Classification of Elements and Periodicity in Properties

🗺️ Chapter Blueprint (What we need to conquer)

We will break this chapter down into **5 manageable milestones**.

- **Milestone 1: The Basics & Nomenclature**
  - Modern Periodic Law vs. Mendeleev's Law.
  - IUPAC naming for elements with Atomic Number (Z) > 100 (e.g., Unnilennium). *This guarantees a 1-mark question!*
- **Milestone 2: Electronic Configuration & Locating Elements**
  - Understanding s, p, d, and f blocks.
  - How to find the Group number, Period number, and Block of any element just by looking at its atomic number.
- **Milestone 3: Physical Properties & Trends (The Core of the Chapter)**
  - Atomic and Ionic Radii (and the magic rule for Isoelectronic species).
  - Ionization Enthalpy ($\Delta_i H$) - *We will focus heavily on the exceptions here (like why Nitrogen has higher IE than Oxygen, or why Beryllium is higher than Boron).*
  - Electron Gain Enthalpy ($\Delta_{eg} H$) - *The classic Fluorine vs. Chlorine trap.*
  - Electronegativity.
- **Milestone 4: Chemical Properties & Trends**
  - Valency/Oxidation states.
  - Metallic vs. Non-metallic character.
  - Nature of Oxides (Basic, Acidic, Amphoteric).
  - Anomalous behavior of 2nd-period elements and Diagonal Relationships.
- **Milestone 5: Advanced Problem Solving**
  - Case-based / Passage-based questions (like guessing the element based on its successive ionization enthalpies, which carries 4-5 marks in your provided PDFs).

💡 How we will work together:

1. **I will explain the concept** in the simplest terms possible, with tricks to remember them.
2. **I will test you** with a few basic questions to ensure it clicked.
3. **We will solve the exact questions from your PDF** to lock in the exam-level confidence.

Milestone 1: The Basics & IUPAC Nomenclature (Z > 100).

Part A: The Two Big Laws

Before we jump into the math and logic, you need to know the historical shift that made the Modern Periodic Table possible. It's super simple:

4. **Mendeleev's Periodic Law:** He arranged elements based on their **Atomic MASS**.
   - *The flaw:* Isotopes (atoms of the same element with different masses) ruined his table.
5. **Modern Periodic Law (Henry Moseley):** He proved that the physical and chemical properties of elements are periodic functions of their **Atomic NUMBERS**.
   - *Why is this brilliant?* Because the Atomic Number tells us the number of electrons, and electrons dictate how an element behaves chemically!

Part B: Naming Elements with Z > 100 (The IUPAC Cheat Code)

When scientists started discovering heavy elements (Atomic Number > 100), they used to fight over what to name them. So, IUPAC made a strict mathematical rule. All you have to do is memorize these roots for numbers 0 to 9:

- **0** = nil (n)
- **1** = un (u)
- **2** = bi (b)
- **3** = tri (t)
- **4** = quad (q)
- **5** = pent (p)
- **6** = hex (h)
- **7** = sept (s)
- **8** = oct (o)
- **9** = enn (e)

The Rule: Just write the root for each digit in order, and add **"-ium"** at the very end. *(Note: If the last digit's root ends in 'i' like 'bi' or 'tri', don't write double 'i'. Just write '-ium' once. Example: bi + ium = bium).*

Example: Atomic Number 104

- 1 = un
- 0 = nil
- 4 = quad
- *Name:* un + nil + quad + ium = **unnilquadium**
- *Symbol:* Take the first letter of each root: **Unq**

🧠 Your Turn! (Testing your knowledge)

These questions are picked directly from the exam PDFs you provided. Grab a piece of paper, use the cheat code above, and tell me the answers to these 3 questions:

Question 1: As per the IUPAC, the element with atomic no. **119** would be named as: (a) Ununseptium (b) Ununennium (c) Unniloctium (d) Unnilseptium

Question 2: The symbol of an element with atomic number **120** as per IUPAC will be: (a) unb (b) unt (c) ubn (d) uun

Question 3: Write the full IUPAC names for the elements with atomic numbers **112** and **123**.

Question 1: So, 1 is "un" and 9 is "enn." Since it's 119, it's like un + un + enn and then you add "ium" at the end. So that would be **(b) Ununennium**. (I remember the 'enn' has two n's, so I hope I got that right).

Question 2: For 120, I have to look at the numbers again. 1 is "u" (un), 2 is "b" (bi), and 0 is "n" (nil). If you put the first letters together, you get U-B-N. So the answer is **(c) ubn**.

Question 3: Okay, for these ones I just have to write the whole thing out:

- **112:** That's 1 (un), 1 (un), and 2 (bi). So it's **Ununbium**. (The teacher said if it ends in 'bi' you don't put two 'i's, so I just put one).
- **123:** That's 1 (un), 2 (bi), and 3 (tri). So it's **Unbitrium**. (Same thing here, just one 'i' at the end).

🚀 Milestone 2: Locating Elements (Group, Period, Block)

This is a classic exam question: *"An element has atomic number X. Find its group, period, and block."*

To do this, you just need to write the electronic configuration. Here is the ultimate step-by-step cheat sheet:

1. Finding the Period Number:

- **The Rule:** The Period number is simply the **highest shell number (n)** in the electronic configuration.
- *Example:* If the configuration ends in $3s^2 3p^5$, the highest number in front of the letters is **3**. So, it's in **Period 3**.

2. Finding the Block (s, p, d, or f):

- **The Rule:** Look at where the **last electron** goes.
- *Example:* In $1s^2 2s^2 2p^3$, the last electron entered the **p** subshell. So, it is a **p-block** element.

3. Finding the Group Number (The only part that requires a tiny bit of math): First, find the block, then apply the matching formula:

- **If it's an s-block element:** Group = Number of valence s electrons (It will always be Group 1 or 2).
- **If it's a p-block element:** Group = 10 + (Number of valence s and p electrons).
- **If it's a d-block element:** Group = (electrons in the ns orbital) + (electrons in the $(n-1)d$ orbital).

🎩 The "Magic Numbers" Shortcut Trick

If you don't want to write the long configuration for huge atomic numbers (like Z = 114), memorize the atomic numbers of the Noble Gases (Group 18): **2 (He), 10 (Ne), 18 (Ar), 36 (Kr), 54 (Xe), 86 (Rn), 118 (Og)**

- *How to use it:* If they ask for Z = 12. You know the closest noble gas *before* 12 is Neon (10).
    - Neon is at the end of Period 2. So, 11 and 12 must start **Period 3**.
    - Since 12 is two steps after 10, it lands in **Group 2**. Simple!

🧠 Your Turn! (Let's test this)

Here are 3 actual questions from your PDF. Use the rules or the shortcut trick to solve them:

Question 1 (From Page 4, Q6): An element has the atomic number **19**. (a) Write its electronic configuration (Hint: nearest noble gas is Ar-18). (b) Tell me its Block, Period, and Group.

Question 2 (From Page 6, Q1b): In terms of the period and group, where would you locate the element with **Z = 114**? *(Hint: Use the Magic Numbers trick! 114 is very close to the last noble gas 118).*

Question 3 (From Page 5, Q10c): Determine the position (Block, Period, and Group) of an element in the periodic table whose atomic number is **32**.

Question 1: (a) Okay, so atomic number is 19. The hint says Argon (Ar) is 18. So I just need to add one more electron after Ar. The next shell after 3p is 4s, right? So it's just **[Ar] $4s^1$**. (Wait, I should probably write it out long-way just in case: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$).

- **Block:** The last electron I wrote went into the 's' thingy, so it's **s-block**.
- **Period:** The biggest number in front of the letters is 4, so it's **Period 4**.
- **Group:** There's only 1 electron in that last 4s part, so it's just **Group 1**.

Question 2: Z = 114. The hint says 118 is a noble gas. My teacher said noble gases are always in Group 18. So if 118 is Group 18... then:

- 117 must be Group 17
- 116 must be Group 16
- 115 must be Group 15
- **114 must be Group 14** And since 118 is the end of the 7th row, 114 is also in **Period 7**. (I think all those super heavy ones are in Period 7).

Question 3: Z = 32. I have to do the configuration again... let me count on my fingers... $1s^2$ (2), $2s^2$ (4), $2p^6$ (10), $3s^2$ (12), $3p^6$ (18), $4s^2$ (20), $3d^{10}$ (30)... okay, I need 2 more to get to 32. So $4p^2$.

- **Block:** The very last electron went into 'p', so it's **p-block**.
- **Period:** The highest shell number I used was 4, so it's **Period 4**.
- **Group:** For p-block, my notes say you take the electrons in the outer shell and add 10? Or maybe it's $12 + p$ electrons? Wait... $10 + (2 \text{ from } 4s) + (2 \text{ from } 4p) = 14$. Yeah, **Group 14**. (Wait, is that the same group as the other one? I guess so).

🔥 Milestone 3: Physical Properties & Trends (The Core of the Chapter)

This is where the real "Chemistry Logic" starts. If you understand the "why" here, you'll never have to memorize the trends. We will split this milestone into chunks. Let's start with the first chunk: **Atomic & Ionic Radii**.

Concept 1: Atomic Radius (Size of the Atom)

Imagine the nucleus as a giant magnet pulling on the outer electrons.

- **Down a Group (Top to Bottom): Size INCREASES.**
  - *Why?* Because you are adding completely new shells (like layers of an onion). More shells = bigger atom. The inner electrons also "shield" the outer ones from the nucleus's pull.
- **Across a Period (Left to Right): Size DECREASES.**
  - *Why?* The number of shells stays the same, but you keep adding more protons to the nucleus. More protons = a stronger magnet pulling the same electron shell closer to the center. This stronger pull is called **Effective Nuclear Charge (**Z_{eff}**)**.

Concept 2: Ionic Radius (Cations vs. Anions) - *Super Important for Exams!*

- **Cation (Positive Ion, e.g.,** Na^+**):** It loses an electron. Now there are more protons than electrons. The nucleus pulls the remaining electrons *much tighter*. Sometimes it even loses an entire shell!
  - **Rule:** A Cation is ALWAYS **smaller** than its parent atom ($Na^+ < Na$).
- **Anion (Negative Ion, e.g.,** Cl^-**):** It gains an electron. Now there are more electrons than protons. The electrons repel each other (push each other away), making the electron cloud expand.
  - **Rule:** An Anion is ALWAYS **larger** than its parent atom ($Cl^- > Cl$).

Concept 3: ISOELECTRONIC SPECIES (The Ultimate Exam Favorite)

"Iso" means same. Isoelectronic species are atoms/ions that have the **exact same number of electrons**.

- *Trick to solve:* Since electrons are the same, just look at the protons (Atomic Number). The one with the **most protons** pulls the hardest, so it will be the **smallest**.
- *Example:* Na^+ (11 protons, 10 electrons) vs F^- (9 protons, 10 electrons). Na^+ has a stronger nucleus (11 protons), so Na^+ is smaller than F^-.

🧠 Your Turn! (Let's apply the logic)

These are heavily tested questions straight from your PDFs.

Question 1 (From Page 4, Q4): Explain why are cations smaller and anions larger in size than their parent atoms? *(Use the logic we just talked about!)*

Question 2 (From Page 4, Q5): Consider the following species: N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+} and Al^{3+} (a) What is common in them? (b) Arrange them in the order of increasing ionic radii (from smallest to largest).

Question 3 (From Page 1, Q8): Consider the isoelectronic series: K^{+}, S^{2-}, Cl^{-}, and Ca^{2+}. The radii of the ions decrease as: (a) $Ca^{2+} < K^{+} > Cl^{-} > S^{2-}$ (b) $Cl^{-} > S^{2-} > K^{+} > Ca^{2+}$ (c) $S^{2-} > Cl^{-} > K^{+} > Ca^{2+}$ (d) $K^{+} > Ca^{2+} > S^{2-} > Cl^{-}$ *(Hint: Arrange them from biggest radius to smallest radius. Who has the weakest pull?)*

Ugh, my head hurts from all these ions, but I think I remember the "pulling" rule my teacher told us. Let me try...

Question 1: Okay, so for **cations**, they lose electrons, right? Like, if you lose an electron, there are fewer electrons for the protons to pull on. So the nucleus (the protons) can pull the leftover electrons in much tighter. Sometimes they even lose a whole shell, so they get way smaller. For **anions**, it's the opposite. They gain extra electrons. Because there are more electrons now, they all start pushing each other away more (my teacher called it repulsion), and the protons can't pull them all in as well. So the "cloud" spreads out and gets bigger.

Question 2: (a) Let me count the electrons for all of them:

- N^{3-}: 7 + 3 = 10
- O^{2-}: 8 + 2 = 10
- F^{-}: 9 + 1 = 10
- Na^{+}: 11 - 1 = 10
- Mg^{2+}: 12 - 2 = 10
- Al^{3+}: 13 - 3 = 10 Wait, they all have **10 electrons**. My teacher called this "isoelectronic" or something like that.

(b) To order them, I just look at the protons (the atomic number). The one with the most protons pulls the hardest and makes the ion the smallest.

- Al has 13 protons (strongest pull $\rightarrow$ smallest)
- Mg has 12
- Na has 11
- F has 9
- O has 8
- N has 7 protons (weakest pull $\rightarrow$ biggest) So, from smallest to largest: $Al^{3+} < Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$

Question 3: This is like the one I just did. They all have 18 electrons (I checked the periodic table: K is 19, S is 16, Cl is 17, Ca is 20).

- Ca^{2+} has 20 protons (Smallest because 20 protons pull 18 electrons really hard).
- K^+ has 19 protons.
- Cl^- has 17 protons.
- S^{2-} has 16 protons (Biggest because 16 protons can't pull 18 electrons very well).

So the order from biggest to smallest (decreasing) would be $S^{2-} > Cl^- > K^+ > Ca^{2+}$. That matches **(c)**. (I have to be careful with the $>$ signs, they always trip me up!)

Teachers *obsess* over this topic because it has rules, but more importantly, it has **exceptions**. And exams are built on exceptions!

🔥 Milestone 3, Part 2: Ionization Enthalpy (The Art of Stealing Electrons)

The Concept: Ionization Enthalpy is simply the **energy required to steal (remove) the outermost electron** from an isolated atom.

- If the atom holds its electron tightly $\rightarrow$ It takes a LOT of energy to steal it (High IE).
- If the atom is big and the electron is far away $\rightarrow$ It's easy to steal (Low IE).

The General Rule:

- **Down a Group:** Size gets bigger. The outer electron is far from the nucleus. It's easy to steal. $\rightarrow$ **IE Decreases.**
- **Across a Period:** Size gets smaller. The nucleus pulls harder. It's tough to steal. $\rightarrow$ **IE Increases.**

🚨 THE EXCEPTIONS (Exam Goldmine! 🚨)

Across a period, IE is *supposed* to always increase. But there are two major speedbumps where the trend breaks because of **Electronic Configuration Stability**.

Exception 1: Group 2 vs. Group 13 (The 's' vs 'p' battle)

- *Example:* Beryllium (Be, Z=4) vs. Boron (B, Z=5).
- Normally, B should have a higher IE than Be. But **Be > B**.
- *Why?* Write the configuration!
  - Be = $1s^2\mathbf{2s^2}$. The $2s$ orbital is completely full and closer to the nucleus. It is very stable.
  - B = $1s^22s^2\mathbf{2p^1}$. It is much easier to pluck off that single, lonely $2p$ electron than to break a full $2s$ pair.
  - *Same applies to Magnesium (Mg) > Aluminum (Al).*

Exception 2: Group 15 vs. Group 16 (The Half-Filled Rule)

- *Example:* Nitrogen (N, Z=7) vs. Oxygen (O, Z=8).
- Normally, O should have a higher IE than N. But **N > O**.
- *Why?* Write the configuration!
  - N = $1s^22s^2\mathbf{2p^3}$. The p-orbital holds 6 electrons. So, p^3 is **exactly half-filled**. Half-filled orbitals are uniquely stable and symmetrical. Nitrogen refuses to give up an electron.

- O = $1s^2 2s^2 \mathbf{2p^4}$. Oxygen has one extra paired electron in the p-orbital. Those paired electrons repel each other, so Oxygen is actually *happy* to lose one to become stable like Nitrogen!
- *Same applies to Phosphorus (P) > Sulfur (S).*

🧠 Your Turn! (The Ultimate Exception Test)

These questions are copied word-for-word from your PDF (Pages 1, 4, and 8). If you can answer these, you will secure at least 3-4 marks on your paper.

Question 1 (From Page 4, Q8): Among the second period elements, why does: (i) Be have a higher first ionization enthalpy ($\Delta_i H_1$) than B? (ii) N have a higher first ionization enthalpy than O? *(Hint: Just explain what we discussed above using their configurations!)*

Question 2 (From Page 4, Q1 & Q9a - A beautiful logic puzzle): The *first* ionization enthalpy of Magnesium (Mg) is higher than Sodium (Na). BUT the *second* ionization enthalpy of Sodium (Na) is VERY MUCH higher than Magnesium (Mg). Explain why. *(Hint: Write the configuration for Na (11) and Mg (12). What happens after you steal the FIRST electron from both? Look at what they turn into!)*

Question 3 (From Page 1, Q6): The correct order of first ionization potential among Be, B, C, N, and O is: (a) B < Be < C < O < N (b) B < Be < C < N < O (c) Be < B < C < N < O (d) Be < B < C < O < N *(Hint: Put them in normal left-to-right order first, then swap the exceptions!)*

Oh man, ionization enthalpy is the one about how much energy it takes to "steal" an electron, right? My teacher kept saying "stable configurations," so I'll try to use that.

Question 1: (i) **Be vs B:** I wrote their configurations in my rough notes. Be is $1s^2 2s^2$ and B is $1s^2 2s^2 2p^1$. In Be, the $2s$ orbital is completely full, which is like a "stable" setup, so it doesn't want to give up an electron. But in B, that extra electron is just sitting in a $2p$ orbital all by itself, and it's a bit further away from the nucleus, so it's easier to snatch it away.

(ii) **N vs O:** This one is the "half-filled" rule. N is $1s^2 2s^2 2p^3$. Since the p-shell has three boxes and N has one electron in each, it's half-full and very happy/stable. O is $1s^2 2s^2 2p^4$. That fourth electron has to share a box with another electron, and they start pushing each other away (repulsion). Because they are already pushing each other, it takes less energy to pull one out of O than it does to break N's stable half-shell.

Question 2: Okay, I like these logic puzzles. Let's look at what happens:

- **First electron:** Mg ($3s^2$) has more protons than Na ($3s^1$), so it pulls harder. Also, Mg's s-shell is full. So, the first jump for Mg is harder.
- **Second electron:**

  - If you take one from Na, it becomes Na^+, which has the configuration of **Neon** ($2p^6$). It's a noble gas now! It's super happy and "locking the doors"—it will take a *massive* amount of energy to steal another one.
  - If you take one from Mg, it becomes Mg^+, which is $[Ne]3s^1$. It has one lonely electron left. It actually *wants* to lose that electron so it can also become like Neon. So, taking the second electron from Na is like trying to rob a bank, but taking it from Mg is much easier.

Question 3: The teacher told us that usually, the energy goes up as you go across the row (Left to Right). So normally it would be: Be < B < C < N < O. But then we have those two "swaps" or exceptions I just explained in Question 1:

1. **B is lower than Be** (because of the $2p^1$ thing).
2. **O is lower than N** (because of the half-filled p^3 thing).

So, if I start with the lowest and swap them into the line: It goes **B** first, then **Be**, then **C**, then **O**, then **N** is the highest. Checking the list... that's **(a) B < Be < C < O < N**. (I hope I didn't mix up the "greater than" signs again!)

This is where examiners lay their favorite trap. But after reading this, you won't fall for it.

🧲 Milestone 3, Part 3: The "Catching" of Electrons

Concept 1: Electron Gain Enthalpy ($\Delta_{eg}H$)

If Ionization Enthalpy is the energy *required* to steal an electron, Electron Gain Enthalpy is the energy **released** when an atom *accepts* an electron.

- Because energy is released, the value is usually **negative** (exothermic). The more an atom "wants" an electron, the more negative the value.
- **Across a Period:** Atoms get closer to a full octet, so they want electrons more. $\rightarrow$ $\Delta_{eg}H$ becomes **more negative**.
- **Down a Group:** Atoms get bigger, the nucleus is far away, so it doesn't pull in new electrons as strongly. $\rightarrow$ $\Delta_{eg}H$ becomes **less negative**.

🚨 **THE ULTIMATE EXAM TRAP: Fluorine vs. Chlorine** 🚨 Normally, since Fluorine is at the top right, it should have the most negative $\Delta_{eg}H$. **BUT IT DOESN'T. Chlorine does.**

- *Why?* Fluorine is **too small**. It has 7 electrons crammed into a tiny $2p$ shell. If you try to shove one more electron in there, the existing electrons scream "NO ROOM!" and repel it.
- Chlorine, however, has a larger $3p$ shell. It has plenty of space to comfortably welcome the new electron.
- *Conclusion:* **Chlorine (Cl)** has the highest negative electron gain enthalpy in the whole periodic table!

Concept 2: Electronegativity (EN)

This sounds similar, but it's different! Electronegativity is the power of an atom **in a chemical bond** to pull the *shared* electrons toward itself. It's like a tug-of-war game between two bonded atoms.

- **Trend:** Increases Left to Right. Decreases Top to Bottom.
- **The King of EN:** Because it's about pulling power (and not about having room to *store* an extra electron), **Fluorine (F)** is the undisputed champion of Electronegativity.

Summary of the Trap:

- Highest Electronegativity = **Fluorine (F)**
- Most Negative Electron Gain Enthalpy = **Chlorine (Cl)**

🧠 Your Turn! (Dodging the Traps)

Here are 3 questions straight from your PDFs (Pages 2, 3, and 5). Let's see you dodge the traps!

Question 1 (From Page 5, Assertion/Reason Q11): *Assertion (A):* Electron gain enthalpy of Fluorine is more negative than that of Chlorine. *Reason (R):* Size of Fluorine atom is very small. (a) Both A and R are correct, R is correct explanation. (b) Both A and R are correct, R is not the correct explanation. (c) A is correct, R is incorrect. (d) A is incorrect, R is correct.

Question 2 (From Page 3, Q8): Noble gases have zero (or slightly positive) electron gain enthalpy values. Explain why. *(Hint: Think about their electron configuration. Do they want a new electron? Where would it even go?)*

Question 3 (From Page 3, Q2): What is the basic difference between electron gain enthalpy and electronegativity? *(Hint: One is about isolated atoms taking a whole electron, the other is about a tug-of-war in a bond. Put it in your own words!)*

Ugh, electron gain enthalpy... this is the one where the energy is usually negative because the atom *releases* energy when it grabs an electron, right? My teacher said it's like a "happiness scale" for getting a gift. Let me try.

Question 1: Okay, I remember this one because it's a trick! Even though Fluorine is at the top of the group, **Chlorine** actually has a more negative electron gain enthalpy. My teacher said it's because Fluorine is so tiny that if you try to shove another electron in there, the electrons that are already there get really "grumpy" and push it away (repulsion).

- So, Assertion (A) is **wrong** (Chlorine is more negative).
- But Reason (R) is **true** (Fluorine *is* really small, that's actually why it's not as negative). The answer is **(d) A is incorrect, R is correct.**

Question 2: Noble gases are like the "rich kids" of the periodic table—they already have everything they need. Their outer shells are completely full (like $2s^2 2p^6$). If you try to give them another electron, they don't have any room left in their current shell. They would have to start a **whole new shell** (like going from $n = 2$ to $n = 3$), which takes a ton of energy. They don't want to do that, so they don't "gain" energy by taking an electron; you'd actually have to force it on them. That's why it's zero or positive.

Question 3: I think of it like this:

- **Electron Gain Enthalpy:** This is for a single, lonely atom (isolated). It's like the atom is just standing there and someone hands it an electron, and we measure how much energy it gives off. It's an actual number you can measure in a lab.
- **Electronegativity:** This is more like a **tug-of-war**. This only happens when two atoms are already in a "relationship" (a bond). It's not about taking the electron

completely; it’s just about which atom pulls the shared electrons closer to its side. It doesn't have a unit like kJ/mol, it’s just a scale from 0 to 4.

🧪 Milestone 4: Chemical Properties & Trends

Now we look at how these elements actually behave in the real world when they react. Since you already know the physical trends (size, pulling power), this will feel like a breeze.

Concept 1: Metallic vs. Non-Metallic Character

- **What makes a good Metal?** A good metal *wants to lose* electrons easily to form positive ions (cations). So, atoms with **low Ionization Enthalpy** are the best metals.
  - *Trend:* Metallic character **increases down a group** (electrons are easier to lose) and **decreases across a period** (nucleus pulls too hard).
  - *The King of Metals:* Francium/Cesium (Bottom Left).
- **What makes a good Non-Metal?** A good non-metal *wants to gain* electrons to form negative ions (anions). So, atoms with **high Electronegativity** are the best non-metals.
  - *Trend:* Non-metallic character **increases across a period** and **decreases down a group**.
  - *The King of Non-Metals:* Fluorine (Top Right).

Concept 2: Nature of Oxides (The Golden Rule)

Elements react with Oxygen to form oxides, but they don't all act the same when you drop them in water:

- **M**etal Oxides are **B**asic (Think **M-B**: Mother Board). Example: Na_2O, MgO.
- **N**on-Metal Oxides are **A**cidic (Think **N-A**: North America). Example: CO_2, SO_2.
- *Trend:* As you go left to right across the periodic table, oxides go from highly Basic $\rightarrow$ Amphoteric (both) $\rightarrow$ Acidic.

Concept 3: The Diagonal Relationship (Anomalous Behavior)

The first element of every group in the s and p blocks (like Li, Be, B) acts like a weirdo. It doesn't behave like the rest of its group.

- *Why?* Because it is exceptionally small, has high electronegativity, and **does not have empty d-orbitals** to expand its bonds.
- Instead of acting like its own group, it acts like the element **diagonally below it to the right**.
  - Lithium (Group 1) acts like Magnesium (Group 2).
  - Beryllium (Group 2) acts like Aluminum (Group 13).
- *The Scientific Reason:* They have a very similar **charge-to-size ratio** (also called polarizing power).

🧠 Your Turn! (Testing Chemical Logic)

Here are 3 exact questions from your question bank (Pages 4, 6, and 7). Let's lock in Milestone 4!

Question 1 (From Page 4, Q10b): Arrange the following elements in their **increasing metallic character**: Br, Al, Mg, K *(Hint: Find them on the periodic table mentally. Who wants to lose an electron the most? Put the least metallic first, most metallic last).*

Question 2 (From Page 6, Q1a & Q1c - Diagonal Relationship): (a) Why do Li and Mg show similarity in their properties? (b) The first member of every group shows anomalous (weird) behavior. Why? *(Hint: Use the exact keywords we just talked about!)*

Question 3 (From Page 7, Case Study Q2): Which of the following groups contains metals, non-metals, as well as metalloids? (a) Group 1 (b) Group 17 (c) Group 14 (d) Group 2 *(Hint: Group 1 and 2 are pure metals. Group 17 are pure non-metals. Which group is right in the middle where the "staircase" of the periodic table is?)*

Ugh, okay, last ones. My brain is getting kind of fried, but I think I can finish this.

Question 1: Okay, metallic character means "how much does it want to lose an electron?" The atoms on the far left are the most metallic, and the ones at the bottom are even more metallic. Looking at where they are:

- **Br** is way over on the right (Group 17), so it's a non-metal. It's the least metallic.
- **Al** is in Group 13.
- **Mg** is in Group 2 (further left than Al).
- **K** is in Group 1 (the furthest left). So, from least metallic to most metallic: **Br < Al < Mg < K**. (K is the most metallic because it's a big Group 1 atom and loses its electron super easily).

Question 2: (a) My teacher called this the **diagonal relationship**. Li and Mg are diagonal to each other on the table. They act the same because they ended up having a similar **ionic size** and the same **charge-density** (or charge-to-radius ratio, I think that's the fancy term).

(b) The first members (the "babies" of the group) are always weird because:

1. They are **really small** in size.
2. They have a **high electronegativity** (they are "greedy" for electrons compared to the bigger ones below them).
3. They **don't have d-orbitals**. The bigger atoms in the group have empty d-orbitals they can use for extra bonds, but the first member doesn't have that room.

Question 3:

- Group 1 and Group 2 are the Alkali and Alkaline Earth things, so they are just **metals**.
- Group 17 are the Halogens, which are all **non-metals**.
- **Group 14** is the one. It starts with Carbon (non-metal), then goes to Silicon and Germanium (which are the metalloids on the "staircase"), and then Tin and Lead at

the bottom are metals. So the answer is **(c) Group 14**. (I remember this because it’s the group with Carbon!)

👑 Milestone 5: Advanced Problem Solving (The 5-Mark Data Tables)

In your exams, the biggest 4 or 5-mark question is usually a **data table puzzle**. They will give you hypothetical elements (like A, B, C, or I, II, III) with their Ionization Enthalpies ($\Delta_i H$) and Electron Gain Enthalpies ($\Delta_{eg} H$), and you have to act like a detective to figure out who they are.

Here is your **Cheat Code** for reading these tables:

1. **The Noble Gas (Least Reactive):**
   - Look for the one with the highest First Ionization Enthalpy ($\Delta_i H_1$) because it refuses to lose an electron.
   - Look for a **Positive (+)** Electron Gain Enthalpy because it absolutely does *not* want a new electron.
2. **The Most Reactive Metal (Alkali Metal - Group 1):**
   - Look for the **lowest** $\Delta_i H_1$ (it desperately wants to lose its 1 valence electron).
   - *The Trapdoor Check:* Look at its $\Delta_i H_2$. It should be **HUGE** (like jumping from 500 to 7000) because once it loses that first electron, it becomes a Noble Gas and locks its doors!
3. **The Alkaline Earth Metal (Group 2):**
   - It wants to lose **two** electrons to form things like $MgCl_2$ (which is M^{2+} or MX_2).
   - So, its $\Delta_i H_1$ will be low, and its $\Delta_i H_2$ will be slightly higher but still reasonable (not a massive 10x jump like Group 1).
4. **The Most Reactive Non-Metal (Halogen - Group 17):**
   - Look for the **most negative** Electron Gain Enthalpy ($\Delta_{eg} H$). It is starving for one electron.
   - Its $\Delta_i H_1$ will also be very high because it wants to *gain*, not lose.

🧠 The Final Boss Challenge (Case Study from your PDF)

This is an exact 5-mark question taken from Page 7 (and Page 29) of your PDFs. Put on your detective hat.

Here is the data table (Values are in kJ/mol):

Element	$\Delta_i H_1$ **(First IE)**	$\Delta_i H_2$ **(Second IE)**	$\Delta_{eg} H$ **(Electron Gain Enthalpy)**
I	520	7300	-60
II	419	3051	-48
III	1681	3374	-328

Element	$\Delta_i H_1$ **(First IE)**	$\Delta_i H_2$ **(Second IE)**	$\Delta_{eg} H$ **(Electron Gain Enthalpy)**
IV	1008	1846	-295
V	2372	5251	**+48**
VI	738	1451	-40

Your Mission: Identify which element (I, II, III, IV, V, or VI) fits the following descriptions:

(a) The least reactive element. *(Hint: Who has a positive electron gain enthalpy?)*

(b) The most reactive metal. *(Hint: Who has the absolute lowest energy needed to steal its first electron?)*

(c) The most reactive non-metal. *(Hint: Who has the most negative number for electron gain enthalpy?)*

(d) The metal which can form a stable binary halide of the formula MX_2 **(where M is the metal and X is a halogen).** *(Hint: To form* MX_2*, the metal must easily lose exactly TWO electrons. Which element has a relatively low first AND second IE, without a massive, crazy jump between them?)*

Oh boy, another table. My eyes are starting to go blurry with all these numbers, but I'll try to use the hints and look at the "jumps" between the values like my teacher said.

(a) The least reactive element: The hint says to look for the positive electron gain enthalpy. Usually, these numbers are negative because atoms "give off" energy when they get an electron, but **Element V** has a **+48**. That means it really, *really* doesn't want an extra electron. Plus, its first ionization energy is 2372, which is the highest on the list! It's basically holding onto its electrons with superglue. So that must be a noble gas, the least reactive one. **Answer: Element V**

(b) The most reactive metal: For this, I need to find the element where it's easiest to steal the first electron (the lowest $\Delta_i H_1$).

- Element I is 520.
- **Element II is 419.**
- The others are all much higher. Since 419 is the lowest number in the first column, it takes the least effort to turn it into an ion. That makes it super reactive. **Answer: Element II**

(c) The most reactive non-metal: The hint says the most negative electron gain enthalpy. I remember my teacher saying a big negative number means the atom is "happiest" to get an electron.

- Element III is -328.
- Element IV is -295. **Element III** has the biggest negative number (-328), so it's the most "greedy" for electrons. That makes it the most reactive non-metal (probably a halogen like Fluorine or Chlorine). **Answer: Element III**

(d) The metal which can form a stable binary halide of the formula MX_2: Okay, MX_2 means the metal needs to lose **two** electrons to be stable (M^{2+}). I have to look at the first two ionization energies:

- In **Element I**, the first is 520 but the second is 7300! That jump is insane. It will lose one electron, but it will *never* let go of a second one. (That's like an Alkali metal).
- In **Element VI**, the first is 738 and the second is 1451. 1451 is bigger, but it's only about double 738. It's not a "crazy" jump. So this atom is okay with losing two electrons to become stable. That fits the MX_2 thing (like $MgCl_2$). **Answer: Element VI**