

Structure 2.1

Answers

IB CHEMISTRY SL

25 Mn Manganese 54.938045	16 S Sulfur 32.065	J	6 C Carbon 12.0107	2 He Helium 4.002602	25 Mn Manganese 54.938045
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Structure 2.1.1

Understandings:

- When metal atoms lose electrons, they form positive ions called cations. When non-metal atoms gain electrons, they form negative ions called anions.

Learning outcomes:

- Predict the charge of an ion from the electron configuration of the atom.

Additional notes:

- The formation of ions with different charges from a transition element should be included.

Linking questions:

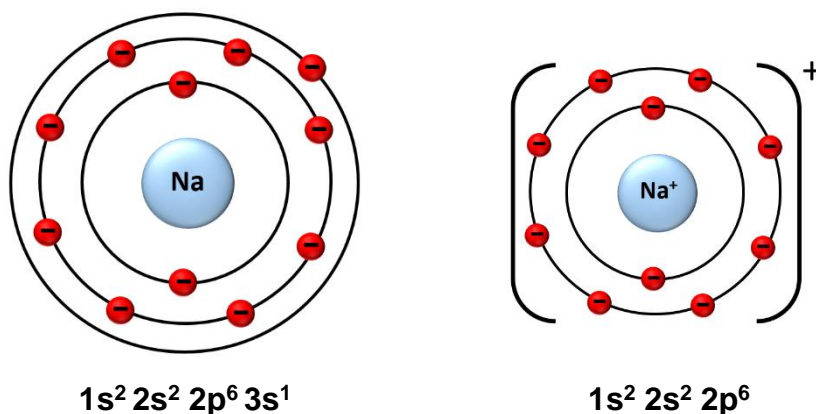
- Structure 3.1 How does the position of an element in the periodic table relate to the charge of its ion(s)
- Structure 1.3 (HL) How does the trend in successive ionisation energies of transition elements explain their variable oxidation states?

Ions

- Neutral atoms have equal numbers of protons and electrons.
- Positive ions (cations) are formed when atoms lose electrons.
- Negative ions (anions) are formed when atoms gain electrons.

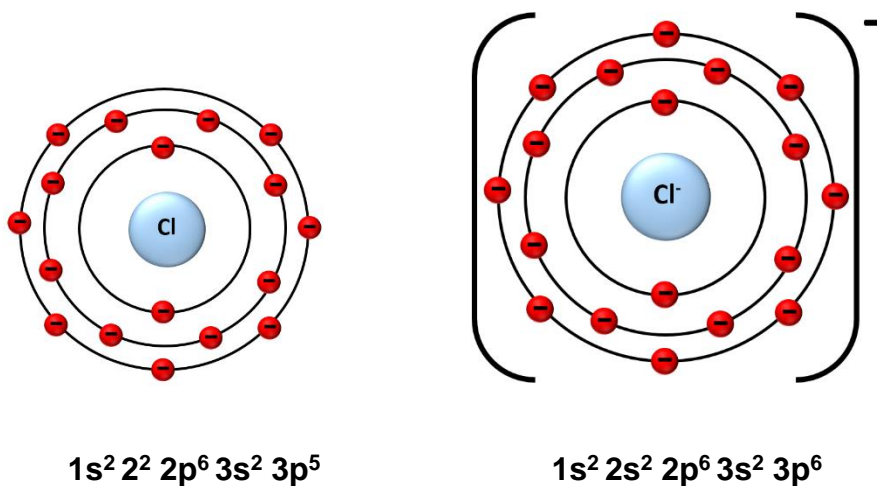
Positive ions

- Positive ions are formed when atoms lose electrons.
- The sodium atom loses its one valence electron to form a positive ion with a 1+ charge, Na^+ .
- It now has one less occupied energy level and the same electron configuration as the noble gas neon, Ne.



Negative ions

- Negative ions are formed when atoms gain electrons.
- The chlorine atom gains one electron to form a negative ion with a 1- charge, Cl⁻.
- It now has a full outer shell of electrons and the same electron configuration as the noble gas argon, Ar.



Exercises:

1. Explain why the sodium atom has a larger radius than the sodium ion.
The sodium atom has 3 occupied energy levels and the sodium ion has 2 occupied energy levels.
The sodium atom has equal numbers of protons and electrons.
The sodium ion has more protons than electrons (11:10), therefore the remaining electrons experience a stronger electrostatic attraction from the protons, making the ion smaller.
2. Explain why the chloride ion has a larger radius than the chlorine atom.
The chloride ion has gained an electron, therefore it has more electrons than protons. The extra electron causes extra repulsion between electrons making the ion bigger.
The electrons in the ion feel a weaker electrostatic attraction from the same number of protons, therefore the electrons are held less tightly (making the ion bigger).
3. Explain why metals tend to lose electrons and form positive ions and non-metals tend to gain electrons and form negative ions.
Metals have low ionisation energy values so they can electrons easily to form positive ions. Non-metals have higher ionisation energy values so they gain electrons to form negative ions.

Structure 2.1.2

Understandings:

- The ionic bond is formed by electrostatic attractions between oppositely charged ions.
- Binary ionic compounds are named with the cation first, followed by the anion. The anion adopts the suffix “ide”.

Learning outcomes:

- Deduce the formula and name of an ionic compound from its component ions, including polyatomic ions.
- Interconvert names and formulas of binary ionic compounds.

Additional notes:

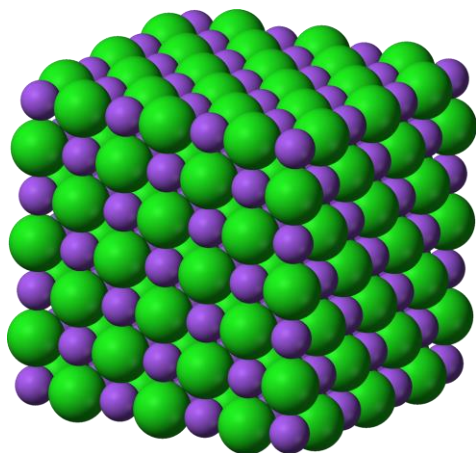
- The following polyatomic ions should be known by name and formula: ammonium, hydroxide, nitrate, hydrogencarbonate, carbonate, sulfate and phosphate.

Linking questions:

- Reactivity 3.2 Why is the formation of an ionic compound from its elements a redox reaction?
- HL Structure 2.2 How is formal charge used to predict the preferred structure of sulfate?
- HL Reactivity 3.1 Polyatomic anions are conjugate bases of common acids. What is the relationship between their stability and the conjugate acid's dissociation constant, K_a ?

Ionic bonding

- An ionic bond is the electrostatic attraction between oppositely charged ions.
- An ionic bond forms between two elements with a difference in electronegativity of equal to or greater than 1.8 units (but there are exceptions).
- Ionic bonds occur between metal and non-metal elements (those elements on the left and right of the periodic table).



Ionic compounds have a lattice structure. The lattice is held together by the electrostatic attraction between the oppositely charged ions. Each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions. Ionic compounds are solids under standard conditions.

Exercises:

1. Describe an ionic bond.
An ionic bond is the electrostatic attraction between oppositely charged ions.

2. An element on the far left and an element on the far right of the periodic table form a chemical bond. Determine the type of bond they would form and explain your reasoning.
An element on the far left is likely to be a metal, an element on the far right is likely to be a non-metal.
Metals and non-metals on opposite ends of the periodic table have a big difference in electronegativity (1.8 units or greater), therefore they form an ionic bond.

3. Describe the lattice structure of NaCl.
The sodium and chloride ions are held together by their opposite charges.
Each sodium ion is surrounded by six chloride ions.
Each chloride ion is surrounded by six sodium ions.

Writing formulae of ionic compounds

List of common ions:

Positive ions (cations)		Negative ions (anions)	
Name	Symbol	Name	Symbol
Hydrogen	H ⁺	Fluoride	F ⁻
Sodium	Na ⁺	Chloride	Cl ⁻
Silver	Ag ⁺	Bromide	Br ⁻
Potassium	K ⁺	Iodide	I ⁻
Lithium	Li ⁺	Hydrogencarbonate	HCO ₃ ⁻
Ammonium	NH ₄ ⁺	Hydroxide	OH ⁻
Barium	Ba ²⁺	Nitrate	NO ₃ ⁻
Calcium	Ca ²⁺	Oxide	O ²⁻
Copper(II)	Cu ²⁺	Sulfate	SO ₄ ²⁻
Magnesium	Mg ²⁺	Carbonate	CO ₃ ²⁻
Zinc	Zn ²⁺	Phosphate	PO ₄ ³⁻
Mercury(I)	Hg ⁺	Nitride	N ³⁻
Lead	Pb ²⁺	Sulfide	S ²⁻
Iron(II)	Fe ²⁺	Phosphide	P ³⁻
Iron(III)	Fe ³⁺	Nitrite	NO ₂ ⁻
Aluminium	Al ³⁺	Sulfite	SO ₃ ²⁻

Exercise: Write the formulae for the following ionic compounds:

- Potassium bromide
- Calcium fluoride
- Beryllium sulfide
- Strontium iodide
- Magnesium nitride
- Ammonium sulfate
- Iron(III) sulfite
- Copper(II) nitrite
- Potassium hydrogencarbonate
- Aluminium sulfate
- Aluminium oxide
- Sodium carbonate
- Copper(II) phosphide
- Zinc phosphate
- Ammonium nitrate
- Mercury(I) nitride
- Iron(II) nitrite
- Barium nitrate
- Iron(II) phosphide
- Calcium hydrogencarbonate

Answers:

1) KBr

2) CaF₂

3) BeS

4) SrI₂

5) Mg₃N₂

11) (NH₄)₂SO₄

12) Fe₂(SO₃)₃

13) Cu(NO₂)₂

14) KHCO₃

15) Al₂(SO₄)₃

6) Al₂O₃

7) Na₂CO₃

8) Cu₃P₂

9) Zn₃(PO₄)₂

10) NH₄NO₃

16) Hg₃N

17) Fe(NO₂)₂

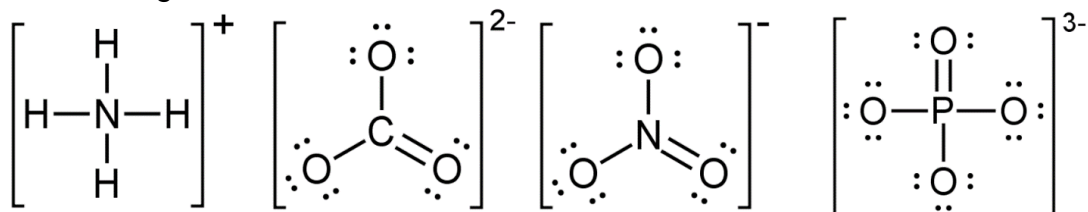
18) Ba(NO₃)₂

19) Fe₃P₂

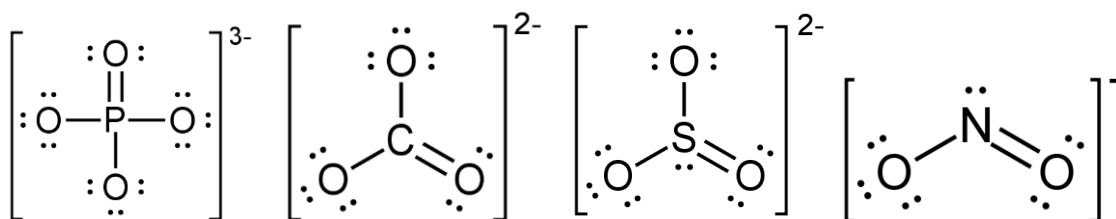
20) Ca(HCO₃)₂

Polyatomic ions

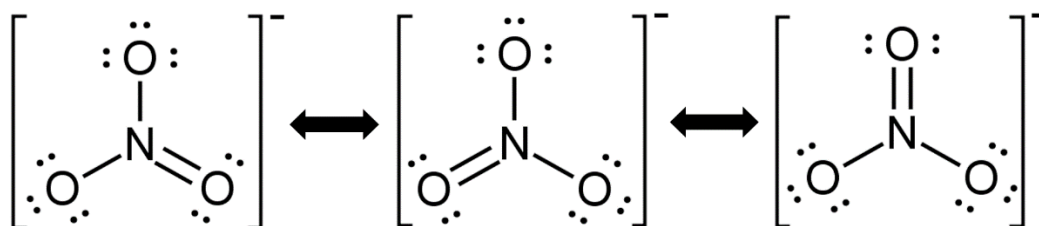
- Polyatomic ions (molecular ions) are ions that consist of two or more atoms bonded together with covalent bonds.



- The atoms in a polyatomic ion are bonded with covalent bonds.
- The bonding between the ions in a compound that contains a polyatomic ion is ionic.
- The geometry of a polyatomic ion depends on the number of electron domains around the central atom.



- Polyatomic ions with more than one position for a multiple bond exist as resonance structures.



- The N-O bonds are equal length and equal strength – intermediate in length and strength between a single and a double bond.

Exercise: State the two types of bonding in an ionic compound containing a polyatomic ion.

An ionic compound that contains a polyatomic ion has two types of bonding; covalent bonds between the atoms in the polyatomic ion and ionic bonding between the constituent ions.

Structure 2.1.3

Understandings:

- Ionic compounds exist as three-dimensional lattice structures, represented by empirical formulas.

Learning outcomes:

- Explain the physical properties of ionic compounds to include volatility, electrical conductivity and solubility.

Additional notes:

- Include lattice enthalpy as a measure of the strength of the ionic bond in different compounds, influenced by ion radius and charge (HL).

Linking questions:

- Structure 3.1 How can lattice enthalpies and the bonding continuum explain the trend in melting points of metal chlorides across period 3?

Properties of ionic compounds

Electrical conductivity

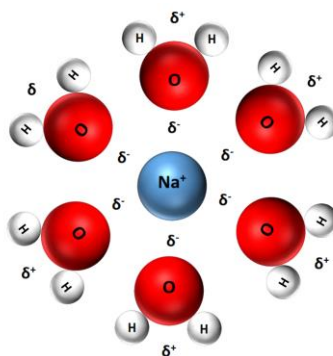
- Ionic compounds do not conduct electricity when solid, because the ions are held in fixed positions by the electrostatic attractions between the ions.
- They only conduct electricity when molten (melted) or dissolved in water (aqueous).
- When molten or dissolved, the ions are free to move and conduct the electric current.

Melting and boiling point

- Ionic compounds have high melting and boiling points due to the strong electrostatic attractions between the oppositely charged ions (NaCl has a melting point of 800 °C).
- The greater the charge and smaller the ionic radius of the ions, the stronger the electrostatic attraction and the higher the melting point.

Solubility

- Ionic compounds are soluble in polar solvents.
- The ions are separated from the lattice structure by the polar water molecules.
- The ions become surrounded by water molecules (hydration) as shown below.



Exercises:

1. Explain the conductivity of ionic compounds when molten or dissolved and when solid.

Solid ionic compounds do not conduct electricity because the ions are held in fixed positions in the lattice structure by strong electrostatic attractions. When molten or dissolved, the ions are free to move and can conduct electricity.

2. Explain the high melting point of ionic compounds.

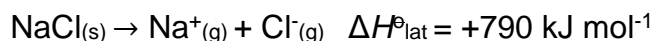
The high melting point of ionic compounds is due to the strong electrostatic attractions between the oppositely-charged ions.

3. Explain why NaF has a higher melting point than KF.

Sodium has a smaller ionic radius than potassium. The smaller ionic radius means that the electrostatic attraction between the ions in sodium chloride is stronger than the electrostatic attraction in potassium fluoride.

Lattice enthalpy ($\Delta H^{\ominus}_{\text{lat}}$)

- Lattice enthalpy ($\Delta H^{\ominus}_{\text{lat}}$) is the enthalpy change when one mole of solid ionic compound is separated into its gaseous ions under standard conditions.



- Lattice enthalpy influences the strength of an ionic bond – higher values of lattice energy correspond to a stronger ionic bond and vice versa.
- There are two factors that affect the value of the lattice enthalpy: the charge on the ions and the ionic radius of the ions.
- Ions with a higher charge have a stronger force of attraction, therefore a higher lattice enthalpy.
- Smaller ions have a stronger force of attraction, therefore a higher lattice enthalpy.

Exercises:

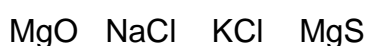
- Which has the higher lattice enthalpy, MgCl_2 or NaCl ?

MgCl_2 has a higher lattice enthalpy than NaCl because the Mg^{2+} ion has a higher charge than the Na^+ ion.

- Which has the higher lattice enthalpy, NaBr or KBr ?

NaBr has a higher lattice enthalpy than KBr because the Na^+ ion has a smaller ionic radius than the K^+ ion.

- Arrange the following in order of increasing magnitude of lattice enthalpy (lowest first).



$\text{KCl} < \text{NaCl} < \text{MgS} < \text{MgO}$