Structure 3.1 Answers

IB CHEMISTRY SL

Structure 3.1.1 and 3.1.2

Understandings:

- The periodic table consists of periods, groups and blocks (3.1.1).
- The period number shows the outer energy level that is occupied by electrons. Elements in a group have a common number of valence electrons (3.1.2).

Learning outcomes:

- Identify the positions of metals, metalloids and non-metals in the periodic table (3.1.1).
- Deduce the electron configuration of an atom up to Z = 36 from the element's position in the periodic table and vice versa (3.1.2).

Additional notes:

- The four blocks associated with the sublevels s, p, d, f should be recognized.
- A copy of the periodic table is available in the data booklet.
- Groups are numbered from 1 to 18.
- The classifications "alkali metals", "halogens", "transition elements" and "noble gases" should be known.

Linking question(s):

- Structure 1.2 How has the organization of elements in the periodic table facilitated the discovery of new elements?

Groups and periods on 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 valence electrons.
- A period is a horizontal row - elements in the same period have the same number of occupied 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** 83.80
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 **Nh** (286)	114 **Fl** (289)	115 **Mc** (288)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (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

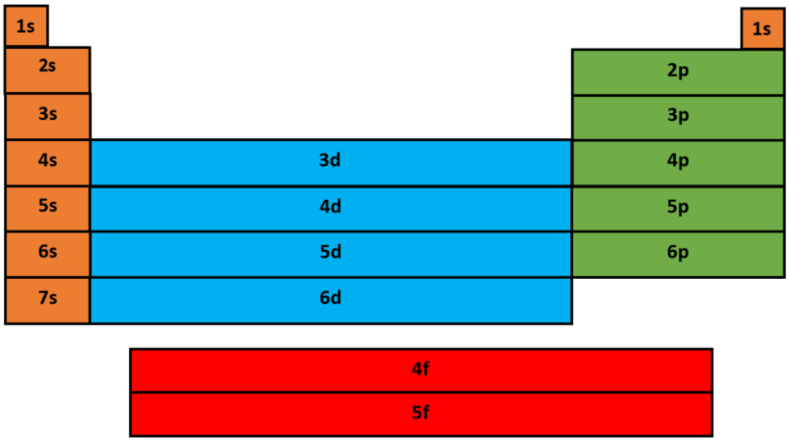
- Group 1 – Alkali metals (Li, Na, K, Rb, Cs, Fr)
- Group 2 - Alkaline Earth metals (Be, Mg, Ca, Sr, Ba, Ra)
- Group 17 – Halogens (salt formers) (F, Cl, Br, I, At, Ts)
- Group 18 – Noble gases (He, Ne, Ar, Kr, Xe, Rn, Og)
- Groups 3 – 11: Transition elements
- La – Lu: Lanthanides (lanthanoids)
- Ac – Lr: Actinides (actinoids)

Exercise: Outline what can be deduced about an element from its group number and period number.

The group number tells us the number of valence electrons and the period number tells us the number of occupied energy levels.

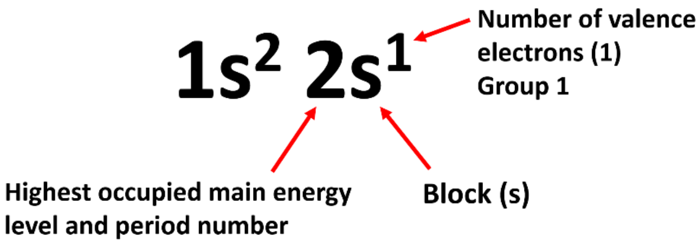
Blocks of the periodic table

- 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 lithium, Li, is $1s^2\ 2s^1$ – what can we determine about its position on the periodic table?



- Lithium is an s-block element in group 1 and period 2.

Exercises: For each element and its electron configuration, determine to which block, group and period it belongs.

1. C $1s^2\ 2s^2\ 2p^4$ Carbon is a p block element in group 16 and period 2.

2. Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$ Magnesium is an s block element in group 2 and period 3.

3. S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ Sulfur is a p block element in group 16 and period 3.

4. Br $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$ Bromine is a p block element in group 17 and period 4.

Structure 3.1.3
Understandings:

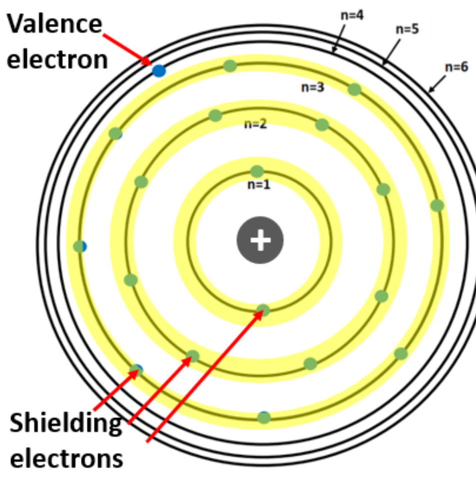
- Periodicity refers to trends in properties of elements across a period and down a group.

Learning outcomes:

- Explain the periodicity of atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.

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 valence electrons do not experience the full attraction from the nucleus as they are shielded by the inner electrons.
- Effective nuclear charge (Z_{eff}) is the net positive charge experienced by the 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.

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

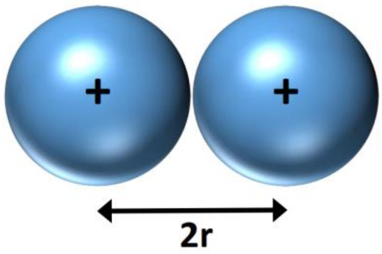
Z is the atomic number
S is the number of shielding electrons

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

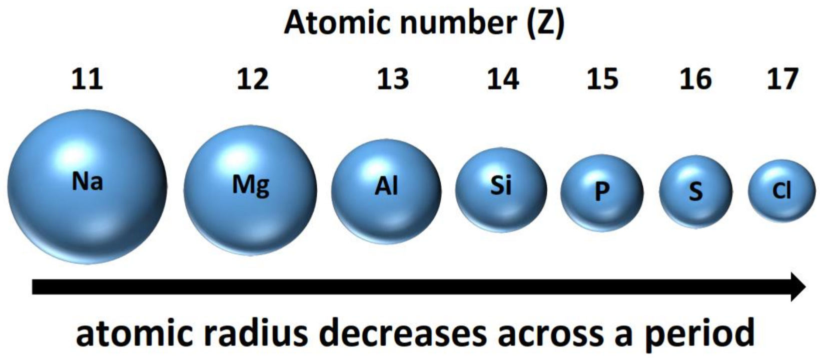
$Z_{eff} = Z - S = 19 - 18 = +1$

Atomic radius

- The atomic radius is measured as half the distance between neighboring nuclei.
- Atomic and ionic radii values can be found in section 10 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 $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

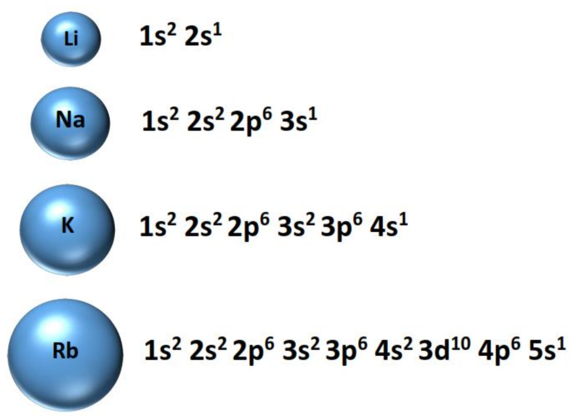
S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

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

Explanation for the trend:

- Nuclear charge increases across a period.
- Electrons are added to the same main energy level (electron shielding remains constant across a period).
- The electrons are pulled closer to the nucleus, therefore, the atomic radius decreases.

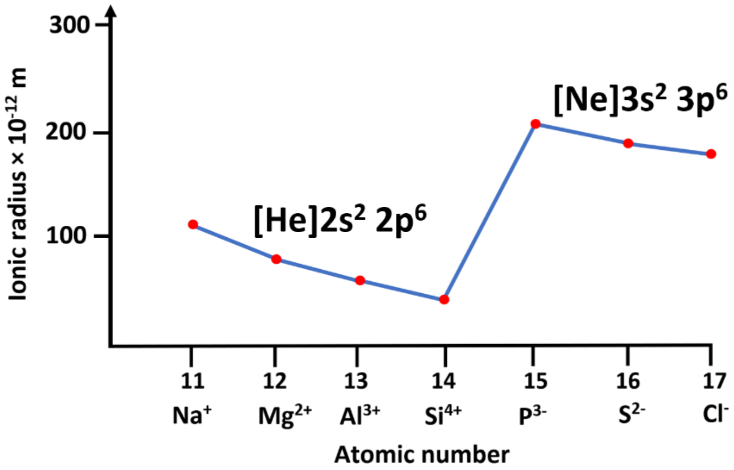
Trend down a group



Explanation for the trend:
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 the positive ions, increases for 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 ions N^{3-} to Al^{3+}.

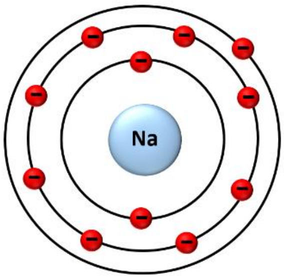
Ion	Atomic number	Electron configuration	Ionic radius (× 10^{-12} m)
N^{3-}	7	$1s^2\ 2s^2\ 2p^6$	146
O^{2-}	8	$1s^2\ 2s^2\ 2p^6$	140
F^-	9	$1s^2\ 2s^2\ 2p^6$	133
Na^+	11	$1s^2\ 2s^2\ 2p^6$	102
Mg^{2+}	12	$1s^2\ 2s^2\ 2p^6$	72
Al^{3+}	13	$1s^2\ 2s^2\ 2p^6$	54

- 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+}.

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

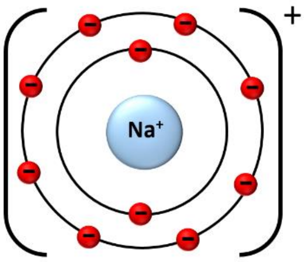
Positive ions (cations)

Sodium atom (160×10^{-12} m)



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

Sodium ion (102×10^{-12} m)

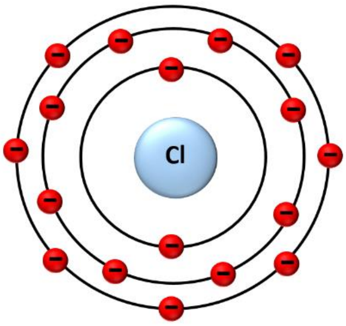


$1s^2\ 2s^2\ 2p^6$

- 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 electrons.

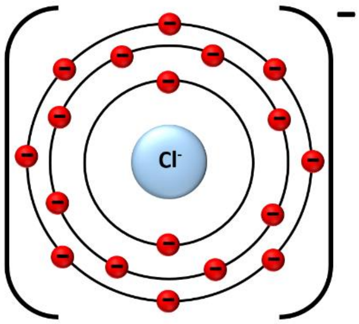
Negative ions (anions)

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 their parent atoms.
- Negative ions have more electrons than protons which results in a decreased attraction between the nucleus and 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 one occupied energy level, compared to two 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.

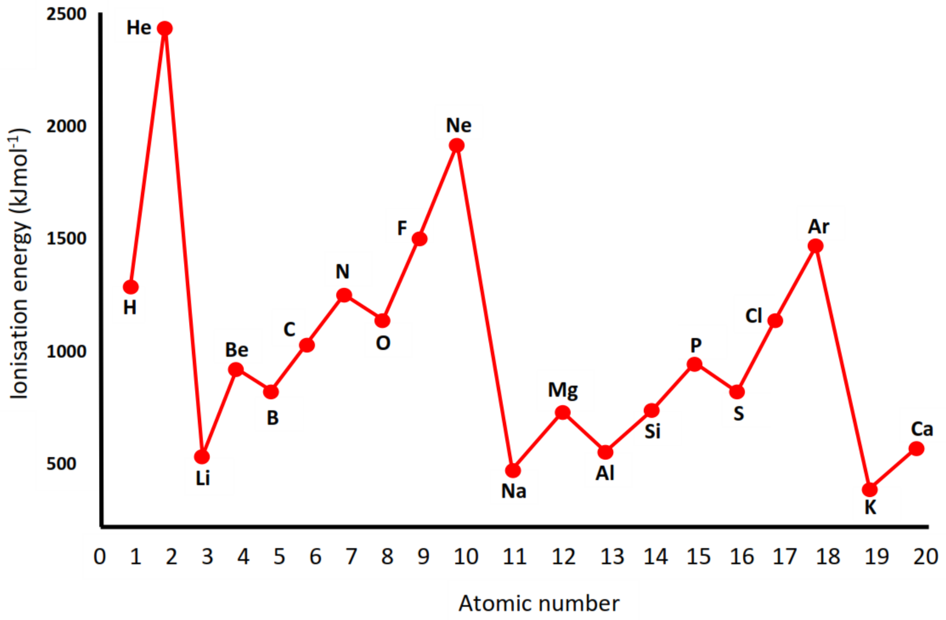
5. Arrange the following in order of increasing atomic radius (smallest first): Cl, Si, Na. Explain your reasoning.
Cl < Si <Na
Na, Si and Cl are all period 3 elements and atomic radius decreases from left to right across a period because of increasing nuclear charge and the electron shielding remains 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 the above ions have the same number of electrons (isoelectronic) and the same number of occupied energy levels (three), 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.

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



Two general trends can be seen from the above graph.

Down a group:

Ionisation energy decreases down a group.

- 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.

Across a period:

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

- As nuclear charge increases across a period, the electrostatic attraction between the nucleus and valence electrons increases.
- 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.

- Note that there are discontinuities in the trend in first ionisation energy across a period - these are covered in more detail in Structure 3.1.7 (HL).

Exercises:

1. Outline what is meant by 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.

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.

Trend and explanation

- 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. 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.

2. 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.

3. Explain why fluorine has the highest electronegativity value.

Fluorine has a high nuclear charge and a small atomic radius – these two factors mean that it has the highest electronegativity value.

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.

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

- Electron affinity generally decreases down a group and increases across a period.
- 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 less energy is released when an electron is added to the atom.

Exercise: State and explain the general trend in electron affinity in the periodic table.
Electron affinity generally becomes less exothermic down a group and more exothermic across a period, although there are exceptions.

Metallic and non-metallic character

- Metallic character is the tendency of an element to lose electrons and form positive ions.
- 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 from left to right across a period and increases down a group in the periodic table.
- Non-metallic character is the tendency of an element to accept electrons and form negative ions.
- Non-metallic character increases from left to right across a period and decreases from top to bottom in the periodic table.

Exercises:

1. Outline the property used to classify elements as metallic or non-metallic.
   Metallic character is how easily an atom loses electrons and non-metallic character is how easily an atom gains electrons.

2. Describe and explain the trend in metallic and non metallic character across a period and down a group.
   Metallic character decreases from left to right across a period and increases down a group in the periodic table. Non-metallic character increases from left to right across a period and decreases from top to bottom in the periodic table.

Structure 3.1.4

Understandings:

- Trends in properties of elements down a group include the increasing metallic character of group 1 elements and decreasing non-metallic character of group 17 elements.

Learning outcomes:

- Describe and explain the reactions of group 1 metals with water, and of group 17 elements with halide ions.

Group 1 elements – 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 atomic radius increases.
- The metallic character of the group 1 metals increases down the group as ionisation energy decreases.

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.
- Alkali metals react by losing their one valence electron to form positive ions.
- The atomic radius increases down a group as the number of occupied energy levels increases.
- Ionisation energy decreases so reactivity increases.

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. Describe and explain the trend in melting point down group 1.
   Melting point decreases down group 1 as atomic radius increases which causes the metallic bond to become weaker.

3. Describe and explain the trend in reactivity down group 1.
   Reactivity increases down the group as a result of the decreasing ionisation energy.

4. Write a balanced chemical equation, complete with state symbols for the reaction of potassium and water.
   $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

5. What pH would you expect the resulting solution to be? Explain your answer.
   The resulting solution has a pH of 12-14 because the metal hydroxide produced in the reaction is a strong base.

Group 17 elements – the halogens

- The halogens (salt formers) are very reactive non-metal elements.
- They are coloured – F_2 is a pale-yellow gas, 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 which results in stronger intermolecular forces between the molecules.
- The halogens are diatomic – they form molecules composed of two atoms bonded together.
- Then non-metallic character of the group 17 elements decreases down the group.

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, the attraction for the electron decreases.

Reaction with group 1 metals

- The halogens react with group 1 metals to form ionic compounds.

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

Exercises:

1. Describe and explain the change in state of the group 17 elements down the group.
   F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid. Molar mass increases down the group which increases the strength of the intermolecular forces between the molecules.
2. Explain the meaning of the term *diatomic*.
   The term 'diatomic' means that a molecule consists of two atoms (sometimes the same atoms, as in the halogens, but they can be different) bonded together.
3. Describe the reactivity of the halogens.
   The reactivity decreases down a group. Fluorine is the most reactive and iodine the least.
4. Write a balanced symbol equation for the reaction of potassium and bromine.
   $2K(s) + Br_2(l) \rightarrow 2KBr(s)$

Structure 3.1.5

Understandings:

- Metallic and non-metallic properties show a continuum. This includes the trend from basic metal oxides through amphoteric to acidic non-metal oxides.

Learning outcomes:

- Deduce equations for the reactions with water of the oxides of group 1 and group 2 metals, carbon and sulfur.

Additional notes:

- Include acid rain caused by gaseous non-metal oxides, and ocean acidification caused by increasing CO_2 levels.

Linking questions:

- Structure 2.1, 2.2 How do differences in bonding explain the differences in the properties of metal and non-metal oxides?

Acid–base character of the period 3 oxides

- The table below shows the acid-base properties of the 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.
- Metals oxides react with water to produce a metal hydroxide.

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$$

- 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 the trend in the acid-base character of the period 3 oxides.
Sodium and magnesium form basic oxides, aluminium oxide is amphoteric and the remaining oxides are acidic. So, the trend changes from basic, to amphoteric to acidic.

2. Write equations for the reactions of Li_2O and MgO with water.
Na2O(s) + $H_2O(l) \rightarrow 2NaOH(aq)$
$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$

Acid deposition

- Unpolluted rainwater is naturally acidic with a pH of 5.6
- The equation below shows that carbon dioxide dissolves in water to form carbonic acid H_2CO_3.

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

- Acid deposition has a pH of less than 5.0

Sources of acidic gases

- Sulfur oxides: combustion of coal that contains sulfur and volcanoes.
- Nitrogen oxides: internal combustion engines and lightning.

Formation of acid rain

Sulfur oxides

- Sulfuric acid (H_2SO_4) and sulfurous acid (H_2SO_3) are formed by the combustion of coal that contains high levels of sulfur.
- Sulfur (S) burns in oxygen (O_2) to form sulfur dioxide (SO_2).

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

- SO_2 dissolves in water to form sulfurous acid (H_2SO_3).

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

- SO_2 can react with O_2 to form sulfur trioxide (SO_3).

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

- SO_3 dissolves in water to form H_2SO_4.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Nitrogen oxides

- Nitric acid (HNO_3) and nitrous acid (HNO_2) are formed from the reaction of nitrogen (N_2) and oxygen (O_2) at high temperatures in internal combustion engines.
- Nitrogen (N_2) and oxygen (O_2) react at high temperatures in internal combustion engines to form nitrogen monoxide (NO).

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

- NO reacts with O_2 to form nitrogen dioxide (NO_2).

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

- NO_2 dissolves in water to form HNO_3 and HNO_2.

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

Exercises:

1. Deduce an equation to show why rainwater is naturally acidic.
   $CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

2. State a natural and anthropogenic source of sulfur oxides and nitrogen oxides.

   Sulfur oxides
   Anthropogenic source: combustion of coal containing sulfur
   Natural source: volcanoes
   Nitrogen oxides
   Anthropogenic source: reaction of nitrogen with oxygen at high temperatures in internal combustion engines
   Natural source: lightning

3. Acid rain has a pH of less than 5.0. Explain how the burning of coal can lead to the formation of acid rain.
   Coal, containing sulfur, is burned in a power station to produce heat.
   The sulfur reacts with oxygen to form sulfur dioxide
   $S(s) + O_2(g) \rightarrow SO_2(g)$
   The sulfur dioxide then reacts with water to form sulfurous acid
   $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$
   The sulfur dioxide can react with oxygen to form sulfur trioxide which then reacts with water to form sulfuric acid.
   $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$
   $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

4. Outline the process responsible for the production of acid rain from nitrogen oxides.

Nitrogen reacts with oxygen at high temperatures inside an internal combustion engine to produce nitrogen monoxide.

$N_2(g) + O_2(g) \rightarrow 2NO(g)$

The nitrogen monoxide reacts with oxygen to form nitrogen dioxide.

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

The nitrogen dioxide then reacts with water to form nitrous and nitric acids.

$2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$

Ocean acidification

- Ocean acidification is the process by which increasing levels of atmospheric CO_2 cause the pH of the oceans to decrease.
- Approximately 30% of anthropogenic carbon dioxide is absorbed by the oceans (carbon sink).
- CO_2 dissolves in sea water.

$$CO_2(g) \rightleftharpoons CO_2(aq)$$

- A heterogeneous equilibrium exists between concentrations of gaseous carbon dioxide in the atmosphere and aqueous carbon dioxide dissolved in the oceans.
- The CO_2 dissolves in sea water to produce carbonic acid (H_2CO_3) which is a weak acid.

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

- Carbonic acid (H_2CO_3) partially dissociates in water to produce $H^+(aq)$.
- Increasing levels of atmospheric CO_2 cause the equilibrium position to shift the right.

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

- The increasing $[H^+]$ causes the pH of the water to decrease.
- Since the beginning of the industrial revolution, the pH of the oceans has decreased by 0.1 pH units.
- Continued acidification of the oceans could have harmful effects on marine organisms.

Exercise: Describe how increasing concentrations of atmospheric CO_2 could decrease the pH of ocean water using an equation to support your answer.

$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$
$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$ **or** $H_2CO_3(aq) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$
Increasing concentrations of CO_2 shift the equilibrium position to the right, decreasing the pH.

Structure 3.1.6

Understandings:

- The oxidation state is a number assigned to an atom to show the number of electrons transferred in forming a bond. It is the charge that atom would have if the compound were composed of ions.

Learning outcomes:

- Deduce the oxidation states of an atom in an ion or a compound.
- Explain why the oxidation state of an element is zero.

Additional notes:

- Oxidation states are shown with a + or – sign followed by the Arabic symbol for the number, e.g., +2, –1.
- Examples should include hydrogen in metal hydrides (–1) and oxygen in peroxides (–1).
- The terms “oxidation number” and “oxidation state” are often used interchangeably, and either term is acceptable in assessment.
- Naming conventions for oxyanions use oxidation numbers shown with Roman numerals, but generic names persist and are acceptable. Examples include nitrate, nitrite, sulfate, sulfite.

Linking questions:

- Reactivity 3.2 How can oxidation states be used to analyse redox reactions?

Oxidation states

- The oxidation state is the hypothetical charge an atom would have if the bonds were assumed to be 100% ionic with no covalent character.
- Oxidation states are written with the + or – first followed by the number (+2, not 2+).

Assigning oxidation states

- The rules to determine the oxidation state are shown in the table below, together with examples.

Rules for determining oxidation states	
1	Elements are assigned an oxidation state of zero. Examples: Fe(s), Cu(s), Zn(s), O_2(g), Br_2(l), Cl_2(g) and N_2(g) are all elements and have oxidation states of zero.
2	The sum of the oxidation states of the atoms in a compound must be equal to zero. Example: In H_2O, the oxidation state of the O is −2 and the H is +1. The sum of the oxidation states is (−2 + (2 × +1)) = 0
3	The charge on an ion is numerically equal to its oxidation state. Examples: The oxidation state of the Mg^{2+} ion is +2. The oxidation state of the S^{2-} ion is −2.
4	Hydrogen in compounds is assigned an oxidation state of +1 except in certain metal hydrides (e.g. NaH) in which it is −1. Examples: In methane, CH_4, the hydrogen has an oxidation state of +1 and the carbon is −4. In NaH, the Na has an oxidation state of +1 and the H is −1.
5	Fluorine in compounds is always assigned an oxidation state of −1.
6	Oxygen in a compound is assigned an oxidation state of −2 unless it is combined with fluorine (for example OF_2) or in a peroxide (H_2O_2). Examples: In OF_2, the F has an oxidation state of −1 and the O is +2. In H_2O_2, the H is +1 and the O is −1.
7	Chlorine in a compound has an oxidation state of −1 unless it is combined with oxygen or fluorine. Example: In Cl_2O, the oxidation state of the Cl is +1.
8	For a polyatomic ion (molecular ion) the sum of the oxidation states must equal the charge on the ion. Example: In the SO_4^{2-} ion, the oxidation state of the S is +6 and the O is −2. The sum of the oxidation states is (+6 + (2 × −2)) = −2.

Oxidation states of polyatomic ions

Name	Formula	Oxidation states
Nitrate ion	NO_3^-	N +5
Nitrite ion	NO_2^-	N +3
Sulfate ion	SO_4^{2-}	S +6
Sulfite ion	SO_3^{2-}	S +4

Oxidation numbers

- Oxidation states can be represented by a Roman numeral **-** note that these are actually called oxidation numbers but are used interchangeably with oxidation state.
- Transition elements such as copper and iron can have variable oxidation states/numbers.

Examples:

Cu_2O copper(I) oxide

CuO copper(II) oxide

$FeCl_2$ Iron(II) chloride

$FeCl_3$ Iron(III) chloride

Exercises:

1. Deduce the oxidation states of the following:

They are all elements, therefore the oxidation state is zero.

$O_{2(g)}$ 0	$S_{8(s)}$ 0	$F_{2(g)}$ 0
$N_{2(g)}$ 0	$Al_{(s)}$ 0	$Fe_{(s)}$ 0
$Cu_{(s)}$ 0	$Mn_{(s)}$ 0	$Na_{(s)}$ 0

2. Deduce the oxidation states of the following ions:

For monoatomic ions, the oxidation state is the same as the charge on the ion.

$F^-_{(aq)}$ -1	$Cl^-_{(aq)}$ -1	$Na^+_{(aq)}$ +1
$Mg^{2+}_{(aq)}$ +2	$Al^{3+}_{(aq)}$ +3	$Fe^{2+}_{(aq)}$ +2
$Cu^+_{(aq)}$ +1	$Mn^{2+}_{(aq)}$ +2	$O^{2-}_{(aq)}$ -2

3. Deduce the oxidation states of the species in bold in the following compounds:

All are neutral compounds, so the sum of the oxidations states is zero.

$CO_{2(g)}$ +4	$NH_{3(g)}$ -3	$H_2SO_{3(aq)}$ +4
$Fe_2O_{3(s)}$ +3	$SiF_{4(g)}$ +4	$KIO_{3(aq)}$ +5
$Na_2SO_{4(s)}$ +6	$H_2O_{2(l)}$ -1	$Na_2Cr_2O_{7(s)}$ +6

4. Deduce the oxidation state of the species in bold in the following polyatomic ions:

The sum of the oxidation states is equal to the charge on the ion.

CO_3^{2-} +4	NO_3^- +5	SO_4^{2-} +6
PO_4^{3-} +5	NO_2^- +3	SO_3^{2-} +4

5. Deduce the oxidation state of the metal ion in the following.

The Roman numerals are oxidation numbers.

iron(II) oxide +2	manganese(IV) oxide +4	manganate(VII) ion +7
chromium(III) oxide +3	copper(I) chloride +1	copper(II) chloride +2