

UNIT 3: INORGANIC CHEMISTRY

**Recommended Prior Knowledge:** Unit 1 (Theoretical Chemistry) and Unit 2 (Physical Chemistry) should have been studied before this unit. The unit builds on the ideas of atomic and molecular structure and bonding developed in Unit 1, and those concerning enthalpy changes, equilibrium, kinetics and redox developed in Unit 2.

**Context:** This self-contained unit can be studied either before or after the two Organic Chemistry Units 4 and 5.

**Outline:** The unit covers the whole of the Inorganic part of the AS course. Concepts developed in Units 1 and 2 (see prior knowledge above) are applied to Period 3 and Groups II and VII of the Periodic Table, as well as to aspects of nitrogen and sulphur chemistry.

section	Learning Outcomes	Suggested Teaching Activities	Lesson allocation (approx.%)	Other references
9.1 (a)	[Across Period 3] describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements.	Graphs could be plotted of atomic radius, ionic radius and melting point from Na to Ar, (and also from Li to Ne), to show periodicity. Conductivity increases from Na to Al, then low as Si, then zero.	4	<a href="#">site 2</a> (AS / Periodic Table)
9.1 (b)	explain qualitatively the variation in atomic radius and ionic radius	Increasing nuclear charge across periodic pulls e <sup>-</sup> in more. In the same period anions contain one more shell than cations.	4	<a href="#">site 5e</a> (Periodic Table: Period 3)
9.1 (c)	interpret the variation in melting point and in electrical conductivity in terms of the presence of simple molecular, giant molecular or metallic bonding in the elements	Relate conductivity to (increasingly more) delocalised e <sup>-</sup> in Na-Al; conduction bands in Si; localised electrons in covalent bonds for the rest. A graphic illustration of the semiconduction of Si is showing the <b>increase</b> in conduction on heating a block of Si over a bunsen (use a battery + bulb or ammeter circuit).		
9.1 (d)	explain the variation in first ionisation energy	Increasing nuclear charge but little extra shielding across period causes general increase in I.E. Slight decreases at Mg-Al, and P-S due to new sub-shell, and repulsion between electrons sharing the same orbital, respectively.		
9.1 (e)	describe the reactions, if any, of the elements with oxygen (to give Na <sub>2</sub> O, MgO, Al <sub>2</sub> O <sub>3</sub> , P <sub>4</sub> O <sub>10</sub> , SO <sub>2</sub> , SO <sub>3</sub> ), chlorine (to give NaCl; MgCl <sub>2</sub> ; Al <sub>2</sub> Cl <sub>6</sub> ; SiCl <sub>4</sub> ; PCl <sub>5</sub> ), and water (Na and Mg only)	Na, Mg, Al, P and S can all be burned in oxygen or air. To avoid contamination by metallic combustion spoons, a good technique with Na is to heat a small piece of Na on a <b>dry</b> brick with a Bunsen until it starts to burn, and then to place over it an inverted gas jar of O <sub>2</sub> . This also works with Cl <sub>2</sub> , producing a <b>white</b> powder of NaCl.	8	<a href="#">site 3</a> (Inorganic / Reactions of elements in Period 3)
9.1 (f)	state and explain the variation in oxidation number of the oxides and chlorides	Revise work done in Unit 2 on oxidation numbers. Maximum O.N. equals the number of electrons in the outer shell.		<a href="#">site 5e</a> (Periodic Table: Period 3)

9.1 (g)	describe the reactions of the above oxides with water [treatment of peroxides and superoxides is <b>not</b> required]	Add universal indicator solution to the oxides prepared above. The trend is from strong alkali (NaOH), through very weak alkali (Mg(OH) <sub>2</sub> ) to neutral (Al <sub>2</sub> O <sub>3</sub> and SiO <sub>2</sub> ) to acidic. Full balanced equations for all reactions.	6	<a href="#">site 5e</a> (Periodic Table: Period 3)
9.1 (h)	describe and explain the acid/base behaviour of Period 3 oxides and hydroxides, including, where relevant, amphoteric behaviour in reaction with sodium hydroxide (only) and acids	This could include some titrations, but also practical work on the dissolving on MgO in acids, the redissolving of Al(OH) <sub>3</sub> in NaOH(aq). Sodium aluminate can be represented as either NaAlO <sub>2</sub> or NaAl(OH) <sub>4</sub> .		<a href="#">site 3</a> (Inorganic / Reactions of elements in Period 3)  <a href="#">site 5e</a> (Periodic Table: Period 3)
9.1 (i)	describe and explain the reactions of the above chlorides with water	Emphasise that NaCl and MgCl <sub>2</sub> merely dissolve (to form pH 7 solutions containing dissociated ions) whereas the rest undergo hydrolysis. With a small amount of water AlCl <sub>3</sub> gives Al <sub>2</sub> O <sub>3</sub> + HCl, but with an excess a solution containing [Al(H <sub>2</sub> O) <sub>5</sub> (OH)] <sup>2+</sup> + H <sup>+</sup> + Cl <sup>-</sup> ions is formed.	2	<a href="#">site 3</a> (Inorganic /Reactions of Chlorides)
9.1 (j)	interpret the variations and trends described above in terms of bonding and electronegativity	Electropositive metals such as Na and Mg form ionically bonded oxides which are basic, and ionically bonded chlorides which are neutral.	2	<a href="#">site 1c</a> (Electronegativity)
9.1 (k)	suggest the types of chemical bonding present in chlorides and oxides from observations of their chemical and physical properties	As the electronegativity of the element increases and becomes similar to that of O or Cl, covalently bonded oxides and chlorides are formed, which are acidic. Al is on the borderline – the structure of the solid is ionic but with a large degree of polarisation, and the oxide is amphoteric. These generalisation can be applied to the oxides and chlorides of other elements such as Ga, Ge, As.		
9.1 (l)	predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity	Properties such as metallic/non-metallic; macro or simple covalent; reactive or non-reactive, forming ionic or covalent oxides and chlorides, m.pt. and b.pt.	4	
9.1 (m)	deduce the nature, possible position in the Periodic Table, and identity of unknown elements from given information of physical and chemical properties	Similar generalisations to those above. Also include Group trends – metals becoming more reactive down their groups, but non-metals less reactive.		
9.2 (a)	[For Group II] describe the reactions of the elements with oxygen and water	The reactions become more vigorous as the Group is descended. Mg ribbon can be reacted with steam, and Mg powder reacts slowly with water.	4	<a href="#">site 1b</a> (Chem of Group II Elements)  <a href="#">site 3</a> (Inorganic /Reactions of Group 1&2)
9.2 (b)	describe the behaviour of the Group II oxides with water	As the solubility of the hydroxides increases from Mg to Ba, so the solutions formed become increasingly alkaline. The exothermicity of the CaO + H <sub>2</sub> O reaction (the “slaking” of lime) should be noted.	2	<a href="#">site 1b</a> (Chem of Group II Elements)  <a href="#">site 3</a> (Inorganic /Reactions of Group 1 &2)

9.2 (c)	describe the thermal decomposition of the nitrates and carbonates of Group II	Decomposition becomes more difficult down the Group, as the polarising power of the cation decreases with increasing radius.	6	<a href="#">site 1b</a> (Chem of Group II Elements)
9.2 (d)	interpret, and make predictions from, the trends in physical and chemical properties of the Group II elements and their compounds	m. pts. and b. pts. of elements show no trend, but atomic and ionic radii increase, and $E^\circ$ values become more negative. Lattice energies of the oxides, and hence their m.pts., decrease down the Group.	4	CIEX 17
9.2 (e)	explain the use of magnesium oxide as a refractory lining material and carbonate as a building material	Due to its high m. pt. and low thermal conductivity. Used in, e.g. blast furnaces (where its basic nature is also an advantage).		
9.2 (f)	describe the use of lime in agriculture	Describe the production of CaO and Ca(OH) <sub>2</sub> from CaCO <sub>3</sub> . Both CaCO <sub>3</sub> and Ca(OH) <sub>2</sub> are used, to decrease the acidity of the soil, and hence to enhance the uptake of essential cations by plants.	4	
9.4 (a)	describe the trends in volatility and colour of chlorine, bromine and iodine	Deepening colour (from pale green to orange-brown to purple) is due to the decreasing energy of the $n - \pi^*$ transition.	2	<a href="#">site 1b</a> (Chem of Group VII Elements)
9.4 (b)	interpret the volatility of the elements in terms of van der Waals' forces	Increased van der Waals' forces are due to the increasing number of (polarisable) electrons in the clouds around the molecules.		
9.4 (d)	describe and explain the reactions of the Group VII elements with hydrogen	The explosive nature of the H <sub>2</sub> + Cl <sub>2</sub> mixture with light; the steady burning of a H <sub>2</sub> jet in a gas jar of Cl <sub>2</sub> ; the Pt-catalysed combination of H <sub>2</sub> + Br <sub>2</sub> or I <sub>2</sub> . The steady decrease in thermal stability of HX (e.g. the effect of a hot nichrome wire plunged into a test tube of the gas), explained in terms of the weakening H-X bond. Use of bond energies to calculate $\Delta H^\circ$ for the reaction.	5	<a href="#">site 1b</a> (Chem of Group VII Elements)
9.4 (e)	(i) describe and explain the relative thermal stabilities of the hydrides, (ii) interpret these relative stabilities in terms of bond energies			
9.4 (c)	describe the relative reactivity of the Group VII elements as oxidising agents	Practical examples could include X <sub>2</sub> + NaBr, NaI etc.; X <sub>2</sub> + Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> . Other reactions include Cl <sub>2</sub> + Fe ( $\rightarrow$ FeCl <sub>3</sub> ) and I <sub>2</sub> + Fe ( $\rightarrow$ FeI <sub>2</sub> )	5	<a href="#">site 1b</a> (Chem of Group VII Elements) CIEX 2, 16
9.4 (f)	describe and explain the reactions of halide ions with (i) aqueous silver ions followed by aqueous ammonia, (ii) concentrated sulphuric acid	Practical work can involve preparing and noting the colours of AgX (white, cream and yellow) and whether they dissolve or not in dil or conc NH <sub>3</sub> . Equations (full and ionic) and complex formation ( $\rightarrow$ [Ag(NH <sub>3</sub> ) <sub>2</sub> ] <sup>+</sup> ). Simple acid-base reaction with Cl <sup>-</sup> ; increasing oxidation of X <sup>-</sup> to X <sub>2</sub> with Br <sup>-</sup> ( $\rightarrow$ SO <sub>2</sub> as byproduct) and with I <sub>2</sub> ( $\rightarrow$ H <sub>2</sub> S and SO <sub>2</sub> byproducts)	4	<a href="#">site 1b</a> (Chem of Group VII Elements) <a href="#">site 3</a> (Inorganic /Reactions of Group VII))
9.4 (h)	describe and interpret in terms of changes of oxidation number the reaction of chlorine with cold, and with hot, aqueous sodium hydroxide	Introduce the idea of disproportionation. The preparation of KClO <sub>3</sub> from the reaction with hot KOH can be attempted, but NaClO <sub>3</sub> is too soluble for easy separation.	4	<a href="#">site 3</a> (Inorganic /Reactions of Group VII))

9.4 (i)	explain the use of chlorine in water purification	The equilibrium $\text{Cl}_2 + \text{H}_2\text{O} = \text{HCl} + \text{HOCl}$ produces too great an acidity, so an alkali is added at the same time.	4	<a href="#">site 12</a> (Search the web/Chlorine in Water)
9.4 (j)	recognise the industrial importance and environmental significance of the halogens and their compounds, {e.g. for bleaches; pvc; halogenated hydrocarbons as solvents, refrigerants and in aerosols}	Mention the discovery by atmospheric chemists of the CFC-caused problem in the $\text{O}_3$ layer, and the replacement of CFCs by hydrocarbons and freons. The potential problems with the incineration of pvc waste; the phasing-out of chlorinated solvents; DDT and BHC residues in the environment. (see also Unit 4)		
9.4 (g)	outline a method for the manufacture of chlorine from brine by a diaphragm cell	See Unit 2. Revise the conditions, the overall equation, and the main uses of the three products.	2	<a href="#">site 12</a> (Search the web/Diaphragm Cell)
9.6 (a)	explain the lack of reactivity of nitrogen	Mention that the strength of the $\text{N}\equiv\text{N}$ affects both the kinetic inertness (large $E_a$ ) and the thermodynamic stability.	2	
9.6 (c)	describe the Haber process for the manufacture of ammonia from its elements, giving essential operating conditions, and interpret these conditions (qualitatively) in terms of the principles of kinetics and equilibria	See Unit 2. Revise the conditions, and the reasons for using them, the overall equation, and the main uses of the ammonia produced.	4	<a href="#">site 6</a> (Home/AS/ Chemistry/ Industrial Inorg.Chem)
9.6 (b)	describe: (i) the formation, and structure, of the ammonium ion (ii) the displacement of ammonia from its salts	$\text{NH}_3$ as a Bronsted base. Neutralisation, titration. $\text{NH}_4^+$ isoelectronic with $\text{CH}_4$ , hence same shape. Warming with $\text{NaOH}(\text{aq})$ and testing for $\text{NH}_3$ with moist red litmus paper – a good test for $\text{NH}_4^+$ .	4	
9.6 (d)	recognise the industrial importance of ammonia and nitrogen compounds derived from ammonia	Ammonium salts, urea and $\text{NH}_3$ itself as fertilisers. Nitric acid, and its use in making nitrates, explosives, dyes, polyamides, polyurethanes, and pharmaceuticals.	4	<a href="#">site 3</a> (Applied Organic Chemistry / Fertilisers)
9.6 (e)	recognise the environmental consequences of the uncontrolled use of nitrate fertilisers	Eutrophication of rivers and lakes, especially in areas of high rainfall, due to algal blooms and subsequent decomposition using up dissolved oxygen.		
9.6 (f)	describe and explain the occurrence, and catalytic removal, of oxides of nitrogen	The production of $\text{NO}$ from $\text{N}_2 + \text{O}_2$ in internal combustion engines. The role of Pt as a heterogeneous catalyst. Key reactions that occur in the catalytic converter are $\text{NO} + \text{CO} \rightarrow \frac{1}{2}\text{N}_2 + \text{CO}_2$ ; $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$ ; $\text{h/c} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ .	6	<a href="#">site 12</a> (Search the web/Oxides of Nitrogen)
9.6 (g)	explain why atmospheric oxides of nitrogen are pollutants, including their catalytic role in the oxidation of atmospheric sulphur dioxide	Formation of peroxyacetylnitrate (PAN) in smog. The homogeneous catalysis that occurs when $\text{SO}_2$ and $\text{NO}_2$ interact.		
9.6 (h)	describe the formation of atmospheric sulphur dioxide from the combustion of sulphur contaminated carbonaceous fuels	The increasing use of flue gas desulphurisation (FGD) in power stations, and low-sulphur road fuels in cars and lorries.		<a href="#">site 12</a> (Search the web/Acid Rain)
9.6 (i)	state the role of sulphur dioxide in the formation of acid-rain and describe the main environmental consequences of acid-rain	$\text{SO}_2 + \frac{1}{2}\text{O}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$ . Catalysed by sunlight <i>or</i> $\text{NO}_2$ . Damage to stonework and mortar on buildings; aggravation of asthma; acidifying lakes and rivers		

9.6 (j)	state the main details of the Contact process for sulphuric acid production	See Unit 2. Revise the conditions, the equations for each stage, and the main uses of the sulphuric acid produced. (detergents, dyes, paints, car batteries, “pickling” metals etc)	4	<a href="#">site 6</a> (Home/AS/Chemistry/Industrial Inorg.Chem)
9.6 (k)	recognise the industrial importance of sulphuric acid			
9.6 (l)	describe the use of sulphur dioxide in food preservation			
		White wine, dried vegetables and fruit.		