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VCE Chemistry ½
Models of Atoms [1.1]
Workbook

Outline:

Learning Objectives:

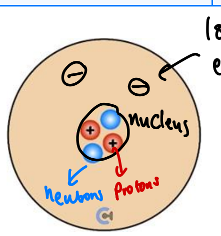
- ☐ **CH12 [1.1.1]** - Describe the Composition of an Atom, and write the Isotopic Symbol of an Element/Ion & use it to identify an Element's/Ion's Atomic and Mass Number
- ☐ **CH12 [1.1.2]** - Describe Bohr's Model of the Atom & draw Shell Model diagrams & apply Emission Spectra to Bohr's Model of the Atom
- ☐ **CH12 [1.1.3]** - Explain Schrodinger's Model of the Atom and identify differences between his Model and Bohr's Model
- ☐ **CH12 [1.1.4]** - Write Electron Configurations of Elements and Ions, in both Ground and Excited States, using both Bohr and Schrodinger Models (including Cu and Cr exceptions and Condensed Notation)

Section A: Atom Structure

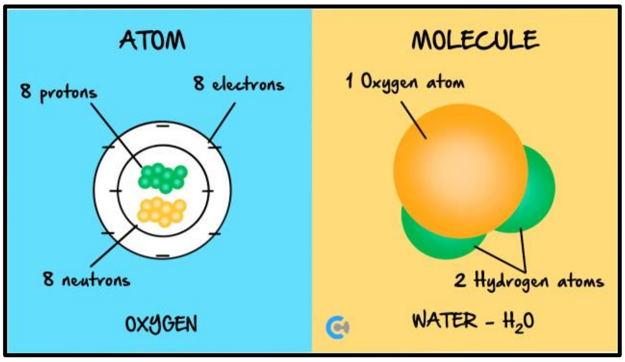
Sub-Section: Atoms Recap

Sub-Atomic Particles

Protons	Neutrons	Electrons
[Positive] / [Negative] / [Neutral]	[Positive] / [Negative] / [Neutral]	[Positive] / [Negative] / [Neutral]
Found in: nucleus	Found in: nucleus	Found in: shells around the nucleus
Size: similar	Size: similar	Size: Very smaller



Atoms and Molecules



- Atomic Number: number of protons
- Mass Number: number of protons + neutrons
- Isotopic Symbol:

mass number
$^{A}_{Z}E$ element symbol
atomic number

Isotopes

- What are the following? isotopes

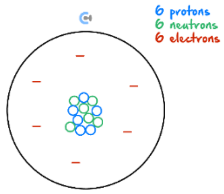
$^{12}_{6}C$ *and* $^{13}_{6}C$

Isotopic Symbol	Worded Representation
$^{12}_{6}C$	carbon-12
$^{13}_{6}C$	Carbon-13

- Isotope Definition:
  - [same] / [different] number of **protons**
  - [same] / [different] number of **neutrons**

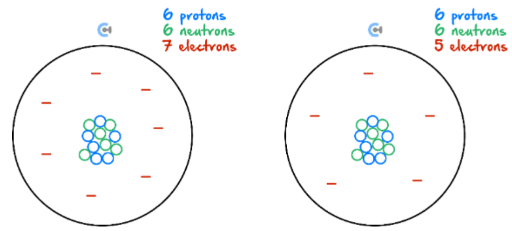
Ions

- Consider Carbon (Atomic Number 6)



- Overall Charge: +6 - 6 = 0 → atom

- If the carbon atom gains or loses an electron:



Cations	Anions
[Positive] / [Negative] Charge	[Positive] / [Negative] Charge
[Gains] / [Loses] Electrons	[Gains] / [Loses] Electrons

Let's have a look at some questions together!

Question 1 Walkthrough.

What is the mass number of Uranium which contains 92 protons and 146 neutrons?

$A = 92 + 146 = 238$

Question 2 Walkthrough.

How many protons, neutrons and electrons are in the following atom?

$^{23}_{11}Na$

11p
11e

$A = 23 = 11 + N$
$N = 12$

Try some questions!

Question 3

Which of the following is not true regarding atomic composition?

A. Atoms are mostly empty space.

B. Electrons are in constant motion around the nucleus.

C. The nucleus takes up a large portion of an atom's size due to its weight.

D. Electrons are roughly 1800 times lighter than neutrons.

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Question 4

Which of the following is true?

A. An element can be made up of multiple atoms.

B. An element can be made up of more than one type of atom.

C. An atom can have multiple nuclei, as long as they are identical.

D. F is an example of a molecule.

Question 5

a. What is the mass number of tin, which contains 50 protons and 70 neutrons?

50 + 70 = 120

b. Find the mass number of aluminium, which has an atomic number of 13 and 14 neutrons.

13 + 14 = 27

Question 6

For each of the following, state whether the substance is a/an **atom**, **molecule** or **ion**. If it is an ion, specify whether it is a **cation** or **anion**.

a. Mg atom	**d.** Cl^- anion
b. H_2O molecule	**e.** CO_2 molecule
c. F_2 molecule	**f.** Na^+ cation

Question 7

For each of the following, state how many **neutrons & electrons** each atom has, and state whether it is a **cation/ anion/ neutral** atom.

a. $^{19}_{9}F^{-}$ 9p 10e⁻ 10n anion	**c.** $^{12}_{6}C$ 6p 6e⁻ 12 = 6 + N N = 6 atom
b. $^{39}_{19}K^{+}$ 19p 18e⁻ 39 = 19 + N N = 20 cation	**d.** $^{13}_{6}C$ 6p 6e⁻ 7 N atom

Question 8 Additional Question.

Give an example of an atom, element, and molecule made from chlorine Cl.

Cl, Cl, Cl_2

Question 9 Additional Question.

State whether $SO_4{}^{2-}$ would be classified as an element, atom, molecule or ion. (Hint: Multiple answers can be correct).

Molecule & ion

Question 10 Additional Question.

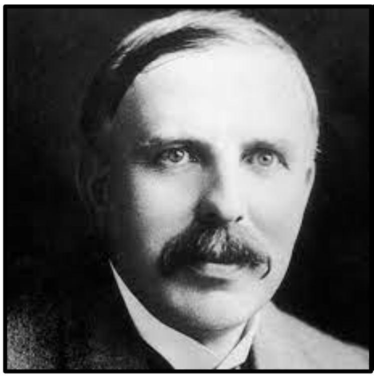
State the number of protons, electrons and neutrons a carbon atom with a mass number of 13 will have, respectively.

Let's have a look at how this model of the atom was discovered!

Sub-Section: Rutherford's Gold Foil experiment

History: Ernest Rutherford's Gold Foil Experiment

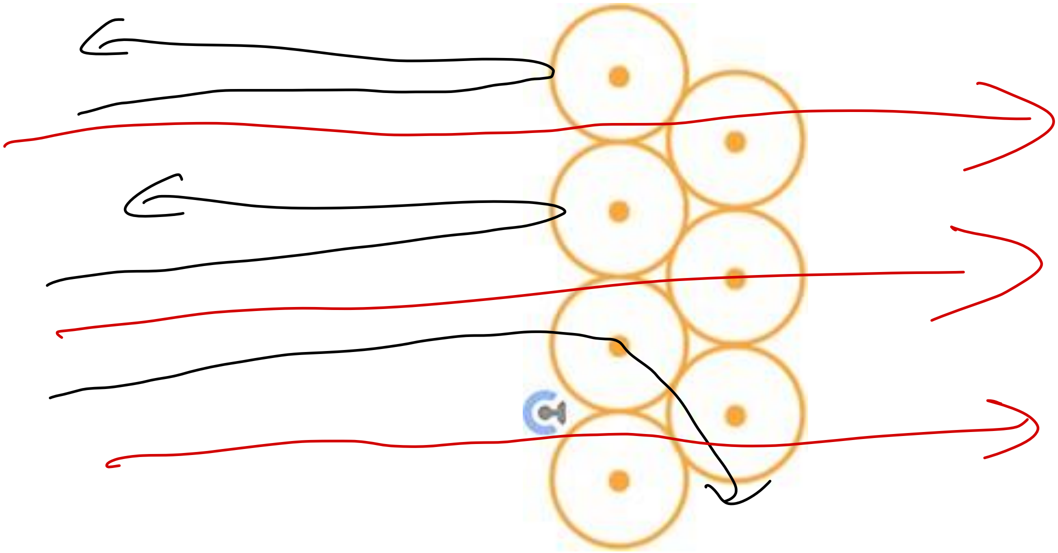
- Ernest Rutherford was famous for his experiment using a sheet of gold in 1908.



Exploration: Rutherford's Gold Foil Experiment

protons & neutrons

- Rutherford fired a beam of alpha particles (helium nuclei) at a thin sheet of gold.
  - An alpha particle has a **[positive]/ [negative]/ [neutral] charge**.



- Some of the particles reflected back. Would you expect this result? **[Yes]/ [No]**
- Most of the particles passed through. Would you expect this result? **[Yes]/ [No]**
- **Conclusion:** This proves that atoms are primarily made of empty space.

Definition

Rutherford's Gold Foil Experiment

- **Alpha particles** were fired at a **thin gold sheet.**
- Some particles **reflected** back, but most particles **passed through.**
- The majority of an atom is **empty space.**

Try a question!

Question 11

Circle the correct alternative from the following:

A. Rutherford's model of the atom explains how electrons exist in shells.

B. Electrons are negatively charged and have the same size as protons.

C. Protons and neutrons are called nucleons because they have similar masses.

D. Neutrons do not have any charge but are found in almost all atoms' nuclei.

Key Takeaways

- Atoms are made up of three subatomic particles.
- Protons are positively charged and are found in the nucleus.
- Neutrons have no charge and are also found within the nucleus.
- Electrons are negatively charged and significantly smaller in size and mass than nucleons.
- Rutherford's Gold Foil experiment saw him firing alpha particles at a thin gold sheet, where some of the particles reflected back, but the majority passed through, proving that atoms are mostly composed of empty space.
- Atoms are identified by their atomic number.
- Atoms can gain or lose electrons to form ions.

<u>Key Takeaways</u>

- The mass number is the number of **nucleons** present in an atom.
- Its formula is given by:

$$A = Z + number\ of\ neutrons$$

- **Isotopic Symbol Representation:**

- **An equal** number of protons, but a **different** number of neutrons.
- Therefore, **different** mass numbers.

Let's have a look at another model of the atom!

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Section B: Bohr's Model of the Atom

Discussion: What did Bohr propose in his model of the atom?

electron configurations

shells, valency

History: Bohr's Model of the Atom

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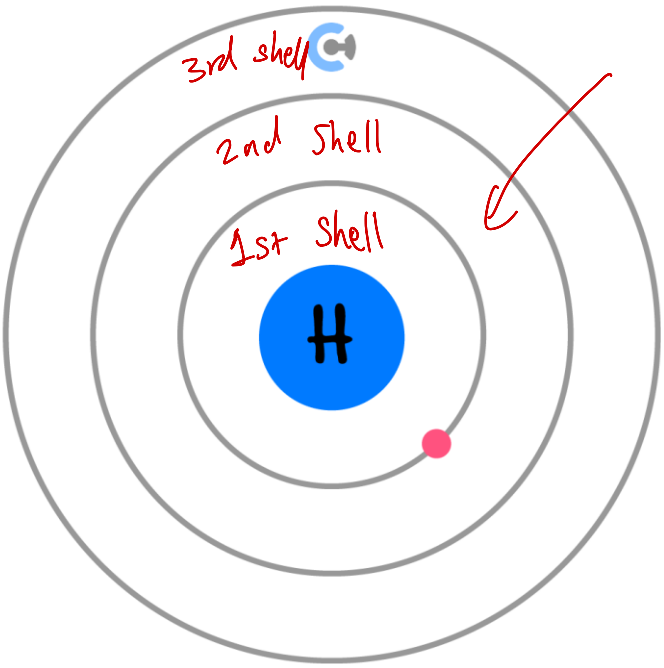
Sub-Section: Bohr's Key Ideas

Exploration: Bohr's Model

- In 1913, Niels Bohr proposed a model of the atom which was as follows:



Electrons found in shells, Not be in between shells

- Key Ideas:
  - Electrons existed at discrete energy levels (called shells).
  - Electrons could move between energy levels.

Electron Shell

- **Definition:**
  - A region of space in the electron cloud of an atom at a discrete energy level, whereby an electron can be found.

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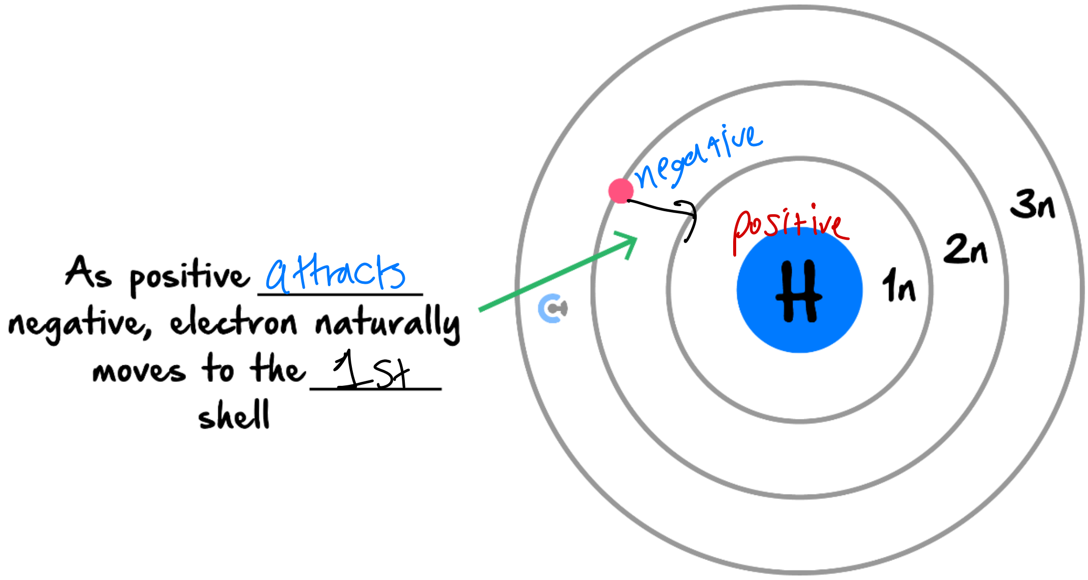
Sub-Section: Movement of Electrons Between Energy Levels

How do electrons move between energy levels?

Exploration: Electron movement between Shells

- What charge does the nucleus have? Positive or negative? *(Label Below)*
- What charge does an electron have? Positive or negative? *(Label Below)*
- A/an **[attractive]**/ **[repelling]** force acts between the nucleus and the electrons. *(Label Below)*
- The electrons will want to move **[towards]**/ **[away]** from the nucleus.
- Therefore, if an electron was already in the second electron shell, it would move to the [first]/ [third] shell. *(Label Below)*



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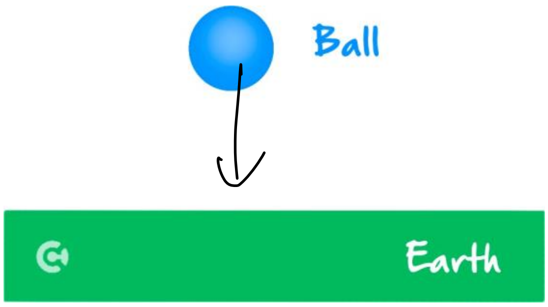
Analogy: Ball

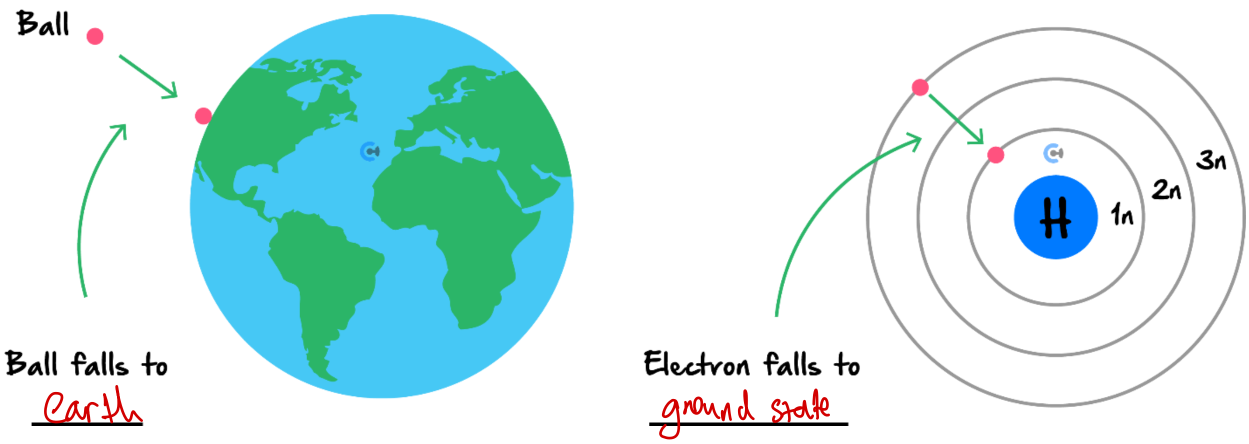
- Think about a ball held in the air which is let go.



- Which way will the ball go?

fall

- If we imagine the Earth like an atom:



- The first electron shell is called the ground state, as it acts similarly to the ground on the **Earth**!

Ground State

- Electrons always try to go the lowest energy level possible.
- The first electron shell is known as **the ground state.**

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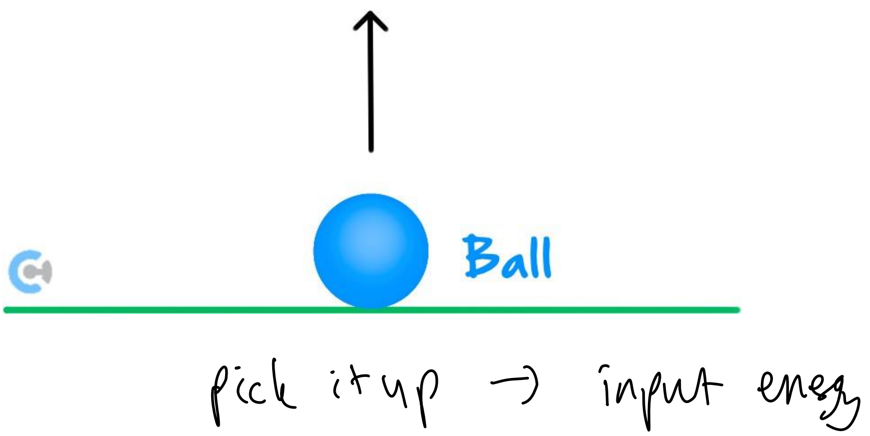
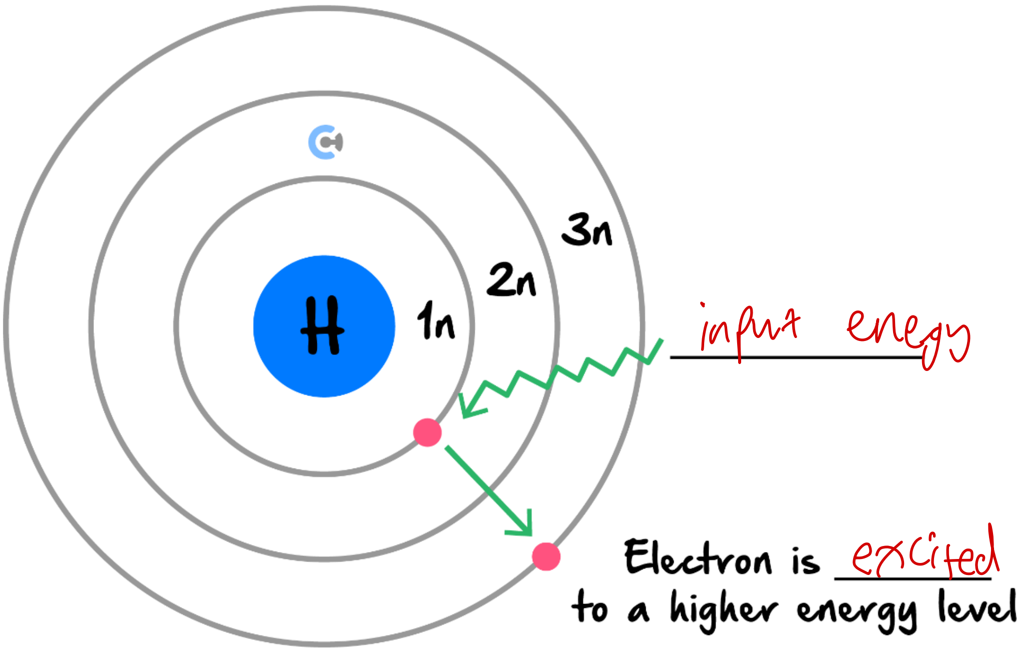
How do we get electrons to move up in an electron shell?

Discussion: If we have a ball sitting on the ground, how do we get it to move up in the air?

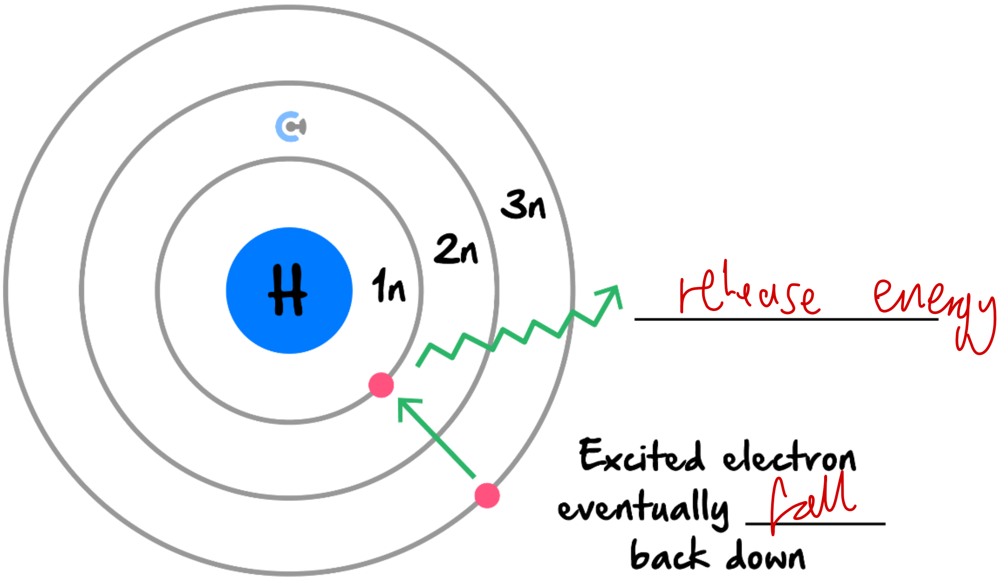


Exploration: Electrons moving up energy levels

- How to move to another higher energy level from the ground state? *(Label Below)*



- However, what must the electron eventually do? *(Label Below)*



Discussion: What type of energy is this energy released as?

lights

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Analogy

- After eating a lot of sugar, you might be super hyper for a while:

- But eventually, you crash and return to your original energy.

- When energy is inputted, electrons are excited to a [higher]/ [lower] energy level.
- Excited electrons eventually drop back down, whereby energy is released, in the form of light.

Extension: LED Lighting

- A large majority of light sources we see today, use this to emit light, light sources like LEDs and fluorescent lighting all produce light due to electrons going down to a lower energy level.

Try a question!

Question 12

Outline the 2 major stages that occur when energy is provided to an atom.

- Electrons are excited to a higher energy level (shell)
- Eventually, electrons fall back down to the ground state, releasing light as it falls

Question 13 Additional Question.

Bohr's model suggests which of the following?

A. Electrons orbit the nucleus randomly within a cloud at continuous energy levels.

B. Electrons exist in shells within the nucleus.

C. Atoms have shells which are discrete energy levels.

D. Electrons are fixed to a particular shell at all times.

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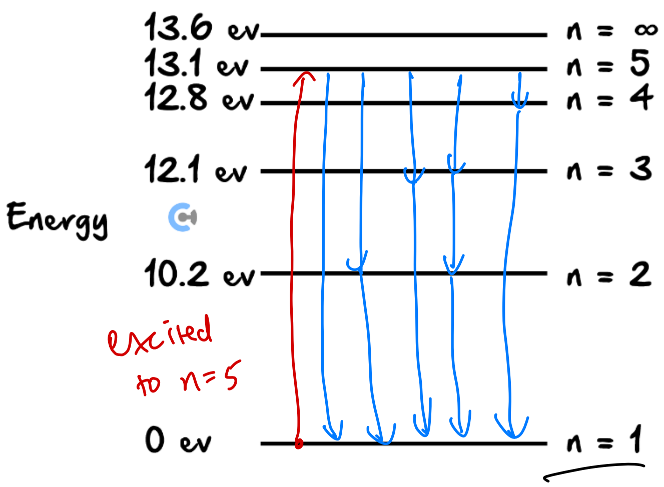
Sub-Section: Emission Spectrum

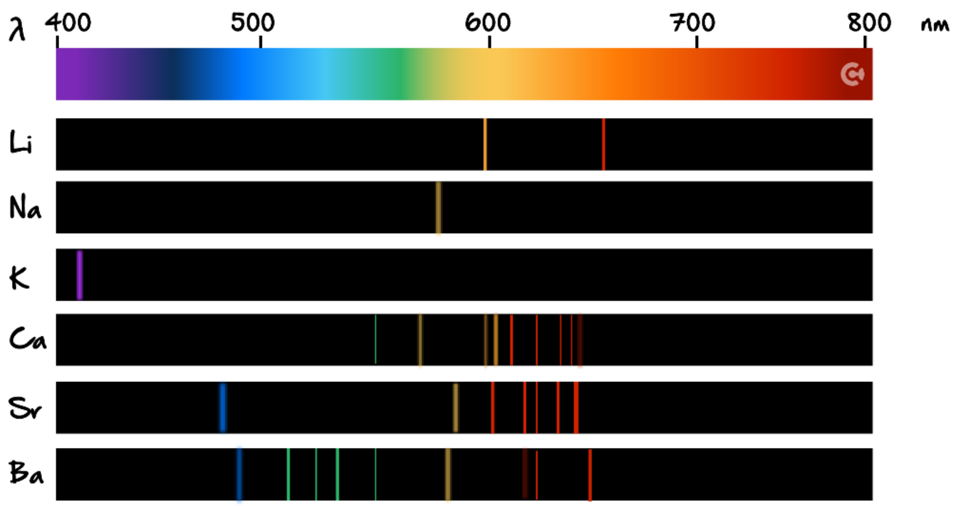
What does the light emitted look like?

Exploration: Returning back to the Ground State

- There are multiple ways an electron can return to the ground state from an excited state.



- Each path the electron takes on the way down has a different energy difference.
- This energy difference leads to different colour of light being emitted.
- Different elements release different spectra of light when their electrons are excited.



- Each element has a unique spectra of light when their electrons return to the ground state after being excited.

Discussion: What happens if we input so much energy that the electron is excited past the last energy shell level?

dissapears

Cation

NOTE: The energy required to completely remove the first electron from an atom is called its first ionisation energy.

ALSO NOTE: The idea of **first ionisation energy** will be covered in the **next lesson**!

Try a question for yourself!

Question 14

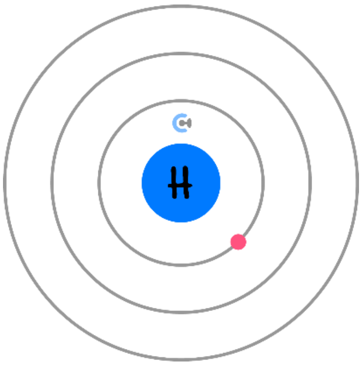
State whether the following statements about the electron shell model of the atom are true or false.

	True	**False**
Electrons can exist between two energy levels.		✓
When electrons occupy the lowest possible energy level, the atom is in the ground state.	✓	
The energy difference between $n = 2$ and $n = 3$ is greater than the energy difference between $n = 3$ and $n = 4$.	✓	
Electrons in atoms in excited states return to the next lowest energy level.	✓	✓
An electron that moves between $n = 5$ and $n = 2$ emits higher energy light than an electron that falls between $n = 4$ and $n = 2$.	✓	

Sub-Section: Electron Configuration

Context

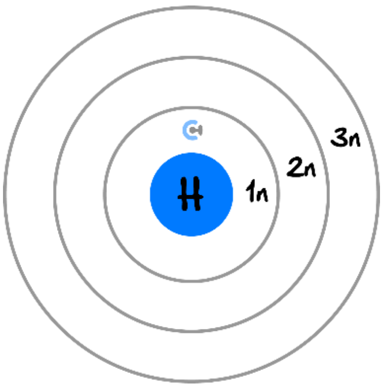
- We know Bohr's Model looks something like this and can have electrons in each shell.



- But how many electrons fit into each energy level?

Exploration: Maximum number of electrons in each shell

- Which shell can hold more electrons between the first and third shell?



- As the electron shell number increases, the maximum number of electrons per shell ____increases____ as the shells become larger.

- **Rule:**

$$\textbf{Maximum Number of Electrons: } 2n^2$$

- n - Shell Number

- Maximum number of electrons that each shell can hold is:

$$\textit{Maximum Number of Electrons}: 2 \times n^2$$

- n - Shell Number

Try a question!

Question 15

How many electrons can each of the following shells hold? (Hint: use the $2n^2$ formula)

a. Shell 1 $2(1)^2 = 2$	**c. Shell 3** $3(2)^2 = 18$
b. Shell 2 $2(2)^2 = 8$	**d. Shell 4** 32

NOTE: Electron shells are filled in order from the nucleus, with the innermost shells fully filled before moving on.

<u>Valence Shell / Valence Electrons</u>

- **Definition:**
  - Valence - Outer.
  - Valence Shell - Outer-most electron shell.
  - Valence Electrons - Electrons in outermost electron shell.

Let's have a look at a question together!

Question 16 Walkthrough.

Write out the electron configuration for each of the following elements:

a. C (carbon) (6 electrons)

6 - 2 = 4

n = 1: 2

n = 2: 8

2, 4

b. Na (sodium)

2, 8, 1

NOTE: Sometimes, we need to look at the **periodic table** to find the number of electrons present.

REMINDER: Don't forget!

- Assuming that the atom has a neutral charge, the **number of electrons** is **equal** to the **atomic number**!

Active Recall

Shell Number	Maximum Number of Electrons
1	
2	
3	
4	

Now try a question for yourself!

Question 17

Write out the electron configuration for each of the following elements:

a. N (Nitrogen) 2, 5	**c. Ar (Argon)** 2, 8, 8
b. S (Sulphur) 2, 8, 6	**d. K (Potassium)** 2, 8, 8, 1

Misconception

"The electron configuration of elements such as potassium (K) is 2, 8, 9."

TRUTH:

- The octet rule states the valence shell can hold a maximum of 8 electrons.
- If the electron shell can hold more than 8 electrons, if it is the valence shell, electrons are added to the next electron shell instead.
- As such, potassium (K) has an electron configuration of:

2, 8, 8, 1

Octet Rule

- **Definition:**
  - The valence electron shell can only hold a maximum of 8 electrons.

Try another similar question!

Question 18

Write out the electron configuration for calcium, which has 20 electrons.

2, 8, 8, 2

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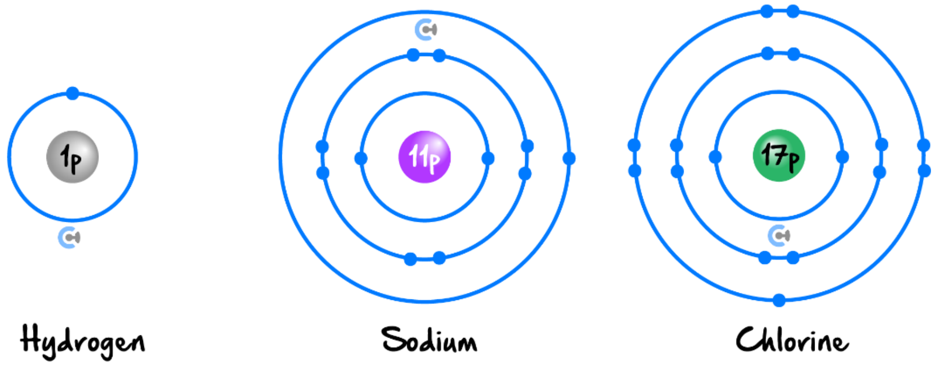
Sub-Section: Shell Model Diagrams

Context

- Electron configurations of atoms are represented by **shell model diagrams**.



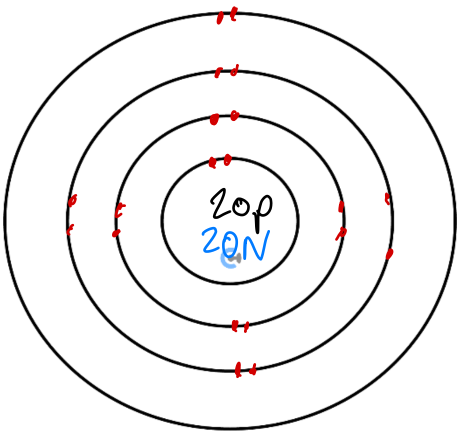
Shell model diagrams

- They show the number of electrons which are present in each electron shell of an atom.

Let's have a look at a question together!

Question 19 Walkthrough.

Draw the shell model diagram for Calcium (Ca).



REMINDER: Don't forget!

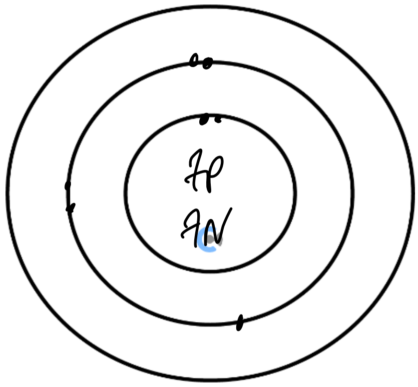
- Pair up the electrons where possible!

Try some questions for yourself!

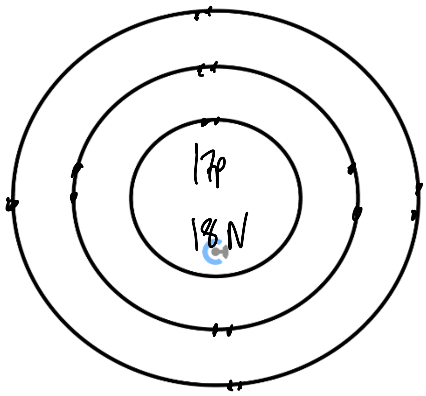
Question 20

Draw the shell model diagram for each of the following elements:

a. Nitrogen (N)



b. Chlorine (Cl)



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Key Takeaways

- Key Ideas from Bohr's Model:
- Electrons exist in **discrete** energy levels (called **shells**).
- Electrons can **move** between energy levels.
- When energy is inputted, electrons are **excited** to a **higher** energy level.
- Electrons eventually drop back down - as they always want to be in the **lowest** energy state - whereby energy is released, in the form of **light**, which is unique to each element.
- Electron shells are filled in order from the nucleus, with the **innermost** shells fully filled before moving on.
- Each shell can hold up to $2n^2$ electrons, where n is the shell number.
- The octet rule states that the valence electron shell can only hold a maximum of **eight** electrons.
- Shell model diagrams show the number of electrons within each **shell**.

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Section C: Schrödinger's Model of the Atom

Context

- Bohr's Model has several shortcomings:
  - Cannot explain why electron shells can only hold $2n^2$ **electrons**.
  - Cannot explain why the **fourth shell accepts two electrons before the third shell** is completely filled (we haven't covered this scenario yet but we'll have a look at it soon!).
- Erwin Schrödinger came up with his new revised model of the atom.

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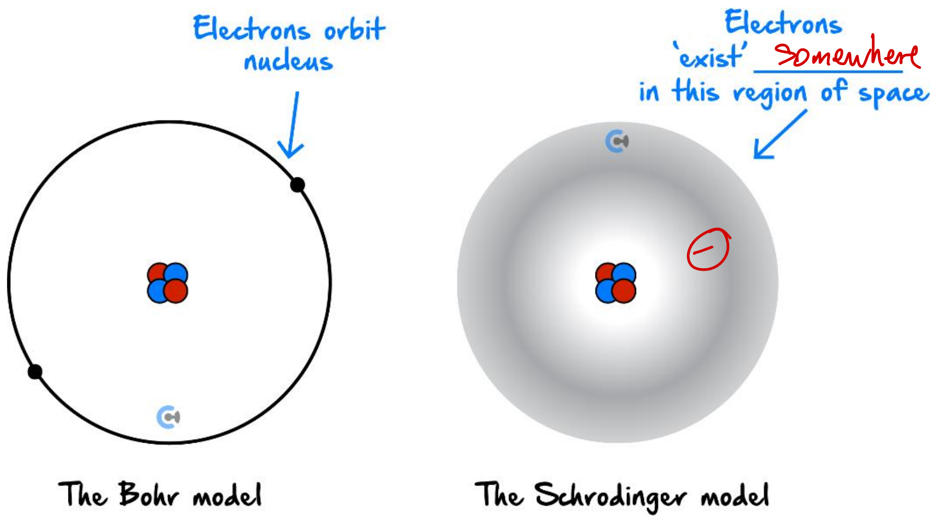
Sub-Section: Key Features of Schrödinger's Model

History: Schrödinger's Model of the atom

- In 1926, Erwin Schrödinger came up with his new revised model of the atom which explained these limitations that the Bohr Model did not explain.

Exploration: Schrödinger's Model vs Bohr's model of the atom

- Main difference:
  - Electrons did not orbit the nucleus like planets but instead had wave-like properties like light.
  - Electrons merely existed ___somewhere___ in a region of space, which was called an ___orbital___.



- He claimed that each orbital can only hold ___two___ electrons.

Orbital

- **Definition:** A region of space in which electrons exist randomly, not in fixed discrete energy levels.
- **Feature**: Can hold up to 2 electrons.

Exploration: Subshells

- Electron shells can be broken down into subshells.
- Each subshell has a different shape and orientation in 3D space. These include the:
  - s - orbital.
  - p - orbital.
  - d - orbital.
  - f - orbital.

NOTE: These subshells are quoted in this order as they are arranged in increasing size.

Let's have a look at some of these subshells in-depth!

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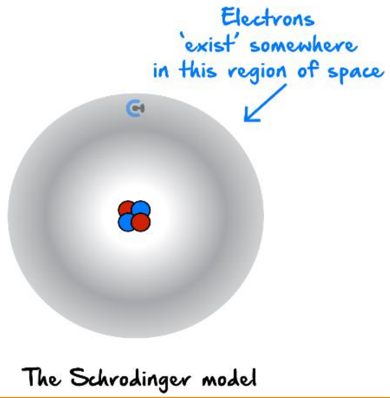
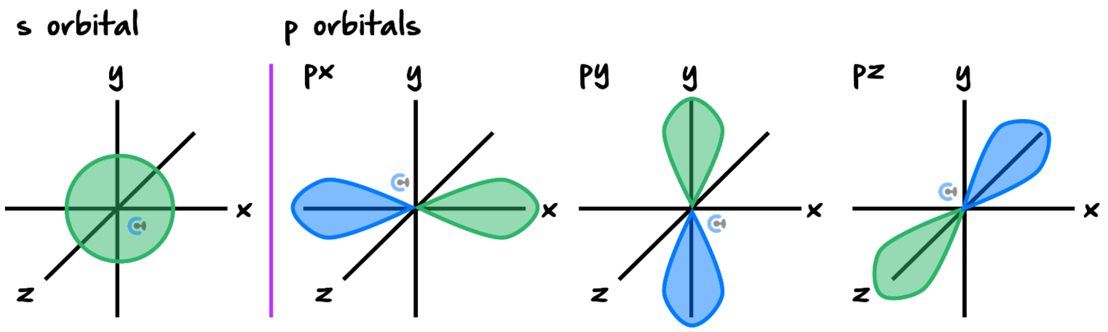
Exploration: s-orbital

- The s-orbital was visualised to be a **sphere**, where the electrons could be anywhere in that space.

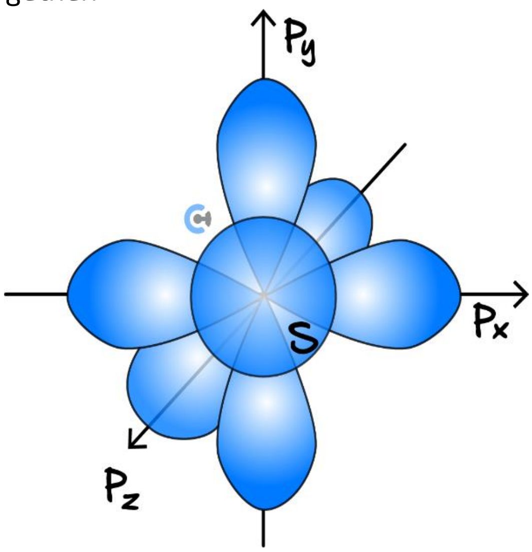


Exploration: p-orbital

- The p-orbitals basically look like two blown-up balloons attached to each other.
- Aligned with either the x, y or z-axes.



- The s and p orbitals fitted together:

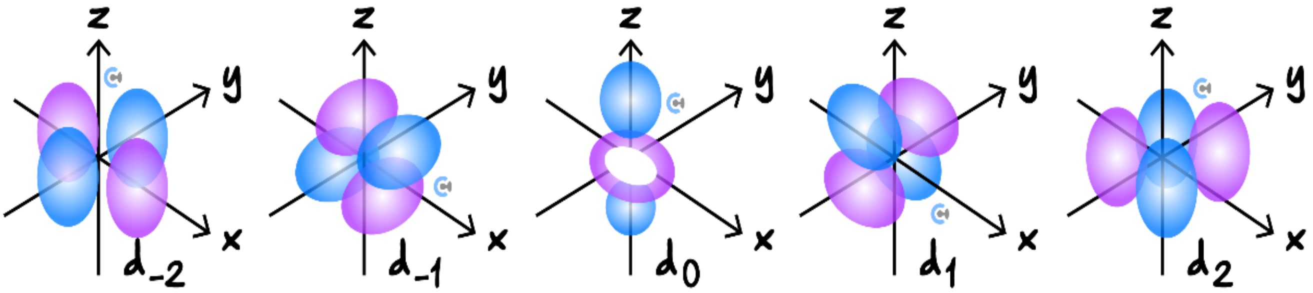


NOTE: The z-axis exists in 3D space, as we live in a 3-dimensional world!

ALSO NOTE: Electrons are assumed to be in constant motion, so their exact location within an orbital is unknown.

Extension: Shape of d-orbitals (not assessed)

- d-orbitals start to look a bit more complicated! There are five d-orbitals:



Orbital capacity

- There are four types of orbitals, s, p, d and f.

Type of Orbital	Number of Orbitals
s	1
p	3
d	5
f	7

Active Recall: How many electrons can fit into each of these orbitals?

2

How do orbitals relate to subshells and shells?

Exploration: Shells, Subshells and Orbitals

- The first electron shell is small, contains s -orbitals.
- The second electron shell is larger, contains s-orbitals and p -orbitals.
- The third electron shell is even larger, contains s-orbitals, p-orbitals and d -orbitals.
- **Conclusion:** As the electron shell increases, it can fit more orbitals!

Shell number (n)	Number of subshells	Subshell symbol	Number of Orbitals	Maximum number of electrons in the subshell	Total Number of electrons in the shell
1	1	s	1	2	2
2	2	s	1	2	8
		p	3	6	
3	3	s	1	2	18
		p	3	6	
		d	5	10	
4	4	s	1	2	32
		p	3	6	
		d	5	10	
		f	7	14	

Bohr

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Subshells

- The number of subshells in a given shell matches the shell number.
- For example, shell 3 can hold 3 subshells.
- Example:

Subshell	Orbital	Electrons
3s	1	2
3p	3	6
3d	5	10

18

Try some questions!

Question 21

Select the correct alternative from the following:

A. The 1st shell can hold s and p orbitals.

B. The d orbital can hold 5 electrons.

C. The p subshell can hold 6 electrons.

D. The 2s subshell has 2 orbitals.

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Question 22

Which of the following is NOT correct regarding the Schrodinger and Bohr Models of the atom?

A. Bohr's model explains the existence of shells.

B. Schrodinger's model states that electrons are in fixed positions within orbitals.

C. Schrodinger's model is the newer one out of the two, explaining more properties of atoms.

D. Bohr's model can be linked to emission spectra of elements.

Question 23 Additional Question.

State how many orbitals and electrons the 5d subshell can hold, respectively:

Orbitals: 5

Electrons: 10

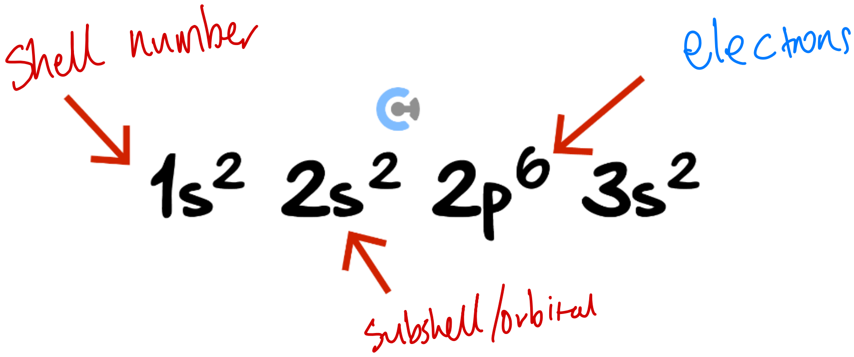
Space for Personal Notes

Sub-Section: Representation of Schrödinger Electron Configurations

How do we write electron configurations according to Schrödinger's model?

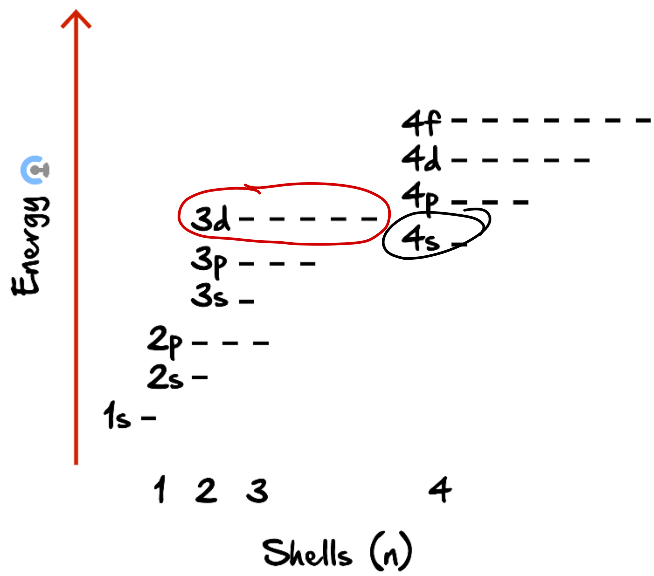
Schrödinger Electron Configurations

- **Representation:**



Exploration: Energy Levels of Sub-shells

- Each of the subshells has slightly different energy levels.

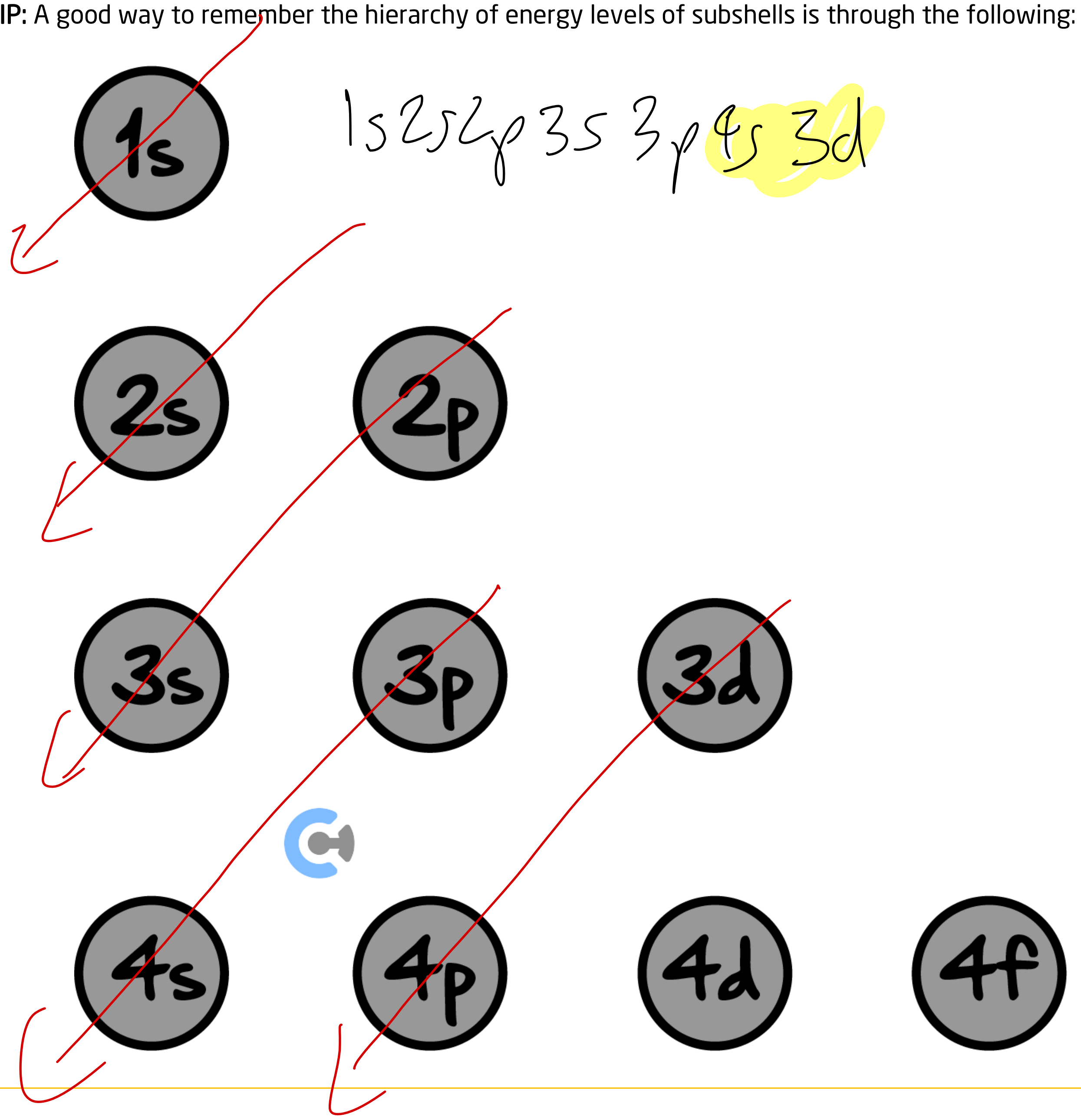


- Electrons are filled up from the lowest energy level to the highest energy level.

Discussion: 4s and 3d Subshell comparison

- The 4s subshell has a **[higher]/[lower]** energy level than 3d.
- Therefore, the 4s subshell will be filled **[before]/[after]** the 3d subshell.

TIP: A good way to remember the hierarchy of energy levels of subshells is through the following:



- Each subshell has different energy levels.
- For electron configuration, start from the lower energy levels and fill up to higher energy levels.

Active Recall: The following subshells can hold:

Type of Orbital	Number of Orbitals	Number of Electrons
s	1	2
p	3	6
d	5	10
f	7	14

Let's try a question together!

Question 24 Walkthrough.

Write Schrödinger's electronic configuration for Potassium (K). 19p , 19e -

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

19 − 2 = 17 − 2 = 15 − 6 = 9 − 2 = 7

7 − 6 = 1

REMINDER: Don't forget the 4s subshell fills **before** the 3d subshell

NOTE: The third electron shell starts to fill up while the fourth electron shell remains at one/two electrons!

TIP: Write Schrödinger's electronic configuration using the **periodic table!**

Try some questions for yourself!

Question 25

Write Schrodinger's electronic configuration for each of the following elements:

a. Cl 17

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

b. K

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

c. V

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

d. Fe

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

$3d^6 4s^2$

NOTE: There are two ways to write 3d and 4s orbitals in electron configurations:

- 4s before 3d.
- 3d before 4s.
- As both are valid, follow what your teacher wants!

Question 26 Additional Question.

Write Schrodinger's electronic configuration for each of the following elements:

a. O

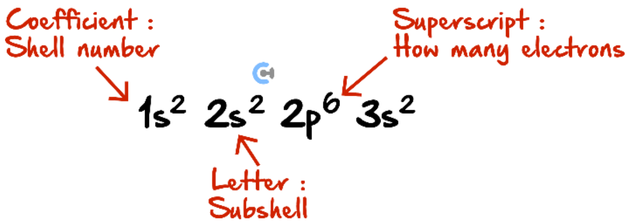
$1s^2 2s^2 2p^4$

b. Al

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

Key Takeaways

- An **orbital** is a region of space in which electrons exist randomly, not in fixed discrete energy levels.
- Each orbital can hold up to $\mathbf{2}$ electrons.
- There are 4 different types of orbitals: $\mathbf{s}, \mathbf{p}, \mathbf{d}$ **and** $\mathbf{f}$, which respectively hold $\mathbf{2}, \mathbf{6}, \mathbf{10}, \mathbf{14}$ electrons.
- Each shell n contains n subshells. For example, shell 2 contains 2 subshells: 2s and 2p.
- Representation:



- When writing out the electron configuration of elements, start from the **lower** energy levels and fill up to higher energy levels.
- 4s subshell fills **before** $\mathbf{3d}$ subshell.

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Section D: Atypical Electron Configurations

Sub-Section: Ionic and Excited State Configurations

Exploration: Excited state configurations

- Default electron configurations of atoms in their ground state, meaning that the electrons fill up the lowest energy levels first.
- If an electron is excited, it will go to a [higher]/[lower] energy state.
- Example: Magnesium

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

[ground] / [excited] state

- How do we know?

Jumped 4th & 5th shells

Excited state electron configurations

If electrons are in higher energy levels without lower ones being filled then...	Excited state (e.g., 4^{th} shell filled before 3^{rd}).
If lower energy subshells are not filled but higher ones are being filled then...	Excited state (e.g., 3p orbital is filled but not 3s).

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Try some questions!

Question 27

a. Find the Schrodinger electron configuration of an aluminium ion, Al^{3+}. → $10e^-$

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

b. The following electron configuration represents an atom or an ion. Is it in an excited state?

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

[Yes] / [No]

c. Is $1s^2 2s^2 2p^6 3s^2\ 3p^3 4s^1$ an excited state configuration?

$3p^4$

[Yes] / [No]

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Sub-Section: Chromium and Copper

Context

- There are two specific exceptions you need to be aware of to the above rules in Chemistry ½.

Exploration: Chromium and Copper

- **Expected configurations:**

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$$\textbf{Chromium}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$$

29

$$\textbf{Copper}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$$

- **Actual Configurations:**

18 19

$$\textbf{Chromium}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6\ 4s^1\ 3d^5$$

18 19

$$\textbf{Copper}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6\ 4s^1\ 3d^{10}$$

- **Result**: Gives chromium a half-filled d-subshell and copper a fully-filled d-subshell.

- **Stability**: Partly filled subshells < Half-filled subshells < Fully filled subshells.

- The combination of half-filled/filled subshell ($3d^5$ and $4s^1$) is **[more] / [less]** stable than one filled subshell and one partly-filled subshell ($3d^4$ and $4s^2$).

Chromium and Copper: Atypical Electron Configurations

Definition

- Their Shrodinger electron configurations are:

$$\text{Chromium}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$$

$$\text{Copper}\ \ 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$$

- Due to an **increased stability** in these forms.

Practice this idea!

Question 28

a. Write the full Schrödinger's electron configuration for Copper (Cu).

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

b. Rewrite this in Bohr's electron configuration.

2, 8, 18, 1

NOTE: As an electron from the fourth electron shell jumps back to the third shell, there is only one electron left in the fourth shell.

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Sub-Section: Condensed Electron Configuration

Definition

- The periodic table arranges all of the chemical elements in terms of increasing atomic number

Periodic table of the elements

Key: 79 (Atomic number) / Au (Symbol of element) / 197.0 (Relative atomic mass) / Gold (Name of element)

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
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 96.0 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 178.5 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 (267) 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 (289) Moscovium	116 Lv (292) 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) Actinium	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

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

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

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Exploration: Condensed Electron Configuration

- The electronic configuration of silicon:

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

- Condensed electronic configuration of silicon:

[Ne] $3s^2 3p^2$

- Examples:

  - Starting with:

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

$[Ne]3s^2 3p^2$	$[Mg]3p^2$
[Correct]/[Incorrect]	[Correct]/[Incorrect]

Definition

Condensed Electron Configuration

- **Use**: To 'get rid' of entire shells-worth of notation, for tidiness and efficiency.
- **Notation**: noble gas in square brackets, as these are the elements with full outer shells which may be condensed.

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Let's have a look at a question together!

Question 29 Walkthrough.

Write the **condensed** electron configuration of the following elements:

a. Boron (B)

$1s^2$ [He] $2s^2 2p^1$

b. Magnesium (Mg)

[Ne] $3s^2$

Try some for yourself!

Question 30

Write the **condensed** electron configuration of the following elements:

a. Lithium (Li)

[He] $2s^1$

b. Potassium (K)

[Ar] $4s^1$

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Contour Check

Learning Objective: [1.1.1] Describe the Composition of an Atom, and write the Isotopic Symbol of an Element/Ion & use it to identify an Element's/Ion's Atomic and Mass Number

Study Design

the definitions of elements, isotopes and ions, including appropriate notation: atomic number; mass number; and number of protons, neutrons and electrons

Key Takeaways

- Atoms are made up of three Subatomic particles.
- Protons are positively **charged** and are found in the nucleus.
- Neutrons have no **charge** and are also found within the nuclew.
- Electrons are negatively **charged** and significantly smaller in size and mass than nucleons.
- Rutherfold's gold experiment saw him firing **alpha particles** at a **thin gold sheet**, where some of the particles **reflected** back, but the majority passed through, proving that atoms are mostly comprised of empty space.
- Atoms are identified by their atomic number.
- Atoms can gain or lose electrons to form ions.
- The mass number is the number of nucleons present in an atom.
- **Isotopic symbol Representation:**

  A → mass #

  Z → atomic #

  E ← symbol

- Isotopes are two or more of the same elements with same number of **protons**, but different number of **neutrons**, and therefore, different mass numbers

Learning Objective: [1.1.2] Describe Bohr's Model of the Atom & draw Shell Model diagrams & apply Emission Spectra to Bohr's Model of the Atom

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

Key Ideas from Bohr's Model:

- Electrons exist in discrete energy levels (called shells)
- Shell model diagrams show the number of electrons within each shell
- Electrons can move between energy levels
- When energy is inputted, electrons are excited to a higher energy level.
- Electrons eventually drop back down - as they always want to be in the lowest energy state – whereby energy is released, in the form of light which is unique to each element

Learning Objective: [1.1.3] Explain Schrodinger's Model of the Atom and identify differences between his Model and Bohr's Model

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

- An orbital is a region of space in which electrons exist randomly, not in fixed discrete energy levels
- Each orbital can hold up to 2 electrons
- There are 4 different types of orbitals: s, p, d, f which respectively hold 2, 6, 10, 14 electrons
- Each shell n contains n subshells. For example, shell 2 contains 2 subshells: 2s and 2p

Learning Objective: [1.1.4] Write Electron Configurations of Elements and Ions, in both Ground and Excited States, using both Bohr and Schrodinger Models (including Cu and Cr exceptions and Condensed Notation)

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

- Electron shells are filled in order from the nucleus, with the innermost shells fully filled before moving on.
- Each shell can hold up to $2n^2$ electrons, where n is the shell number
- The octet rule states that the valence electron shell can only hold a maximum of 8 electrons.
- Schrodinger representation: shell # → ; # of e⁻

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

orbital

- When writing out the electron configuration of elements, start from the lowest energy levels
- 4s subshell fills before 3d subshell
- If electrons are found to be in higher energy levels - **shells or subshells** - without the lower ones being filled first, you can assume that they are in the excited state.
- Cu and Cr Shrodinger electron configurations are
  - Chromium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
  - Copper: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
- This is due to an increased stability in these forms
- Condensed notation always has a noble gas in square brackets, as these are the elements with full outer shells