

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 or in between 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 sulfur chemistry.

AO	Learning Outcomes	Suggested Teaching Activities	Learning Resources
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 falls to a low at Si, then zero for the rest.	C+H 15.1-15.3 M 87 site 4 (period 3) site 7 (periodic trends) site 17 (the periodic table)
9.1 (b)	Explain qualitatively the variation in atomic radius and ionic radius	Increasing nuclear charge across periodic pulls the electron shells in more. In the same period anions contain one more shell than cations.	C+H 15.3 M 88 R+N 9.4 R+N 9.5
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 .	Melting point should be related to structure (see sections 3k and 4e). Conductivity should be related to (increasingly more) delocalised e ⁻ 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 increase in conduction on heating a block of Si over a Bunsen (use a battery + bulb or ammeter circuit in a demonstration of this). See also section 2i. 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 (d)	Explain the variation in first ionisation energy		R+N 9.6
9.1 (e)	describe the reactions, if any, of the elements with oxygen (to give Na ₂ O, MgO, Al ₂ O ₃ , P ₄ O ₁₀ , SO ₂ , SO ₃), chlorine (to give NaCl; MgCl ₂ ; Al ₂ Cl ₆ ; SiCl ₄ ; PCl ₅), and water (Na and Mg only)	Na, Mg, Al, P and S should all be burned in gas jars filled with 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 dry brick with a Bunsen until it starts to burn, and then to place over it an inverted gas jar of O ₂ . This also works with Cl ₂ , producing a white powder of NaCl. Revise work done in Unit 2 on oxidation numbers. Maximum O.N. equals the number of electrons in the outer shell.	C+H 15.4-15.5 M 89 R+N 9.7 P(AS) 17, 18, 19
9.1 (f)	state and explain the variation in oxidation number of the oxides and chlorides		R+N 9.8
9.1 (g)	describe the reactions of the above oxides with water [treatment of peroxides and superoxides is not required]	Add universal indicator solution to the oxides prepared above. The trend is from strong alkali (NaOH), through very weak alkali (Mg(OH) ₂) to neutral (Al ₂ O ₃ and SiO ₂) to acidic. Students should work out the full balanced equations for all reactions.	C+H 15.4 M 89 R+N 9.9
9.1 (h)	describe and explain the acid/base behaviour of Period 3 oxides and	This could include some titrations, but also practical work on the dissolving on MgO in acids, and the precipitation and redissolving of Al(OH) ₃ in NaOH(aq). Sodium aluminate can be represented as either NaAlO ₂ or NaAl(OH) ₄ .	site 3 (the nature of oxides) site 16 (periodicity)

	hydroxides, including, where relevant, amphoteric behaviour in reaction with sodium hydroxide (only) and acids		
9.1 (i)	describe and explain the reactions of the above chlorides with water	Show by experiment that NaCl and MgCl ₂ merely dissolve (to form pH 7 solutions containing dissociated ions) whereas the rest undergo hydrolysis. With a small amount of water AlCl ₃ gives Al ₂ O ₃ + HCl(g), but with an excess a solution containing [Al(H ₂ O) ₅ (OH)] ²⁺ + H ⁺ + Cl ⁻ ions is formed.	C+H 15.5 M 89 R+N 9.11 site 3 (aluminium chloride + water)
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. 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.	C+H 15.1 M 88-89 R+N 9.12
9.1 (k)	suggest the types of chemical bonding present in chlorides and oxides from observations of their chemical and physical properties	These generalisation can be applied to the oxides and chlorides of other elements such as Ga, Ge, As.	R+N 9.13
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.	C+H 15.1 M 89
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. A classroom team-game of “predict the properties” of a series of elements drawn out of a hat encourages the application of knowledge.	R+N 9.14, 9.15
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. Show Mg ribbon reacting with steam and Mg powder reacting slowly with water (the hydrogen can be collected using an inverted funnel + inverted test tube)	C+H 16.1-16.2 M 93 R+N 10a.5 site 7 (periodic trends)
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 ₂ O reaction (the “slaking” of lime) should be demonstrated and contrasted with the virtual non-reaction of MgO + H ₂ O.	C+H 16.2 M 93 R+N 10a.7 P(AS) 20
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. Heating samples of the (anhydrous) nitrates will show the relative ease of production of NO ₂	C+H 16.4 M 93 R+N 10a.8 P(AS) 20 site 16 (periodicity)
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 ^o values become more negative. Lattice energies of the oxides, and hence their m.pt.s., decrease down the Group.	C+H 16.3-16.5 M 93
9.2 (e)	explain the use of magnesium oxide as a refractory lining material and calcium	Due to its high m. pt. and low thermal conductivity. Used in, e.g. blast furnaces (where its basic nature is also an advantage). Students could search their local areas for examples of	site 8 (Group II)

	carbonate as a building material	limestone or marble being used in buildings or for statues.	
9.2 (f)	describe the use of lime in agriculture	Describe the production of CaO and Ca(OH) ₂ from CaCO ₃ . Both CaCO ₃ and Ca(OH) ₂ are used, to decrease the acidity of the soil, and hence to enhance the uptake of essential cations by plants. Investigate local sources of "garden lime" and find out whether it is CaCO ₃ or Ca(OH) ₂ , and what it is used for.	
9.4 (a) 9.4 (b)	describe the trends in volatility and colour of chlorine, bromine and iodine interpret the volatility of the elements in terms of van der Waals' forces	Deepening colour (from pale green to orange-brown to purple) is due to the decreasing energy of the n - π* transition. Prepare solutions of the halogens in hexane to demonstrate these colours. Increased van der Waals' forces are due to the increasing number of (polarisable) electrons in the clouds around the molecules.	C+H 17.1 M 101 R+N 12.1
9.4 (c)	describe the relative reactivity of the Group VII elements as oxidising agents	Practical examples could include X ₂ + NaBr, NaI etc.; X ₂ + Na ₂ S ₂ O ₃ . Other reactions include Cl ₂ + Fe (→ FeCl ₃) and I ₂ + Fe (→ FeI ₂)	C+H 17.4 M 101 R+N 12.2 P(AS) 21, 23
9.4 (d) 9.4 (e)	describe and explain the reactions of the Group VII elements with hydrogen (i) describe and explain the relative thermal stabilities of the hydrides, (ii) interpret these relative stabilities in terms of bond energies	The explosive nature of the H ₂ + Cl ₂ mixture with light; the steady burning of a H ₂ jet in a gas jar of Cl ₂ ; the Pt-catalysed combination of H ₂ + Br ₂ or I ₂ . The steady decrease in thermal stability of HX (e.g. demonstrate 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 ΔH° for the reaction.	C+H 17.2 M 101 R+N 12.3, 12.4 site 21a
9.4 (f)	describe and explain the reactions of halide ions with (i) aqueous silver ions followed by aqueous ammonia, (ii) concentrated sulfuric 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 ₃ . Equations (full and ionic) and complex formation (→ [Ag(NH ₃) ₂] ⁺). Simple acid-base reaction with Cl ⁻ ; increasing oxidation of X ⁻ to X ₂ with Br ⁻ (→ SO ₂ as byproduct) and with I ₂ (→ H ₂ S and SO ₂ byproducts) Students should use oxidation numbers to construct balanced equations for these reactions.	C+H 17.3-17.4 M101 R+N 12.5, 12.6 P(AS) 22 site 21a
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.	C+H 20.4 M 84 R+N 6a.2 site 3 (the membrane cell)
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 ₃ from the reaction with hot KOH can be attempted, but NaClO ₃ is too soluble for easy separation. Students should use oxidation numbers to construct balanced equations for these reactions.	C+H 17.2, 17.4 M 101 R+N 12.7
9.4 (i)	explain the use of chlorine in water purification	The equilibrium Cl ₂ + H ₂ O = HCl + HOCl produces too great an acidity, so an alkali is added at the same time. Visit the local waterworks.	R+N 12.8
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 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). Internet searches for other uses	C+H 17.1 R+N 12.9

9.6 (a)	explain the lack of reactivity of nitrogen	Mention that the strength of the N≡N affects both the kinetic inertness (large E _a) and the thermodynamic stability.	M 97 R+N 14.1
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 (AO 6b). Revise the conditions, and the reasons for using them, the overall equation, and the main uses of the ammonia produced.	C+H 20.2 M 83 R+N 7a.7 site 16 (manufacture of ammonia)
9.6 (b)	describe: (i) the formation, and structure, of the ammonium ion (ii) the displacement of ammonia from its salts	NH ₃ as a Bronsted base. Neutralisation, titration. NH ₄ ⁺ isoelectronic with CH ₄ , hence same shape. Warming with NaOH(aq) and testing for NH ₃ with moist red litmus paper – a good test for NH ₄ ⁺ .	C+H 3.12, 6.3 M 97 R+N 14.2
9.6 (d)	recognize the industrial importance of ammonia and nitrogen compounds derived from ammonia	Ammonium salts, urea and NH ₃ itself as fertilisers (students could calculate the %N in each). Nitric acid, and its use in making nitrates, explosives, dyes, polyamides, polyurethanes, and pharmaceuticals. Internet search for the formulae and uses of these products.	C+H 20.2 M 97 R+N 14.3
9.6 (e)	recognize 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. Discuss local examples.	R+N 14.4
9.6 (f)	describe and explain the occurrence, and catalytic removal, of oxides of nitrogen (in IC engines and power stations)	The production of NO from N ₂ + O ₂ in internal combustion engines. The role of Pt as a heterogeneous catalyst. Key reactions that occur in the catalytic converter are NO + CO → ½N ₂ + CO ₂ ; CO + ½O ₂ → CO ₂ ; h/c + O ₂ → CO ₂ + H ₂ O.	C+H 15.4 R+N 14.5
9.6 (g)	explain why atmospheric oxides of nitrogen are pollutants, including their catalytic role in the oxidation of atmospheric sulfur dioxide	Formation of peroxyacetyl nitrate (PAN) in smog. The homogeneous catalysis that occurs when SO ₂ and NO ₂ interact.	
9.6 (h)	describe the formation of atmospheric sulfur dioxide from the combustion of sulfur contaminated carbonaceous fuels	The increasing use of flue gas desulfurisation (FGD) in power stations, and low-sulfur road fuels in cars and lorries.	R+N 14.6
9.6 (i)	state the role of sulfur dioxide in the formation of acid-rain and describe the main environmental consequences of acid-rain .	SO ₂ + ½O ₂ + H ₂ O → H ₂ SO ₄ . Catalysed by sunlight or NO ₂ . Damage to stonework and mortar on buildings; aggravation of asthma; acidifying lakes and rivers. Investigate local examples.	C+H 15.4 R+N 14.7
9.6 (j)	state the main details of the Contact process for sulfuric acid production	See Unit 2. Revise the conditions, the equations for each stage, and the main uses of the sulfuric acid produced. (detergents, dyes, paints, car batteries, “pickling” metals etc)	C+H 20.2 M 83
9.6 (k)	recognize the industrial importance of sulfuric acid	White wine, dried vegetables and fruit. Extension project could be a study of food labels or an internet search.	R+N 7a.7, 14.8
9.6 (l)	describe the use of sulfur dioxide in food preservation		R+N 14.9