

Structure 3.1

IB CHEMISTRY SL

<p>25 Mn Manganese 54.938045</p>	<p>16 S Sulfur 32.065</p>	<p>J</p>	<p>6 C Carbon 12.0107</p>	<p>2 He Helium 4.002602</p>	<p>25 Mn Manganese 54.938045</p>
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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																	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)	

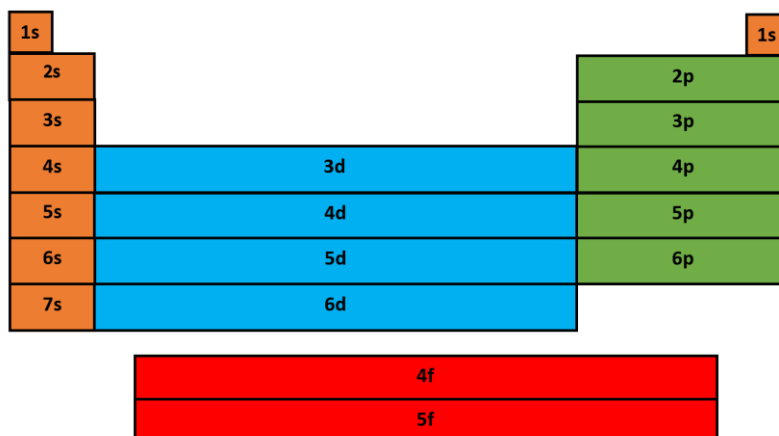
Names of the groups in the periodic table

- 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.

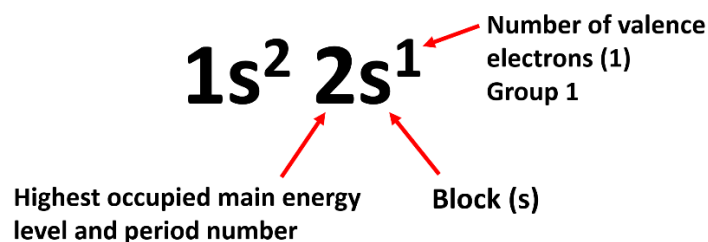
Blocks of the periodic table

- 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 $1s^2 2s^2 2p^4$
2. Mg $1s^2 2s^2 2p^6 3s^2$
3. S $1s^2 2s^2 2p^6 3s^2 3p^4$
4. Br $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Structure 3.1.3

Understandings:

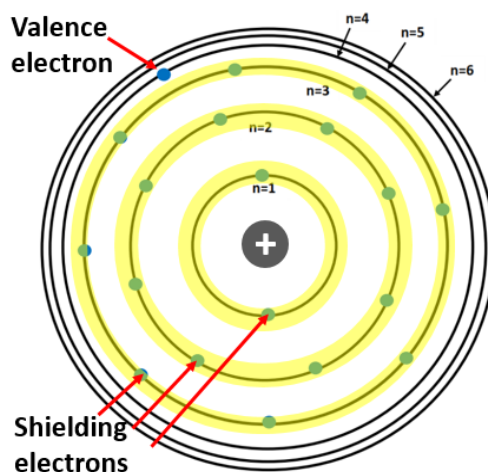
- Periodicity refers to trends in properties of elements across a period and down a group.

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.

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

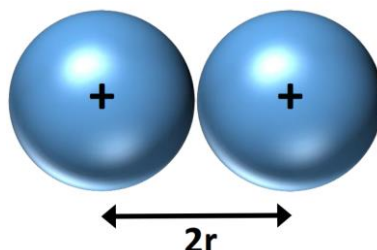
Z is the atomic number

S is the number of shielding electrons

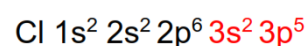
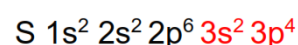
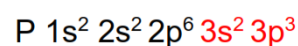
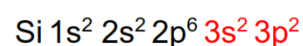
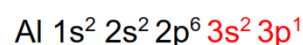
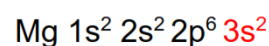
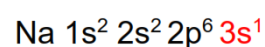
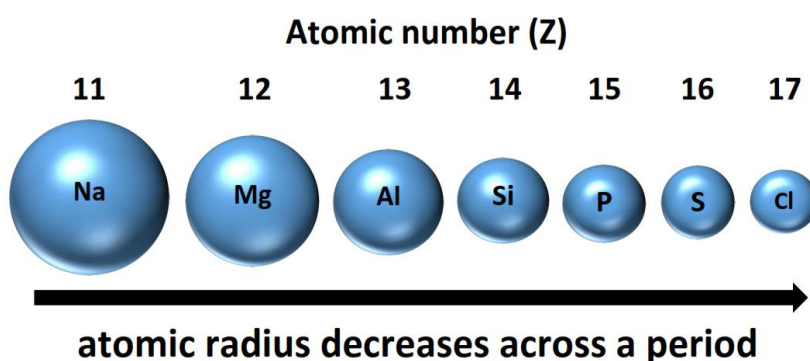
Exercise: Determine the effective nuclear charge for the above atom (potassium).

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.



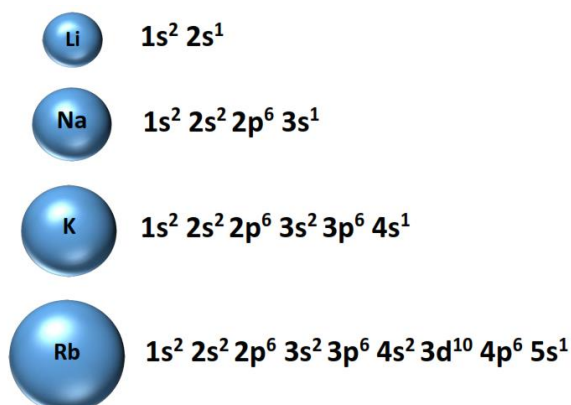
Trend across a period



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



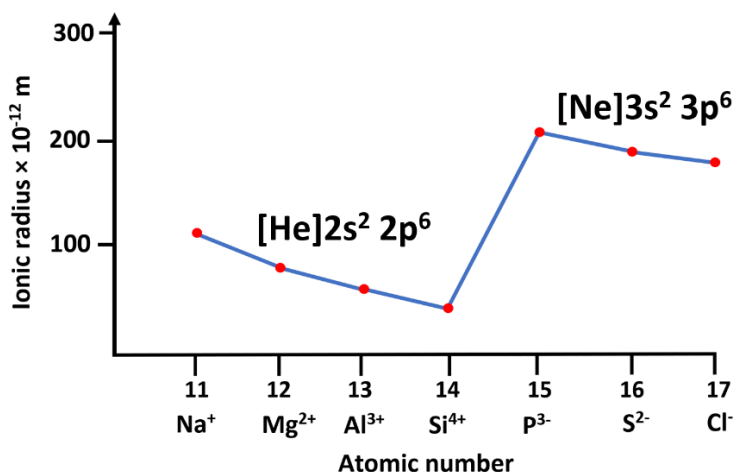
Explanation for the trend:

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.



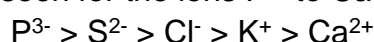
- The first four positive ions have two occupied energy levels and the negative ions have three occupied energy levels, therefore, the ionic radius increases.

Trend for isoelectronic ions

- The table below shows data for the ions N³⁻ to Al³⁺.

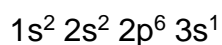
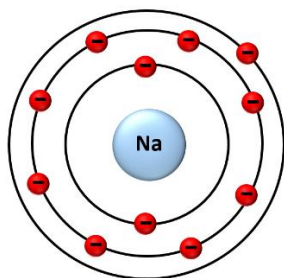
Ion	Atomic number	Electron configuration	Ionic radius ($\times 10^{-12}$ m)
N ³⁻	7	1s ² 2s ² 2p ⁶	146
O ²⁻	8	1s ² 2s ² 2p ⁶	140
F ⁻	9	1s ² 2s ² 2p ⁶	133
Na ⁺	11	1s ² 2s ² 2p ⁶	102
Mg ²⁺	12	1s ² 2s ² 2p ⁶	72
Al ³⁺	13	1s ² 2s ² 2p ⁶	54

- 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 P³⁻ to Ca²⁺.

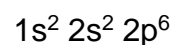
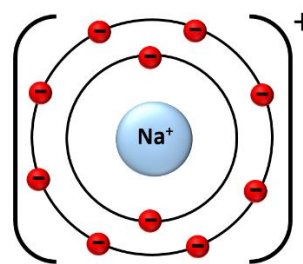


Positive ions (cations)

Sodium atom (160×10^{-12} m)



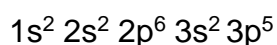
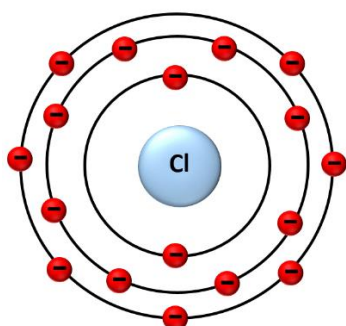
Sodium ion (102×10^{-12} m)



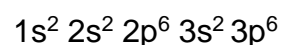
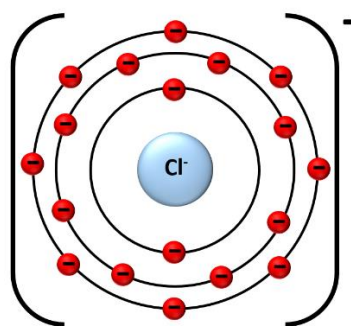
- Positive ions are smaller than their parent atoms.
- Positive ions have fewer occupied energy levels than their parent atoms.
- Positive ions have more protons than electrons which results in an increased attraction between the nucleus and electrons.

Negative ions (anions)

Chlorine atom (100×10^{-12} m)



Chloride ion (181×10^{-12} m)



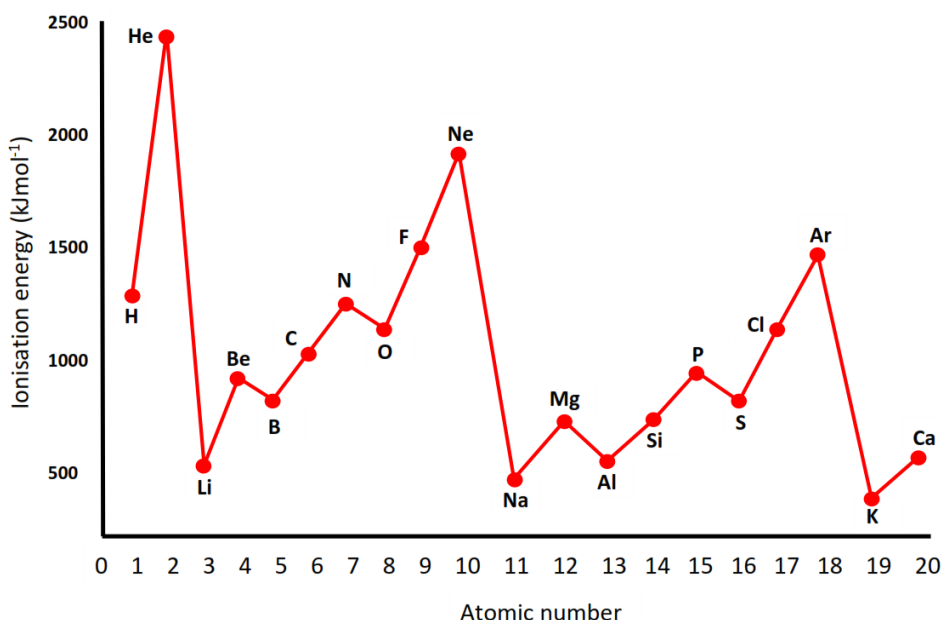
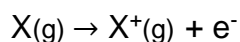
- Negative ions are bigger than their parent atoms.
- Negative ions have more electrons than protons which results in a decreased attraction between the nucleus and electrons.
- The addition of electrons to the outer energy level causes extra repulsion between the valence electrons, making the ion larger.

Exercises:

1. State and explain the trend in atomic radius down a group.
2. State and explain the trend in atomic radius across a period.
3. State and explain which is smaller, the lithium atom or the lithium ion (Li^+).
4. State and explain which is bigger, the fluorine atom or the fluoride ion (F^-).
5. Arrange the following in order of increasing atomic radius (smallest first): Cl, Si, Na. Explain your reasoning.
6. Arrange the following in order of increasing radius: Mg^{2+} , Na^+ , O^{2-} , F^- , N^{3-} , Al^{3+} . Explain your reasoning.

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.



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).

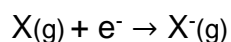
- As nuclear charge increases across a period, the electrostatic attraction between the nucleus and valence electrons increases.
- Atomic radius decreases across a period – valence electrons are closer to the nucleus which results in a stronger attraction between the nucleus and the valence electrons.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

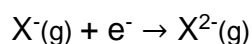
1. Outline what is meant by first ionisation energy.
2. State and explain the trend in ionisation energy across period 3.
3. State and explain the trend in ionisation energy down group 1.

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.



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



- 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.

Metallic and non-metallic character

- The 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.
- The 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.

2. Describe and explain the trend in metallic and non metallic character across a period and down a group.

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

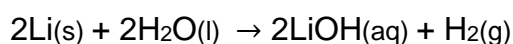
- Alkali metals are very reactive metals – they are shiny when cut, but quickly react with oxygen in air to form a layer of oxide.
- They are stored in oil to prevent the reaction with air.
- They have low densities – Li, Na and K float on water.
- They have low melting points – the melting point decreases down the group as the atomic radius increases.
- The metallic character of the group 1 metals increase 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).



Exercises:

1. Explain why the alkali metals are stored in oil.
2. Describe and explain the trend in melting point down group 1.
3. Describe and explain the trend in reactivity down group 1.
4. Write a balanced chemical equation, complete with state symbols for the reaction of potassium and water.
5. What pH would you expect the resulting solution to be? Explain your answer.

Group 17 elements – the halogens

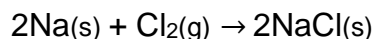
- The halogens (salt formers) are very reactive non-metal elements.
- They halogens are coloured – F₂ is a pale-yellow gas, Cl₂ is a yellow-green gas, Br₂ is a reddish-brown liquid, I₂ is a purple solid.
- They show a change from gases (F₂, Cl₂), to liquid (Br₂) to solid (I₂) as the molar mass increases down the group.
- The halogens are diatomic – they form molecules composed of two atoms bonded together.
- The non-metallic character of the group 17 elements decreases down the group as ionisation energy decreases.

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.



Exercises:

1. Describe the change in state of the group 17 elements down the group.
2. Explain the meaning of the term *diatomic*.
3. Describe the reactivity of the halogens.
4. Write a balanced symbol equation for the reaction of potassium and bromine.

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 CO₂ levels.

Linking questions:

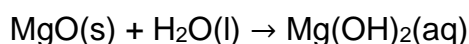
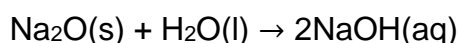
- Structure 2.1, 2.2 How do differences in bonding explain the differences in the properties of metal and non-metal oxides?

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 ₂ O(s)	MgO(s)	Al ₂ O ₃ (s)	SiO ₂ (s)	P ₄ O ₁₀ (s) P ₄ O ₆ (s)	SO ₃ (l) SO ₂ (g)	Cl ₂ O ₇ (l) Cl ₂ O(g)
Acid–base Character	Basic		Amphoteric	Acidic			

- Metal oxides are basic.
- Metals oxides react with water to produce a metal hydroxide.



- 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.

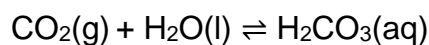
Exercises:

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

2. Write equations for the reactions of Li₂O and CaO with water.

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 .



- Acid deposition has a pH of less than 5.0

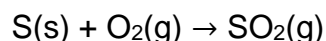
Sources of acidic gases

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

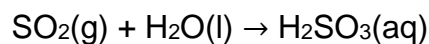
Formation of acid rain

Sulfur oxides

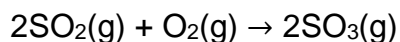
- Sulfuric acid (H_2SO_4) and sulfurous acid (H_2SO_3) are formed by the combustion of coal that contains high levels of sulfur.
- Sulfur (S) burns in oxygen (O_2) to form sulfur dioxide (SO_2).



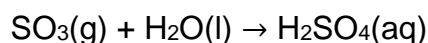
- SO_2 dissolves in water to form sulfurous acid (H_2SO_3).



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

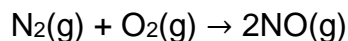


- SO_3 dissolves in water to form H_2SO_4 .

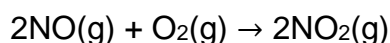


Nitrogen oxides

- Nitric acid (HNO₃) and nitrous acid (HNO₂) are formed from the reaction of nitrogen (N₂) and oxygen (O₂) at high temperatures in internal combustion engines.
- Nitrogen (N₂) and oxygen (O₂) react at high temperatures in internal combustion engines to form nitrogen monoxide (NO).



- NO reacts with O₂ to form nitrogen dioxide (NO₂).



- NO₂ dissolves in water to form HNO₃ and HNO₂.

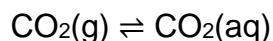


Exercises:

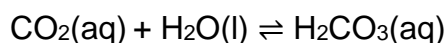
1. Deduce an equation to show why rainwater is naturally acidic.
2. State a natural and anthropogenic source of sulfur oxides and nitrogen oxides.
3. Acid rain has a pH of less than 5.0. Explain how the burning of coal can lead to the formation of acid rain.
4. Outline the process responsible for the production of acid rain from nitrogen oxides.

Ocean acidification

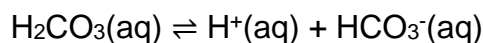
- Ocean acidification is the process by which increasing levels of atmospheric CO₂ cause the pH of the oceans to decrease.
- Approximately 30% of anthropogenic carbon dioxide is absorbed by the oceans (carbon sink).
- CO₂ dissolves in sea water.



- A heterogeneous equilibrium exists between concentrations of gaseous carbon dioxide in the atmosphere and aqueous carbon dioxide dissolved in the oceans.
- The CO₂ dissolves in sea water to produce carbonic acid (H₂CO₃) which is a weak acid.



- Carbonic acid (H₂CO₃) partially dissociates in water to produce H⁺(aq).
- Increasing levels of atmospheric CO₂ cause the equilibrium position to shift the right.



- 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 CO₂ could decrease the pH of ocean water using an equation to support your answer.

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 ₂ O, the oxidation state of the O is –2 and the H is +1. The sum of the oxidation states is $(-2 + (2 \times +1)) = 0$
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 ₂ O ₂). Examples: In OF ₂ , the F has an oxidation state of –1 and the O is +2. In H ₂ O 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 ₂ O, 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 $(+6 + (2 \times -2)) = -2$.

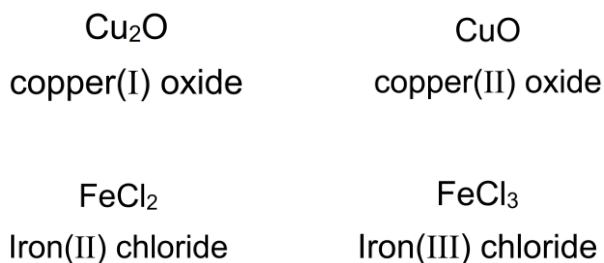
Oxidation states of polyatomic ions

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:

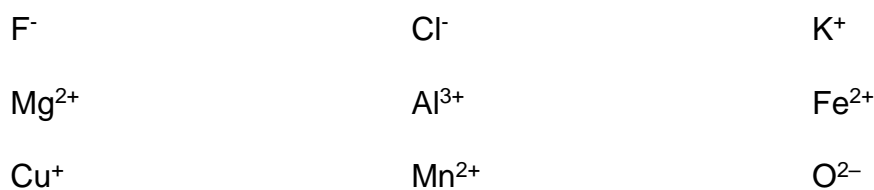


Exercises:

1) Deduce the oxidation states of the following.



2) Deduce the oxidation states of the following ions.



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



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



5) Deduce the oxidation number of the metal ion in the following:

iron(II) oxide

manganese(IV) oxide

manganate(VII) ion

chromium(III) oxide

copper(I) chloride

copper(II) chloride