Periodicity SL (answers)

IB CHEMISTRY SL

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

Syllabus objectives

3.1 Periodic table

Understandings:

- The periodic table is arranged into four blocks associated with the four sub-levels—s, p, d, and f.
- The periodic table consists of groups (vertical columns) and periods (horizontal rows).
- The period number (*n*) is the outer energy level that is occupied by electrons.
- The number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table.
- The periodic table shows the positions of metals, non-metals and metalloids.

Applications and skills:

- Deduction of the electron configuration of an atom from the element's position on the periodic table, and vice versa.

Guidance:

- The terms alkali metals, halogens, noble gases, transition metals, lanthanoids and actinoids should be known.
- The group numbering scheme from group 1 to group 18, as recommended by IUPAC, should be used.

3.2 Periodic trends

Understandings:

- Vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.
- Trends in metallic and non-metallic behaviour are due to the trends above.
- Oxides change from basic through amphoteric to acidic across a period.

Applications and skills:

- Prediction and explanation of the metallic and non-metallic behaviour of an element based on its position in the periodic table.
- Discussion of the similarities and differences in the properties of elements in the same group, with reference to alkali metals (group 1) and halogens (group 17).
- Construction of equations to explain the pH changes for reactions of Na_2O, MgO, P_4O10_4, and the oxides of nitrogen and sulfur with water.

Guidance:

- Only examples of general trends across periods and down groups are required.
- For ionization energy the discontinuities in the increase across a period should be covered.
- Group trends should include the treatment of the reactions of alkali metals with water, alkali metals with halogens and halogens with halide ions.

Introduction to the periodic table

- Elements in the periodic table are arranged in groups and periods.
- A group is a vertical column in the periodic table.
- Elements in the same group have the same number of electrons in the outer energy level (same number of valence electrons).
- A period is a horizontal row.
- Elements in the same period have the same number of occupied main energy levels.
- Elements on the periodic table are arranged in order of increasing atomic number.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.01						Atomic number **Element** Relative atomic mass											2 **He** 4.00
2	3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.90
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
6	55 **Cs** 132.91	56 **Ba** 137.33	57 † **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.20	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89 ‡ **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (269)	107 **Bh** (270)	108 **Hs** (269)	109 **Mt** (278)	110 **Ds** (281)	111 **Rg** (281)	112 **Cn** (285)	113 **Uut** (286)	114 **Uuq** (289)	115 **Uup** (288)	116 **Uuh** (293)	117 **Uus** (294)	118 **Uuo** (294)

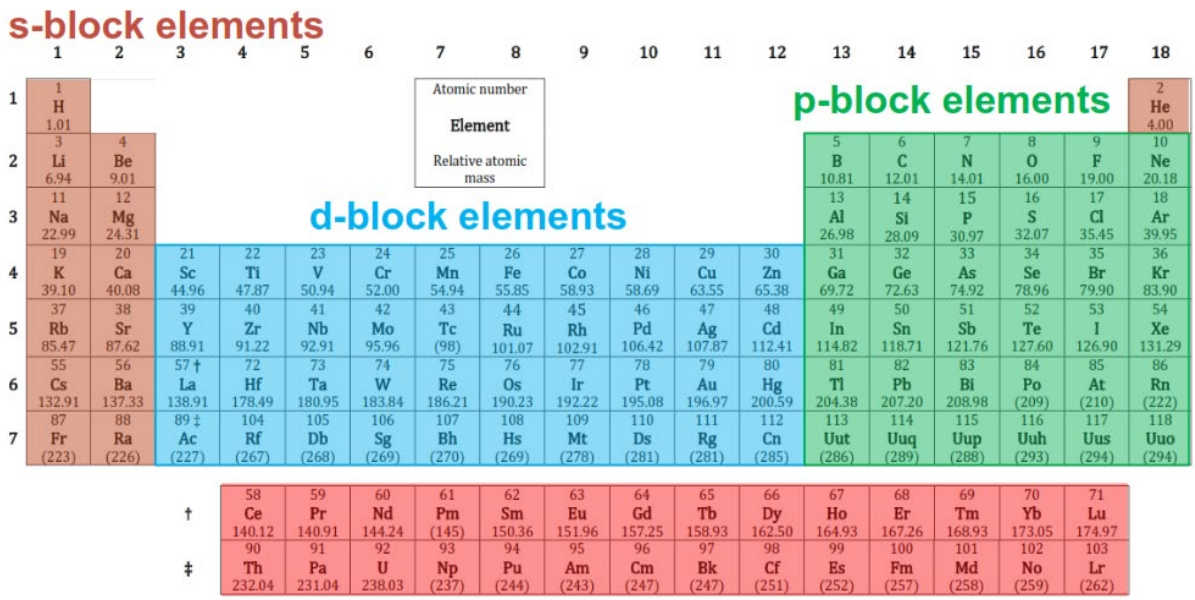
†	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
‡	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Names of the groups in the periodic table

- Group 1 – Alkali metals (Li, Na, K, Rb, Cs, Fr)
- Group 2 - Alkaline Earth metals (Be, Mg, Ca, Sr, Ba, Ra)
- Group 17 – Halogens (salt formers) (F, Cl, Br, I, At)
- Group 18 – Noble gases (He, Ne, Ar, Kr, Xe, Rn)
- Groups 3 – 12: Transition metals (elements)
- La – Lu: Lanthanides (lanthanoids)
- Ac – Lr: Actinides (actinoids)

Electron configurations and the periodic table

- The electron configuration of an element can be deduced from its position on the periodic table and vice-versa.
- The periodic table is divided into blocks (s, p, d, f).
- The block in which an element is located tells us which sub-level is in the process of being filled.



Example:

- The electron configuration of Li is $1s^2\ 2s^1$
- Lithium is an s block element in group 1 and period 2.
- Li has 2 occupied energy levels with the valence shell (n=2) containing 1 electron in the 2s sub-level.

Exercise:

Deduce the electron configuration of Ca, S and Kr from their positions in the periodic table.

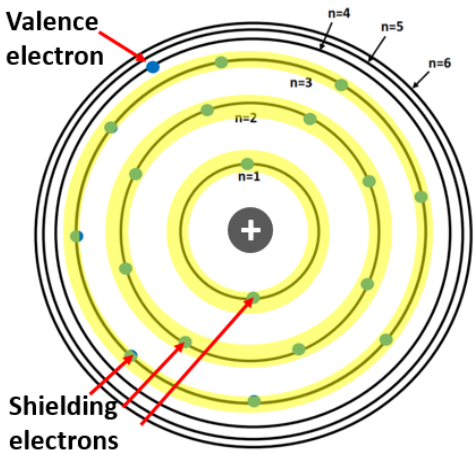
Ca Group 2 period 4

S Group 16 period 3

Kr Group 18 period 4

Electron shielding

- Electron shielding occurs when the inner (shielding) electrons shield the outer (valence) electrons from the full attraction of the nucleus.
- The valence electron(s) require less energy to remove than the inner electrons.



- Electron shielding remains constant across a period (left to right) because the number of shielding electrons is the same across a period.
- Electron shielding increases down a group because the number of shielding electrons increase down a group.

Nuclear charge and effective nuclear charge (Z_{eff})

- The nuclear charge of an atom is given by the atomic number (the number of protons in the nucleus) and increases by one between successive elements in the periodic table, as a proton is added to the nucleus.
- The outer valence electrons do not experience the full attraction from the nucleus as they are shielded from the nucleus and repelled by inner electrons.
- Effective nuclear charge (Z_{eff}) is the net positive charge experienced by valence electrons.
- The effective nuclear charge is less than the actual nuclear charge.
- The effective nuclear charge can be approximated by the following equation, where Z is the atomic number and S is the number of shielding electrons.

$$Z_{eff} = Z - S$$

Z is the atomic number
S is the number of shielding electrons

Exercise: Determine the effective nuclear charge for the above atom (potassium).

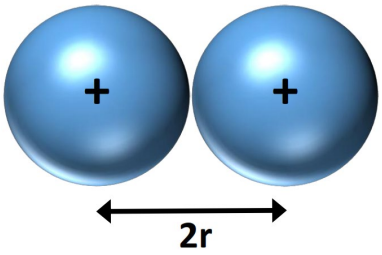
$Z_{eff} = Z - S$

$Z_{eff} = 19 - 18$

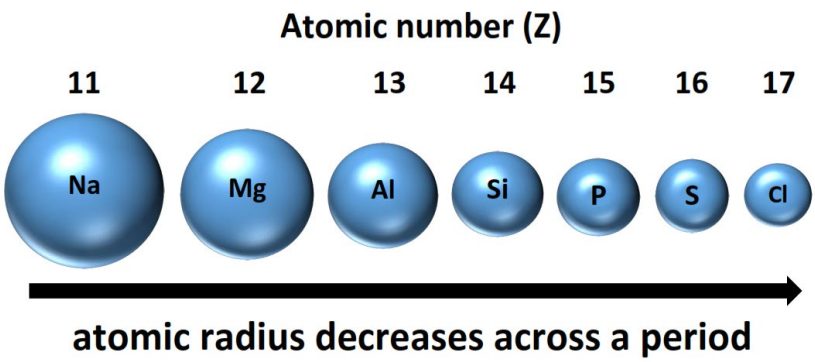
$Z_{eff} = 1+$

Atomic radius

- The atomic radius (van der Waals radius) is measured as half the distance between neighboring nuclei.
- Atomic and ionic radii values can be found in section 9 of the data booklet.



Trend across a period



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

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

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

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

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

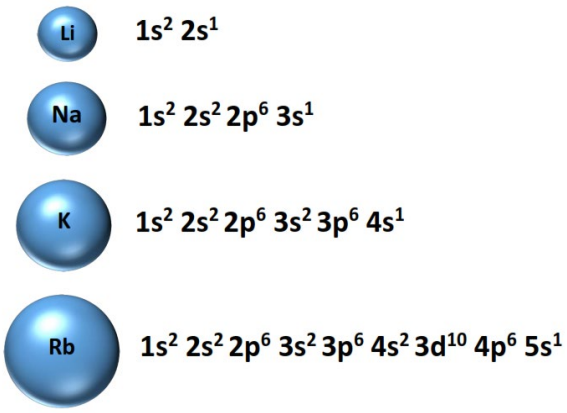
S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Explanation for the trend:

- Nuclear charge increases across a period.
- Electron shielding remains constant across a period.
- The electrons are pulled closer to the nucleus, therefore, the atomic radius decreases.

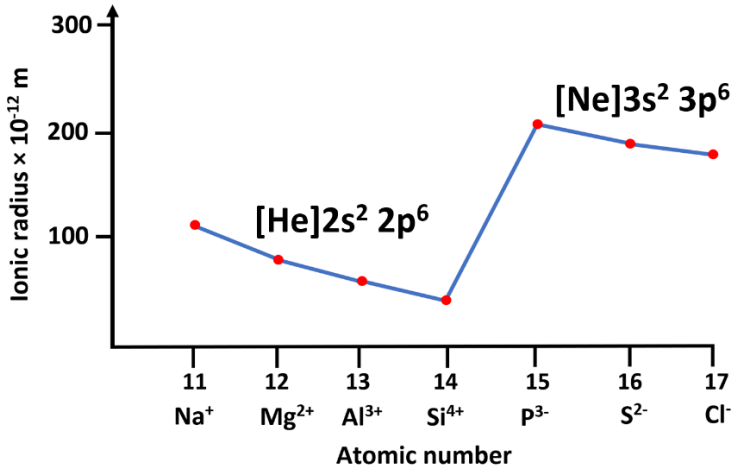
Trend in atomic radius down a group



The atomic radius increases down a group because the number of occupied energy levels increases.
This results in a weaker attraction between the nucleus and the valence electrons.

Trends in ionic radius

- Ionic radius increases down a group because of the increasing number of occupied energy levels.
- Ionic radius decreases across a period for positive ions.
- It then increases as we get to the negative ions and then decreases again.



- The first four positive ions have two occupied energy levels and the negative ions have three occupied energy levels, therefore, the ionic radius increases.

Trend for isoelectronic ions

- The table below shows data for the ion N^{3-} to Al^{3+}.

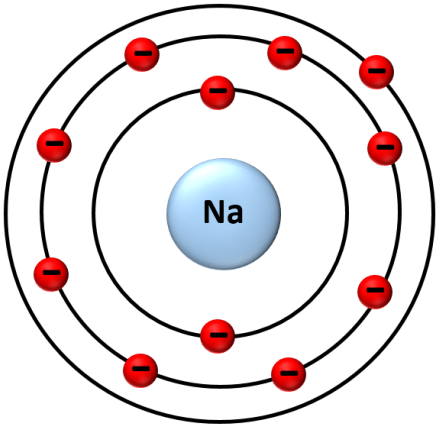
Ion	Atomic number	Electron configuration	Ionic radius ($\times 10^{-12}$ m)
N^{3-}	7	$1s^2\ 2s^2\ 2p^6$	146
O^{2-}	8	$1s^2\ 2s^2\ 2p^6$	140
F^-	9	$1s^2\ 2s^2\ 2p^6$	133
Na^+	11	$1s^2\ 2s^2\ 2p^6$	102
Mg^{2+}	12	$1s^2\ 2s^2\ 2p^6$	72
Al^{3+}	13	$1s^2\ 2s^2\ 2p^6$	54

- All six ions are isoelectronic (have the same electron configuration).
- The number of protons increases but the number of electrons remains the same.
- The attraction between the nucleus and electrons increases, which causes the ionic radius to decrease.
- The same trend can be seen for the ions P^{3-} to Ca^{2+}.

$$P^{3-} > S^{2-} > Cl^- > K^+ > Ca^{2+}$$

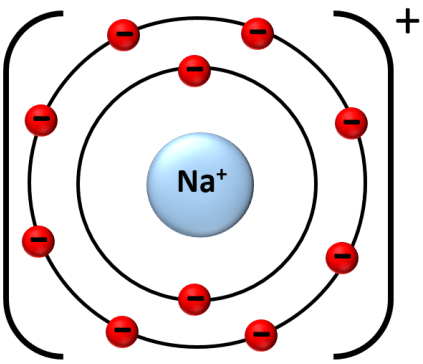
Positive ions (cations)

Sodium atom (160×10^{-12} m)



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

Sodium ion (102×10^{-12} m)

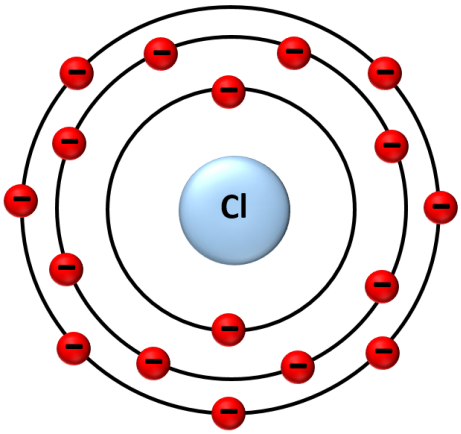


$1s^2\ 2s^2\ 2p^6$

- Positive ions are smaller than their parent atoms.
- Positive ions have fewer occupied energy levels than their parent atoms.
- Positive ions have more protons than electrons which results in an increased attraction between the nucleus and the valence electrons.
- The valence electrons are held more tightly (making the positive ion smaller).

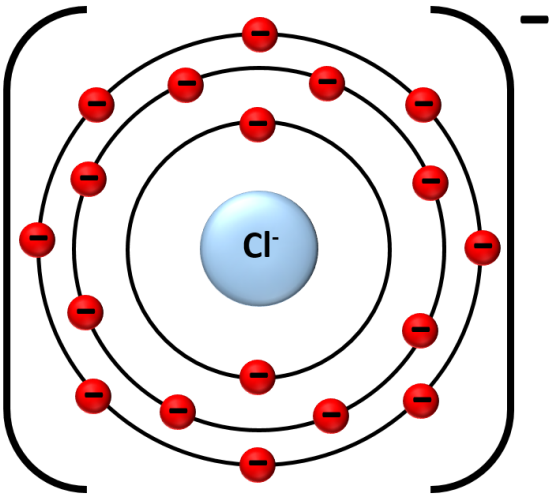
Negative ions (anions)

Chlorine atom (100×10^{-12} m)



$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Chloride ion (181×10^{-12} m)



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

- Negative ions are bigger than the parent atoms.
- Negative ions have more electrons than protons which results in a decreased attraction between the nucleus and the valence electrons.

Exercises:

1) State and explain the trend in atomic radius down a group.
Atomic radius increases down a group as the number of occupied energy levels increases.

2) State and explain the trend in atomic radius across a period.
Atomic radius decreases across a period as the nuclear charge increases and the electrons are added to the same main energy level (electron shielding remains constant).

3) State and explain which is smaller, the lithium atom or the lithium ion (Li^+).
The lithium ion is smaller than the lithium atom. The lithium ion has 1 occupied energy level, compared to 2 occupied energy levels for the lithium atom. The lithium ion has more protons than electrons which increases the attraction between the nucleus and valence electrons, making the ion smaller.

4) State and explain which is bigger, the fluorine atom or the fluoride ion (F^-).
The fluoride ion is bigger than the fluorine atom. The ion has more electrons than protons which decreases the attraction between the nucleus and valence electrons, making the ion bigger than the atom.

5) Arrange the following in order of increasing atomic radius (smallest first): Cl, Si, Na. Explain your reasoning.

Cl < Si <Na

Na, Si and Cl are all period 3 elements – atomic radius decreases from left to right across a period because of increasing nuclear charge and the electron shielding remains constant.

6) Arrange the following in order of increasing radius: Mg^{2+}, Na^+, O^{2-}, F^-, N^{3-}, Al^{3+}. Explain your reasoning.

$$Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$$

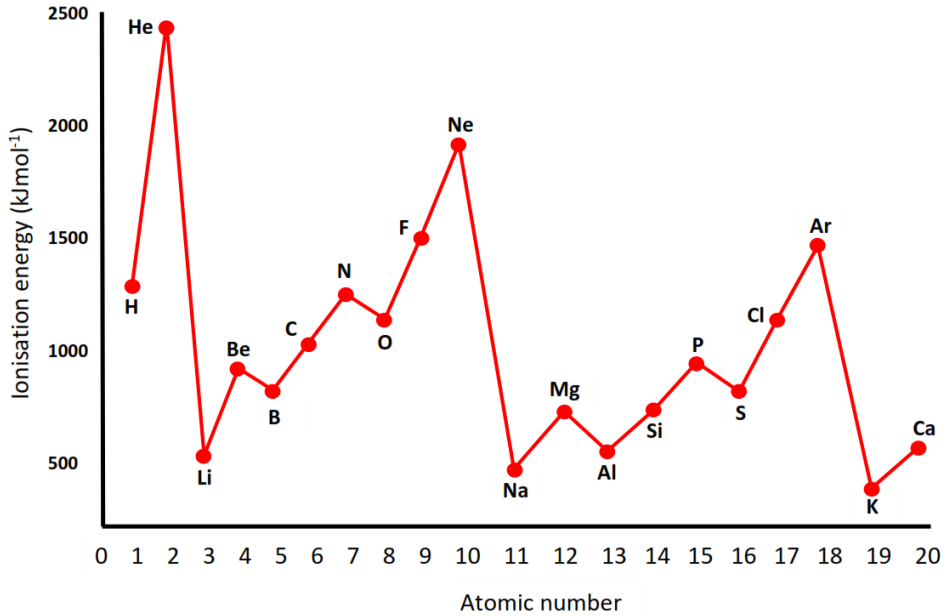
All ions have the same number of electrons (isoelectronic) and the same number of occupied energy levels (3), but different numbers of protons.
The Al^{3+} ion has the greatest number of protons and the N^{3-} ion the least number of protons. This results in an increased attraction between the nucleus and the outer electrons so the ionic radius decreases.

Ionisation energy

- The first ionisation energy of an element is the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous ions.

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

- First ionisation energies are a measure of the attraction between the nucleus and the outermost electrons.



Two general trends can be seen from the above graph.

Down a group:

Ionisation energy decreases down a group.

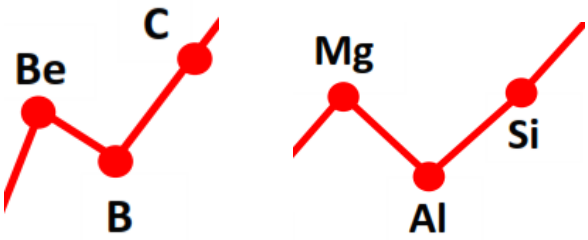
1. As the number of occupied energy levels increases, the valence electrons are further from the nucleus which results in a weaker electrostatic attraction (between nucleus and valence electrons).
2. Increased electron shielding by the inner electrons reduces the electrostatic attraction between nucleus and valence electrons.

Across a period:

Ionisation energy increases across a period (left to right).

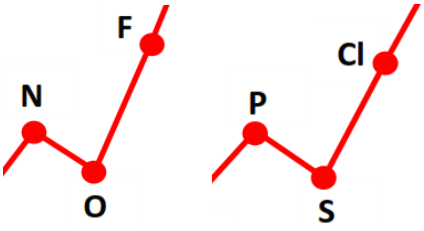
1. As nuclear charge increases across a period, the electrostatic attraction between the nucleus and valence electrons increases.
2. Atomic radius decreases across a period – valence electrons are closer to the nucleus which results in a stronger attraction between the nucleus and the valence electrons.

Discontinuities in ionisation energy across a period (Be to B and Mg to Al)

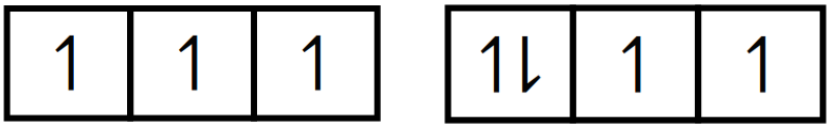


- Be has the electronic configuration $1s^2\ 2s^2$
- B has the electronic configuration $1s^2\ 2s^2\ 2p^1$
- Electrons in p orbitals are of higher energy and further from the nucleus than electrons in s orbitals, therefore they require less energy to remove.
- The same explanation can be applied for 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 – O and P – S



- N has the electronic configuration $1s^2\ 2s^2\ 2p^3$
- O 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) Define first ionisation energy.

The first ionisation energy is the energy required to remove 1 mole of electrons from 1 mole of atoms in the gaseous state (to form one mole of gaseous 1+ ions).

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

2) State and explain the trend in ionisation energy across period 3.

Ionisation energy increases across a period because of increasing nuclear charge and decreasing atomic radius. This results in a stronger electrostatic attraction between the positive nucleus and the valence electrons.

3) State and explain the trend in ionisation energy down group 1.

Ionisation energy decreases down a period because of increasing number of occupied energy levels – the valence electrons are further from nucleus which results in a weaker electrostatic attraction between the nucleus and the valence electrons. There is also an increase in electron shielding down a group.

4) Explain the reason for the decrease in ionisation energy between Mg and Al.

Mg has the electronic configuration $1s^2\,2s^2 2p^6\,3s^2$, Al has the electronic 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.

5) Explain the reason for the decrease in ionisation energy between P and S.

P has the electronic configuration $1s^2\,2s^2 2p^6\,3s^2\,3p^3$. S has the electronic 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.

Electronegativity

- Electronegativity is a measure of the ability of an atom to attract a bonding pair of electrons.
- Electronegativity is measured on a relative scale called the Pauling scale which assigns fluorine a value of 4.0 and francium a value of 0.7
- Electronegativity values of elements can be found in section 8 of the IB data booklet.

First ionization energy (kJ mol^{-1}) / Electron affinity (kJ mol^{-1}) (2nd EA / kJ mol^{-1}) / Element / Electronegativity																		
1312 −73 **H** 2.2																		2372 **He**
520 −60 **Li** 1.0	900 **Be** 1.6											801 −27 **B** 2.0	1086 −122 **C** 2.6	1402 **N** 3.0	1314 −141 (+753) **O** 3.4	1681 −328 **F** 4.0	2081 **Ne**
496 −53 **Na** 0.9	738 **Mg** 1.3											578 −42 **Al** 1.6	787 −134 **Si** 1.9	1012 −72 **P** 2.2	1000 −200 (+545) **S** 2.6	1251 −349 **Cl** 3.2	1520 **Ar**
419 −48 **K** 0.8	590 −2 **Ca** 1.0	633 −18 **Sc** 1.4	659 −8 **Ti** 1.5	651 −51 **V** 1.6	653 −64 **Cr** 1.7	717 **Mn** 1.6	762 −15 **Fe** 1.8	760 −64 **Co** 1.9	737 −112 **Ni** 1.9	745 −119 **Cu** 1.9	906 **Zn** 1.6	579 −41 **Ga** 1.8	762 −119 **Ge** 2.0	944 −78 **As** 2.2	941 −195 **Se** 2.6	1140 −325 **Br** 3.0	1351 **Kr**
403 −47 **Rb** 0.8	549 −5 **Sr** 1.0	600 −30 **Y** 1.2	640 −41 **Zr** 1.3	652 −88 **Nb** 1.6	684 −72 **Mo** 2.2	702 −53 **Tc** 2.1	710 −101 **Ru** 2.2	720 −110 **Rh** 2.3	804 −54 **Pd** 2.2	731 −126 **Ag** 1.9	868 **Cd** 1.7	558 −29 **In** 1.8	709 −107 **Sn** 2.0	831 −101 **Sb** 2.0	869 −190 **Te** 2.1	1008 −295 **I** 2.7	1170 **Xe** 2.6
376 −46 **Cs** 0.8	503 −14 **Ba** 0.9	538 −45 **La** 1.1	659 −1 **Hf** 1.3	728 −31 **Ta** 1.5	759 −79 **W** 1.7	756 −14 **Re** 1.9	814 −106 **Os** 2.2	865 −151 **Ir** 2.2	864 −205 **Pt** 2.2	890 −223 **Au** 2.4	1007 **Hg** 1.9	589 −36 **Tl** 1.8	716 −35 **Pb** 1.8	703 −91 **Bi** 1.9	812 −183 **Po** 2.0	−270 **At** 2.2	1037 **Rn**
393 −47 **Fr** 0.7	509 −10 **Ra** 0.9	499 −34 **Ac** 1.1																

- Electronegativity increases from left to right across a period for two reasons; the increase in nuclear charge and the decrease in atomic radius.
- Electronegativity decreases down a group because of increasing atomic radius (bonding electrons are further from the attraction of the nucleus).

Exercises:

1) Define electronegativity.
Electronegativity is a measure of the attraction of an atom for a bonding pair of electrons. Electronegativity is measured on the Pauling scale, which assigns fluorine the highest value and francium the lowest.

2) State and explain the trend in electronegativity across a period.
Electronegativity increases across a period (from left to right) because of increasing nuclear charge and decreasing atomic radius which results in an increased electrostatic attraction between the nucleus and bonding electrons.

3) State and explain the trend in electronegativity down a group.
Electronegativity decreases down a group because of increasing atomic radius - the bonding electrons are further from the nucleus, therefore there is a weaker attraction between the nucleus and the valence electrons.

Electron affinity

- The first electron affinity is the energy released when one mole of electrons is added to one mole of gaseous atoms to form one mole of 1^- ions.

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

- The second electron affinity corresponds to the addition of one mole of electrons to one mole of gaseous 1^- ions.

$$X^-_{(g)} + e^- \rightarrow X^{2-}_{(g)}$$

First ionization energy (kJ mol^{-1}) — Electron affinity (kJ mol^{-1}) (2nd EA / kJ mol^{-1}) — Element — Electronegativity

1312 −73 H 2.2																	2372 He
520 −60 Li 1.0	900 Be 1.6											801 −27 B 2.0	1086 −122 C 2.6	1402 N 3.0	1314 −141 (+753) O 3.4	1681 −328 F 4.0	2081 Ne
496 −53 Na 0.9	738 Mg 1.3											578 −42 Al 1.6	787 −134 Si 1.9	1012 −72 P 2.2	1000 −200 (+545) S 2.6	1251 −349 Cl 3.2	1520 Ar
419 −48 K 0.8	590 −2 Ca 1.0	633 −18 Sc 1.4	659 −8 Ti 1.5	651 −51 V 1.6	653 −64 Cr 1.7	717 Mn 1.6	762 −15 Fe 1.8	760 −64 Co 1.9	737 −112 Ni 1.9	745 −119 Cu 1.9	906 Zn 1.6	579 −41 Ga 1.8	762 −119 Ge 2.0	944 −78 As 2.2	941 −195 Se 2.6	1140 −325 Br 3.0	1351 Kr
403 −47 Rb 0.8	549 −5 Sr 1.0	600 −30 Y 1.2	640 −41 Zr 1.3	652 −88 Nb 1.6	684 −72 Mo 2.2	702 −53 Tc 2.1	710 −101 Ru 2.2	720 −110 Rh 2.3	804 −54 Pd 2.2	731 −126 Ag 1.9	868 Cd 1.7	558 −29 In 1.8	709 −107 Sn 2.0	831 −101 Sb 2.0	869 −190 Te 2.1	1008 −295 I 2.7	1170 Xe 2.6
376 −46 Cs 0.8	503 −14 Ba 0.9	538 −45 La 1.1	659 −1 Hf 1.3	728 −31 Ta 1.5	759 −79 W 1.7	756 −14 Re 1.9	814 −106 Os 2.2	865 −151 Ir 2.2	864 −205 Pt 2.2	890 −223 Au 2.4	1007 Hg 1.9	589 −36 Tl 1.8	716 −35 Pb 1.8	703 −91 Bi 1.9	812 −183 Po 2.0	−270 At 2.2	1037 Rn
393 −47 Fr 0.7	509 −10 Ra 0.9	499 −34 Ac 1.1															

Electron affinity decreases down a group

- The increased nuclear charge down a group is offset by increased electron shielding.
- The greater the distance between the nucleus and the outer energy level, the weaker the electrostatic attraction and the less energy is released when an electron is added to the atom.

Exercises:

1) Define first electron affinity.

The first electron affinity is the energy released when one mole of electrons is added to one mole of gaseous atoms to form one mole of 1^- ions.

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

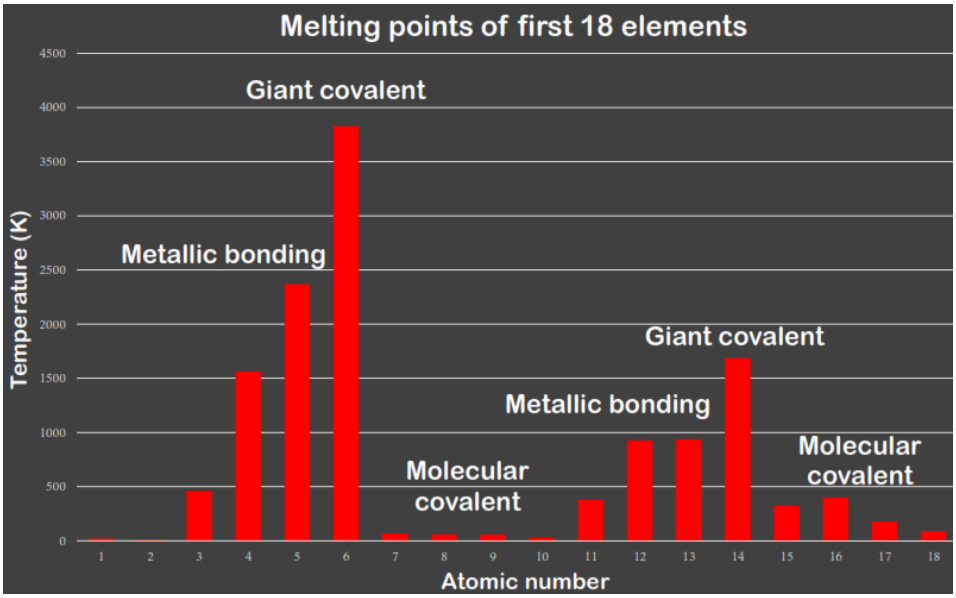
2) State and explain the trend in electron affinity down a group.

The greater the distance between the nucleus and the outer energy level, the weaker the electrostatic attraction and the less energy is released when an electron is added to the atom. The second factor is increased electron shielding as the number of occupied energy levels increases.

3) Explain why the second electron affinity of oxygen is endothermic.

The second electron affinity of oxygen is positive (endothermic) because of the repulsion when an electron is added to a negative ion.

Melting point



- Melting point depends on the type of bonding (covalent, ionic or metallic) and structure (ionic lattice, molecular covalent, giant covalent, or metallic structures).
- Melting point increases across a period as the strength of the metallic bond increases.
- It reaches a peak at carbon and silicon, both of which have giant covalent structures.
- Melting points then decrease with elements that have molecular covalent structures.

1) What are two factors that determine the melting point of an element?
The two factors that determine the melting point of an element are the type of bonding (metallic, covalent or ionic) and the structure (ionic lattice, giant covalent, molecular covalent or metallic structure).

2) Describe and explain the general trend in melting point in the metals in periods 2 and 3.
The melting point of the 3 first metals in periods 2 and 3 increases as the strength of the metallic bond increases.

3) Explain the reason for the peak in melting point in C (carbon) and Si (silicon).
The reason for the peak in melting point in C and Si is related to their structures – they both have giant covalent structures which have high melting points.

4) Describe and explain the trend in melting point for the elements after C and Si in periods 2 and 3.
After C and Si in periods 2 and 3, there is a big decrease in melting point – this is related to the structures – the elements N to F and P to Cl have simple molecular structures with weak intermolecular forces between their molecules.

Metallic character

- The metallic character of an element can be defined as how easily an atom loses electrons.
- Metals tend to lose their outer electrons to form positive ions and non-metals tend to gain electrons to form negative ions.
- Metallic character decreases as you move from left to right across a period in the periodic table; increasing nuclear charge and decreasing atomic radius across a period means the outer electrons are held more tightly.
- Metallic character increases down a group in the periodic table as the outer electrons become easier to remove as the atomic radius increases.

Exercises:

1) State the property used to classify elements as metallic.
   The ease at which they lose electrons to form positive ions. Metals tend to have low ionisation energies which means the outer electrons are easy to remove.

2) Describe and explain the trend in metallic character across a period and down a group.
   Metallic character decreases across a period as nuclear charge increases and atomic radius decreases. Metallic character increases down a group as atomic radius increases and the outer electrons require less energy to remove.

Acid – base character of the period 3 oxides

Complete the table below showing the acid-base properties of period 3 oxides:

Formula and state at room temperature	**$Na_2O(s)$**	**$MgO(s)$**	**$Al_2O_3(s)$**	**$SiO_2(s)$**	**$P_4O_{10}(s)$/ $P_4O_6(s)$**	**$SO_3(l)$/ $SO_2(g)$**	**$Cl_2O_7(l)$/ $Cl_2O(g)$**
Acid–base Character	Basic		Amphoteric	Acidic			

- Metal oxides are basic.
- Aluminium oxide, an ionic oxide with some covalent character is amphoteric (can act as both an acid and a base).
- The remaining oxides have acidic properties.

Exercises:

1) Describe and explain the trend in the bonding of the oxides across period 3.

The bonding changes from ionic to covalent across period 3. The type of bonding is related to the difference in electronegativity between oxygen and the group 3 element. A difference of greater than 1.8 units on the Pauling scale indicates an ionic bond. As the difference in electronegativity decreases, the bonding changes to covalent.

2) Describe the trend in acid-base character of the period 3 oxides.

Na and Mg form basic oxides, Al_2O_3 is amphoteric and the remaining elements (Si, P,S and Cl) form acidic oxides.

3) Write equations for the reactions of Na_2O, MgO, SO_3, P_4O_{10} and NO_2 with water.

Basic:

$Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)}$

Acidic:

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

$P_4O_{10(s)} + H_2O_{(l)} \rightarrow 4H_3PO_{4(aq)}$

$2NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)} + HNO_{2(aq)}$

Chemical properties of group 1 and group 17 elements

- The chemical properties of an element are determined by its electron configuration.
- Elements of the same group have similar chemical properties as they have the same number of electrons in their outer shells.

Group 1 – the alkali metals

- Alkali metals are very reactive metals – they are shiny when cut, but quickly react with oxygen in air to form a layer of oxide.
- They are stored in oil to prevent the reaction with air.
- They have low densities – Li, Na and K float on water.
- They have low melting points – the melting point decreases down the group as the metallic bond weakens.
- Alkali metals react by losing their one valence electron to form positive ions.

3 **Li** 6.94
11 **Na** 22.99
19 **K** 39.10
37 **Rb** 85.47
55 **Cs** 132.91
87 **Fr** (223)

- Reactivity increases down the group.
- The atomic radius increases down a group as the number of occupied energy levels increases.
- Less energy is required to remove the outer electron (ionisation energy decreases).

Reactions with water

- The group 1 metals react with water to form hydrogen gas and the metal hydroxide.
- The resulting solution is alkaline (pH 12-14)

$$2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$$

Exercises:

1) Explain why the alkali metals are stored in oil.
The alkali metals are stored in oil to prevent the reaction with the oxygen in the air.

2) Which of the 3 alkali metals have a density of less than 1 gcm^{-3}?
Lithium, sodium and potassium all have densities lower than 1 gcm^{-3} (to float on water, a substance must have a density of less than the density of water which is 1 gcm^{-3}).

3) Describe and explain the trend in melting point down group 1.
Melting point decreases down group 1 as the metallic bond get weaker.

4) Describe and explain the trend in reactivity down group 1.

- Reactivity increases down the group.
- The distance between the nucleus and the outer electron increases down a group as the number of occupied energy levels increases (atomic radius increases).
- Less energy is required to remove the outer electron (ionisation energy decreases)

5) Write a balanced chemical equation, complete with state symbols for the reaction of potassium and water.

$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$

6) What pH would you expect the resulting solution to be? Explain your answer.
The resulting solution would be pH 12-14.

Group 17 – the halogens

- The halogens (salt formers) are very reactive non-metal elements.
- They are coloured – Cl_2 is a yellow-green gas, Br_2 is a reddish-brown liquid, I_2 is a purple solid.
- They show a change from gases (F_2, Cl_2), to liquid (Br_2) to solid (I_2) as the molar mass increases down the group.
- The halogens are diatomic – they form molecules composed of two atoms.

9 **F** 19.00
17 **Cl** 35.45
35 **Br** 79.90
53 **I** 126.90
85 **At** (210)

- Reactivity decreases down the group.
- The halogens react by gaining one electron to form negative ions.
- As the atomic radius increases down the group (more occupied energy levels) the attraction for the extra electron decreases.
- The increasing number of occupied energy levels also increases the electron shielding between the nucleus and the outer energy level.

Reaction with group 1 metals

- The halogens react with group 1 metals to form ionic compounds. The halogen atom gains one electron from the group 1 metal to form a halide ion (X^-)
- The resulting ions both have the electron configuration of a noble gas.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Exercises:

1) Why are the group 17 elements also known as the halogens?
The term halogens mean 'salt formers'. The halogens from salts with the group 1 metals.

2) Describe the change in state of the group 17 elements down the group (under standard conditions).
F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid (under standard conditions). Molecular mass increases down the group which increases the strength of the intermolecular forces between the molecules.

3) Explain the meaning of the term *diatomic*.
The term 'diatomic' means that a molecule consists of 2 atoms (usually the same atoms, as in the halogens) bonded together.

4) Describe and explain the reactivity of the halogens down group 17.
The reactivity decreases down a group. As the atomic radius increases down the group (more occupied energy levels) the attraction for the extra electron decreases. The increasing number of occupied energy levels also increases the electron shielding between the nucleus and the outer energy level. Fluorine has a small atomic radius and weak electron shielding between the nucleus and the outer energy level which results in a strong attraction for an extra electron.

5) Write a balanced symbol equation for the reaction of potassium and bromine.
$2K_{(s)} + Br_{2(l)} \rightarrow 2KBr_{(s)}$