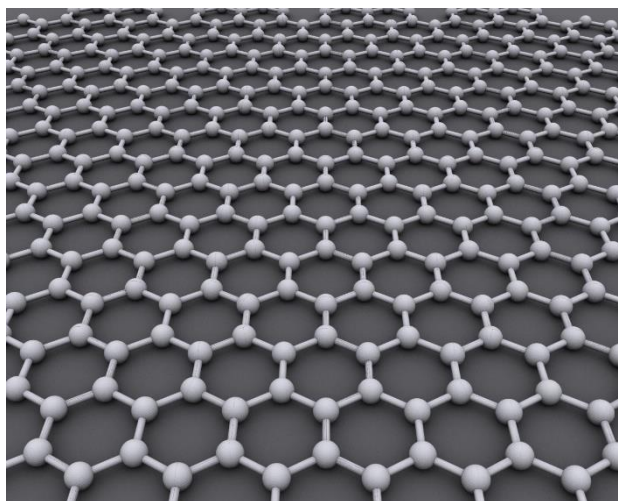**Master**

Kirkcaldy High School

Higher Chemistry

Unit 1 – part 1

Chemical changes and structure

Name: _____

Class: _____

Teacher: _____

Assessment Page

Homework

Homework title	Date	Mark/Total Mark
N5 General Bonding Recap		**/**
H Structure and Bonding in the first 20 elements	.	**/**

Notes/comments

Check tests

Test title	Date	Mark/Total Mark
Unit 1: Key area 1ai		**/**
Unit 1: Key area 1aii		**/**

Notes/comments

Structure and bonding in the first 20 elements

Overarching question(s) for this topic

- What is the structure and bonding in the first 20 elements?

- How do the bonding and structure of the first 20 elements influence their melting and boiling points, and what trends can be observed in these properties?

▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬ ▬

Atomic Number, groups and periods

Atomic Number:

The atomic number is the number of **protons** in the **nucleus** of an atom. It defines the element and its position in the Periodic Table.

If the atom is **neutral**, the atomic number also represents the number of **electrons** in the element.

Groups:

Vertical columns in the Periodic Table are called **groups**.

Elements in the same group have the same number of **outer** shell electrons (**valence** electrons), leading to **similar** chemical properties.

- The number of **outer** electrons corresponds to the **group number** for Groups 1-7.

Periods:

- Horizontal rows in the Periodic Table are called **periods**.

- Elements in a period are arranged in order of **increasing** atomic number.

- Moving across a period from left to right, elements transition from **metals** to **non-metals**, and their properties change from **metallic** to **non-metallic**.

Using your knowledge from National 5, label group 1, 2, 7 and 0/8 below. Use your data booklet to write in the symbols for the elements for each group you are labelling. Draw in the step that shows the transition from metal to non-metal elements.

Key

Atomic number

Name of element

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
1 Hydrogen							2 Helium
3 Lithium	4 Beryllium	5 Boron	6 Carbon	7 Nitrogen	8 Oxygen	9 Fluorine	10 Neon
11 Sodium	12 Magnesium	13 Aluminium	14 Silicon	15 Phosphorus	16 Sulfur	17 Chlorine	18 Argon
19 Potassium	20 Calcium	31 Gallium	32 Germanium	33 Arsenic	34 Selenium	35 Bromine	36 Krypton
37 Rubidium	38 Strontium	49 Indium	50 Tin	51 Antimony	52 Tellurium	53 Iodine	54 Xenon
55 Caesium	56 Barium	81 Thallium	82 Lead	83 Bismuth	84 Polonium	85 Astatine	86 Radon

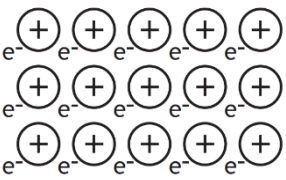**Label key**: alkali metals ☐ alkali earth metals ☐ halogens ☐ noble gases ☐

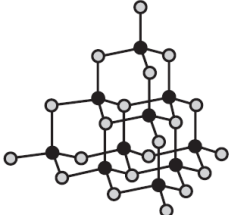---
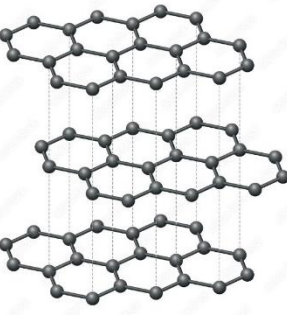
Summarise the reactivity (including the trend down the group) of group 1 elements:

_____

_____

Summarise the reactivity (including the trend down the group) of group 7 elements:

_____

_____

Summarise the reactivity of group 0/8 elements:

_____

Bonding and Structure of the First 20 Elements

Summarise N5 structure and bonding. Annotate the diagrams where necessary.

Metallic Lattice	Covalent Network	Covalent Molecular	Monatomic

C(**diamond**)

C(**graphite**)

Metallic Lattice: Describe the structure of a metallic lattice and state their general properties:

_____

_____

Covalent Network: Describe the structure of a covalent network and state their general properties (and the exception):

_____

_____

_____

Covalent Molecules: Describe the structure of covalent molecules and state their general properties:

_____

_____

Monatomic: Describe the structure of covalent molecules and state their general properties:

_____

_____

Using the melting point and boiling point data, label the first 20 elements with their structure and bonding. Add in the 'step'. Remember, room temperature is around 25 °C when deciding the state of the element. Add a 'D' to the diatomic elements.

.

Key

Atomic number
Name of element
Melting point (°C)
Boiling point (°C)

Group 1	Group 2		Group 3	Group 4	Group 5	Group 6	Group 7	Group 0
1 Hydrogen −259 −253								2 Helium −271 −269
3 Lithium 181 1342	4 Beryllium 1287 2471*		5 Boron 2077 4000	6 Carbon †3825	7 Nitrogen −210 −196	8 Oxygen −219 −183	9 Fluorine −220 −188	10 Neon −249 −246
11 Sodium 98 883	12 Magnesium 650 1090		13 Aluminium 660 2519	14 Silicon 1414 3265	15 Phosphorus 44 281	16 Sulfur 115 445	17 Chlorine −101 −34	18 Argon −189 −186
19 Potassium 63 759	20 Calcium 842 1484							

Label key: metallic ☐ covalent network ☐ monatomic ☐

covalent molecular gas ☐ covalent molecular solid ☐

Covalent Molecular Solids

Phosphorous, sulfur and carbon(fullerene) are all covalent molecular **solids**. Like the diatomic elements, these elements have a **specific** number of atoms in each molecule. Phosphorous has **4** atoms each molecule, sulfur has **8** atoms each molecule and carbon(fullerene) has **60** atoms in each molecule.

Element	Phosphorous	Sulfur	Carbon (fullerene)
Formula	P_4	S_8	C_{60}
Structure			

Explaining the trends in melting and boiling points in the elements

Covalent Bonding (recap)

A covalent bond forms when two **positive nuclei** are held together by their **common attraction** for a **shared pair** of **electrons**.
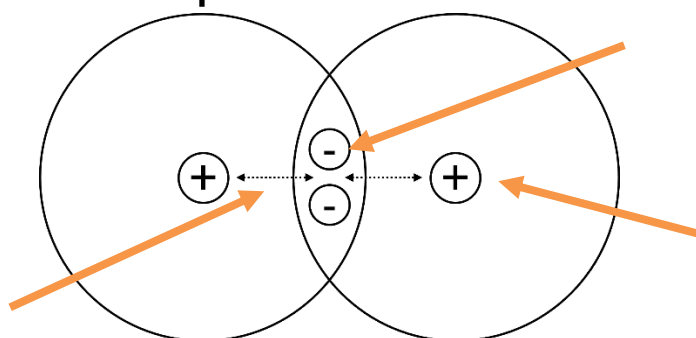
These bonds are very **strong** are therefore hard to **break**.

Covalent networks

In covalent networks atoms are bonded by strong covalent bonds in a continuous network.

- To melt a covalent network, these strong **covalent** bonds must be **broken**. This requires a **large** amount of energy, so will only melt at **very high** temperatures.

Covalent molecular elements

Although molecules do contain strong **intramolecular covalent** bonds, it is **not** these bonds that break when a molecule melts or boils. It is the weak **intermolecular** forces that break.
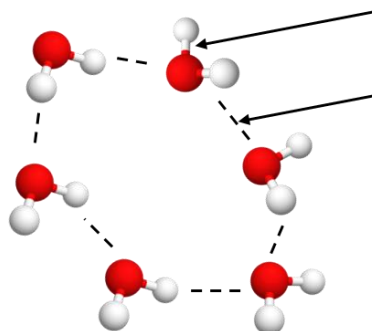
From here, we will often refer to **intermolecular** forces as **Van der Waals** forces.

There are different types of Van der Waals, however, covalent molecular elements only contain one - **London Dispersion Forces** (**LDF**).

London Dispersion Forces

LDFs are **weak** intermolecular forces that arise due to **temporary dipoles** created when **electrons** are **unevenly** distributed around a molecule or atom.

- The electrons become unevenly distributed when a **neutral** atom/molecule approaches another atom/molecule and there is **attraction** between the **nucleus** of one atom and the **electrons** of another.

- **All atoms** experience LDFs, but they are especially significant in covalent molecular and monatomic elements.
  - In metals and covalent networks, LDFs do exist but are **insignificant** compared to their strength **dominant** bonding forces.

Space for teacher explanation:

How do LDFs arise?
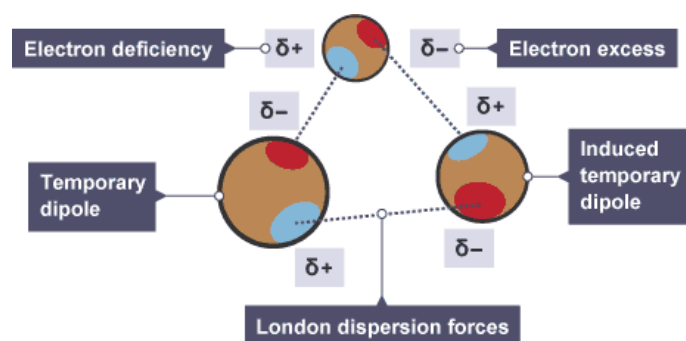
LDFs form as a result of **electrostatic** attraction between **temporary** dipoles and **induced** dipoles caused by **movement** of **electrons** in atoms and molecules.

Increasing the strength of LDF

As the number of **electrons** in an atom/molecule **increases**, the strength of LDF **increases**. As the **strength** of LDF **increases**, the melting/boiling point of the molecule **increases**.

Questions - LDF (in your jotter)

1. a) State the trend in the boiling point as you go down group 0/8. (1 mark)

b) Explain fully, in terms of structure and the type of van der Waals forces present, why the boiling points of the group 0/8 elements show this trend. (3 mark)

2. a) State the trend in the boiling point as you go down group 7. (1 mark)

b) Explain fully, in terms of structure and the type of van der Waals forces present, why the boiling point of chlorine is higher than that of argon. (3 mark)

3. a) State the boiling points of nitrogen and fluorine.

b) Write the formula for nitrogen and fluorine and count the total number of electrons in both molecules.

c) Explain fully, in terms of structure and the type of van der Waals forces present, why the boiling point of fluorine is higher than that of nitrogen. (3 marks)

4. The boiling point of chlorine is much higher than that of argon. Explain fully, in terms of structure and the type of van der Waals forces present, why the boiling point of chlorine is higher than that of argon. (3 marks)

**5.** The melting point of sulfur is much higher than that of phosphorus.

Explain fully, in terms of the structures of sulfur and phosphorus molecules and the intermolecular forces between molecules of each element, why the melting point of sulfur is much higher than that of phosphorus. (3 marks)

Summary of melting point and boiling point trends

Complete the table below to summarise the type of bonding broken when different types of bonding/structure are melted/boiled.

Type of Bonding/structure	Covalent Networks	Covalent Molecules	Monatomic
Melting/Boiling point (high/low)			
Name of bond broken when melting/boiling			

Flash Cards:
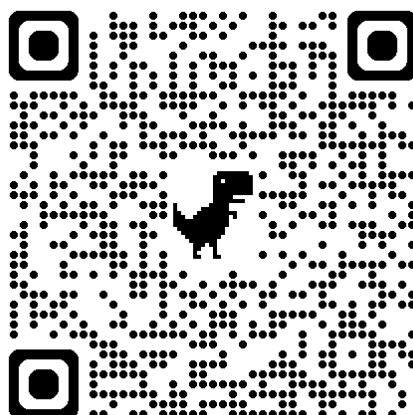
[Structure and Bonding in the first 20 elements](#)

Extension questions:
Hodder and Gibson page 21-35 (questions on page 35)
Blue Chemcord book page 7-9
SCHOLAR

Past Papers

S & B 1st 20	2015	2016	2017	2018	2019	2020	2022	2023
MC	1		4				1	
S2	1.a-b (i) 1.b. (iii)	3. a. 3. c.	1.a. 1.c.	1.b. 2.c.(i)	3.	2a.		2.a.(i)

TEAMS: Check Test – Unit 1: Key Area 1ai

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1. Silicon tetrachloride can be used to make silicon nitride (Si_3N_4), a compound found in many cutting tools. Silicon nitride has a melting point of 1900 °C and does not conduct electricity when molten.
   a. **Explain fully**, in terms of structure and bonding, why silicon nitride has a high melting point. (2 marks)

2. Phosphine (PH_3) is used as an insecticide in the storage of grain. Phosphine can be produced by the reaction of water with aluminium phosphide

   $$AlP(s) + 3H_2O(\ell) \rightarrow PH_3(g) + Al(OH)_3(aq)$$

   a. **State** the type of bonding and structure in phosphine (1 mark)

3. Sulfur and oxygen contain London dispersion forces.
   a. **Explain** how London dispersion forces arise. (2 marks)
   b. **Explain** fully why the London dispersion forces in sulfur are stronger than those in oxygen. (2 marks)

4. Hydrogen gas has a boiling point of –253 °C.
   a. **Explain** clearly why hydrogen is a gas at room temperature. In your answer you should name the intermolecular forces involved and indicate how they arise. (2 marks)

5. Volcanoes produce a variety of molten substances, including sulfur and silicon dioxide.
   a. Complete the table to show the strongest type of attraction that is broken when each substance melts. (2 marks)

Substance	Melting point (^0C)	Strongest type of attraction broken when substance melts
Sulfur	113	
Silicon dioxide	1610	

6. The melting point of sulphur is much higher than that of phosphorus.
   a. **Explain fully**, in terms of the structures of sulphur and phosphorus and the intermolecular forces between molecules of each element, why the melting point of sulphur is much higher than that of phosphorus. (3 marks)

7. The melting point of non-metal elements depends on structure and bonding. **Using your knowledge of chemistry**, comment on this statement. (3 marks)

Periodicity

Overarching question(s) for this topic

- How do the properties of elements in the Periodic Table change across periods and down groups?

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Atomic Structure and Electron Arrangements

Atoms are the basic units of matter, consisting of a nucleus (containing protons and neutrons) and electrons that orbit the nucleus.

- **Protons** are **positively** charged particles found in the nucleus.

- **Neutrons** are **neutral** particles found in the nucleus.

- **Electrons** are **negatively** charged particles that orbit the nucleus in regions called electron **shells**.

Electron Shells and Electron Arrangement

Electron shells (energy levels) are layers around the nucleus where electrons are likely to be found.

Electron arrangements: These show the number of electrons in each shell, for the first 20 elements they follow the **2,8,8 rule.**

Questions
*Using your data booklet, write the electron arrangement for the following.*

Atom	Electron Arrangement	Ion	Electron Arrangement
H		P^{3-}	
He		S^{2-}	
Li		Al^{3+}	
C		Na^{+}	
Ne		Mg^{2+}	
S		Cl^{-}	
O		K^{+}	

Understanding the trends

Nuclear Charge

- **Nuclear charge** is the total charge of the **protons** in the nucleus, equal to the **atomic** number of the element.
    - A **higher** nuclear charge means a **greater positive** charge in the nucleus, attracting electrons more **strongly**.
    - **Effective** nuclear charge is the net positive charge experienced by an electron in an atom

Label the diagram below
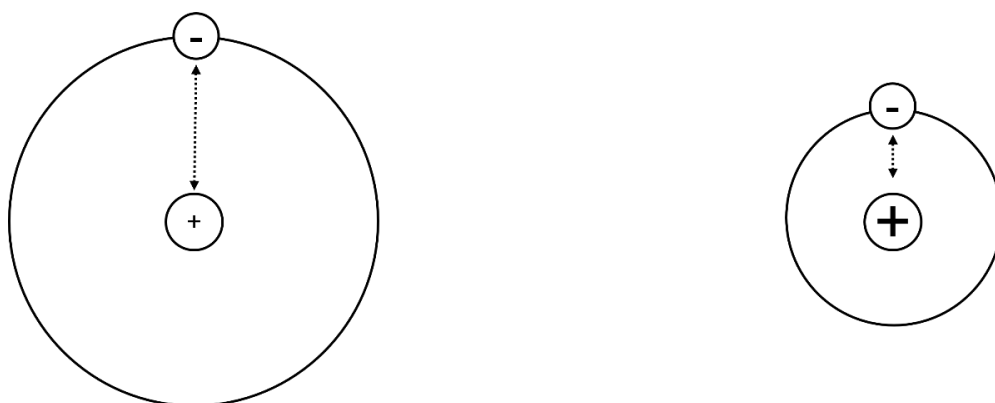
Shielding Effect

- The shielding effect occurs when inner-shell electrons **reduce** the **effective nuclear** charge felt by **outer**-shell electrons.
    - Inner electrons repel outer electrons, weakening the attraction between the nucleus and outer electrons.
        - The **greater** the number of electron **shells** there are, the **greater** the shielding effect.
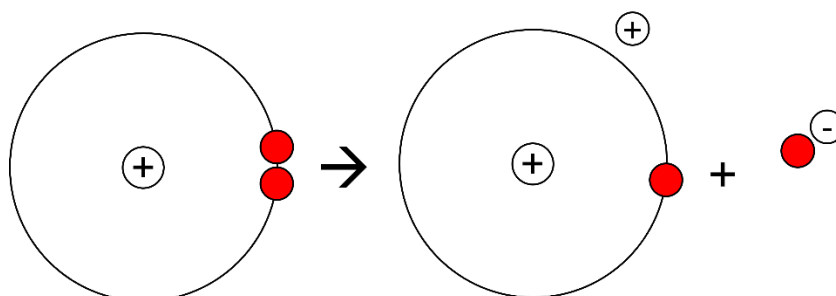
Label the diagram below
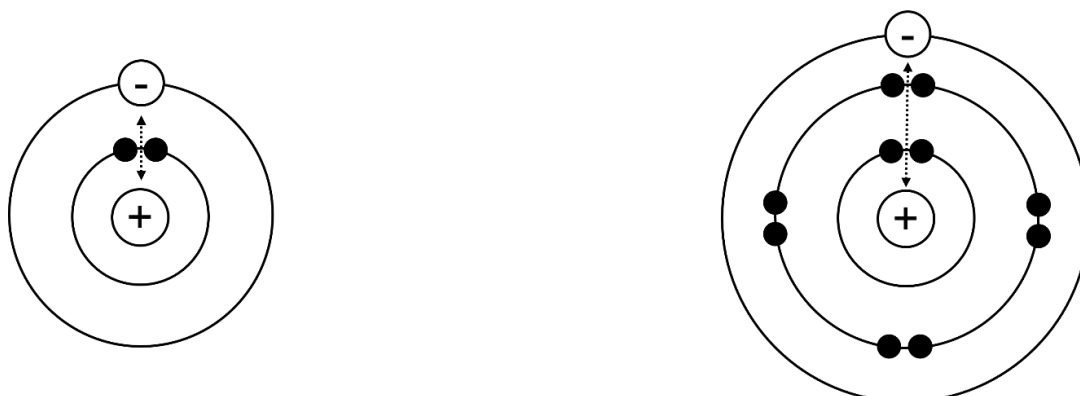
Definitions of the properties of atoms

Covalent radius is a measure of the size of an atom that forms part of a covalent bond. It is half the distance between the nuclei of two identical atoms covalently bonded together. (values on page 7 of your DB). *Label the diagram.*
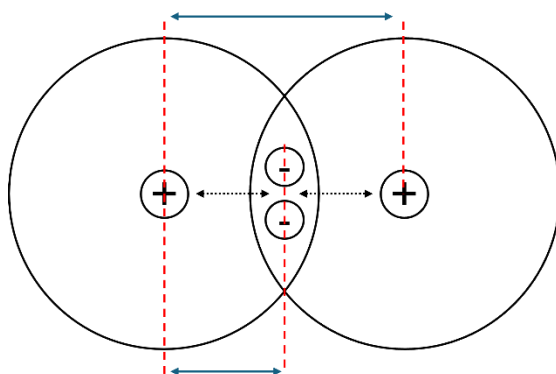
Ionisation energy is the energy required to remove a mole of electrons from a mole of atoms in the gaseous state. (values on page 12 of your DB). *Label the diagram.*

Electronegativity is the ability of an atom to attract electrons in a covalent bond. *Label the diagram.*
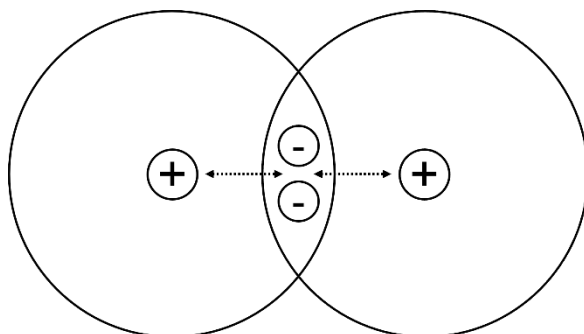
These are measured with the Pauling Scale (values on page 12 of your DB).

Questions (using the data booklet)

Using your data booklet, write the values for each of the trends in the tables.

Table 1: Elements of period 2.

Element	Li	Be	B	C	N	O	F
Covalent Radius (pm)							
First Ionisation Energy (kJ/mol)							
Electro-negativity							

Table 2: Elements of group 1.

Element	Li	Na	K	Rb	Cs
Covalent Radius (pm)					
First Ionisation Energy (kJ/mol)					
Electro-negativity					

Analyse your information for trends, describe the patterns emerging.

_____

_____

_____

_____

15

Using the data booklet, draw a scatter graph of the covalent radius for the first 20 elements. Label any trends you notice on the graph or below.

Atomic number vs. Covalent Radius (pm)

Covalent Radius (pm) vs. Atomic Number
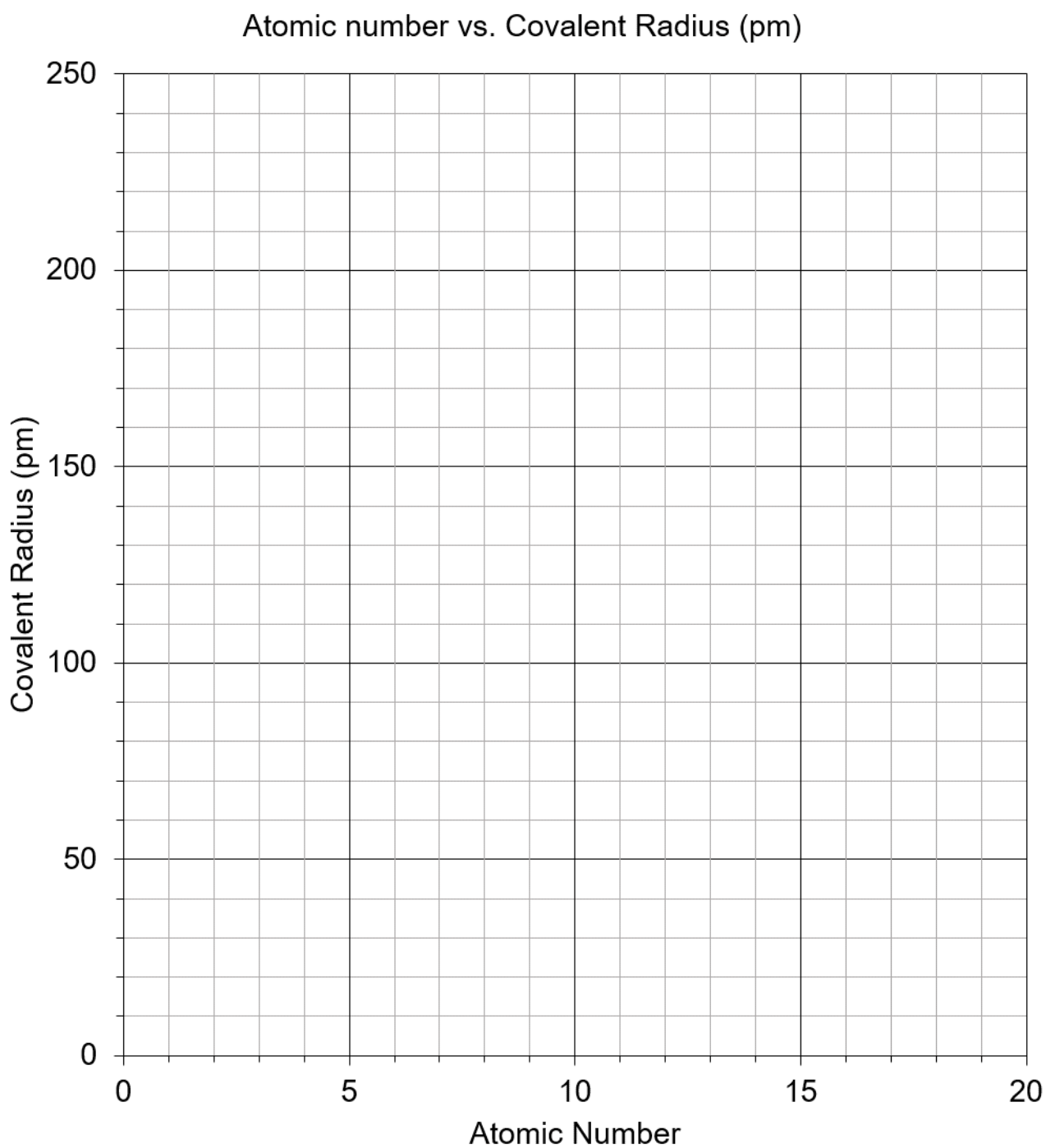
Using the data booklet, draw a scatter graph of the first and second ionisation energies for the first 20 elements. Both ionisation energies will be on the graph, make sure to add a key. Label any trends you notice on the graph or below.

First and second Ionisation Energies for the first 20 elements

Ionisation Energy (kJ/mol) vs Atomic Number
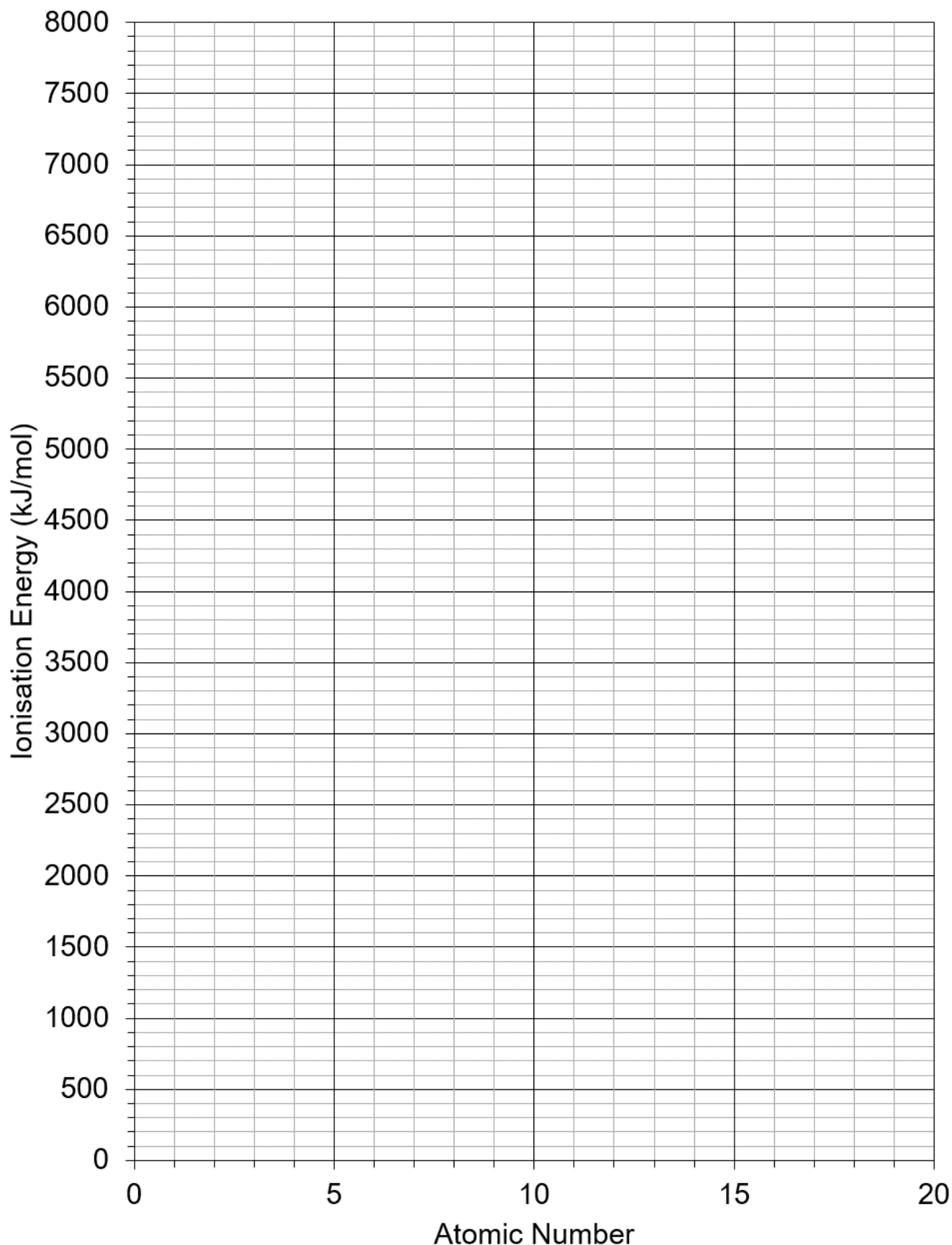
The influence of Nuclear Charge and Shielding

Covalent Radius
Becomes …

Larger down a group:
- Due to more electron shells, increased **shielding** reduces the **effective nuclear** charge felt by outer electrons, causing the atom to be **larger**.

Smaller across a period:
- Nuclear charge **increases** while **shielding** remains relatively **constant**, pulling electrons closer to the nucleus and **reducing** atomic size.

Ionisation Energy
Becomes …

Decreases down a group:
- Increased shielding **reduces** the **effective nuclear** charge on outer electrons, making them easier to **remove**.

Increases across a period:
- Increasing nuclear charge pulls electrons closer to the nucleus, making them harder to remove.

Electronegativity
Becomes …

Decreases down a group:
- Increased **shielding** reduces the **effective nuclear** charge, making it **harder** for the nucleus to attract **bonding** electrons.

Increases across a period:
- Increasing nuclear charge **attracts** electrons more strongly, increasing electronegativity.

Table for summary

Summarise the trends in the table below.

Atomic Structure Trend	Periodic Trend	Effect on Covalent Radius	Effect on Ionisation Energy	Effect on Electro-negativity
No. of Protons increase	**Nuclear Charge**	**Smaller**	**Higher**	**Increases**
No. of Electron Shells Increase	**Shielding Effect**	**Larger**	**Lower**	**Decreases**

Questions – definitions and trends (in your jotter)

1. **Define** nuclear charge.

2. **Explain** how nuclear charge affects the attraction between the nucleus and electrons.

3. **Define** the shielding effect.

4. **Explain** why the shielding effect increases down a group in the periodic table.

5. **Define** covalent radius.

6. **Describe** how the covalent radius changes as you move down a group.

7. **Explain** why the covalent radius decreases across a period.

8. **Define** ionisation energy.

9. **Explain** why ionisation energy generally decreases down a group.

10. **Explain** why ionisation energy increases across a period.

11. **Define** electronegativity.

12. **Describe** how electronegativity changes down a group and explain why.

13. **Explain** why electronegativity increases across a period.

14. **Describe** how nuclear charge and shielding effect together influence the covalent radius of an atom.

15. **Explain** how ionisation energy is affected by both nuclear charge and shielding effect.

16. **Describe** the relationship between electronegativity, nuclear charge, and shielding effect.

Using your knowledge of the trends and information from the data booklet, determine the corners of the periodic table with the lowest and highest covalent radius, ionisation energy, electronegativity and add these labels to the periodic table. Add in the trends for nuclear charge and the shielding effect. You will then use the to help explain the trends.

Group 1

Group 1	Group 2
1 Hydrogen	
3 Lithium	4 Beryllium
11 Sodium	12 Magnesium
19 Potassium	20 Calcium
37 Rubidium	38 Strontium
55 Caesium	56 Barium

Key

Atomic number
Name of element

21 Scandium	22 Titanium	23 Vanadium	24 Chromium	25 Manganese	26 Iron	27 Cobalt	28 Nickel	29 Copper	30 Zinc
39 Yttrium	40 Zirconium	41 Niobium	42 Molybdenum	43 Technetium	44 Ruthenium	45 Rhodium	46 Palladium	47 Silver	48 Cadmium
57 Lanthanum	72 Hafnium	73 Tantalum	74 Tungsten	75 Rhenium	76 Osmium	77 Iridium	78 Platinum	79 Gold	80 Mercury

Group 3	Group 4	Group 5	Group 6	Group 7
5 Boron	6 Carbon	7 Nitrogen	8 Oxygen	9 Fluorine
13 Aluminium	14 Silicon	15 Phosphorus	16 Sulfur	17 Chlorine
31 Gallium	32 Germanium	33 Arsenic	34 Selenium	35 Bromine
49 Indium	50 Tin	51 Antimony	52 Tellurium	53 Iodine
81 Thallium	82 Lead	83 Bismuth	84 Polonium	85 Astatine

Questions – explaining trends in the periodic table (in your jotter)

1. **Explain** why the *first ionisation energy* of lithium (Li) is higher than that of sodium (Na).

2. **Explain** why the *covalent radius* of chlorine (Cl) is smaller than that of bromine (Br).

3. **Explain** why the *electronegativity* of fluorine (F) is higher than that of iodine (I).

4. **Explain** why the *first ionisation energy* of magnesium (Mg) is lower than that of beryllium (Be).

5. **Explain** why the *first ionisation energy* of oxygen (O) is higher than that of nitrogen (N).

6. **Explain** why the *covalent radius* of sodium (Na) is larger than that of aluminium (Al).

7. **Explain** why the *electronegativity* of chlorine (Cl) is higher than that of sulfur (S).

8. **Explain** why the *first ionisation energy* of neon (Ne) is higher than that of fluorine (F).

9. **Explain** why the *first ionisation energy* of lithium (Li) is lower than that of fluorine (F).

10. **Explain** why the *covalent radius* of potassium (K) is larger than that of aluminium (Al).

11. **Explain** why the *electronegativity* of carbon (C) is higher than that of Calcium (Ca).

12. **Explain** why the *first ionisation energy* of sodium (Na) is lower than that of fluorine (F).
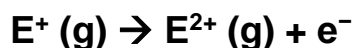
Writing Ionisation Energy

The ionisation energies (I.E.) can be written as equations.

Generic examples of these are shown in page 12 of the DB:

The first ionisation energy for an element E refers to the reaction:

$$E(g) \rightarrow E^+(g) + e^-$$
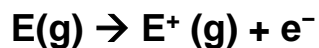
the second ionisation energy refers to:

$$E^+(g) \rightarrow E^{2+}(g) + e^-$$

Questions:

Complete the table by writing the equation for the ionisation energy given.

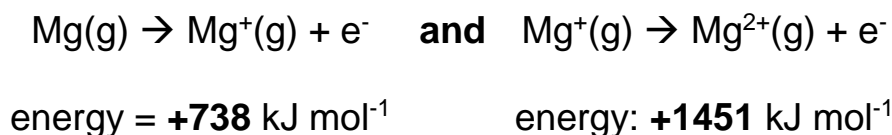| Ionisation Energy	Equation
First ionisation energy of Sodium	$Na(g) \rightarrow Na^+(g) + e^-$
First ionisation energy of Caesium	
Second ionisation energy of Chlorine	$Cl^+(g) \rightarrow Cl^{2+} + e^-$
Second ionisation energy of Beryllium	
Third ionisation energy of Magnesium	
Third ionisation energy of Aluminium	
Fourth ionisation energy of Carbon	
Fourth ionisation energy of Silicon	

Complete the table by stating the ionisation energy given from the equation.

Ionisation Energy	Equation
	$K(g) \rightarrow K^+(g) + e^-$
	$Ca^{2+}(g) \rightarrow Ca^{3+}(g) + e^-$
	$Sr^+(g) \rightarrow Sr^{2+} + e^-$
	$Sn^{3+}(g) \rightarrow Sn^{4+} + e^-$

Combining Ionisation Energies

It is possible to combine ionisation energies to calculate the total energy required to remove multiple moles electrons from an atom. This process involves adding the individual ionisation energies for each electron removed. *The energies for each ionisation energy can be found on page 12 of the data booklet.*

For example

$$Mg(g) \rightarrow Mg^+(g) + e^- \quad \textbf{and} \quad Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$$

$$\text{energy} = \textbf{+738} \text{ kJ mol}^{-1} \qquad \text{energy: } \textbf{+1451} \text{ kJ mol}^{-1}$$

Combine to form

$$Mg(g) \rightarrow Mg^{2+}(g) + \textbf{2e}^-$$

$$\text{Total I.E.} = \text{ first I.E. + Second I.E.}$$
$$= (\textbf{+738}) + (\textbf{+1451})$$
$$= \underline{\textbf{+2189}} \text{ kJ mol}^{-1}$$

Questions

Complete the table to identify the ionisation energies involved and calculate the total ionisation energies.

Equation	Ionisation energies involved	Total ionisation energy
$Mg(g) \rightarrow Mg^{2+}(g) + \textbf{2e}^-$	**1st and 2nd of Mg**	**(+738)+(+1451) = +2189 kJ mol^{-1}**
$Ca(g) \rightarrow Ca^{2+}(g) + \textbf{2e}^-$		
$C^{2+}(g) \rightarrow C^{4+}(g) + \textbf{2e}^-$		
$Mg^+(g) \rightarrow Mg^{3+}(g) + \textbf{2e}^-$		
$Al(g) \rightarrow Al^{3+}(g) + \textbf{3e}^-$		
$B(g) \rightarrow B^{3+}(g) + \textbf{3e}^-$		
$Si(g) \rightarrow Si^{4+}(g) + \textbf{4e}^-$		

Understanding increases in ionisation energies

A. Write the first four I.E.'s *for Aluminium and the associated energies for each (remember the energies are **positive - endothermic**).*
B. *Complete the electron shell diagram with the electron arrangement and label the diagram to show which electron would be removed for each* I.E.*.*
C. *Sketch a graph with the* I.E. *and write an explanation the smaller increases and the large increase between two of the ionisation energies.*
D. *Look at other elements (e.g. Na and Mg) to see which I.E. shows a large increase and write a generic rule for this.*

A.

First Ionisation Energy:

Al(g) → Al⁺(g) + e⁻

Energy: _____ kJ mol⁻¹

Second Ionisation Energy:

Al⁺(g) → Al²⁺(g) + e⁻

Energy: _____ kJ mol⁻¹

Third Ionisation Energy:

Al²⁺(g) → Al³⁺(g) + e⁻

Energy: _____ kJ mol⁻¹

Fourth Ionisation Energy:

Al³⁺(g) → Al⁴⁺(g) + e⁻

Energy: _____ kJ mol⁻¹

B.

aluminium

C.

Energy (kJ mol⁻¹)

| 1ˢᵗ | 2ⁿᵈ | 3ʳᵈ | 4ᵗʰ |
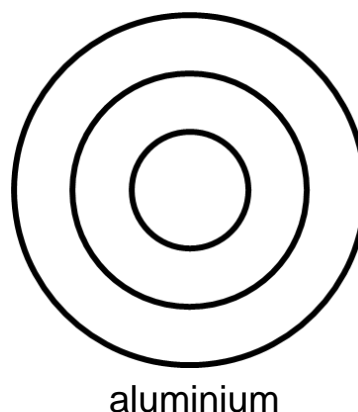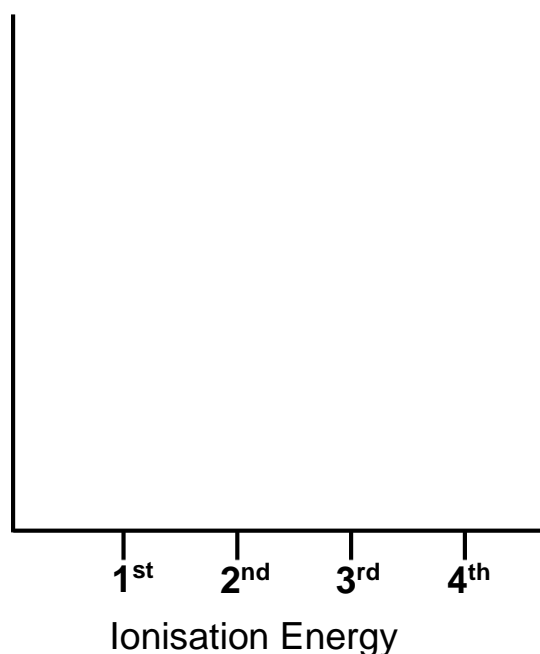
Ionisation Energy

D.
Rule for increases in I.E.

_____

_____

_____

_____

Explaining large Increases in Ionisation Energy

A significant increase in ionisation energy occurs when an electron is removed from an **inner** shell after the **outer** shell electrons have been **removed**. This is due to the **increased effective nuclear charge** experienced by inner electrons and the **stability** of **filled** electron shells.

In Aluminium for example, the large increase between the **third** and **fourth** ionisation energies indicates the removal of an electron from a **stable**, **inner** shell after the **valence** electrons have been removed. This large increase highlights the **stability** and the **stronger** attraction of **inner** electrons to the **nucleus**.
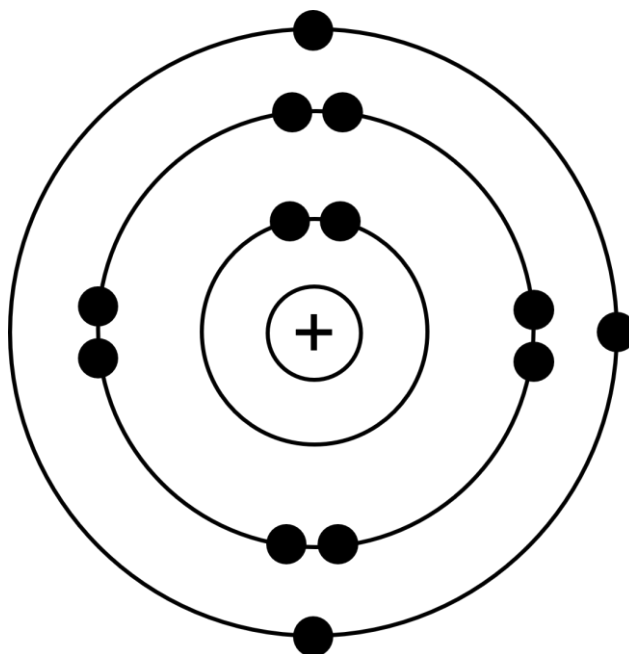
Ionic Radius

Ionic Radius: The radius of an ion, which can be either a **positively** charged ion or a **negatively** charged ion. (these can be found on page 18 of the DB)

- The ionic radius can differ significantly from the covalent radius due to the **gain** or **loss** of electrons.

Using page 7 and 18, complete the following table and write conclusions based on the trends you notice, selecting the appropriate elements for each to compare.

Element	Covalent radius (pm)	Ionic Formula	Ionic Radius (pm)
Li		Li^+	
Na			
K			
Be			
B			
C			
N			
O			
F			

Conclusions:

I.   **State** the trend in ionic radius down a group

_____

II.   **State** the trend in ionic radius across a period.

_____

III.   **State** the general trend shown between the size of an atom (covalent radius) and its positive ion (ionic radius)

_____

IV.   **State** the general trend shown between the size of an atom (covalent radius) and its negative ion (ionic radius)

_____

   Challenge – Explain all of these trends using your knowledge of chemistry.

Factors Affecting Ionic Radius

Electron Shell Completion or Removal:

When an atom gains or loses electrons to form an ion, the number of electron **shells** can change, affecting the ionic radius.

Positively charged ions

- Formed by the **loss** of electrons.

- They typically have a **smaller** radius than their neutral atoms because they lose an **electron shell**, resulting in a **reduced** size.

Negatively charged ions

- Formed by the **gain** of electrons.

- They typically have a **larger** radius than their neutral atoms because the **addition** of electrons **increases** the shielding effect.

Nuclear Charge:

- A **higher** nuclear charge pulls electrons **closer** to the nucleus, **reducing** the ionic radius.

Shielding Effect:

- **Increased** shielding **reduces** the effective nuclear charge felt by **outer** electrons, resulting in a **larger** ionic radius.

- Conversely, less shielding leads to a **stronger** attraction between the nucleus and electrons, resulting in a **smaller** ionic radius.

Questions – Comparing Ionic and atomic radii

1. **Explain** why the ionic radius of Mg^{2+} is smaller than the atomic radius of Mg.

2. **Explain** why the ionic radius of F^- is larger than the atomic radius of F.

3. **Explain** why the ionic radius of Ca^{2+} is smaller than the atomic radius of Ca.

4. **Explain** why the ionic radius of O^{2-} is larger than the atomic radius of O.

5. a) **Compare** the ionic radius of phosphide ion (P^{3-}) to the ionic radius of sulfur ion (S^{2-}).

   b) **Explain** this observation.

6. a) **Compare** the ionic radius of the Aluminium ion (Al^{3+}) to the ionic radius of phosphorous (P^{3-}).

   b) **Explain** this observation.

Flash Cards
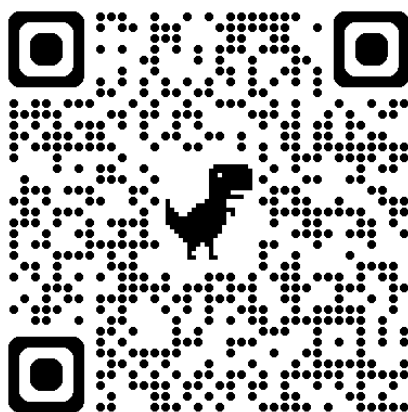
[Periodicity]

Extension questions:
Hodder and Gibson pages 37-44 (questions on pages 43-44)
Blue Chemcord book page 10-14
SCHOLAR

Past Papers

Periodicity	2015	2016	2017	2018	2019	2020	2022	2023
MC	2,3	1	3	4,5		2,6	2	
S2	1.b(ii)	2.	1	1	2	1.a.	2a-b	1

TEAMS: Check Test – Unit 1: Key Area 1aii

Past Paper Questions

1. **Explain** why the covalent radius of sulfur is smaller than that of phosphorus. (1 mark)

2. The covalent radius is a measure of the size of an atom.

    a. **Explain** why covalent radius decreases across the period from sodium to chlorine. (1 mark)

    b. **Explain** fully why the covalent radius of sodium is larger than the

3. ionic radius of sodium (2 marks).

    a. **Explain** why the first ionisation energy decreases going down Group 1. (1 mark)

    b. **Explain** fully why the second ionisation energy is much greater than the first ionisation energy for Group 1 elements. (2 marks)

4. Ionisation energy changes across the period.

    a. **Explain** why the first ionisation energy increases across the period. (1 mark)

    b. **Write an equation**, including state symbols, for the second ionisation energy of magnesium. (1 mark)

5. The table shows the values for the first four ionisation energies of aluminium.

Ionisation energies (kJ mol^{-1})			
First	Second	Third	Fourth
578	1817	2745	11 577

6. **Explain** why there is a large difference between the third and fourth ionisation energies. (1 mark)

7. Ionic radius is a measure of the size of an ion.

    a. **Explain** fully why the ionic radius of phosphorus is greater than the ionic radius of aluminium. (2 marks)