

## 1.1 Module 1: Atoms and Reactions

This module provides candidates with a knowledge and understanding of atomic structure and the chemical ideas that underpin the study of quantitative chemistry:

### 1.1.1 Atoms

- atomic structure;
- relative masses.

### 1.1.2 Moles and Equations

- the mole;
- reacting masses and equations.

### 1.1.3 Acids

- acids and bases;
- salts.

### 1.1.4 Redox

- oxidation number;
- redox reactions.

## 1.1.1 Atoms

### Context and exemplification

### Assessable learning outcomes

#### Atomic structure

- The mass of an electron can be assumed to be 1/2000th the mass of a proton.

#### How Science Works 1, 7a:

- Modern development of the structure of the atom; the changing accepted view of the structure of the atom; acceptance (and

Candidates should be able to:

- (a) describe protons, neutrons and electrons in terms of relative charge and relative mass;
- (b) describe the distribution of mass and charge within an atom;
- (c) describe the contribution of protons and neutrons to the nucleus of an atom, in terms of atomic (proton) number and mass (nucleon) number;

rejection) of different theories for the structure (d) deduce the numbers of protons, neutrons and electrons in:  
of the atom from the Greeks, Dalton,  
Thompson and Rutherford, Moseley, et al.

- (i) an atom given its atomic and mass number,
- (ii) an ion given its atomic number, mass number and ionic charge;
- (e) explain the term *isotopes* as atoms of an element with different numbers of neutrons and different masses;

#### Relative masses

- For simple molecules, the term relative molecular mass will be used.
  - For compounds with giant structures, the term *relative formula mass* will be used.
- (f) state that  $^{12}\text{C}$  is used as the standard measurement of relative masses;
- (g) define the terms *relative isotopic mass* and *relative atomic mass*, based on the  $^{12}\text{C}$  scale;
- (h) calculate the relative atomic mass of an element given the relative abundances of its isotopes;

#### How Science Works 3:

- Use of spreadsheets in calculating relative atomic masses from data;
  - Definitions of relative molecular mass and relative formula mass will not be required.
- (i) use the terms *relative molecular mass* and *relative formula mass* and calculate values from relative atomic masses.

#### 1.1.2 Moles and Equations

##### Context and exemplification

##### Assessable learning outcomes

##### The mole

Candidates should be able to:

- (a) explain the terms:
  - (i) *amount of substance*,
  - (ii) *mole* (symbol ‘mol’), as the unit for amount of substance,
  - (iii) the *Avogadro constant*,  $N_A$ , as the number of particles per mole ( $6.02 \times 10^{23} \text{ mol}^{-1}$ );
- (b) define and use the term *molar mass* (units g mol $^{-1}$ ) as the mass per mole of a substance;

##### Empirical and molecular formulae

- (c) explain the terms:

- (i) *empirical formula* as the simplest whole number ratio of atoms of each element present in a compound,
  - (ii) *molecular formula* as the actual number of atoms of each element in a molecule;
- (d) calculate empirical and molecular formulae, using composition by mass and percentage compositions;

<b>Chemical equations</b>	(e) construct balanced chemical equations for reactions studied and for unfamiliar reactions given reactants and products;
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<b>Calculation of reacting masses, mole concentrations and volumes of gases</b>	(f) carry out calculations, using amount of substance in mol, involving: (i) mass, (ii) gas volume, (iii) solution volume and concentration;  (g) deduce stoichiometric relationships from calculations;  (h) use the terms <i>concentrated</i> and <i>dilute</i> as qualitative descriptions for the concentration of a solution.
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### 1.1.3 Acids

<b>Context and exemplification</b>	<b>Assessable learning outcomes</b>
<b>Acids and bases</b>	Candidates should be able to: (a) explain that an acid releases $H^+$ ions in aqueous solution; (b) state the formulae of the common acids: hydrochloric, sulfuric and nitric acids; (c) state that common bases are metal oxides, metal hydroxides and ammonia; (d) state that an alkali is a soluble base that releases $OH^-$ ions in aqueous solution; (e) state the formulae of the common alkalis: sodium hydroxide, potassium hydroxide and aqueous ammonia;
<b>Salts</b>	(f) explain that a salt is produced when the $H^+$ ion of an acid is replaced by a metal ion or $NH_4^+$ ; (g) describe the reactions of an acid with carbonates, bases and alkalis, to form a salt; (h) explain that a base readily accepts $H^+$ ions from an acid: eg $OH^-$ forming $H_2O$ ; $NH_3$ forming $NH_4^+$ ; (i) explain the terms <i>anhydrous</i> , <i>hydrated</i> and <i>water of crystallisation</i> ; (j) calculate the formula of a hydrated salt from given percentage composition, mass composition or experimental data; (k) perform acid–base titrations, and carry out structured titrations.

**Context and exemplification****Assessable learning outcomes****Oxidation number**

- Candidates will not be expected to use oxidation numbers in peroxides or metal hydrides.
- Ionic equations will only be required in Group 7 chemistry.

Candidates should be able to:

- apply rules for assigning oxidation number to atoms in elements, compounds and ions;
- describe the terms *oxidation* and *reduction* in terms of:
  - electron transfer,
  - changes in oxidation number;
- use a Roman numeral to indicate the magnitude of the oxidation state of an element, when a name may be ambiguous, eg nitrate(III) and nitrate(V);
- write formulae using oxidation numbers;

**Redox reactions**

- For nitric acid, reactions of metals are not expected.
- Ionic equations not required.

(e) explain that:

- metals generally form ions by losing electrons with an increase in oxidation number to form positive ions,
- non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions;
- describe the redox reactions of metals with dilute hydrochloric and dilute sulfuric acids;
- interpret and make predictions from redox equations in terms of oxidation numbers and electron loss/gain.

**Practical Skills are assessed using OCR set tasks. The practical work suggested below may be carried out as part of skill development. Centres are not required to carry out all of these experiments:**

- Making up a standard solution.
- NaOH or Na<sub>2</sub>CO<sub>3</sub>/HCl titration.
- NaOH/H<sub>2</sub>SO<sub>4</sub> to illustrate difference in stoichiometry.
- Titration involving a dilution – citric acid in lime juice cordial.
- Determination of the percentage of water of crystallisation in a hydrated salt.
- Determination of the relative atomic mass of an unknown metal by gas collection.
- Determination of the concentration of lime water.
- Determination of the relative formula mass of washing soda by titration.
- Reactions of the bases, alkalis and carbonates with acids.
- Preparation of salts from an acid and a base, eg copper(II) sulfate, ammonium sulphate.
- Reactions of metals with acids.

In addition to the aims of the scheme, this module provides candidates with a knowledge and understanding of chemical ideas that underpin the study of inorganic chemistry:

## 1.2.1 Electron Structure

- ionisation energies;
- energy levels, shells, sub-shells, orbitals and electron configuration.

## 1.2.2 Bonding and Structure

- ionic bonding;
- covalent bonding;
- the shapes of simple molecules and ions;
- electronegativity and polarity;
- intermolecular forces.

## 1.2.1 Electron Structure

Context and exemplification	Assessable learning outcomes
<b>Ionisation energies</b>	Candidates should be able to: <ul style="list-style-type: none"> <li>(a) Define the terms <i>first ionisation energy</i> and <i>successive ionisation energy</i>;</li> <li>(b) Explain that ionisation energies are influenced by nuclear charge, electron shielding and the distance of the outermost electron from the nucleus;</li> <li>(c) predict from successive ionisation energies of an element:               <ul style="list-style-type: none"> <li>(i) the number of electrons in each shell of an atom,</li> <li>(ii) the group of the element;</li> </ul> </li> </ul>
How Science Works 1:	
	<ul style="list-style-type: none"> <li>(d) state the number of electrons that can fill the first four shells;</li> <li>(e) describe an orbital as a region that can hold up to two electrons, with opposite spins;</li> <li>(f) describe the shapes of s and p orbitals;</li> <li>(g) state the number of:               <ul style="list-style-type: none"> <li>(i) orbitals making up s-, p- and d-sub-shells,</li> <li>(ii) electrons that occupy s-, p- and d-sub-shells;</li> </ul> </li> </ul>
How Science Works 1, 7a:	
	<ul style="list-style-type: none"> <li>(h) describe the relative energies of s-, p- and d-orbitals for the shells 1, 2, 3 and the 4s and 4p orbitals;</li> </ul>

- (i) deduce the electron configurations of:
- (i) atoms, given the atomic number, up to  $Z = 36$ ,
- (ii) ions, given the atomic number and ionic charge, limited to s and p blocks up to  $Z = 36$ ;
- (j) classify the elements into s, p and d blocks.

### 1.2.2 Bonding and Structure

#### Context and exemplification

#### Assessable learning outcomes

##### Ionic bonding

Candidates should be able to:

- (a) describe the term *ionic bonding* as electrostatic attraction between oppositely-charged ions;
- (b) construct 'dot-and-cross' diagrams, to describe ionic bonding;
- (c) predict ionic charge from the position of an element in the Periodic Table;
- (d) state the formulae for the following ions:  $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{SO}_4^{2-}$  and  $\text{NH}_4^+$ ;

##### Covalent bonding and dative covalent (coordinate) bonding

- (e) describe the term *covalent bond* as a shared pair of electrons;
- (f) construct 'dot-and-cross' diagrams to describe:
  - (i) single covalent bonding, eg as in  $\text{H}_2$ ,  $\text{Cl}_2$ ,  $\text{HCl}$ ,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{CH}_4$ ,  $\text{BF}_3$  and  $\text{SF}_6$ ,
  - (ii) multiple covalent bonding, eg as in  $\text{O}_2$ ,  $\text{N}_2$  and  $\text{CO}_2$ ,
  - (iii) dative covalent (coordinate) bonding, eg as in  $\text{NH}_4^+$ ,
  - (iv) molecules and ions analogous to those specified in (i), (ii) and (iii);

##### The shapes of simple molecules and ions

- (g) explain that the shape of a simple molecule is determined by repulsion between electron pairs surrounding a central atom;
- (h) state that lone pairs of electrons repel more than bonded pairs;
- (i) explain the shapes of, and bond angles in, molecules and ions with up to six electron pairs (including lone pairs) surrounding a central atom, eg as in:
  - (i)  $\text{BF}_3$  (trigonal planar),
  - (ii)  $\text{CH}_4$  and  $\text{NH}_4^+$  (tetrahedral),
  - (iii)  $\text{SF}_6$  (octahedral),
  - (iv)  $\text{NH}_3$  (pyramidal),

##### How Science Works 3:

- Three-dimensional molecular shapes.

	<p>(v) <math>\text{H}_2\text{O}</math> (non-linear),</p> <p>(vi) <math>\text{CO}_2</math> (linear);</p>
	<p>(j) predict the shapes of, and bond angles in, molecules and ions analogous to those specified in (i);</p>
<b>Electronegativity and bond polarity</b>	<p>(k) describe the term <i>electronegativity</i> as the ability of an atom to attract the bonding electrons in a covalent bond;</p> <p>(l) explain that a permanent dipole may arise when covalently-bonded atoms have different electronegativities, resulting in a polar bond;</p>
<b>Intermolecular forces</b>	<p>(m) describe intermolecular forces based on permanent dipoles, as in hydrogen chloride, and induced dipoles (van der Waals' forces), as in the noble gases;</p> <p>(n) describe <i>hydrogen bonding</i>, including the role of a lone pair, between molecules containing <math>-\text{OH}</math> and <math>-\text{NH}</math> groups, ie as in <math>\text{H}_2\text{O}</math>, <math>\text{NH}_3</math> and analogous molecules;</p> <p>(o) describe and explain the anomalous properties of <math>\text{H}_2\text{O}</math> resulting from hydrogen bonding, eg:</p> <ul style="list-style-type: none"> <li>(i) the density of ice compared with water,</li> <li>(ii) its relatively high freezing point and boiling point;</li> </ul>
<b>Metallic bonding</b>	<p>(p) describe <i>metallic bonding</i> as the attraction of positive ions to delocalised electrons;</p>
• No details of cubic or hexagonal packing required.	
<b>Bonding and physical properties</b>	<p>(q) describe structures as:</p> <ul style="list-style-type: none"> <li>(i) giant ionic lattices, with strong ionic bonding, ie as in <math>\text{NaCl}</math>,</li> <li>(ii) giant covalent lattices, ie as in diamond and graphite,</li> <li>(iii) giant metallic lattices,</li> <li>(iv) simple molecular lattices, ie as in <math>\text{I}_2</math> and ice;</li> </ul> <p>(r) describe, interpret and/or predict physical properties, including melting and boiling points, electrical conductivity and solubility in terms of:</p> <ul style="list-style-type: none"> <li>(i) different structures of particles (atoms, molecules, ions and electrons) and the forces between them,</li> </ul>

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- (ii) different types of bonding (ionic bonding, covalent bonding, metallic bonding, hydrogen bonding, other intermolecular interactions);
  - (s) deduce the type of structure and bonding present from given information.
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**Practical Skills are assessed using OCR set tasks. The practical work suggested below may be carried out as part of skill development. Centres are not required to carry out all of these experiments:**

- Bonding, structure and physical properties of substances (polarity, electrical conductivity, boiling points, solubility in polar and non-polar solvents).
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### 1.3 Module 3: The Periodic Table

In addition to the aims of the scheme, this module provides candidates with a knowledge and understanding of chemical ideas that underpin the study of inorganic chemistry:

#### 1.3.1 Periodicity

- the Periodic Table;
- trends in physical properties.

#### 1.3.2 Group 2

- redox reactions of Group 2 metals;
- Group 2 compounds.

#### 1.3.3 Group 7

- redox reactions of Group 7 elements;
- halide tests.

### 1.3.1 Periodicity

Context and exemplification	Assessable learning outcomes
<b>The structure of the Periodic Table in terms of groups and periods</b> How Science Works 1, 7a, 7b: <ul style="list-style-type: none"><li>Development of the Periodic Table from Döbereiner, Newlands, Mendeleev, Moseley, Seaborg, <i>et al.</i></li></ul>	Candidates should be able to: <ul style="list-style-type: none"><li>(a) describe the Periodic Table in terms of the arrangement of elements:<ul style="list-style-type: none"><li>(i) by increasing atomic (proton) number,</li><li>(ii) in periods showing repeating trends in physical and chemical properties,</li><li>(iii) in groups having similar physical and chemical properties;</li></ul></li><li>(b) describe <i>periodicity</i> in terms of a repeating pattern across different periods;</li><li>(c) explain that atoms of elements in a group have similar outer shell electron configurations, resulting in similar properties;</li></ul>
<b>Periodicity of physical properties of elements</b> <ul style="list-style-type: none"><li>Periodic trends in ionisation energies will consider s and p blocks only. No consideration of the periodic decreases between Groups 2 and 3, and 5 and 6, will be tested.</li></ul>	<ul style="list-style-type: none"><li>(d) describe and explain the variation of the first ionisation energies of elements shown by:<ul style="list-style-type: none"><li>(i) a general increase across a period, in terms of increasing nuclear charge,</li><li>(ii) a decrease down a group in terms of increasing atomic radius and increasing electron shielding outweighing increasing nuclear charge;</li></ul></li></ul> <p>[See also unit F321: 1.2.1(a)–(c)]</p> <ul style="list-style-type: none"><li>(e) for the elements of Periods 2 and 3:<ul style="list-style-type: none"><li>(i) describe the variation in electron</li></ul></li></ul>

- configurations, atomic radii, melting points and boiling points;
- (ii) explain variations in melting and boiling points in terms of structure and bonding;
- (f) interpret data on electron configurations, atomic radii, first ionisation energies, melting points and boiling points to demonstrate periodicity.

### 1.3.2 Group 2

#### Context and exemplification

#### Assessable learning outcomes

##### Redox reactions of Group 2 metals

Candidates should be able to:

- (a) describe the redox reactions of the Group 2 elements Mg → Ba:
- (i) with oxygen,
  - (ii) with water;
- (b) explain the trend in reactivity of Group 2 elements down the group due to the increasing ease of forming cations, in terms of atomic size, shielding and nuclear attraction;

##### Reactions of Group 2 compounds

- No explanation of thermal decomposition required.

- (c) describe the action of water on oxides of elements in Group 2 and state the approximate pH of any resulting solution;
- (d) describe the thermal decomposition of the carbonates of elements in Group 2 and the trend in their ease of decomposition;
- (e) interpret and make predictions from the chemical and physical properties of Group 2 elements and compounds;
- (f) explain the use of  $\text{Ca}(\text{OH})_2$  in agriculture to neutralise acid soils; the use of  $\text{Mg}(\text{OH})_2$  in some indigestion tablets as an antacid.

### 1.3.3 Group 7

#### Context and exemplification

#### Assessable learning outcomes

##### Characteristic physical properties

Candidates should be able to:

- (a) explain, in terms of van der Waals' forces, the trend in the boiling points of  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$ ;

##### Redox reactions and trends in reactivity of Group 7 elements and their compounds

- (b) describe the redox reactions, including ionic equations, of the Group 7 elements  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$  with other halide ions, in the presence of an organic solvent, to illustrate the relative reactivity of Group 7 elements;

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- (c) explain the trend in reactivity of Group 7 elements down the group from the decreasing ease of forming negative ions, in terms of atomic size, shielding and nuclear attraction;
  - (d) describe the term *disproportionation* as a reaction in which an element is simultaneously oxidised and reduced, illustrated by:
    - (i) the reaction of chlorine with water as used in water purification,
    - (ii) the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach,
    - (iii) reactions analogous to those specified in (i) and (ii);

How Science Works 6a, 6b:

- Health benefits of chlorine use in water;
  - Ethical implications of adding chlorine to public water supplies (also fluorine in drinking water).
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- (e) interpret and make predictions from the chemical and physical properties of the Group 7 elements and their compounds;

- (f) contrast the benefits of chlorine use in water treatment (killing bacteria) with associated risks (hazards of toxic chlorine gas and possible risks from formation of chlorinated hydrocarbons);

**Characteristic reactions of halide ions**

- Complexes with ammonia are not required.

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- (g) describe the precipitation reactions, including ionic equations, of the aqueous anions  $\text{Cl}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$  with aqueous silver ions, followed by aqueous ammonia;

- (h) describe the use of the precipitation reactions in (g) as a test for different halide ions.

**Practical Skills are assessed using OCR set tasks. The practical work suggested below may be carried out as part of skill development. Centres are not required to carry out all of these experiments:**

- Reactions of some Group 2 metals with oxygen and water.
  - Action of water on Group 2 oxides and testing pH of resulting solutions.
  - Thermal decomposition of Group 2 carbonates.
  - Halogen displacement reactions.
  - Testing for the presence of halide ions in solution using silver nitrate.
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