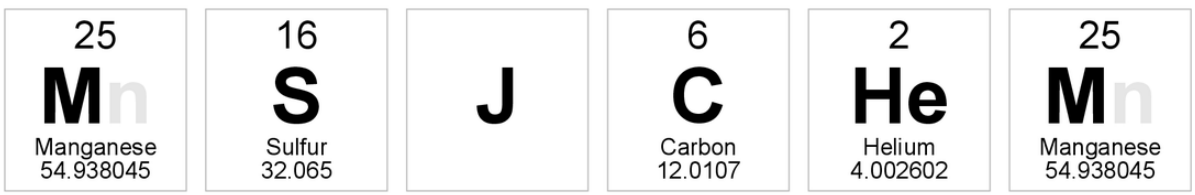
Structure 2.1

IB CHEMISTRY SL



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:

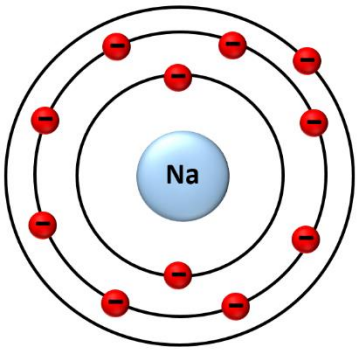
- Structure 3.1 How does the position of an element in the periodic table relate to the charge of its ion(s)
- Structure 1.3 (HL) How does the trend in successive ionisation energies of transition elements explain their variable oxidation states?

Ions

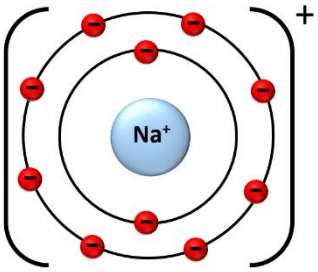
- Neutral atoms have equal numbers of protons and electrons.
- Positive ions (cations) are formed when atoms lose electrons.
- Negative ions (anions) are formed when atoms gain electrons.

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.



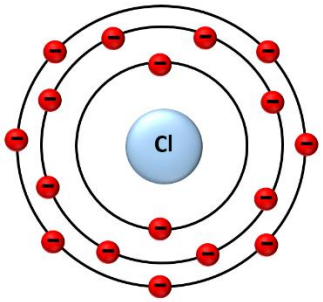
$1s^2\ 2s^2\ 2p^6\ 3s^1$

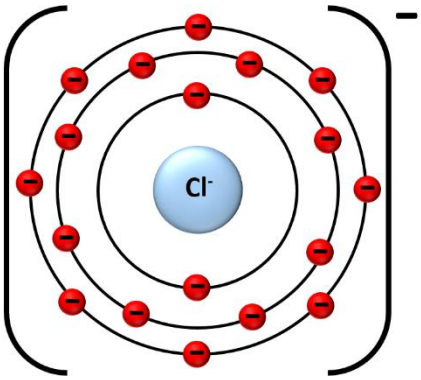


$1s^2\ 2s^2\ 2p^6$

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.





$1s^2\ 2^2\ 2p^6\ 3s^2\ 3p^5$ **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$**

Exercises:

1. Explain why the sodium atom has a larger radius than the sodium ion.

2. Explain why the chloride ion has a larger radius than the chlorine atom.

3. Explain why metals tend to lose electrons and form positive ions and non-metals tend to gain electrons and 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:

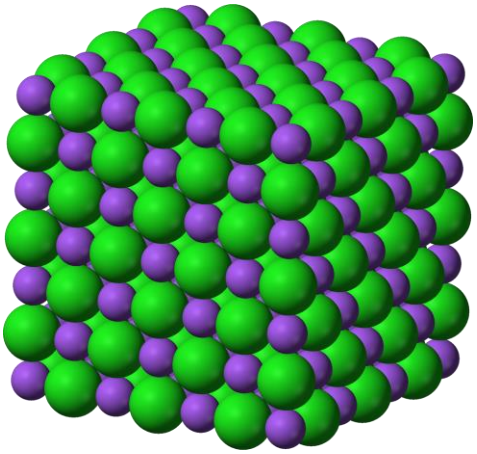
- The following polyatomic ions should be known by name and formula: ammonium, hydroxide, nitrate, hydrogencarbonate, carbonate, sulfate and phosphate.

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.

2. An element on the far left and an element on the far right of the periodic table form a chemical bond. Describe the type of bond they would form and explain your reasoning.

3. Describe the lattice structure of NaCl.

Writing formulae of ionic compounds

List of common ions

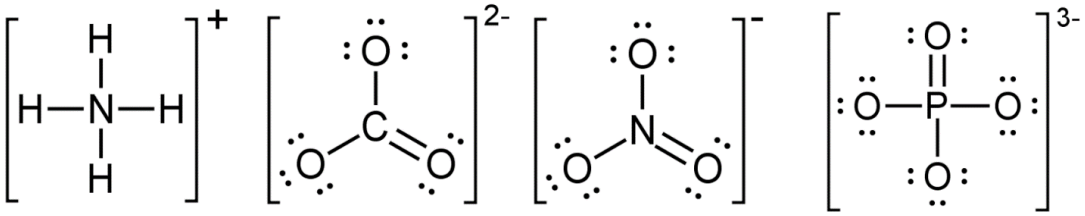
Positive ions (cations)		Negative ions (anions)	
Name	**Symbol**	**Name**	**Symbol**
Hydrogen		Fluoride	
Sodium		Chloride	
Silver		Bromide	
Potassium		Iodide	
Lithium		Hydrogencarbonate	
Ammonium		Hydroxide	
Barium		Nitrate	
Calcium		Oxide	
Copper(I)		Sulfate	
Copper(II)		Carbonate	
Magnesium		Phosphate	
Zinc		Nitride	
Mercury(I)		Sulfide	
Lead		Phosphide	
Iron(II)		Nitrite	
Iron(III)		Sulfite	
Aluminium			

Exercise: Write the formulae for the following ionic compounds:

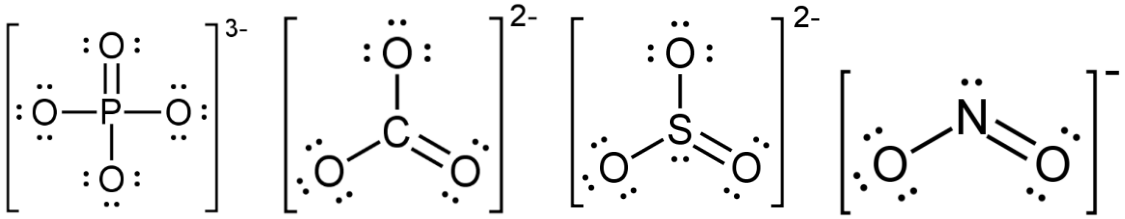
1) Potassium bromide
2) Calcium fluoride
3) Beryllium sulfide
4) Strontium Iodide
5) Magnesium nitride
6) Aluminium oxide
7) Sodium carbonate
8) Copper(II) phosphide
9) Zinc phosphate
10) Ammonium nitrate
11) Ammonium sulfate
12) Iron(III) sulfite
13) Copper(II) nitrite
14) Potassium hydrogencarbonate
15) Aluminium sulfate
16) Mercury(I) nitride
17) Iron(II) nitrite
18) Barium nitrate
19) Iron(II) phosphide
20) Calcium hydrogencarbonate

Polyatomic ions

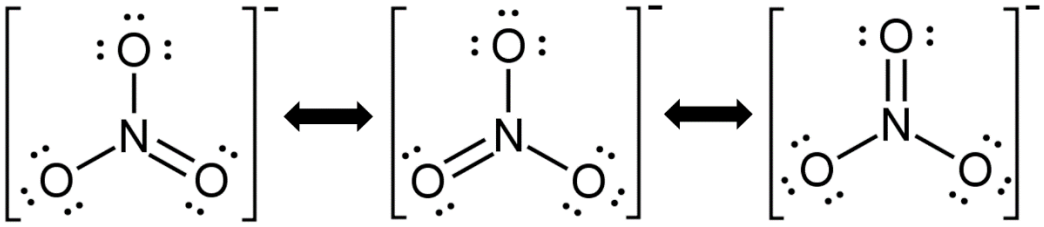
- Polyatomic ions (molecular ions) are ions that consist of two or more atoms bonded together with covalent bonds.



- 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.

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

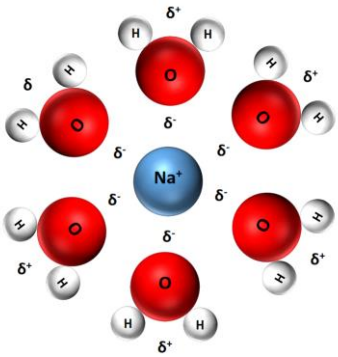
- Ionic compounds do not conduct electricity when solid, because the ions are held in fixed positions by the electrostatic attractions between the ions.
- They only conduct electricity when molten (melted) or dissolved in water (aqueous).
- When molten or dissolved, the ions are free to move and conduct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

2. Explain the high melting point of ionic compounds.

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

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

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

$$NaCl_{(s)} \rightarrow Na^{+}_{(g)} + Cl^{-}_{(g)} \quad \Delta H^{\ominus}_{lat} = +790 \text{ kJ mol}^{-1}$$

- 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:

1. Which has the higher lattice enthalpy, $MgCl_2$ or NaCl?

2. Which has the higher lattice enthalpy, NaBr or KBr?

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

MgO NaCl KCl MgS