# Periodic Table

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**9.1:**
**The Periodic Table**

Periodic table provides invaluable information regarding the properties of elements as well as its classification process

**9.2:**
**Periodic trends**

The periodic table is dominated by metals. However, ironically, our Earth is composed of non-metals at a higher percentage.

As we go from the left to the right, the metallic properties of an element decreases.

What does metallic properties indicate. Here, I am not talking about melting point, density or of that sort.

When metallic property is mentioned, it refers to the ease of an element to lose electron(s) to become stable.

Henceforth, referring back to the 3rd Chapter of this syllabus, we know that the fewer valence electrons an element has, the easier it is to lose electrons.

As an example, chlorine has 7 valence electron. It makes no sense for chlorine to lose 7 e in order to become stable if they can just gain one.

Therefore, if you look at the periodic table, the further right you go, the more valence electrons an element has and hence, the lesser their metallic properties.

Now, if you also realize, **elements in the same group have the same number of valence electrons** whereas **elements in the same period have same number of shells.**

**9.3:**
**Group properties**

We will be taking an in depth look at the properties of Group 1 and Group 7

elements.

Group 1:

- Elements in group 1 have **similar chemical properties (similar, not identical)**
- Elements in group 1 show a gradual change in physical properties
- Down the group, melting point decreases
- Down the group, density decreases
- Down the group, reactivity increases
- Down the group, elements become 'softer.'

All metals in group 1 generally are soft, have low density, react vigorously with water and are good electrical and thermal conductors.

It is important to understand several of these properties and trends. We know that the number of valence electrons and element has dictates its chemical properties. This is why elements in group 1 have similar chemical properties.

You might ask, why are they not the same like the case with isotopes.

The most accurate explanation is out of the context of this syllabus. However, for understanding, it is great to know.

First and foremost, unlike isotopes, the elements in group 1 only have the same valence electrons and not electrons. This means that as we go down the group, the number of electrons increases.

Per logic, if the number of electrons is increasing, it means the element itself is getting larger. In other words, its atomic radius is increasing. Now, with that being said, we apply logical reasoning again. We know that the center of an atom is the nucleus is positively charged and that it is pulling the electrons orbiting it. Following this common sense, the further an electron is from the nucleus, the easer it is to be 'pulled away.'

Reactivity, which is a chemical property, varies in group 1. As you go down the group, the number of electrons increases therefore, the valence electron is getting further from the nucleus. This causes the valence electron to be lost easier and that is why reactivity increases down the group.

Group 7:

- Reactivity decreases down the group
- Diatomic molecules
- Coloured and the color turns darker down the group

- Boiling point increases down the group
- Show similar chemical properties and a trend in physical properties

Fluorine -- Pale yellow
Chlorine -- Yellowish green
Bromine -- Reddish brown
Iodine -- Purplish black
Astatine -- Black

The reason why reactivity decreases down the group demonstrates a similar logic with that of group 1 metals.

We know that halogens, having 7 valence electrons, will definitely want to gain 1 e to become stable

Notice how the valence electron of fluorine is closest to the nucleus. This is why it is the most reactive.

Logically, we know that electrons are negatively charged. Therefore, for fluorine to gain 1 e, the nucleus needs to attract it. Therefore, the further your valence electrons are, the harder it is for the nucleus to attract the 1 e it is about to gain.

With that being said, we can demonstrate this understanding by predicting what we can see in reactions involving halides.

$2KI + Cl_2 \rightarrow 2KCl + I_2$

Here, chlorine displaces iodide as it is more reactive. We will see a purplish black solid being formed.

Similar, if we have this reaction

$KF + Cl_2 \rightarrow$

There will be no reaction as fluoride is more reactive compared to chlorine. Hence, no change will be observed.

Notice how F, being the most reactive, is the **easiest to be reduced.** Now, we discussed in previous chapters that a species which is reduced in a chemical reaction is known as an oxidizing agent. Hence, Fluorine is the strongest oxidizing agent.
Hence, when dealing with reactions involving unknown halides, take note of

which species is being oxidized or reduced. Through this information, you can predict which one is the more reactive halide.

**9.4:**
**Transition elements**

Transition elements are found in the middle of the periodic table. These metals exhibit physical and chemical properties we usually are accustomed to.

Physical properties:
1) Good conductor of electricity and heat
2) High density
3) High melting and boiling point
4) Strong and hard

Chemical properties:
1) Have variable oxidation state (reason why is not required)
2) Can act as catalyst in chemical reactions (Iron is used in Haber Process)
3) Form coloured compounds ($CuSO_4$ is blue in aqueous solution)

Notice how, regardless whether a specific metal is in group 1 or 2 or are classified as a transition element, they are always able to conduct electricity.

Hence, when answering exam questions where a table of unknown elements are given and their properties are given, if a substance does not conduct electricity, it is definitely a non-metal.

The reason why metals can always conduct electricity goes back to its structure itself. Metallic bonding is the electrostatic attraction between positive ions and delocalized sea of electrons. The latter means that the electrons are mobile and free to move.

**9.5:**
**Noble gases**

Noble gases are located in the right-most group. They are known as group 8 or group 0. They are inert which basically means they are unreactive. All noble gases are unreactive due to their full valence shell. This also explains why they are all monoatomic. This piece of information is useful during exams as if you see a diatomic gas, noble gases are definitely ruled out.

Uses of noble gases:

1) Helium is used in hot-air balloons. Previously, hydrogen was used as both have very low densities relative to air. However, hydrogen is no longer used due to the fact it reacts easily and hence can ignite to start a fire
2) Neon is used in lights (colorful)
3) Argon in used in welding of metals and also lamps for a inert atmosphere
4) Xenon and krypton all have different natural colors and hence are used in lights as well

**Exam tips:**

1) I'd say as a general tip to memorize all these trends as do come out quite often. However, if you are able to conceptually understand the reason behind the trend, the whole thing becomes so much easier
2) Memorize at least 1 use of noble gas
3) Memorize the difference between transition metals and the metals in the 1st two groups
4) Thoroughly understand reactions involving halides and the visible changes if any
5) Understand that in a group, there is a **trend in physical properties and similar chemical properties**