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VCE Chemistry ½

Trends in the Periodic Table [1.2]

Workbook

Outline:

The Periodic Table Pg 3-15	Ionisation Energy & Effective Nuclear Charge Pg 16-34
➤ Electron Configurations using the Periodic Table	➤ First Ionisation Energy Across a Period
➤ Valence Electrons in the Periodic Table	➤ Effective Nuclear Charge or Core Charge
	➤ First Ionisation Energy Along a Group
	➤ Metallic Character
	Electronegativity and Atomic Radius Pg 35-42
	➤ Non-Metallic Character

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!

Periodic table of the elements

1 H 1.0 Hydrogen																	2 He 4.0 Helium		
3 Li 6.9 Lithium	4 Be 9.0 Beryllium	Atomic number Relative atomic mass										5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon		
11 Na 23.0 Sodium	12 Mg 24.3 Magnesium	79 Au 197.0 Gold										Symbol of element Name of element		13 Al 27.0 Aluminium	14 Si 28.1 Silicon	15 P 31.0 Phosphorus	16 S 32.1 Sulfur	17 Cl 35.5 Chlorine	18 Ar 39.9 Argon
19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton		
37 Rb 85.5 Rubidium	38 Sr 87.6 Strontium	39 Y 88.9 Yttrium	40 Zr 91.2 Zirconium	41 Nb 92.9 Niobium	42 Mo 95.9 Molybdenum	43 Tc (98) Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon		
55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57-71 Lanthanoids		72 Hf 178.3 Hafnium	73 Ta 180.9 Tantalum	74 W 183.8 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.0 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (210) Polonium	85 At (210) Astatine	86 Rn (222) Radon	
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The value in brackets indicates the mass number of the longest-lived isotope

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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 \rightarrow 2

2

b. N \rightarrow 7

2, 5

c. Mg \rightarrow 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 is that of the block the element is in.

s-block

Periodic table of the elements

1 H 1.0 Hydrogen												2 He 4.0 Helium					
3 Li 6.9 Lithium	4 Be 9.0 Beryllium											10 Ne 20.2 Neon					
11 Na 23.0 Sodium	12 Mg 24.3 Magnesium											18 Ar 39.9 Argon					
d-block													36 Kr 83.8 Krypton				
19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton
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57	58	59	60	61	62	63	64	65	66	67	68	69	70	71
La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
138.9	140.1	140.9	144.2	(145)	150.4	152.0	157.3	158.9	162.5	164.9	167.3	168.9	173.1	175.0
Lanthanum	Cerium	Praseodymium	Neodymium	Promethium	Samarium	Europium	Gadolinium	Terbium	Dysprosium	Holmium	Erbium	Thulium	Ytterbium	Lutetium

89	90	91	92	93	94	95	96	97	98	99	100	101	102	103
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Mn	No	Lr
(227)	232.0	231.0	238.0	(237)	(244)	(243)	(247)	(247)	(251)	(252)	(257)	(258)	(259)	(262)
Actinium	Thorium	Protactinium	Uranium	Neptunium	Plutonium	Americium	Curium	Berkelium	Californium	Einsteinium	Fermium	Mendelevium	Nobelium	Lanthanum

f-block

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

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.
i. Sulphur (S)



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

2, 8, 6

- b. Iron (Fe)



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. $1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow 17$

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 out 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

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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 ($4s^1 3d^5$)
Copper ($4s^1 3d^{10}$)

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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Groups & Valence Electrons

Periodic table of the elements

1

2

3-12

3

4

5

6

7

8

1

2

13

14

15

16

17

18

1 H 1.0 Hydrogen																	2 He 4.0 Helium				
3 Li 6.9 Lithium	4 Be 9.0 Beryllium															5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon
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Atomic number

Relative atomic mass

Symbol of element

Name of element

71

Au

197.0

Gold

Transition Metals

57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium
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Transition Metals

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

- 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

c. Group 17

7

b. Group 16

6

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.

5d or 6s

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 \rightarrow 8 valence e^-

He only has $2e^-$ \rightarrow 2 valence e^-

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



- He has different properties \rightarrow 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?



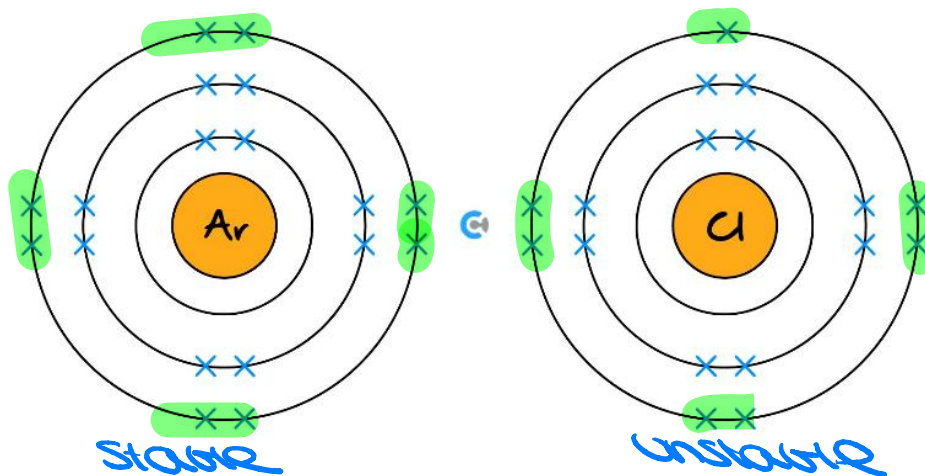
Atoms need $8e^-$ to fill outer shell

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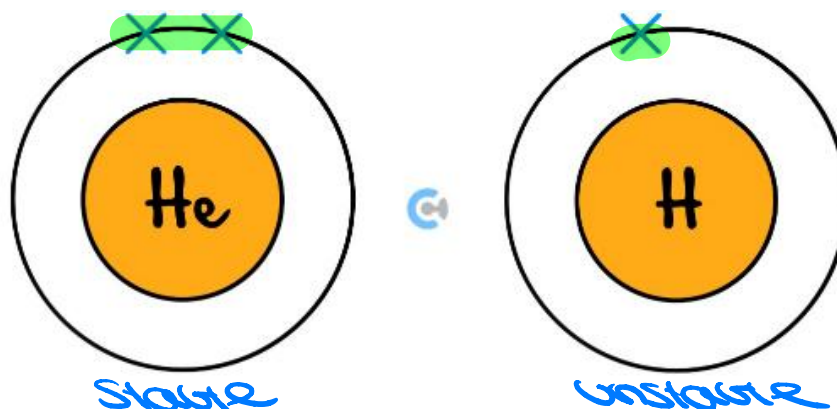


Exploration: Octet Rule in Atoms

- 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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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**:

s-block

Periodic table of the elements

p-block

Atomic number
Relative atomic mass
Symbol of element
Name of element

d-block

1 H 1.0 Hydrogen																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																		</
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f-block

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

- ✓ 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

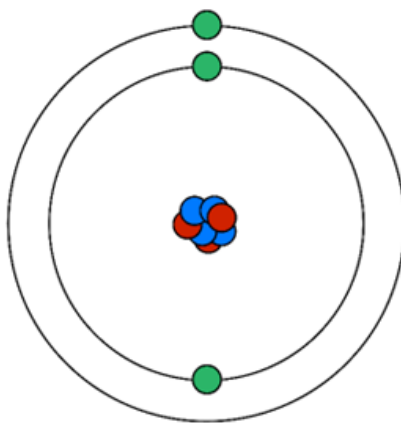


- **Definition:** The first ionisation energy is the energy required to remove one electron from an element in a gaseous state.

Why is it called the first ionisation energy?

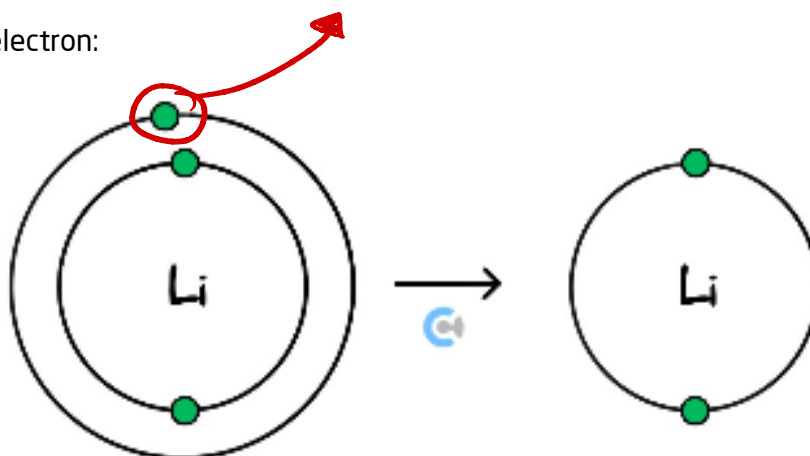
Exploration: First Ionisation Energy

- 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!

Space for Personal Notes

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 7.0 Li Lithium	4 9.0 Be Beryllium	5 10.8 B Boron	6 12.0 C Carbon	7 14.0 N Nitrogen	8 16.0 O Oxygen	9 19.0 F Fluorine	10 20.1 Ne Neon

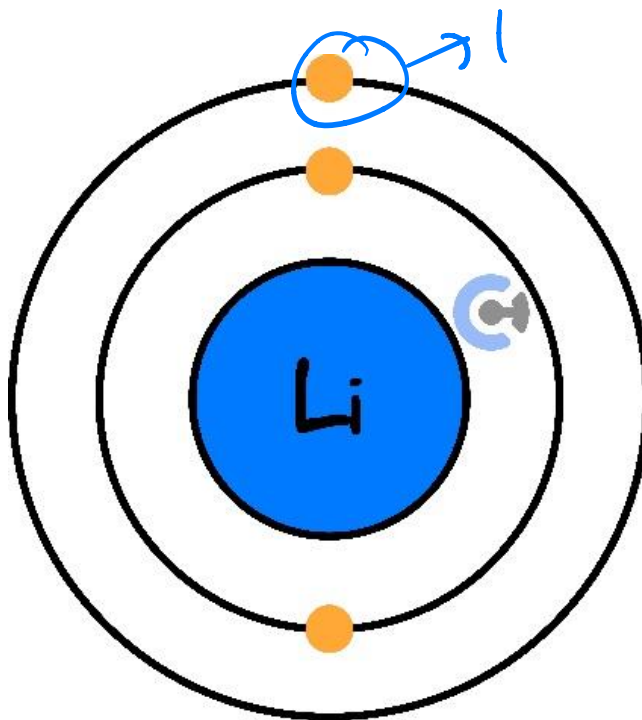
Shell Model
Configuration

- 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?

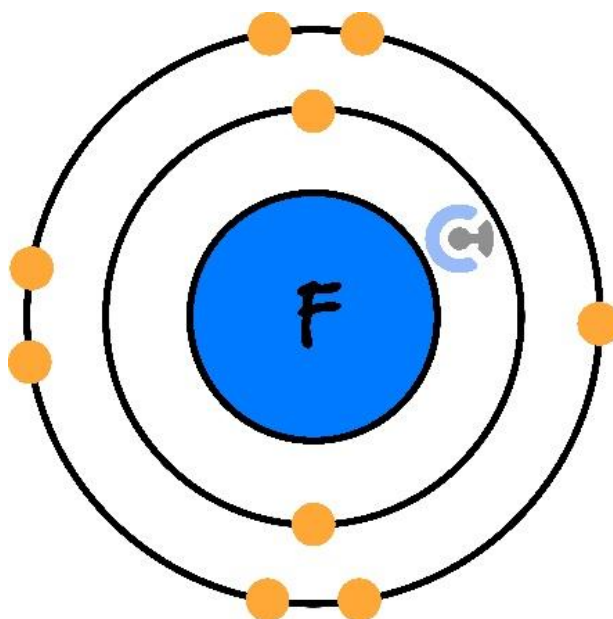
lose e^-

- 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

Space for Personal Notes



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

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 to 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.

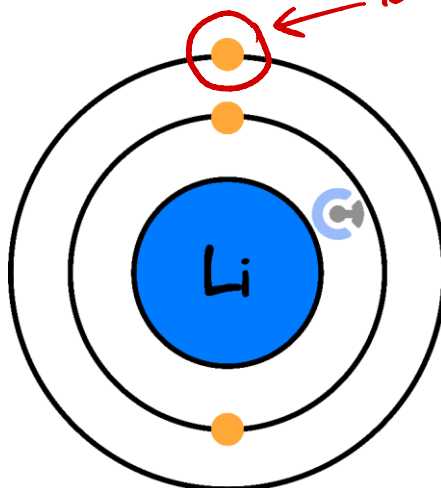
Space for Personal Notes

Sub-Section: Effective Nuclear Charge or Core Charge

Exploration: Effective Nuclear Charge/Core Charge

➤ Lithium atom has 3 protons & 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

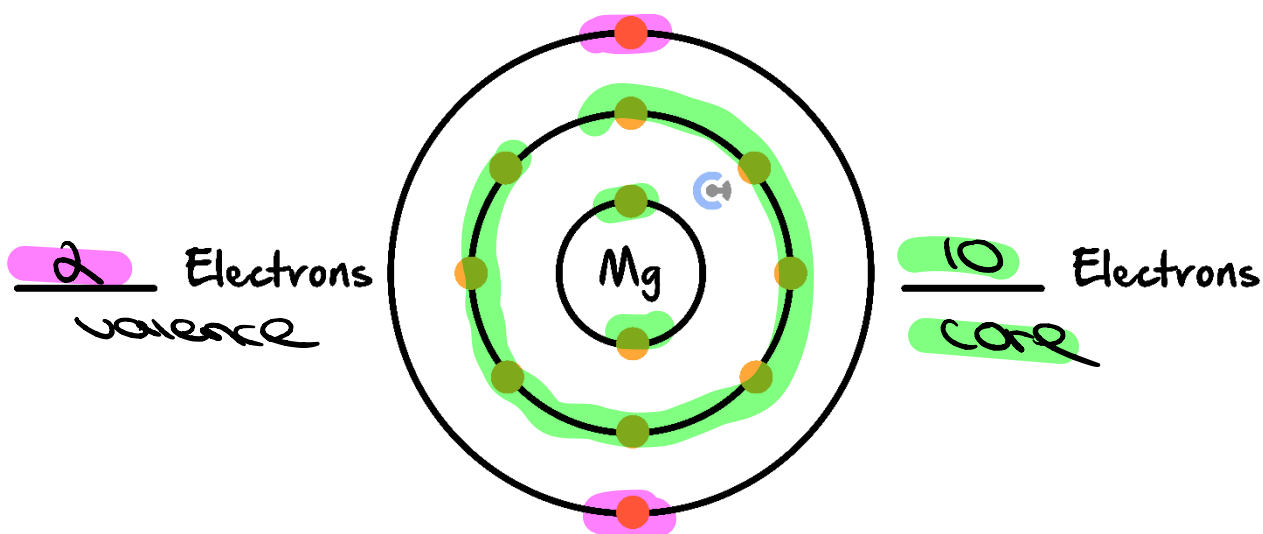
➤ Definition:

- 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?

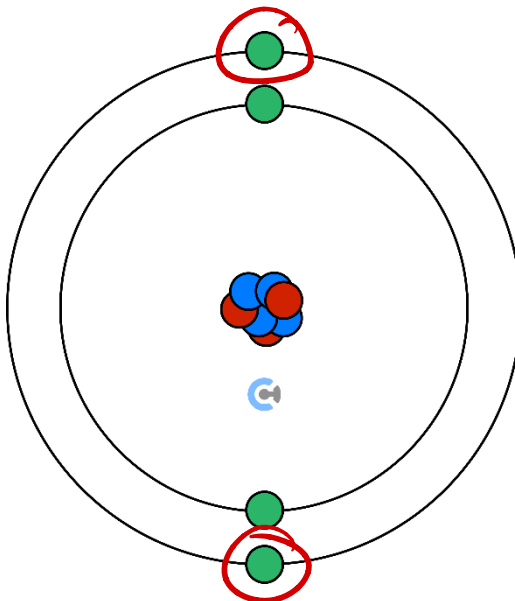
Exploration: Core Electrons & Valence Electrons

- 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:

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

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?



Exploration: Finding Effective Nuclear Charge Quickly



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



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
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) where as 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.



Space for Personal Notes

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, whereas antimony has a core charge +5. This means e^- more attracted to Sb \therefore ↑ 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!

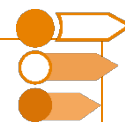


What happens to the trend down a group?



Space for Personal Notes

Sub-Section: First Ionisation Energy Along a Group



Exploration: First Ionisation Energy Along a Group

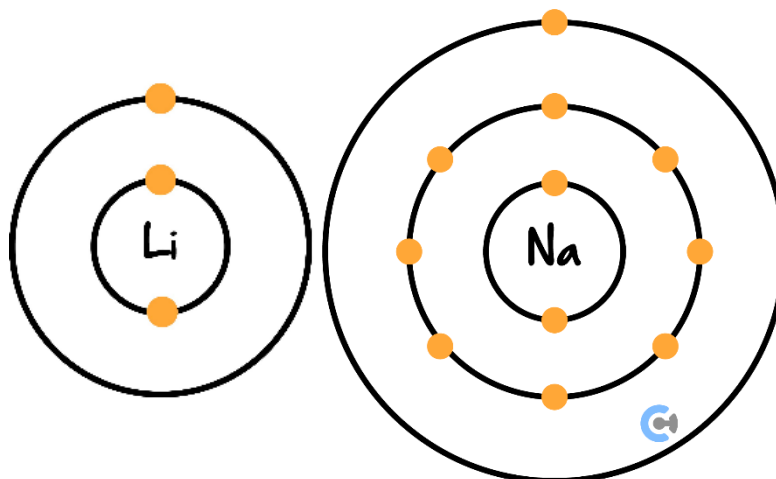


➤ Group 1 elements:

<u>Alkali Metals</u>	<u>Atomic Number</u>	<u>Complete Electronic Configuration</u>
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]**.

First Ionisation Energy Down a Group



- 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~~

Space for Personal Notes

Question 16 (3 marks)

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

- Sulphur has 3rd shells whereas Calcium has 4th shells
- Also sulphur has a core charge of +6 whereas Calcium is +2
- Thus sulphur has ↑ F.I.E.

Space for Personal Notes

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**.

Space for Personal Notes

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 electrons 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), Beryllium (Be) and Boron (B) are all in the same period:

Lithium (Li)	Beryllium (Be)	Boron (B)
Effective Nuclear Charge: <u>+1</u>	Effective Nuclear Charge: <u>+2</u>	Effective Nuclear Charge: <u>+3</u>
Electronegativity: <u>low</u>	Electronegativity:	Electronegativity: <u>high</u>
Atomic Radius: <u>large</u>	Atomic Radius: <u>medium</u>	Atomic Radius: <u>small</u>

Electronegativity and Atomic Radius Across a Period

- 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

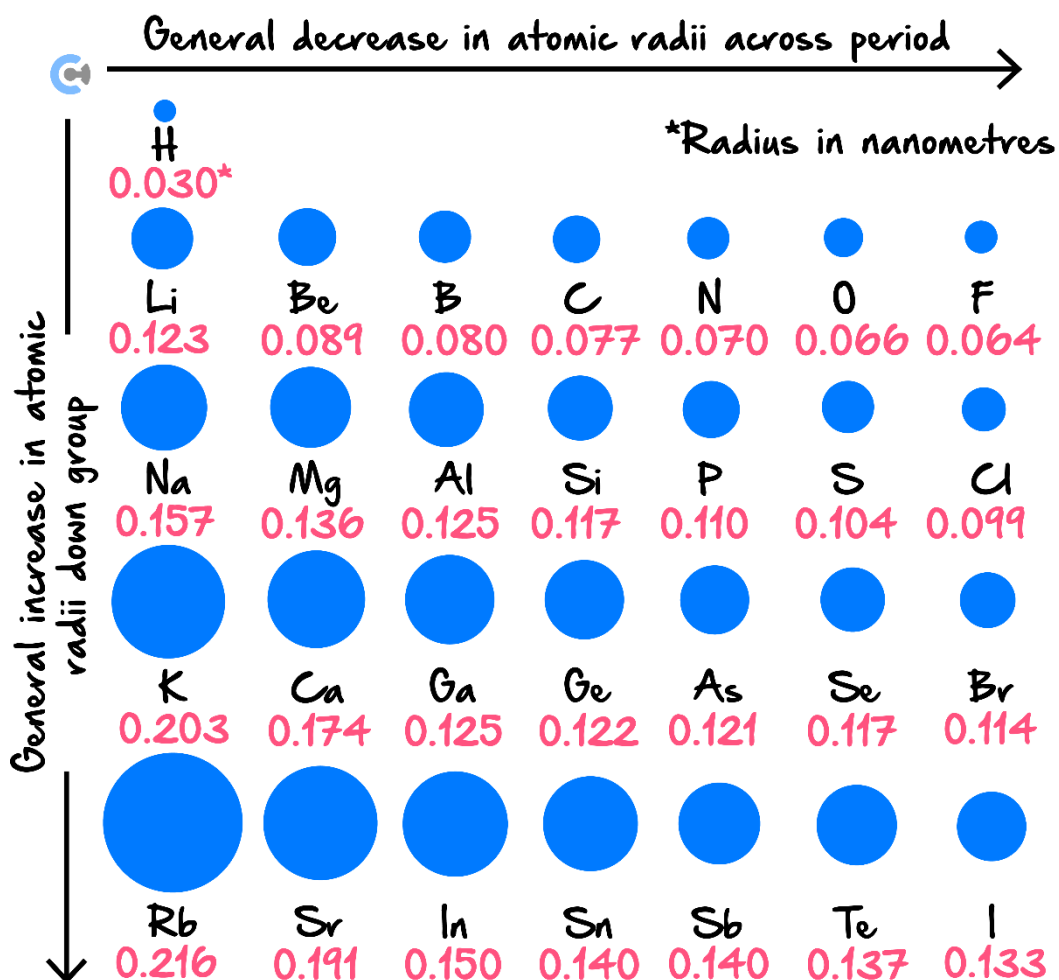
- Down the group, the effective nuclear charge [increases] / [stays the same] / [decreases].
- The distance of the valence shell from the nucleus [increases] / [stays the same] / [decreases].
- As such, the atomic radius [increases] / [stays the same] / [decreases] as the valence shell further increases.
- 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. →

	1	2	G	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						

Electronegativity
decreases down a group. ↓

Atomic radii in nanometres of selected elements



Properties Going Down a Group

Effective Nuclear Charge	Atomic Radius	Electronegativity
Stays the same	Increases	Decreases

Space for Personal Notes

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)

fluorine

- 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.

• Fluorine 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 increases
 • This means e^- are more attracted to nucleus
 • e^- shells are closer $\rightarrow \therefore$ atomic radius decreases

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

• Going down a group no. e^- shells increases
 • Atomic radius \uparrow

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):

s-block

Periodic table of the elements

p-block

d-block

57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.1	71 Lu 175.0
Lanthanum	Cerium	Praseodymium	Neodymium	Promethium	Samarium	Europium	Gadolinium	Terbium	Dysprosium	Holmium	Erbium	Thulium	Ytterbium	Lutetium

81 Au (229)	80 Tl 283.0	81 Pb 281.0	82 Bi 288.0	83 Po (239)	84 At (244)	85 As (243)	86 Cs (249)	87 Ba (249)	88 La (251)	89 Ce (252)	90 Pr (259)	91 Nd (258)	92 Pm (254)	93 U (262)
Aurum	Thallium	Lead	Bismuth	Polonium	Astatine	Antimony	Cesium	Barium	Lanthanum	Cerium	Praseodymium	Neodymium	Neptunium	Plutonium

f-block

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

- 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 shells, 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

- Electronegativity 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.o 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**.

Periodic table of the elements

1 H 1.0 Hydrogen																	3 Li 6.9 Lithium	4 Be 9.0 Beryllium																	5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon																																																																																																																																											
																		Atomic number																		Relative atomic mass																		Symbol of element																		Name of element																																																																																																											
																		74 Au 197.0 Gold																		75 Cu 63.5 Copper																		76 Zn 65.4 Zinc																		77 Ga 69.7 Gallium																		78 Ge 72.6 Germanium																		79 As 74.9 Arsenic																		80 Se 79.0 Selenium																		81 Br 79.9 Bromine																		82 Kr 83.8 Krypton																	
11 Na 23.0 Sodium	12 Mg 24.3 Magnesium																	13 Al 27.0 Aluminium	14 Si 28.1 Silicon	15 P 31.0 Phosphorus	16 S 32.1 Sulfur	17 Cl 35.5 Chlorine	18 Ar 39.9 Argon																			19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton																																																																																																																								
37 Rb 85.5 Rubidium	38 Sr 87.6 Strontium	39 Y 88.9 Yttrium	40 Zr 91.2 Zirconium	41 Nb 92.9 Niobium	42 Mo 95.9 Molybdenum	43 Tc 98.0 Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon																			55 Cs 132.9 Cesium	56 Ba 137.3 Barium	57-71 Lanthanoids	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.8 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.0 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (210) Polonium	85 At (210) Astatine	86 Rn (222) Radon																																																																																																																														
87 Fr (223) Francium	88 Ra (226) Radium	89-103 Actinoids	104 Rf (261) Rutherfordium	105 Db (262) Dubnium	106 Sg (266) Seaborgium	107 Bh (264) Bohrium	108 Hs (269) Hassium	109 Mt (268) Meitnerium	110 Ds (271) Darmstadtium	111 Rg (272) Roentgenium	112 Cn (285) Copernicium	113 Nh (280) Nihonium	114 Fl (289) Flerovium	115 Mc (288) Moscovium	116 Lv (293) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson																			57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium																																																																																																																																	
89 Ac (227)	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np (237) Neptunium	94 Pu (244) Plutonium	95 Am (243) Americium	96 Cm (247) Curium	97 Bk (247) Berkelium	98 Cf (251) Californium	99 Es (252) Einsteinium	100 Fm (257) Fermium	101 Md (258) Mendelevium	102 No (259) Nobelium	103 Lr (262) Lawrencium																			57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium																																																																																																																																				

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

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 stays the same going down a group.
- Effective nuclear charge increases 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.