

PTable WebQuestGoal

- Use **Ptable.com** to explore the organization of the periodic table:
 - By **atomic number, periods, and groups**
 - To locate **metals, nonmetals, and metalloids**
 - To identify key **named groups**: Groups 1, 2, 17, 18
- Apply your observations by **coloring and labeling a blank periodic table**

Materials

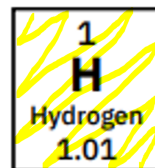
- Computer or device with internet access
- Access to <https://Ptable.com> (**Properties view**)
- Blank periodic table handout (black & white)
- Colored pencils or markers
- Pencil and paper for notes and observations

Instructions & Tasks**1. Atomic Number, Periods, and Groups**

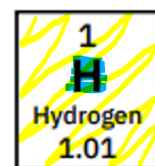
- Look at your blank periodic table and complete these:
 - The periodic table lists elements in order of _____
 - Periods are the _____
 - Groups are the _____

2. Metals, Nonmetals, and Metalloids

- On Ptable.com:
 - Use the side or top menus/legend to locate **Metals, Nonmetals, Metalloids** (color-coded)
 - Hover over or click elements to see their category
- On your handout:
 - **Color-code the boxes** for each category
 - Record your color choices in the key

**3. Phases of Elements**

- On Ptable.com:
 - Locate **Solid, Liquid, Gas, Unknown** categories (color-coded)
 - Hover over or click elements to see which phase they are in
- On your handout:
 - **Color the element symbols** to indicate their phase
 - Record your colors in the key (*Solid colors are already provided*)

**4. Important Named Groups**

- Identify and examine these groups on Ptable.com:
 - **Group 1** (far-left column, exclude Hydrogen if desired)
 - **Group 2** (next to Group 1)
 - **Group 17**
 - **Group 18**
 - **Transition metals** (center block)
- On your handout: **Label each special group**

Periods

Groups

Periodic Table Structure

Color Code Key

Metals:

Nonmetals:

Metalloids:

Solids: Black

Liquids:

Gases:

Source: CRC Handbook of
Chemistry and Physics, 104th ed.,
2023–2024, CRC Press

58	Ce	59	Pr	60	Nd	61	Pm	62	Sm	63	Eu	64	Gd	65	Tb	66	Dy	67	Ho	68	Er	69	Tm	70	Yb	71	Lu
140.12	Cerium	140.91	Praseodymium	144.24	Neodymium	Promethium	Samarium	150.36	Europtium	151.96	Gadolinium	157.25	Terbium	158.93	Dysprosium	162.50	Hoium	164.93	Erbium	167.26	Thulium	168.93	Ytterbium	173.04	Lutetium	174.97	
90	Th	91	Pa	92	U	93	Np	94	Pu	95	Am	96	Cm	97	Bk	98	Cf	99	Es	100	Fm	101	Md	102	No	103	Lr
232.04	Thorium	231.04	Protactinium	238.03	Uranium	Neptunium	Plutonium	Americium	Curium	Berkelium	Californium	Einsteinium	Fermium	Mendelevium	Nobelium	Lawrencium											

The periodic table lists elements in order of

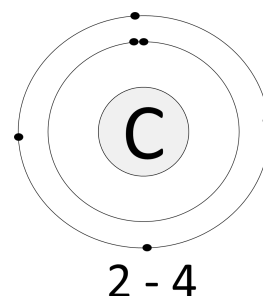
Periods are the

Groups are the

Electron Configurations and the Periodic Table: Bohr Models for Periods 2 and 3

1. Bohr Models

- Your teacher will provide **Bohr model diagrams** for elements in:
 - **Period 2:** Li → Ne
 - **Period 3:** Na → Ar
- For each element:
 - Count the number of electrons in **each ring (energy level)**
 - Write the **electron configuration** (example: Carbon → 2 - 4)



2. Discovering Patterns Across a Period

- Compare all elements in **Period 2**.
- **Observation tasks:**
 - How many energy levels (rings) do the elements have?
 - What changes as you move **left to right** across the period? (*Hint: check the number of electrons in the outermost ring*)
- Write a **short statement** about the pattern you notice in the space under the models.

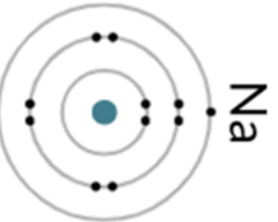
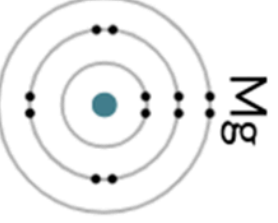
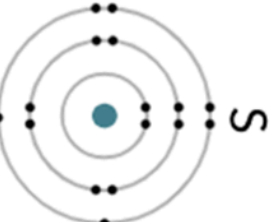
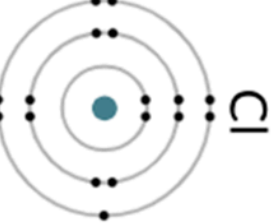
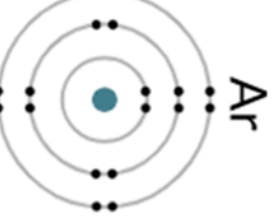
3. Discovering Patterns Down a Group

- Compare elements **Lithium (Li) vs. Sodium (Na)** or **Fluorine (F) vs. Chlorine (Cl)**.
- **Observation tasks:**
 - What happens to the number of energy levels (rings) as you move **down a group**?
 - What stays the same? (*Hint: look at the outermost electrons — valence electrons*)
- Write a **short statement** about the pattern you observe in the space under the models.

4. Drawing Lewis Structures (Dot Diagrams)

- **Step 1:** Write the element symbol. (*If drawing generic diagrams for a group, use "X" as a placeholder.*)
- **Step 2:** Determine the number of **valence electrons**:
 - For main-group elements, this equals the **ones digit of the group number** (e.g., Group 14 → 4 valence electrons)
 - Or, check the last number in the electron configuration
- **Step 3:** Place dots around the symbol to represent valence electrons:
 - Start with **one dot on each side** (top, bottom, left, right)
 - Pair them only **after each side has one dot**
 - Maximum of **two dots per side**
 - Ensure the total number of dots equals the number of valence electrons
- **Step 4:** On your periodic table handout, draw **dot diagrams for elements in each group**, using X if needed

Periods 2 and 3 Bohr Models

	<u>1</u>	<u>2</u>	<u>16</u>	<u>17</u>	<u>18</u>
	Li 	Be 	O 	F 	Ne 
<u>Electron Configurations</u>	2-1				
	Na 	Mg 	S 	Cl 	Ar 
<u>Electron Configurations</u>	3				
<u>Lewis Dot Diagrams</u>	X	X	X	X	X

Patterns: what happens in terms of electron configurations;

- 1) As you go left to right across a period?
- 1) As you go top to bottom down a group?

Properties of special groups

The periodic table is arranged so that elements in the same column (called _____) share similar properties. Some groups are especially important because their elements are very reactive or very stable. These “special groups” help explain how elements behave in chemical reactions and why they are useful in everyday life.

In this handout, we'll look at:

- **Group 1: Alkali Metals** – the most reactive metals
- **Group 2: Alkaline Earth Metals** – strong, reactive metals
- **Group 17: Halogens** – colorful, reactive and dangerous nonmetals
- **Group 18: Noble Gases** – safe, stable, and almost unreactive

Group	Name	Elements	Physical Properties	Chemical Properties
1	Alkali Metals	Li, Na, K, Rb, Cs, Fr	Soft, low density, shiny (tarnish quickly), good conductors	Very reactive, especially with water; form $+1$ ions; reactivity $\uparrow$ down group
2	Alkaline Earth Metals	Be, Mg, Ca, Sr, Ba, Ra	Harder, denser, higher melting points than Group 1	Reactive (less than Group 1); form $+2$ ions; basic oxides/hydroxides
17	Halogens	F, Cl, Br, I, At, Ts	Nonmetals; gases (F_2 , Cl_2), liquid (Br_2), solid (I_2); distinct colors	Very reactive nonmetals; form -1 ions; reactivity $\downarrow$ down group; strong oxidizers
18	Noble Gases	He, Ne, Ar, Kr, Xe, Rn, Og	Colorless, odorless gases; very low boiling points; glow when excited	Very unreactive (inert); full valence shells; only heavier ones form rare compounds

Understanding the Properties Using Electron Configurations

In order to fully understand why these groups have their properties, we should take a look at electron configurations of their atoms and the electron configurations of the ions they form when they react. Complete the chart below, filling in the electron configurations for the ions

Atom	Atom Configuration	Ion	Ion Charge	Ion Configuration	Ionic Dot
Na	2-8-1	Na ⁺	+1		
Mg	2-8-2	Mg ²⁺	+2		
Al	2-8-3	Al ³⁺	+3		
Cl	2-8-7	Cl ⁻	-1		
O	2-6	O ²⁻	-2		
S	2-8-6	S ²⁻	-2		
K	2-8-8-1	K ⁺	+1		
Ca	2-8-8-2	Ca ²⁺	+2		
Ne	2-8	n/a	n/a		

Observations & Patterns

- The metals tend to _____ electrons and form _____ charged ions
- The nonmetals tend to _____ electrons and form _____ charged ions
- Both metals and nonmetals tend to end up with _____ valence electrons when they become ions.

Reflection Question

1. Inertness of Noble Gases:

- Explain why **Group 18 (noble gases)** are generally **unreactive**, using **valence electron configurations**.

Periodic Trends

Patterns of the Periodic Table and Isoelectronic Series

We are going to examine the **periodic table** and discover the patterns that result from its organization. The word “**periodic**” means “**appearing or occurring at regular intervals**”, and we will see how this applies to the elements.

To understand these patterns, we need to recall:

- **Bohr models** and energy levels
- How **opposite charges interact**

From your reference table, the **force of attraction** depends on:

- The **charges** of the particles (q_1 and q_2)
- The **distance squared** (d^2) between them

$F \propto \frac{q_1 q_2}{d^2}$ *qualitative analysis	F = coulombic force q = charge d = distance of separation
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Key points:

- Greater charges → _____ attraction
- Greater distance → _____ attraction

In atoms:

- **Protons (+)** and **electrons (-)** are attracted to each other because they have _____ charges.
- More protons or electrons, or closer distances, → _____ attraction

Lets take a look at “**Isoelectronic series**”. Iso means ‘the _____’ and ‘electronic’ refers to the electron configuration. An Isoelectronic series is a series of atoms and/or ions that have the same number of electrons, and therefore, the same electron configuration. Let's look at the following isoelectronic series

Isoelectronic Series with 10 electrons: Atomic radius vs. Ionic Radius.



You can see how the size of the atoms change as they become ions; the **nonmetals that gain electrons** become bigger and the **metals that lose electrons get smaller**. All of these

ions have the same number of electrons, but the number of PROTONS and therefore its atomic numbers are different. Below is a table that lists their atomic numbers and their radii.

Ion	Radius (pm)	Atomic Number
N^{3-}	146	7
O^{2-}	140	8
F^{-}	133	9
Na^{+}	98	11
Mg^{2+}	79	12
Al^{3+}	57	13

Reflection questions:

1. Explain, in terms of **protons, electrons, energy levels, and forces of attraction**, why the **radius gets smaller** as the atomic number increases in this isoelectronic series.
2. Why can atoms become **heavier** (more protons and neutrons) and **smaller** at the same time?

What Patterns Do You Notice?

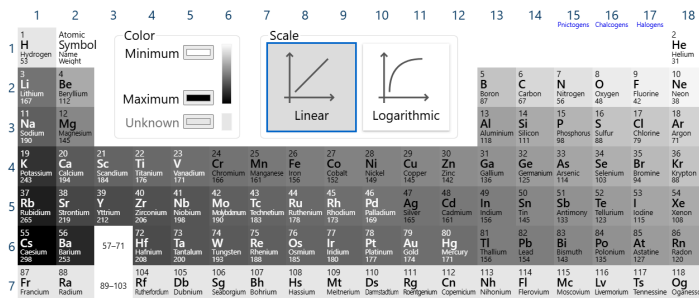
The periodic table is called “periodic” because many properties of elements change in predictable patterns.

Using the color-coded periodic tables provided, identify the general trend for each property.

Atomic Radius

Atomic radius describes the _____ of an atom.

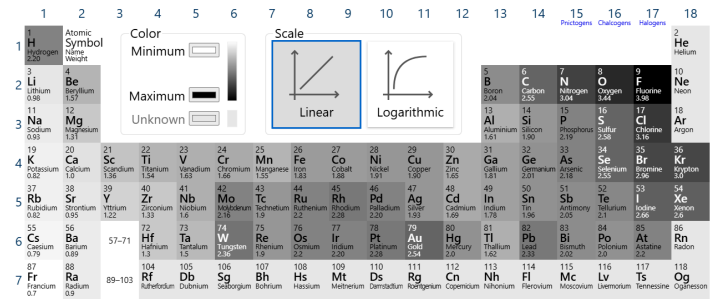
- Across a period: Left → Right
 - Atomic radius generally _____.
- Down a group: Top ↓ Bottom
 - Atomic radius generally _____.



Electronegativity

Electronegativity describes an atom's ability to _____ electrons when it is bonded to another atom.

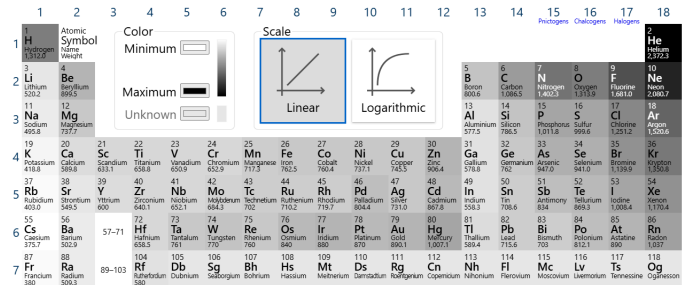
- Across a period: Left → Right
 - Electronegativity generally _____.
- Down a group: Top ↓ Bottom
 - Electronegativity generally _____.



Ionization Energy

Ionization energy is the amount of energy required to _____ an electron from an atom.

- Across a period: Left → Right
 - Ionization energy generally _____.
- Down a group: Top ↓ Bottom
 - Ionization energy generally _____.



Why? Explaining Periodic Trends

We have already identified the patterns in the periodic table. Now let's look at **WHY** these trends occur.

The Big Idea: Attraction

- The nucleus contains positively charged _____.
- Electrons have a _____ charge.
- Opposite charges _____ each other.

Therefore, the positively charged nucleus _____ the negatively charged electrons. The strength of this attraction is important in explaining periodic trends.

Two things we need to consider are:

1. **Number of protons** — More protons means a _____ positive charge in the nucleus and a stronger pull on electrons.
2. **Principal energy levels (PELs)** — Electrons in higher PELs are _____ from the nucleus.

Across a Period →

As you move from **left to right across a period**:

- Number of protons: _____
- Number of occupied PELs: _____
- **More protons + same number of PELs =** _____ **attraction**

What does that mean for;

- **Atomic Radius:** attractive force increases, so the electrons will be _____ attracted to the nucleus and pulled in closer, making the atomic radius _____ left to right
- **Electronegativity:** Attractive force increases, so how strongly it pulls on electrons when bonding will _____.
- **Ionization energy:** Attractive force increases, so how difficult it is to remove an electron from an atom will _____

Down a Group ↓

As you move from **top to bottom down a group**:

- Number of protons: _____
- Number of occupied PELs: _____
- **More PELs, Further from the nucleus =** _____ **attraction**

What does that mean for;

- **Atomic Radius:** attractive force decreases, so the electrons will be _____ attracted to the nucleus and further away from the nucleus, making the atomic radius _____ top to bottom
- **Electronegativity:** Attractive force Decreases, so how strongly it pulls on electrons when bonding will _____.
- **Ionization energy:** Attractive force decreases, so how difficult it is to remove an electron from an atom will _____

Summary Table: use arrows to indicate what direction each property increases

Ionization
Energy ↑

Electro-
negativity ↑

Atomic
Radius ↑

hydrogen 1 H 1.0079																	helium 2 He 4.0026						
lithium 3 Li 6.941	beryllium 4 Be 9.0122																	boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180
sodium 11 Na 22.990	magnesium 12 Mg 24.305																	aluminum 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
potassium 19 K 39.098	calcium 20 Ca 40.078	scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80						
rubidium 37 Rb 85.468	strontium 38 Sr 87.62	yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29						
cesium 55 Cs 132.91	barium 56 Ba 137.33	57-70 * lanthanum 57 La 138.91		hafnium 71 Hf 178.49	tantalum 72 Ta 180.95	tungsten 73 W 183.84	rhenium 75 Re 186.21	osmium 76 Os 190.23	iridium 77 Ir 192.22	platinum 78 Pt 195.08	gold 79 Au 196.97	mercury 80 Hg 200.59	thallium 81 Tl 204.38	lead 82 Pb 207.2	bismuth 83 Bi 208.98	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]					
francium 87 Fr [223]	radium 88 Ra [226]	89-102 * * actinium 89 Ac [227]		hafnium 103 Hf [263]	tantalum 104 Ta [263]	tungsten 105 W [263]	rhenium 106 Re [263]	osmium 107 Os [263]	iridium 108 Ir [263]	platinum 109 Pt [263]	gold 110 Au [263]	mercury 111 Hg [263]	thallium 112 Tl [263]	lead 113 Pb [263]	bismuth 114 Bi [263]	polonium 115 Po [263]	astatine 116 At [263]	radon 117 Rn [263]					

* Lanthanide series

lanthanum 57 La 138.91	cerium 58 Ce 140.12	praseodymium 59 Pr 140.91	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.96	gadolinium 64 Gd 157.25	terbium 65 Tb 158.93	dysprosium 66 Dy 162.50	holmium 67 Ho 164.93	erbium 68 Er 167.26	thulium 69 Tm 168.93	ytterbium 70 Yb 173.04
actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	eskeium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

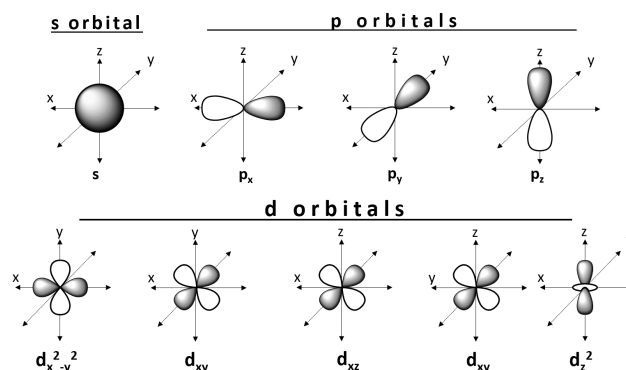
HONORS EXTENSION: Why Bigger Energy Levels Hold More Electrons & Quantum Configurations

1. Energy Levels and Sublevels

- Each main energy level (1, 2, 3, 4, ...) can be divided into **sublevels** labeled **s, p, d, f**:
 - 1st energy level → s only
 - 2nd energy level → s and p
 - 3rd energy level → s, p, and d
 - 4th energy level → s, p, d, and f
- Orbitals:** Each sublevel contains orbitals, and each orbital can hold **2 electrons**:
 - s = 1 orbital → _____ electrons
 - p = 3 orbitals → _____ electrons
 - d = 5 orbitals → _____ electrons
 - f = 7 orbitals → _____ electrons
- Total electrons per energy level:
 - 1st energy level: s only → _____ electrons
 - 2nd: s + p → _____ + _____ = _____ electrons
 - 3rd: s + p + d → _____ + _____ + _____ = _____ electrons
 - 4th: s + p + d + f → _____ + _____ + _____ + _____ = _____ electrons
- Conclusion:** Bigger energy levels hold _____ electrons because they contain more _____, which increases the number of _____, and therefore the number of electrons that can fit.

2. What is an Orbital?

- An **orbital** is a region around the nucleus where there is a high probability of finding an electron.
- Think of it as a “**home**” or “**parking space**” for up to 2 electrons.
- Shapes of orbitals depend on the sublevel:
 - s → sphere-shaped (1 orbital)
 - p → dumbbell-shaped (3 orientations: x, y, z)
 - d → more complex (5 orientations)
 - f → even more complex (7 orientations)
- Task:** Sketch the shapes of **s and p orbitals** next to the s- and p-block labels on your reference table.



3. Quantum Electron Configurations

- A **quantum configuration** shows how electrons are arranged in an atom's orbitals.
- Format: **[principal energy level] [sublevel] ^ [number of electrons]**

Parts of a configuration (example: $1s^2$):

- **1** → Principal energy level ($n = 1, 2, 3 \dots$)
- **s** → Sublevel (s, p, d, f)
 - s → 1 orbital → 2 electrons max
 - p → 3 orbitals → 6 electrons max
 - d → 5 orbitals → 10 electrons max
 - f → 7 orbitals → 14 electrons max
- **²** → Superscript = number of electrons in that sublevel

Example: Carbon → $1s^2 2s^2 2p^2$

- $1s^2$ → 2 electrons in 1st energy level, s sublevel
- $2s^2$ → 2 electrons in 2nd energy level, s sublevel
- $2p^2$ → 2 electrons in 2nd energy level, p sublevel

4. The Aufbau Principle

- **Definition:** Electrons fill orbitals starting with the _____ before moving to higher energy.
- **Sequence:** $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \dots$
 - Some higher-energy orbitals (like 4s) fill _____ lower ones (like 3d) due to energy overlap.

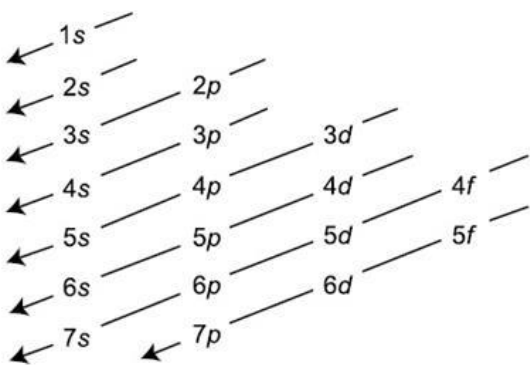
Analogy – Grocery Store Checkout:

- Shoppers (electrons) always pick the **shortest line (lowest-energy orbital)** first.
- Even if another line is “higher-numbered” (like 4s), it may be faster than a crowded lower line (3d).
- **Task:** Use the **Aufbau Principle diagram** to:
 1. Fill in the superscripts for each sublevel (maximum electrons)
 2. Complete the blank periodic table with orbitals being filled (1s, 2s, 2p, etc.)
 3. Label the **s-block**, **p-block**, **d-block**, and **f-block** regions

5. Reflection

- **Question:** How does the Aufbau Principle help explain the unusual shape of the periodic table?
- Write a short answer in your handout based on the orbital filling patterns and the block structure.

Aufbau Principle diagram



1s ¹																		1s ²	
2s ¹	2s ²											2p ¹							

1. The periodic table is divided into regions known as the **s-block**, **p-block**, **d-block**, and **f-block**. Based on your completed table, label these regions accordingly.
2. Reflect on the layout of the periodic table: **How does the Aufbau Principle help explain the unusual shape of the periodic table?**