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VCE Chemistry ½
Trends in the Periodic Table [1.2]
Workbook

Outline:

Learning Objectives:

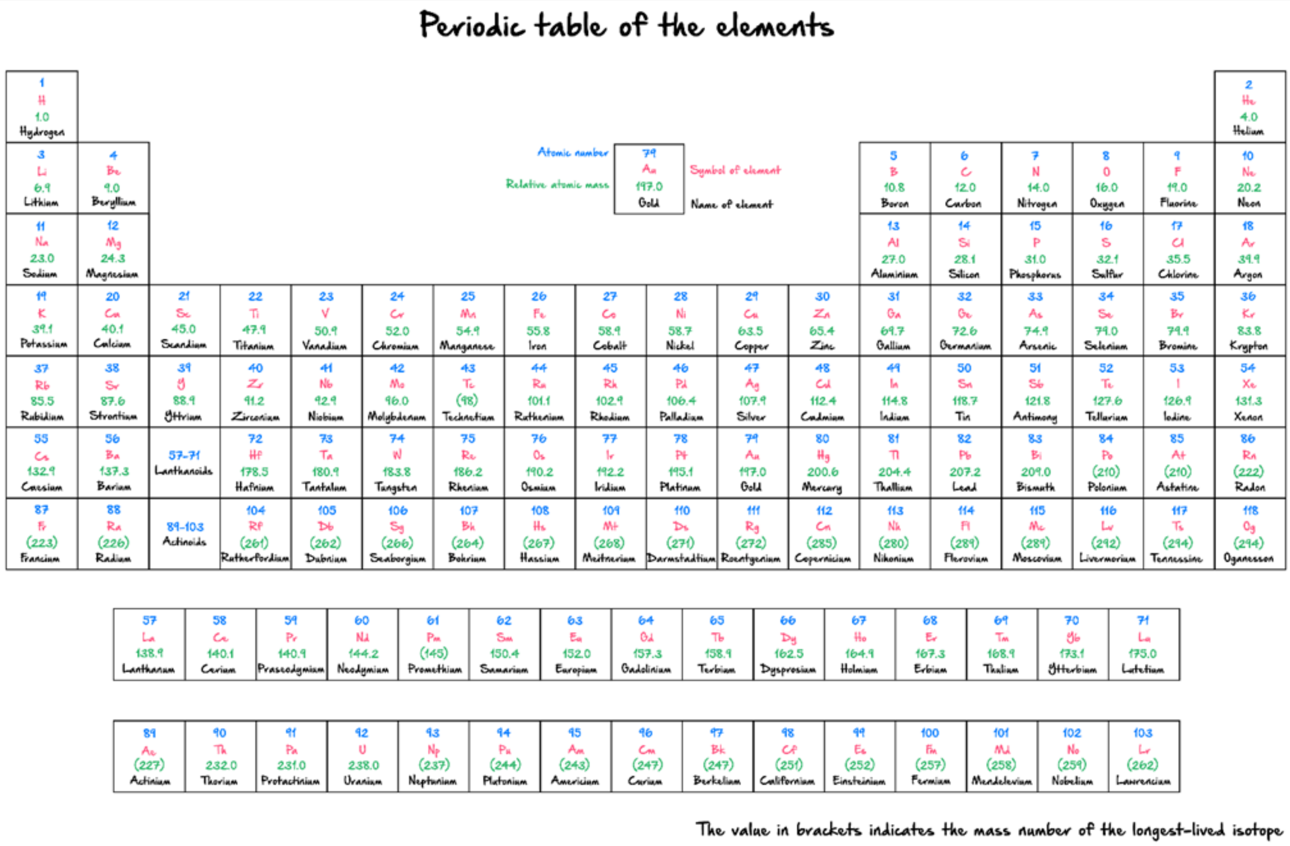
- ☐ **CH12 [1.2.1]** - Explain why the periodic table is arranged the way it is, with respect to blocks, periods and groups.
- ☐ **CH12 [1.2.2]** - Explain what the terms 'electronegativity, 'atomic radius', 'first ionisation energy', 'metallic character' and 'non-metallic character' mean, and explain how they vary across a period and down a group.
- ☐ **CH12 [1.2.3]** - Find the effective nuclear/core charge of an element, explain how it varies across a period and down a group, and apply it to other trends observed in the periodic table.

Section A: The Periodic Table

Context

- Dimitri Mendeleev is known as the founding father of the periodic table, proposing its prototype back in 1869!

- This is what it looks like today!



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Sub-Section: Electron Configurations using the Periodic Table

Let's have a look at why it's arranged the way it is!

The Periodic Table

- Contains all 118 known elements, arranging them in terms of **increasing** atomic number.

Active Recall

- There are 7 horizontal rows in the periodic table called periods.
- The period number represents how many electron shells the element has.

Let's have a look at why it's arranged the way it is!

Question 1

What is the shell model electron configuration of each of the following and state which period it appears in in the periodic table?

a. He → 2

2

b. N → 7

2,5

c. Mg → 12

2,8,2

d. Ca

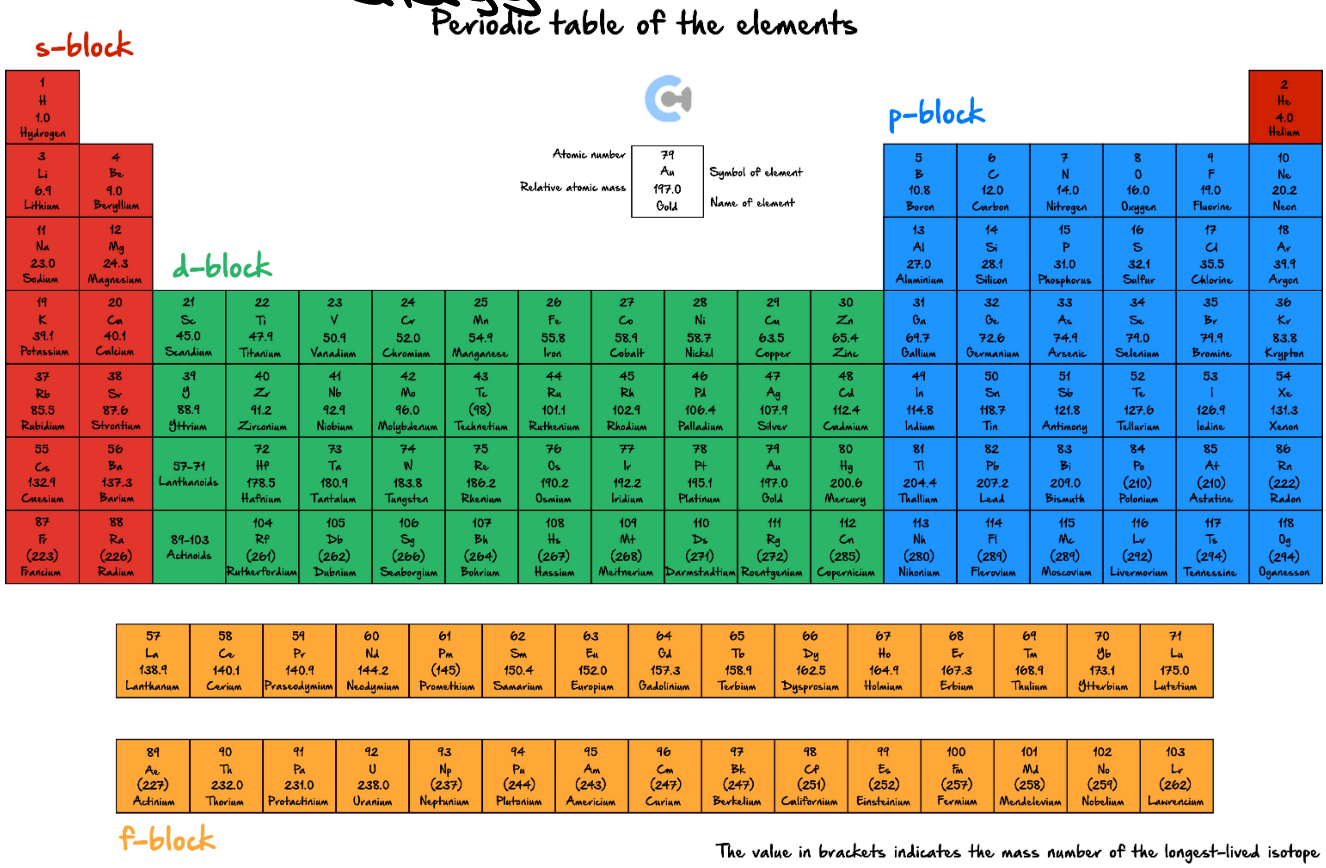
2,8,8,2

Let's have a look at the blocks in the periodic table!

Exploration: Separating Periodic Table into Blocks

- The periodic table can be separated into **blocks**:

- The subshell of the high energy ______________ is that of the block the element is in.



NOTE: The elements found in the d-block are known as ______________________.

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Question 2 Walkthrough.

Write Schrödinger's electronic configuration for each of the following elements, **by using the periodic table**:

a. (16)

i. Sulphur (S)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

2 8 6

ii. Hence, write the shell model for sulphur from Schrödinger's electronic configuration.

2, 8, 6

b. Iron (Fe) 26

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

TIP: Using the periodic table makes writing Schrödinger's electronic configuration much easier!

Question 3

Identify the element by looking at the electron configuration.

a. $1s^2 2s^2 2p^6 3s^2 3p^5$ → 17

Cl

b. $1s^2 2s^2 2p^3$

N

Question 4 (5 marks)

Explain why the periodic table includes the following, with reference to shells and subshells:

a. Two elements in the first period. (1 mark)

- elements in 1st period → filling 1st e⁻ shell
- 1 e⁻ shell only holds 2e⁻

b. Eight elements in the second period. (2 marks)

- in the second period → filling our 2nd e⁻ shell
- the 2 e⁻ shell holds 8e⁻
- ∴ it represents 8 elements.

c. No transition elements in the first three periods. (2 marks)

- transition metals in D block → filling their 3D subshells.
- before this 4s subshell is filled, & only at the start of the 4th period can we start filling 3D.

NOTE: The d-block orbitals start on the 3rd electron shell so their period and the subshells they fill up are **offset by 1** (e.g. the 4th period elements fill up the $3d$ orbitals).

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What happens to transition metals as we go across the period on the periodic table?

Active Recall: What does 'valence electrons' mean?

Outer shell e^-

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Sub-Section: Valence Electrons in the Periodic Table

Exploration: Transition Metal Electron Configurations

Periodic table of the elements

The value in brackets indicates the mass number of the longest-lived isotope

- As we go across the period from elements 21 (Scandium) to 30 (Zinc), what happens to Schrodinger's electronic configuration?

Scandium: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$

Titanium: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

Zinc: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

- What happens to the number of valence electrons?

2

- Are there any exceptions? (Hint: think back to last week!)
  - Chromium (4s¹ 3d⁵)
  - Copper (4s¹ 3d¹⁰)

What happens to non-metals as we go across the period?

Active Recall: Which block are non-metals found in the periodic table?

P-block

Exploration: Non-Metal Electron Configurations

- As we go across the period from group 13 to group 18, the p-orbital goes from ________________.
- The number of valence electrons will [increase] / [decrease] from ________________.

Groups

- **Definition:**
  - The 18 vertical columns of elements are called group________.
  - Groups 1-2 and 13-18 tell the number of valence electrons

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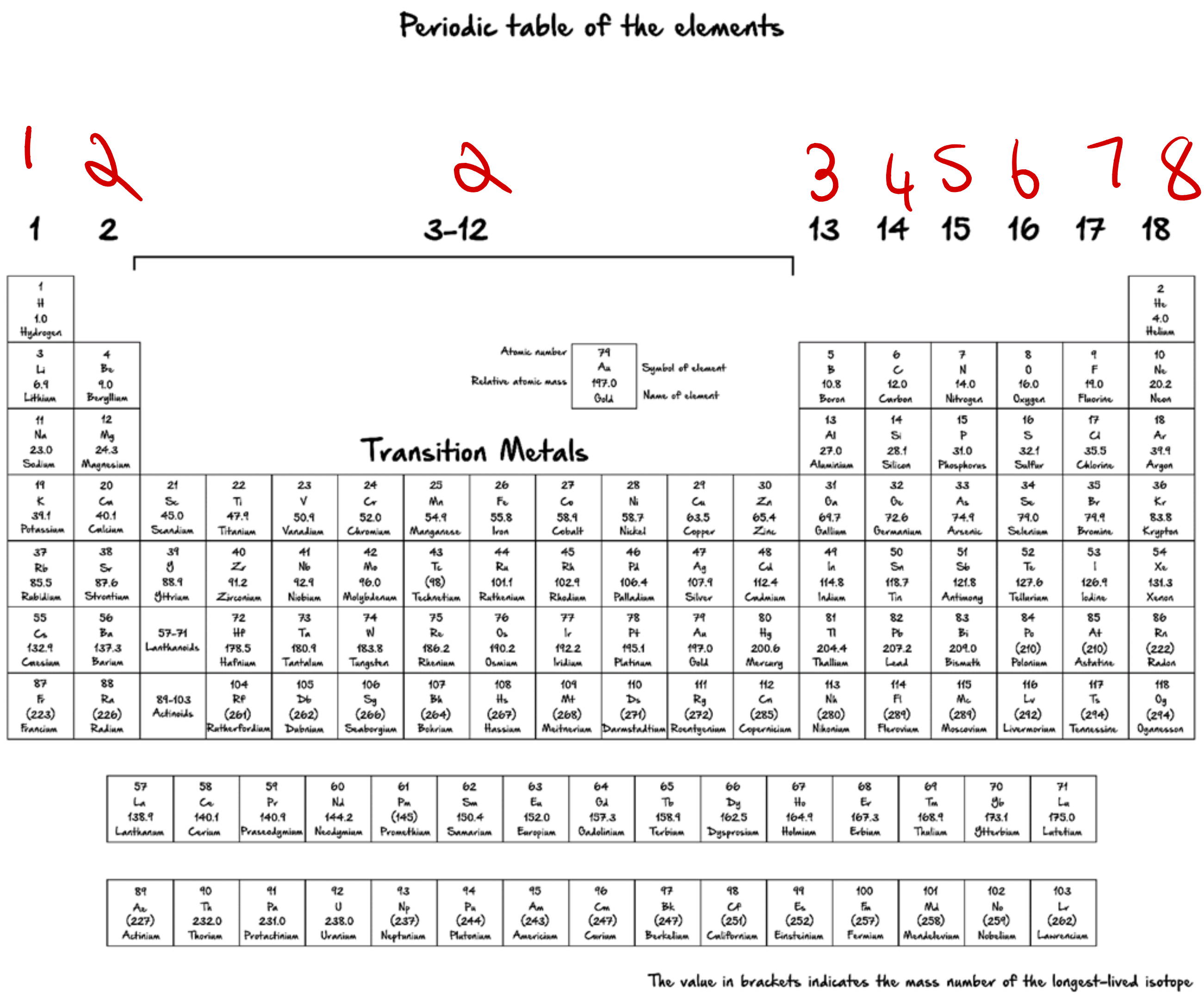
Definition

Groups & Valence Electrons



- Across any period in the d-block, the number of valence electrons remains the same at two.
- The only exceptions are group 6 and 11 (chromium/copper exceptions).
- In the s and p-blocks, the number of valence electrons corresponds to the group no..

TIP: With the exception of groups 3-12 (transition metals), the last number represents the number of valence electrons (e.g. group 15 has 5 valence electrons).

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Try some questions!

Question 5

How many valence electrons do elements in each of the following groups have?

a. Group 1

1

b. Group 16 – 10

6

c. Group 17

7

Question 6

By **only using the location** of the following elements on the periodic table, state the:

- Number of electron shells it has.
- Number of valence electrons it has.
- Subshell of the highest energy it has. → block

a. Cobalt

4, 2, d

b. Chlorine

3, 7, p

Question 7

All members of a group have the same number of valence electrons except for one group. Which group is that and which element is the exception to the trend?

Group 18 → 8 valence e^-

He only has $2e^-$ → 2 valence e^-

Discussion: Why isn't helium classified in group 2 and the *s*-block of the periodic table?

- He has different properties → non-metals
- He has a full outer shell

Active Recall: What does the Octet Rule State? How many electrons do atoms want to have in the valence shell?

Atoms need $8e^-$ to fill outer shell

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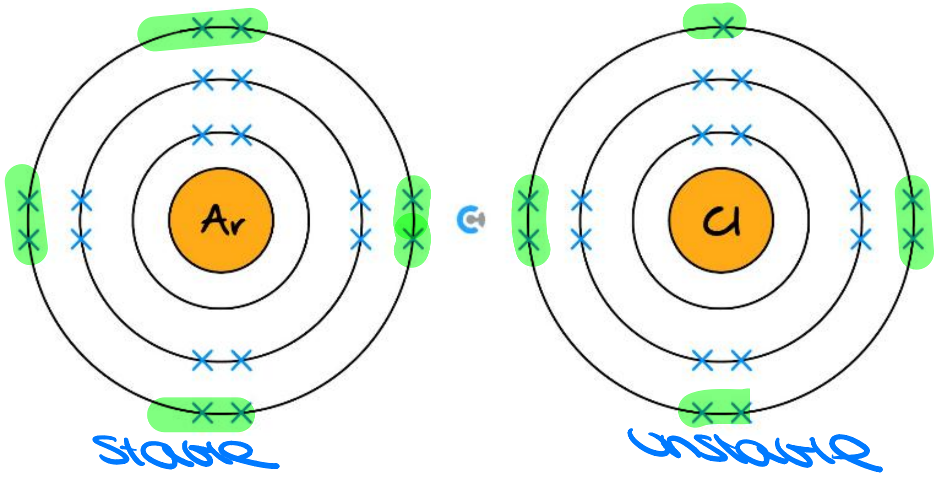
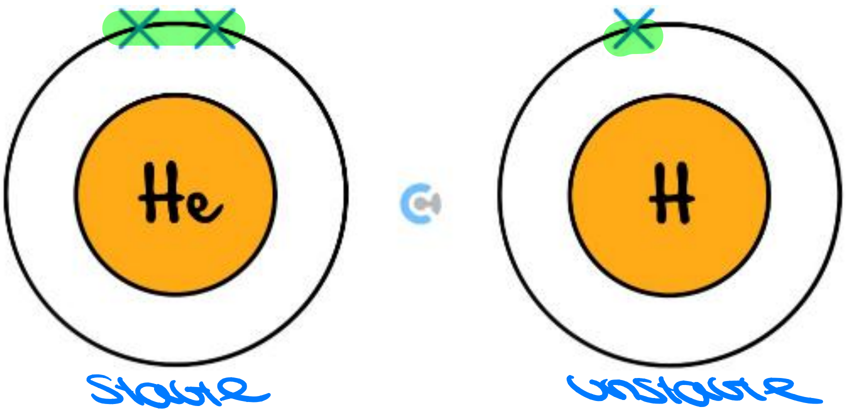
Exploration: Octet Rule in Atoms

- Consider the below atoms and their stabilities: (Label Below)



- Consider Helium (He) and Hydrogen (H):



- The maximum number of electrons that can fit in the valence shell is 2, 8.
- Atoms with an already full outer shell are considered as being [unstable] / [stable].
- **Which group in the periodic table has a full outer shell?**

18

- This group is known as the noble gases and Helium falls within this group.

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Definition

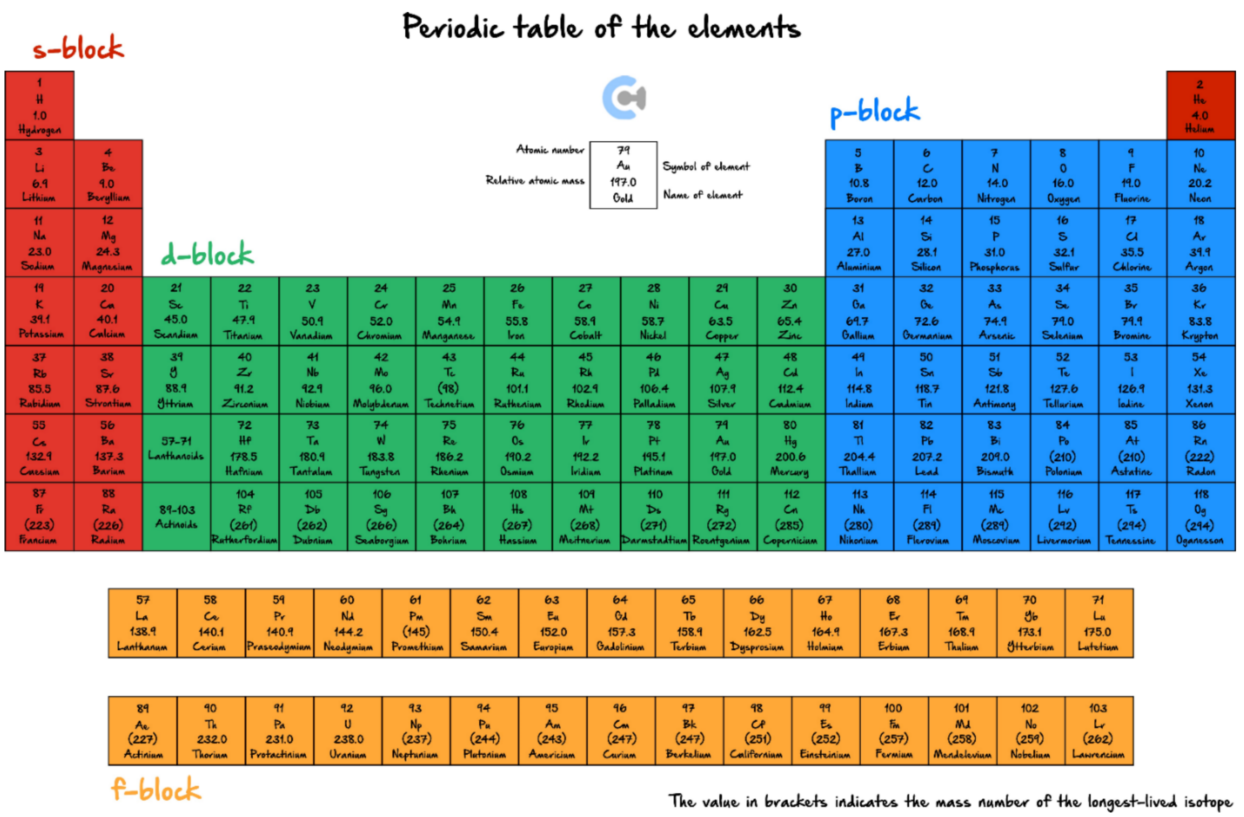
Noble Gases

- Elements in **group 18** are known as the noble gases. They do **not** participate in any type of intramolecular bonding, as they have **full outer shells.**

NOTE: Noble gases are gases.

Key Takeaways

- There are 7 horizontal rows in the periodic table called **periods**.
- The period number represents how many **shells** the element has.
- The periodic table can be separated into **blocks**:



- The 18 vertical columns of elements are called **groups**
- Only groups **1-2 & 13-18** tell the number of **valence electrons.**
- Elements in group 18 are known as **noble gases** and are **unreactive** due to full outer shells.

Section B: Ionisation Energy & Effective Nuclear Charge

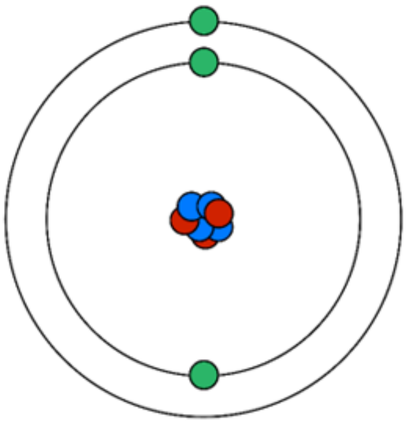
First Ionisation Energy

- **Definition:** The first ionisation energy is the energy required to remove one electron from an element in a gaseous state.

Why is it called the first ionisation energy?

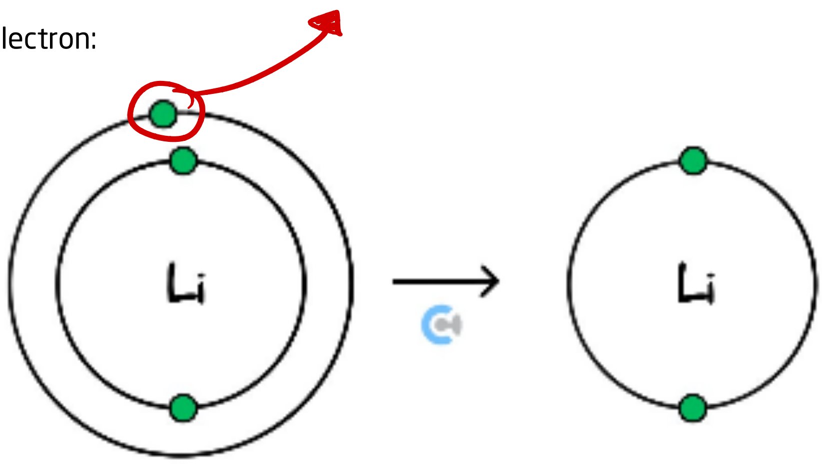
Exploration: First Ionisation Energy

- Lithium (Li):



Protons	Electrons	Overall Charge
3	3	+3 −3 = 0

- After losing an electron:



Protons	Electrons	Overall Charge
3	2	+3 − 2 = +1

- As it's charged, we call it an ___ion___, more specifically a ___cation___.

First Ionisation Energy

- The first ionisation energy is the **energy required** to remove the **first** electron, or the energy required to **ionise** the atom.

NOTE: It is called the **first** ionisation energy as it is the energy required to remove the first electron.

Extension: Second Ionisation Energy

- While there is a second ionisation energy, it is not covered in VCE Chemistry!

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Sub-Section: First Ionisation Energy Across a Period

Discussion: Does every element have the same first ionisation energy?

How does the first ionisation energy change across a period?

Exploration: Period 2 Element Trends

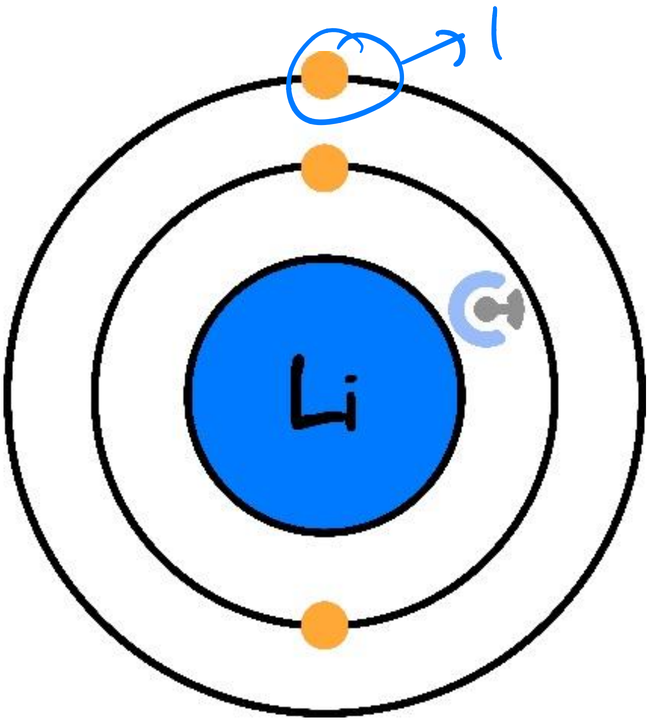
- What are the Shell Model electron configurations for each of the elements? (Label Below)

Group	1	2	13	14	15	16	17	18
Period 2	3 Li Lithium 7.0	4 Be Beryllium 9.0	5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.1
Shell Model Configuration								

- According to the Octet rule, the most stable element here is __neon__.
- As such, __neon__ is the most difficult to remove one electron from.
- Neon (Ne) has the __largest__ first ionisation energy as it already has a full outer shell and thus will __resist__ giving up another electron.

Let's break this idea down further!

Exploration: Lithium (Li)



- What are **two ways** it can achieve a full outer shell?

lose 1e⁻
gain 7e⁻

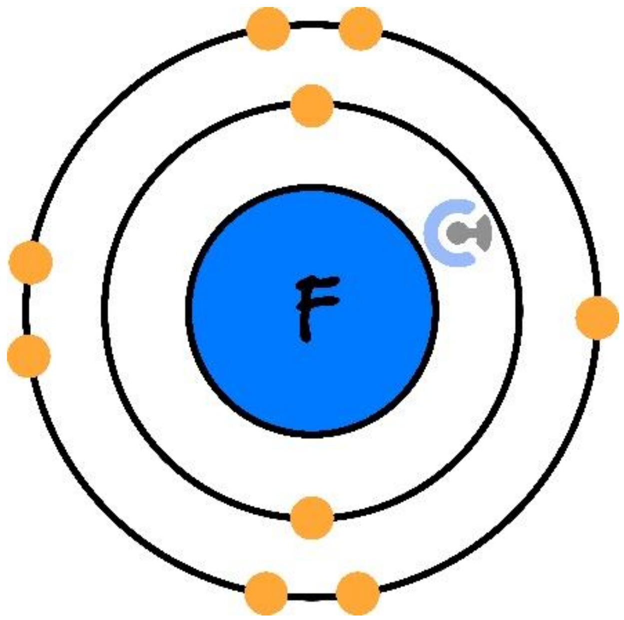
- Which of these two is the easiest?

lose e⁻

- High or low first ionisation energy?

low

Exploration: Fluorine (F)



- What are **two ways** it can achieve a full outer shell?

gain 1e⁻
lose 7e⁻

- Are losing electrons more or less likely?

gain 1e⁻

- High or low first ionisation energy?

high

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Discussion: What happens to the first ionisation energy as we go across the period?

Group	1	2	13	14	15	16	17	18
Period 2	3 Li Lithium 7.0	4 Be Beryllium 9.0	5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.1
Trend	increase →							

First Ionisation Energy Across a Period

- The first ionisation energy **increases** across the period, as it becomes easier to gain electrons instead of losing them to obtain a **full outer shell**.

Let's have a look at a question together!

Question 8 (2 marks) **Walkthrough.**

a. State whether oxygen or carbon is more likely to have a higher first ionisation energy.

Oxygen

b. Explain whether sodium or aluminium is more likely to have a lower first ionisation energy. (2 marks)

- Sodium
- Sodium has 1 valence e^- → wants lose that e^- → low resistance.

Try a question for yourself!

Question 9

A sealed vessel contains an element, which is known to be either arsenic (As) or bromine (Br). Using selenium (Se) for reference, it is found that more energy is required (compared to selenium) to completely remove a single electron from its electron cloud.

Identify which element is inside the sealed vessel.

Bromine → further across period → more valence e^- → more resistance to removal

Question 10 Additional Question.

Predict whether calcium or titanium would have a lower first ionisation energy.

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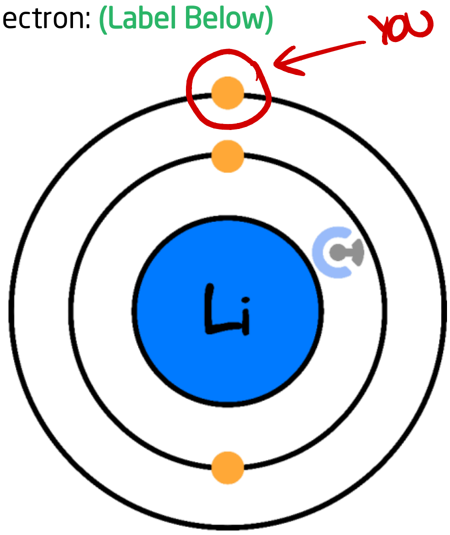
Sub-Section: Effective Nuclear Charge or Core Charge

Exploration: Effective Nuclear Charge/Core Charge

- Lithium atom has 3 proton & **3 electrons**.
- Suppose **you** are the valence electron: (Label Below)



- Considering the atom above:

Attraction/Repulsion to Nucleus	Attraction/Repulsion to Other Electrons	'Net" Attraction/Repulsion	Effective Nuclear Attraction
+3	−2	+3 −2	+1

NOTE: This 'effective nuclear attraction' is named the effective nuclear charge or core charge!

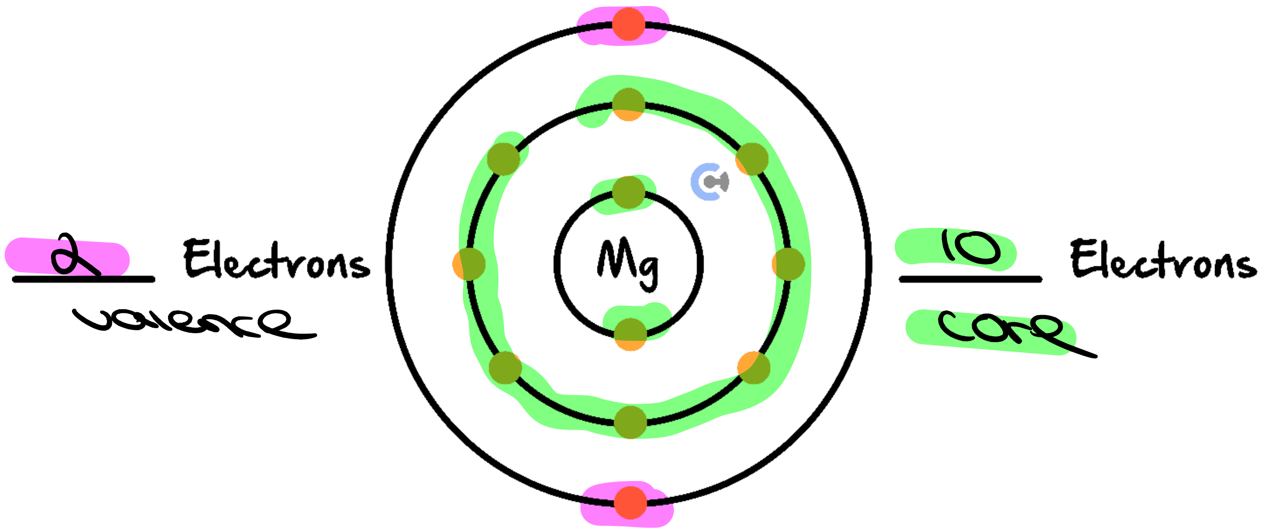
Effective Nuclear Charge/Core Charge

- **Definition:**
  - The effective nuclear charge is the attractive force 'felt' by the valence electrons.
  - First ionisation energy and other trends can be discussed as **effective nuclear charge**.

How do we find effective nuclear charge?

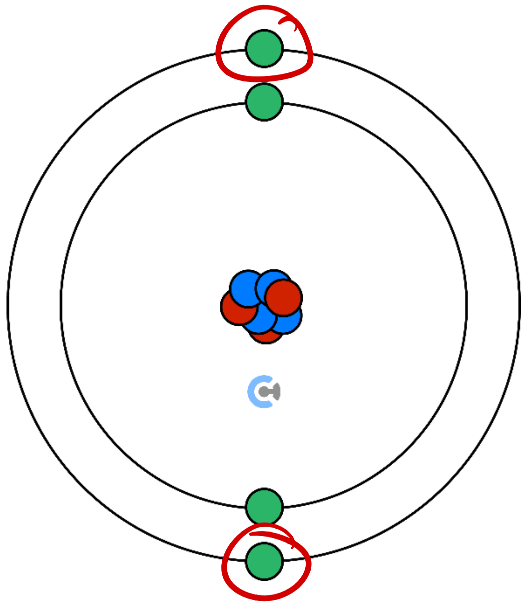
Exploration: Core Electrons & Valence Electrons

- **Valence electrons:** Electrons in the __valence__ shell. (Label Below)
- **Core electrons:** Electrons in __other__ shells. (Label Below)



Exploration: Calculating Effective Nuclear Charge

- Beryllium (Be) has __4__ protons & 4 **electrons**.



- Fill in:

Attraction to Nucleus	Repulsion to Other Electrons	Effective Nuclear Charge
+4	−2	+2

- Operations:

$$+4 - 2 \rightarrow +2$$

- Formulae:

$$Effective\ nuclear\ charge = no. of\ protons - no. of\ core\ electrons$$

NOTE: Core charge is the amount of Shielding provided to valence electrons from the nucleus by core electrons. **The greater the shielding, the lower the core charge.**

Is there a quicker way of finding the core charge?

<u>Exploration</u>: Finding Effective Nuclear Charge Quickly

Group	1	2	13	14	15	16	17	18
Period 2	3 Li Lithium 7.0	4 Be Beryllium 9.0	5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.1
Shell Model Configuration	2,1	2,2	2,3	2,4	2,5	2,6	2,7	2,8
	+1	+2	+3	+4	+5			

Discussion: Is there a quicker way to find the effective nuclear charge?

no # valence e⁻

Core Charge Calculation

- Effective nuclear charge can be found by counting the number of valence electrons, which is the group number!

Let's bring this idea back to the trend!

Exploration: How does effective nuclear charge change across the period?

Group	1	2	13	14	15	16	17	18
Period 2	3 Li Lithium 7.0	4 Be Beryllium 9.0	5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.1
Effective nuclear charge	+1	+2	+3					

- Across the period, the effective nuclear charge of the atoms [increases] / [decreases].
- As effective nuclear charge increases, electrons are [more] / [less] attracted to the nucleus.
- Hence, it is [easier] / [harder] to remove an electron from the atom.
- First ionisation energy [increases] / [decreases] as effective nuclear charge increases.

NOTE: Use effective nuclear charge to justify whether the first ionisation energy is high or low!

Let's have a look at a question together!

Question 11 Walkthrough.

Explain whether sodium (Na) or silicon (Si) would have a greater first ionisation energy.

- Si
- Silicon has a higher core charge (+4) whereas Na has core charge (+1)
- Thus e^- are more attracted to the nucleus of silicon

Sample Response: First Ionisation Energy

- As the effective nuclear charge of x is higher, it feels a **greater attraction** to the nucleus.
- Therefore, electrons are **harder to remove** from the atom.
- **More energy is required** to remove them from the atom.
- They have a **greater** first ionisation energy.

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Your turn!

Question 12

State the effective nuclear charge of each of the following atoms:

a. Rubidium (Rb)

+1

b. Iodine (I)

+7

c. Tin (Sn)

+4

d. Rank these elements in terms of decreasing first ionisation energy.

I , Sn, Rb

Question 13 (3 marks)

By referring to effective nuclear charge, explain whether tin (Sn) or antimony (Sb) has a higher first ionisation energy

Sn has a core charge of +4, where as antimony has a core charge +5. This means e⁻ more attracted to Sb ∴ ↑ first ionisation energy

Question 14 (2 marks) **Additional Question.**

Explain why neon has a higher first ionisation energy than sodium.

TIP: Regardless of what the question says, it's easier to refer to effective nuclear charge to justify your answer!

What happens to the trend down a group?

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Sub-Section: First Ionisation Energy Along a Group

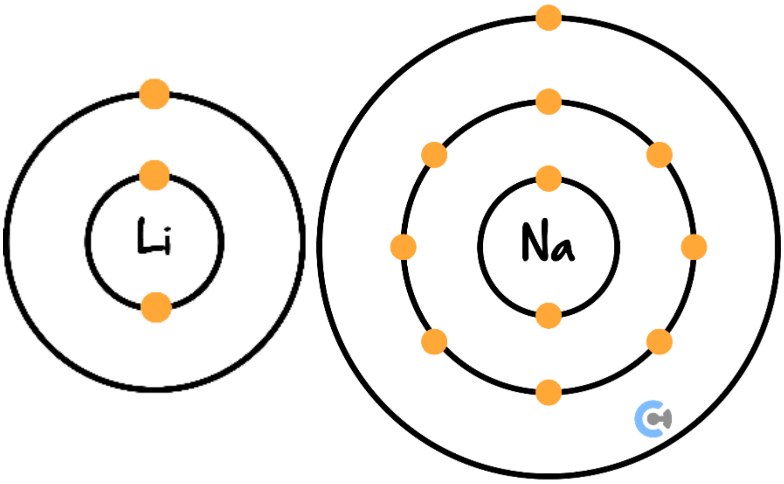
Exploration: First Ionisation Energy Along a Group

- Group 1 elements:

Alkali Metals	Atomic Number	Complete Electronic Configuration
Li	3	2, 1
Na	11	2, 8, 1
K	19	2, 8, 8, 1
Rb	37	2, 8, 18, 8, 1
Cs	55	2, 8, 18, 18, 8, 1
Fr	87	2, 8, 18, 32, 18, 8, 1

- The difference between each of the group 1 elements?

- Lithium and sodium:



- As the number of electron shells increases (as we go down a group), the distance between the nucleus and the valence electrons **[increases]** / **[decreases]**.
- The strength of attraction between the nucleus and valence electrons **[strengthens]** / **[weakens]**.
- Thus, it would be **[easier]** / **[harder]** to remove an electron, and so the first ionisation energy as we go down the group **[increases]** / **[decreases]**.

<u>First Ionisation Energy Down a Group</u>

Definition

- First ionisation energy **[increases]/[decreases]** down the **group** as the valence electrons are located **further** from the nucleus, and thus feel a **weaker** pull.

Try some questions!

Question 15

For each of the following sets, rank them in terms of increasing the first ionisation energy.

a. Al, In, B In, Al, B	**c.** F, O, Br Br, O, F
b. Mg, Cl, Si	**d.** Na, K, Ca

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Question 16 (3 marks)

Determine whether Calcium (Ca) or Sulphur (S) has a greater first ionisation energy and explain why.

- Sulphur has 3e⁻ shells whereas calcium has 4 e⁻ shell
- Also sulphur has a core charge of +6 whereas calcium is +2
- Thus sulphur has ↑ F.I.E.

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Sub-Section: Metallic Character

Metallic Character

- The metallic character is the tendency of an element to **lose electrons** and form cation

Active Recall: What concept is linked to losing an electron and becoming an ion?

first ionisation energy

Discussion: Which element has the greatest metallic character?

Francium

NOTE: Metallic character is directly linked to the first ionisation energy!

Try a question!

Question 17

Which of the following is correct?

A. Potassium has a higher first ionisation energy than calcium.

B. Sodium has a greater metallic character than caesium.

C. Hydrogen has a greater metallic character than beryllium.

D. Rubidium has a greater metallic character than magnesium.

NOTE: Metallic character increases as first ionisation energy ________________.

Key Takeaways

- ☑ The first ionisation energy is the energy required to **remove** one electron from an element.
- ☑ Across the period, the first ionisation energy **increases**.
- ☑ The **effective nuclear charge** is the attractive force 'felt' by the valence electrons.
- ☑ Effective nuclear charge/core charge can be found by counting the number of **valence** electrons.
- ☑ First ionisation energy **decreases** down the **group** as valence electrons are further from the nucleus and feel a **weaker** pull.
- ☑ Metallic character **increases** as the first ionisation energy **decreases.**

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Section C: Electronegativity and Atomic Radius

Let's have a look at some other trends in the periodic table!

Electronegativity

- **Definition:**
  - Electronegativity is the ability of an atom to attract electron toward itself.

Atomic Radius

- **Definition:**
  - The atomic radius is the size of the atom or the distance between the centre of the nucleus to valence electrons.

NOTE: Electronegativity and atomic radius are very closely related to each other.

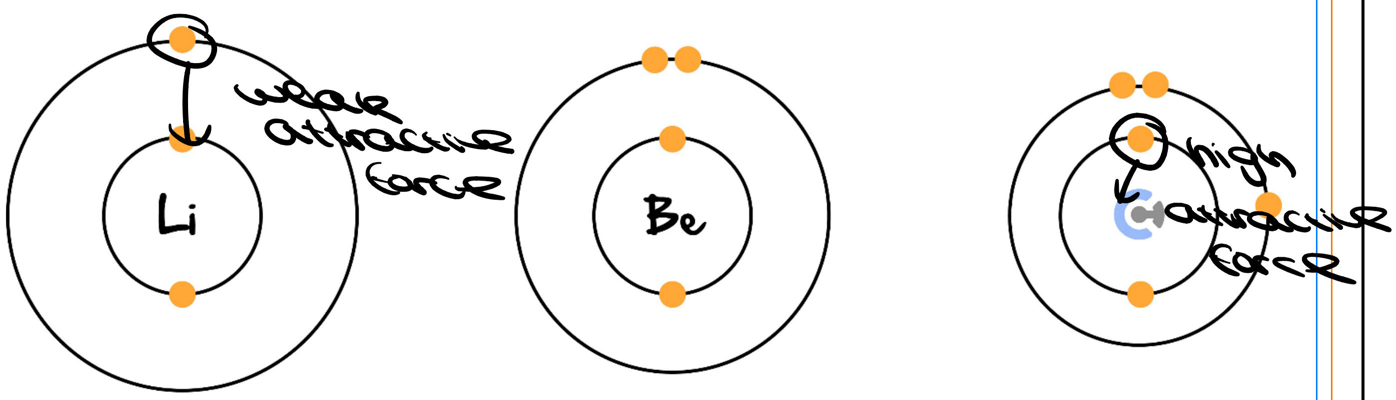
Discussion: What else does atomic radius link to?

- no. shells
- core charge

Exploration: Going Across the Period

- Across the period, the effective nuclear charge **[increases]/ [decreases].**
- If effective nuclear charge increases, the nucleus pulls in electrons more **[strongly] / [weakly].**
- This is the exact definition of electronegativity.

- As effective nuclear charge increases, electronegativity [increases] / [decreases].
- As electrons are pulled closer in, the atomic radius [increases] / [decreases].
- Lithium (Li), Berylium (Be) and Boron (B) are all in the same period:

Lithium (Li)	Berylium (Be)	Boron (B)
Effective Nuclear Charge: +1	Effective Nuclear Charge: +2	Effective Nuclear Charge: +3
Electronegativity: low	Electronegativity:	Electronegativity: high
		
Atomic Radius: large	Atomic Radius: medium	Atomic Radius: small

Electronegativity and Atomic Radius Across a Period

Definition

- Electronegativity **increases** across the period as **effective nuclear charge increases.**
- As electronegativity increases, electrons are pulled closer to the nucleus, **decreasing the atomic radius** of the atom.

Let's take a look at how these trends vary down a group!

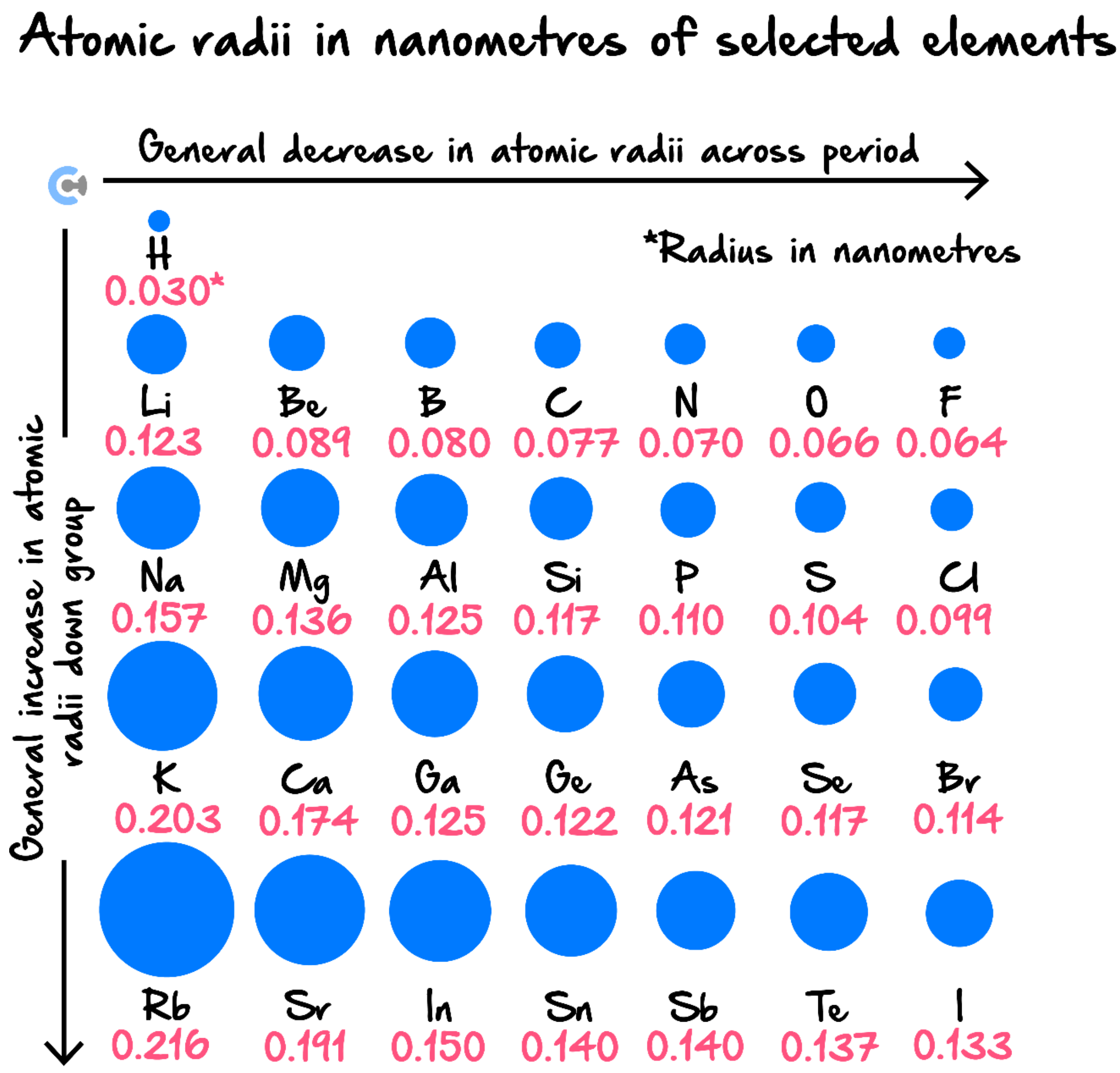
Exploration: Going Down the Group

- Down the group, the effective nuclear charge **[increases] / [stays the same] / [decreases]**.
- The distance of the valence shell from the nucleus **[increases] / [stays the same] / [decreases]**.
- As such, the atomic radius **[increases] / [stays the same] / [decreases]** as the valence shell further increases.
- Consequently, the pull on the valence electrons by the nucleus becomes **[stronger] / [weaker]**.
- As such, electronegativity **[increases] / [decreases]** down the group.

Electronegativity increases across a period.

Electronegativity decreases down a group.

1	2	13	14	15	16	17
Li	Be	B	C	N	O	F
1.0	1.6	2.0	2.6	3.0	3.4	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.3	1.6	1.9	2.2	2.6	3.2
K	Ca	Ga	Ge	As	Se	Br
0.8	1.0	1.8	2.0	2.2	2.6	3.0
Rb	Sr	In	Sn	Sb	Te	I
0.8	1.0	1.8	2.0	2.1	2.1	2.7
Cs	Ba	Tl	Pb	Bi	Po	At
0.8	0.9	2.0	2.3	2.0	2.0	2.2
Fr	Ra					
0.7	0.9					



Definition

Properties Going Down a Group

Effective Nuclear Charge	Atomic Radius	Electronegativity
Stays the same	Increases	Decreases

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Let's have a look at a question together!

Question 18 (4 marks) **Walkthrough.**

a. Predict which element is expected to be the most electronegative by looking at the periodic table. (1 mark)

flourine

b. Justify your reasoning, with respect to relevant phenomena. (3 marks)

· He → is not very e- neg → has a full outer shell

· Doesn't want more e⁻ → low attractive force.

· flourine has 7 valence e⁻ → wants 1e⁻

therefore it has a high e-neg

Sample Response: Electronegativity

- As we go across a period, core charge **increases**, so electrons have a **greater** attraction to the nucleus.
- Furthermore, as we go up a group, the **atomic radius** decreases as we have fewer shells, so the attraction between the nucleus and electrons **increases.**
- Therefore, electronegativity **increases.**

Space for Personal Notes

Try some questions!

Question 19 (4 marks)

a. Going across the periodic table, the numbers of protons and electrons increase. Why then does the size of the atoms decrease? (2 marks)

· Going left to right core charge increase
· This means e⁻ are more attracted to nucleus
· e⁻ shell are closer → ∴ atomic radius decrease.

b. Explain the trend in atomic radius going down a group. (2 marks)

· Going down a group no. e⁻ shells increase
· Atomic radius ↑

Question 20

For each of the following pairs of elements, state which element is more electronegative.

a. K, Ca

Ca

b. Be, Ca

Be

c. Cl, Br Cl	

Question 21 (3 marks) **Additional Question.**

a. Rank the following three atoms in terms of increasing atomic size: (1 mark)

Pb, Bi, Cs

____

b. Explain why this is the case. (2 marks)

____

____

____

____

Space for Personal Notes

Sub-Section: Non-Metallic Character

Non-Metallic Character

- The non-metallic character is the tendency of an element to **gain electrons** and form anion

Active Recall: What do we call the tendency of an element to attract/gain electrons?

electronegativity

Discussion: Which element has the greatest non-metallic character?

fluorine

NOTE: Non-metallic character is directly linked to electronegativity!

Try a question!

Question 22

Which of the following is correct?

A. Silicon is more electronegative than chlorine.

B. Helium is the most electronegative element.

C. Iodine has a greater non-metallic character than aluminium.

D. Oxygen has a weaker non-metallic character than nitrogen.

NOTE: Non-metallic character increases as electronegativity ________________.

Contour Check

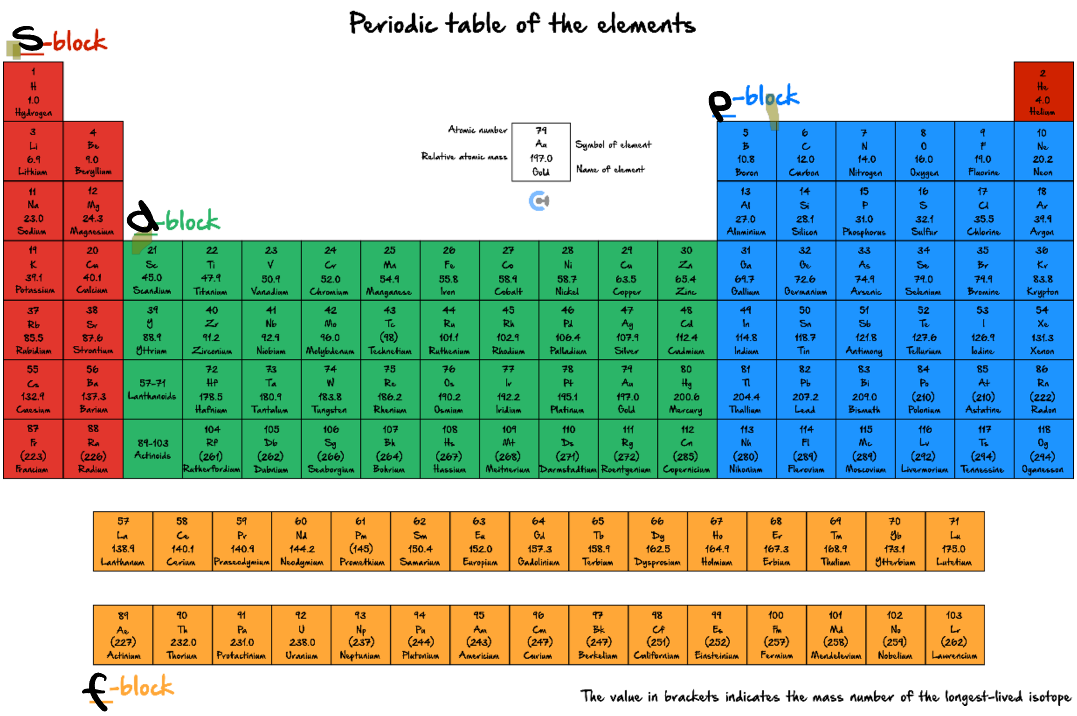
Learning Objective: [1.2.1] - Explain Why the Periodic Table is Arranged the Way it is, With Respect to Blocks, Periods, and Groups

Study Design

The periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic, and non-metallic character and reactivity) of elements.

Key Takeaways

- There are 7 horizontal rows in the periodic table called periods.
- The period number represents how many shells the element has.
- The periodic table can be thought of as being separated into the following blocks (label below):



- The 18 vertical columns of elements are called groups.
- The groups (only for groups 1-2 & 13-18) tell the number of valence e⁻.
- Elements in group 18 are known as the noble gases.
- They have full outer shell, and are therefore, stable / inert

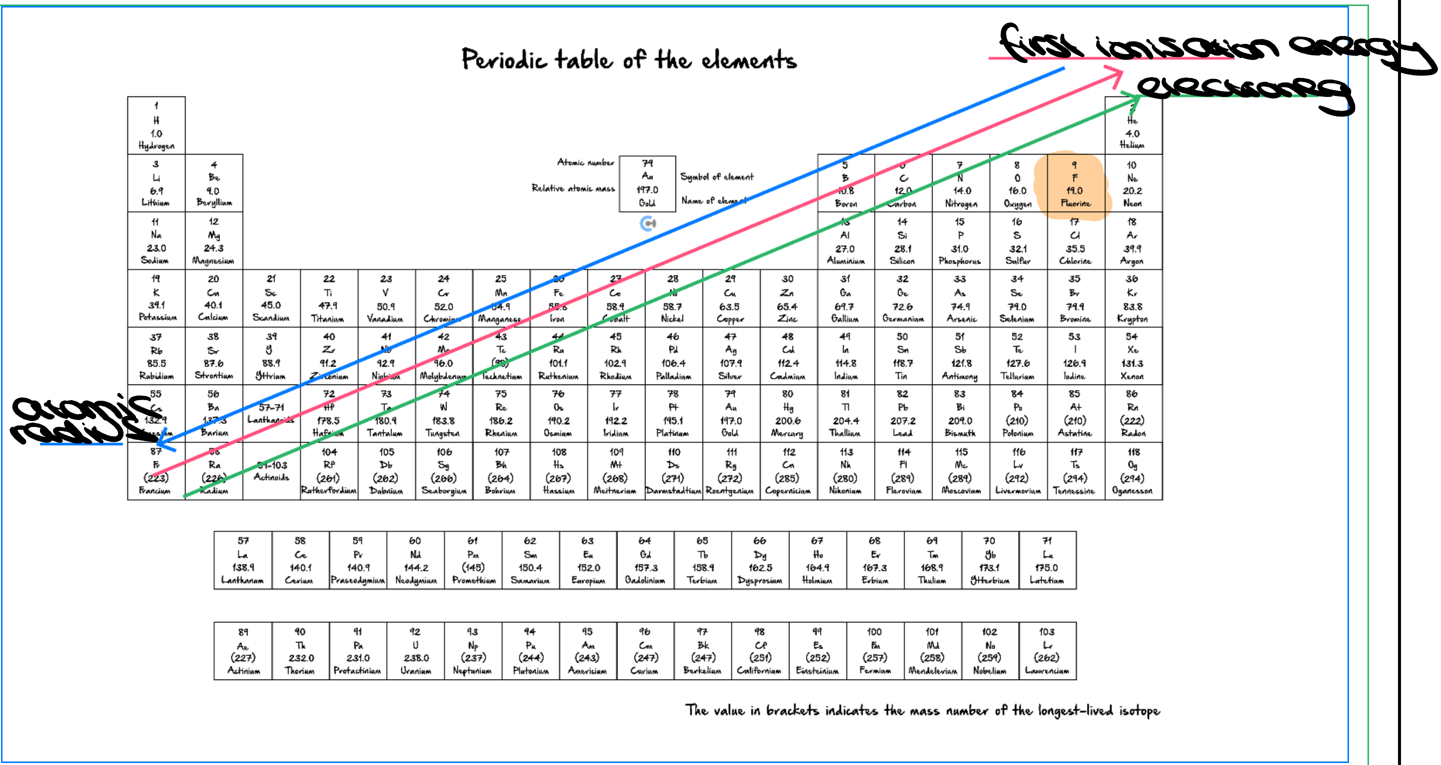
Learning Objective: [1.2.2] - Explain What the Terms 'Electronegativity, 'Atomic Radius', 'First Ionisation Energy', 'Metallic Character' and 'Non-Metallic Character' Mean, and Explain How They Vary Across a Period and Down a Group

Study Design

The periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic and non-metallic character and reactivity) of elements.

Key Takeaways

- ☐ Electroneg______ is the ability of an atom to attract electrons toward itself.
- ☐ Electronegativity decrease______ as we go down a group.
- ☐ Electronegativity increase______ as we go across the period.
- ☐ The non-metallic is the tendency of an element to gain electrons and form anions.
- ☐ Non-metallic character increase as electronegativity **increases.**
- ☐ As electronegativity increases across a period, electrons are pulled closer to the nucleus, effectively decreasing the atomic radius of the atom.
- ☐ The atomic radius is a measurement of the size of the atom.
- ☐ Atomic radius increase as we go down a group.
- ☐ The first i.e______ is the energy required to remove one electron from an element.
- ☐ As we go across the period, the first ionisation energy increase.
- ☐ First ionisation energy decrease down the group as the valence electrons are located further from the nucleus, and thus feel a **weaker** pull.
- ☐ The metallic nature______ is the tendency of an element to lose electrons and form cations.
- ☐ Metallic character increase as first ionisation energy **decreases.**



Learning Objective: [1.2.3] Find the Effective Nuclear/Core Charge of an Element, Explain How It Varies Across a Period and Down a Group, and Apply It to Other Trends Observed in the Periodic Table

Study Design

The periodic table as an organisational tool to identify patterns and trends in, and relationships between, the structures (including shell and subshell electronic configurations and atomic radii) and properties (including electronegativity, first ionisation energy, metallic, and non-metallic character and reactivity) of elements.

Key Takeaways

- The effective nuclear charge is a measure of the attractive force 'felt' by the valence electrons.
- Effective nuclear charge/core charge can be found by simply counting the number of valence electrons.
- Effective nuclear charge stay the same going down a group.
- Effective nuclear charge increase across a period.
- The core charge can be thought of as the amount of shielding provided to the valence electrons from the nucleus **by the core electrons**.