{(g)}$.

2. (a) 5 orbitals.

(b) 18 electrons: 3s holds 2 (1 orbital), 3p holds 6 (3 orbitals) and 3d holds 10 (5 orbitals); $2 + 6 + 10 = 18$, which equals $2n^2 = 2 \times 3^2$.

3. P: $1s^2 2s^2 2p^6 3s^2 3p^3$.

Outer electrons: 3s $[\uparrow\downarrow]$ 3p $[\uparrow][\uparrow][\uparrow]$.

The three 3p electrons occupy separate orbitals with parallel spins (single occupation before pairing), so there are 3 unpaired electrons.

4. Cr: $[\text{Ar}] 3d^5 4s^1$. Cu: $[\text{Ar}] 3d^{10} 4s^1$.

The 3d and 4s sub-shells are very close in energy. A half-filled $3d^5$ or completely filled $3d^{10}$ sub-shell, with one electron in 4s, is a more stable (lower energy) arrangement than $3d^4 4s^2$ or $3d^9 4s^2$.

5. Remove the 4s electrons first, then 3d.

(a) Mn is $[\text{Ar}] 3d^5 4s^2$, so Mn^{2+} is $[\text{Ar}] 3d^5$.

(b) Ni is $[\text{Ar}] 3d^8 4s^2$, so Ni^{2+} is $[\text{Ar}] 3d^8$.

(c) V is $[\text{Ar}] 3d^3 4s^2$; remove two 4s electrons and one 3d electron, so V^{3+} is $[\text{Ar}] 3d^2$.

Check: V^{3+} has $23 - 3 = 20$ electrons, and $18 + 2 = 20$.

6. (a) $10 + 2 + 5 = 17$ electrons, so the atom is chlorine, Cl.

(b) $18 + 5 = 23$ electrons; with a 3+ charge the atom has 26 electrons, so $Z = 26$: the ion is Fe^{3+} .

(c) 10 electrons with a 2- charge means the atom has 8 electrons, so $Z = 8$: the ion is O^{2-} .

7. An s orbital is spherical, centred on the nucleus. A p orbital is dumbbell shaped, with two lobes

on opposite sides of the nucleus; the three p orbitals lie along the x, y and z axes, at right angles to each other.

A 3p orbital is larger (its electrons are on average further from the nucleus) and higher in energy than a 2p orbital. The shape is the same.

8. N is $1s^2 2s^2 2p^3$, with one electron in each 2p orbital. O is $1s^2 2s^2 2p^4$, so two of its 2p electrons share an orbital.

The electron removed from oxygen is one of this pair. It experiences repulsion from the other electron in the same orbital (spin-pair repulsion), so less energy is needed to remove it, even though oxygen has a greater nuclear charge than nitrogen.

9. Potassium's outer electron is in the 4s orbital, in the fourth shell; sodium's is in 3s, in the third shell.

The potassium outer electron is further from the nucleus and is shielded by more inner shells of electrons.

Although potassium has a greater nuclear charge (19 compared with 11), the greater distance and shielding mean that the attraction between the nucleus and the outer electron is weaker. Less energy is needed to remove it.

10. Look at the ratio of each value to the one before: $1577/787 \approx 2.0$, $3232/1577 \approx 2.0$, $4356/3232 \approx 1.3$, $16091/4356 \approx 3.7$, then $19805/16091 \approx 1.2$ and similar.

The large jump is between the fourth and fifth ionisation energies. The first four electrons are removed from the outer shell; the fifth must come from an inner shell, closer to the nucleus and less shielded, so much more energy is needed.

Y has four outer electrons, so it is in Group 14.

11. Na is $1s^2 2s^2 2p^6 3s^1$. The first electron is removed from the 3s orbital in the third shell.

The second electron must be removed from the 2p sub-shell in the second shell, which is much closer to the nucleus and is shielded only by the 1s electrons.

It is also being removed from a positive ion, with 11 protons attracting only 10 electrons.

So the attraction is much stronger and far more energy is needed (roughly ten times as much).

12. (a) Isoelectronic species have the same number of electrons (and the same electronic configuration). Both Na^+ and Ne have 10 electrons, $1s^2 2s^2 2p^6$.

(b) Na^+ needs more energy. It has 11 protons compared with 10 in neon, while the electron being removed is in the same 2p sub-shell with the same shielding. The greater nuclear charge attracts the electron more strongly (and Na^+ is smaller), so more energy is needed to remove it.

(Removing an electron from Na^+ is the second ionisation energy of sodium.)