

# Chapter 6 - The Periodic Table and Bonding

## PERIODIC TABLE OF ELEMENTS

1   2   3   4   5   6   7   8   9   10   11   12   13   14   15   16   17   18																																																																															
1 <b>H</b> Hydrogen 1.008	<div>Atomic # Symbol Name Weight</div>																2 <b>He</b> Helium 4.0026																																																														
3 <b>Li</b> Lithium 6.94	4 <b>Be</b> Beryllium 9.0122	<div><div>C Solid</div><div>Hg Liquid</div><div>H Gas</div><div>Rf Unknown</div></div> <div><div>Metals</div><div>Alkali metals</div><div>Alkaline earth metals</div><div>Lanthanoids (Lanthanides)</div><div>Actinoids (Actinides)</div><div>Transition metals</div><div>Post-transition metals</div><div>Metalloids</div><div>Nonmetals</div><div>Other nonmetals</div><div>Noble gases</div></div> <div><div>Prictogens</div><div>Chalcogens</div><div>Halogens</div></div>																10 <b>Ne</b> Neon 20.180																																																													
11 <b>Na</b> Sodium 22.990	12 <b>Mg</b> Magnesium 24.305	13 <b>Al</b> Aluminium 26.982	14 <b>Si</b> Silicon 28.085	15 <b>P</b> Phosphorus 30.974	16 <b>S</b> Sulfur 32.06	17 <b>Cl</b> Chlorine 35.45	18 <b>Ar</b> Argon 39.948	19 <b>K</b> Potassium 39.098	20 <b>Ca</b> Calcium 40.078	21 <b>Sc</b> Scandium 44.956	22 <b>Ti</b> Titanium 47.867	23 <b>V</b> Vanadium 50.942	24 <b>Cr</b> Chromium 51.996	25 <b>Mn</b> Manganese 54.938	26 <b>Fe</b> Iron 55.845	27 <b>Co</b> Cobalt 58.933	28 <b>Ni</b> Nickel 58.693	29 <b>Cu</b> Copper 63.546	30 <b>Zn</b> Zinc 65.38	31 <b>Ga</b> Gallium 69.723	32 <b>Ge</b> Germanium 72.630	33 <b>As</b> Arsenic 74.922	34 <b>Se</b> Selenium 78.971	35 <b>Br</b> Bromine 79.904	36 <b>Kr</b> Krypton 83.798	37 <b>Rb</b> Rubidium 85.468	38 <b>Sr</b> Strontium 87.62	39 <b>Y</b> Yttrium 88.906	40 <b>Zr</b> Zirconium 91.224	41 <b>Nb</b> Niobium 92.906	42 <b>Mo</b> Molybdenum 95.95	43 <b>Tc</b> Technetium (98)	44 <b>Ru</b> Ruthenium 101.07	45 <b>Rh</b> Rhodium 102.91	46 <b>Pd</b> Palladium 106.42	47 <b>Ag</b> Silver 107.87	48 <b>Cd</b> Cadmium 112.41	49 <b>In</b> Indium 114.82	50 <b>Sn</b> Tin 118.71	51 <b>Sb</b> Antimony 121.76	52 <b>Te</b> Tellurium 127.60	53 <b>I</b> Iodine 126.90	54 <b>Xe</b> Xenon 131.29	55 <b>Cs</b> Caesium 132.91	56 <b>Ba</b> Barium 137.33	57–71	72 <b>Hf</b> Hafnium 178.49	73 <b>Ta</b> Tantalum 180.95	74 <b>W</b> Tungsten 183.84	75 <b>Re</b> Rhenium 186.21	76 <b>Os</b> Osmium 190.23	77 <b>Ir</b> Iridium 192.22	78 <b>Pt</b> Platinum 195.08	79 <b>Au</b> Gold 196.97	80 <b>Hg</b> Mercury 200.59	81 <b>Tl</b> Thallium 204.38	82 <b>Pb</b> Lead 207.2	83 <b>Bi</b> Bismuth 208.98	84 <b>Po</b> Polonium (209)	85 <b>At</b> Astatine (210)	86 <b>Rn</b> Radon (222)	87 <b>Fr</b> Francium (223)	88 <b>Ra</b> Radium (226)	89–103	104 <b>Rf</b> Rutherfordium (267)	105 <b>Db</b> Dubnium (268)	106 <b>Sg</b> Seaborgium (269)	107 <b>Bh</b> Bohrium (270)	108 <b>Hs</b> Hassium (271)	109 <b>Mt</b> Meitnerium (278)	110 <b>Ds</b> Darmstadtium (281)	111 <b>Rg</b> Roentgenium (282)	112 <b>Cn</b> Copernicium (285)	113 <b>Nh</b> Nihonium (286)	114 <b>Fl</b> Flerovium (289)	115 <b>Mc</b> Moscovium (290)	116 <b>Lv</b> Livermorium (293)	117 <b>Ts</b> Tennessine (294)	118 <b>Og</b> Oganesson (294)
For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.																																																																															
57 <b>La</b> Lanthanum 138.91	58 <b>Ce</b> Cerium 140.12	59 <b>Pr</b> Praseodymium 140.91	60 <b>Nd</b> Neodymium 144.24	61 <b>Pm</b> Promethium (145)	62 <b>Sm</b> Samarium 150.36	63 <b>Eu</b> Europium 151.96	64 <b>Gd</b> Gadolinium 157.25	65 <b>Tb</b> Terbium 158.93	66 <b>Dy</b> Dysprosium 162.50	67 <b>Ho</b> Holmium 164.93	68 <b>Er</b> Erbium 167.26	69 <b>Tm</b> Thulium 168.93	70 <b>Yb</b> Ytterbium 173.05	71 <b>Lu</b> Lutetium 174.97	89 <b>Ac</b> Actinium (227)	90 <b>Th</b> Thorium 232.04	91 <b>Pa</b> Protactinium 231.04	92 <b>U</b> Uranium 238.03	93 <b>Np</b> Neptunium (237)	94 <b>Pu</b> Plutonium (244)	95 <b>Am</b> Americium (243)	96 <b>Cm</b> Curium (247)	97 <b>Bk</b> Berkelium (247)	98 <b>Cf</b> Californium (251)	99 <b>Es</b> Einsteinium (252)	100 <b>Fm</b> Fermium (257)	101 <b>Md</b> Mendelevium (258)	102 <b>No</b> Nobelium (259)	103 <b>Lr</b> Lawrencium (260)																																																		

### Group

Group	Name
Group 1	Alkali Metals
Group 2	Alkaline Earth Metals
Group 17	Halogens
Group 18	Noble Gases

- The vertical columns of Periodic Table
- There're **18 groups** (Group 1 - 18) arranged according to the number of valence electrons.
- All members of the same group have the **same number of valence electrons**.
- Members of a group have **similar chemical properties**. Their physical properties may show gradual change when descending group.

### Period

- The horizontal rows are called periods.
- There're **7 periods**.
- Period number is indicated by **number of filled electron shells**.
- All elements in the same period have the **same number of filled electron shells**.

# Pattern in the Periodic Table - Group

## Group 1 (Group IA, alkali metals)

### Physical Properties

1. **Soft metals** which can be easily cut with knife
  - Number of delocalized electron is lesser compared to other metals, the metallic bond is weaker
2. **Silvery** solids with **shiny** surface
3. **Lower melting / boiling point** compared to other metals.
4. **Lower density** compared to other metals
5. Good **conductivity** of heat and electricity

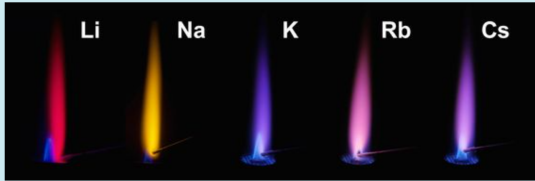
### Trend in physical properties when going down the group

Characteristic	Trend	Reason
1. Atomic size	Increases	The number of electron shells increases when descending the group.
2. Density	Increases	Mass increases faster than the increase in radius.
3. Melting point / Boiling point	Decreases	<ul style="list-style-type: none"> <li>- Atomic size increases.</li> <li>- The distance between delocalized electron and positive atom is increased.</li> <li>- Metallic bond between atoms become weaker.</li> <li>- Less heat energy is needed to overcome the attraction force between delocalized electron and positive metal atoms.</li> </ul>
4. Electropositivity	Increases	<ul style="list-style-type: none"> <li>- Electropositivity: a measurement of the ability of an atom to lose an electron and form a cation.</li> <li>- Atomic size increases.</li> <li>- The attractive force between the positive nucleus and the single valence electron becomes weaker.</li> <li>- The atom loses electron more easily.</li> </ul>

### Chemical Properties

- Each group 1 elements have **similar chemical properties** because of all having 1 valence electron.
- All group 1 elements are **very reactive metals**, they will **release a valence electron** to form 1+ ion to **achieve octet electron arrangement**.
  - $M \rightarrow M^+ + e^-$  (M = Li, Na, K, Rb, Cs, Fr)
- Their **chemical reactivity** is different. It depends on how easily it can **donate its valence electron**.
- The **reactivity** of group 1 elements **increases when going down the group**, because:
  - *The number of shells increases, so the **atomic size increases***
  - *The valence electron in the outermost shell gets further away from nucleus, causing **attractive force** of proton to valence electron becoming **weaker**.*
  - *Thus, the valence electron can be released **easily**.*

## Chemical Reactions

Chemical Reactions of Group 1 Elements			
	Alkali Metals	Observation	Equation
Group 1 metals react with water to produce alkali and hydrogen gas	Lithium	Reacts readily with water	$2\text{Li}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{LiOH}_{(aq)} + \text{H}_{2(g)}$
	Sodium	Reacts vigorously; Catches fire	$2\text{Na}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{NaOH}_{(aq)} + \text{H}_{2(g)}$
	Potassium	Reacts violently; Catches fire; then explodes	$2\text{K}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{KOH}_{(aq)} + \text{H}_{2(g)}$
Group 1 metals can be oxidized in air to form white metal oxides	Lithium	Tarnish in air due to the formation of dull oxide on the surface	$4\text{Li}_{(s)} + \text{O}_{2(g)} \longrightarrow 2\text{Li}_2\text{O}_{(s)}$
	Sodium	Tarnish in air due to the formation of dull oxide on the surface	$4\text{Na}_{(s)} + \text{O}_{2(g)} \longrightarrow 2\text{Na}_2\text{O}_{(s)}$
	Potassium	Tarnish in air due to the formation of dull oxide on the surface	$4\text{K}_{(s)} + \text{O}_{2(g)} \longrightarrow 2\text{K}_2\text{O}_{(s)}$
Group 1 metals burn in sufficient oxygen with characteristic flame colors	 <p>Flame tests are used to preliminarily identify the presence of metal ions in a compound. Not all metal ions give flame colors.</p>		
	Lithium	Bright red flame	
	Sodium	Yellow flame	
	Potassium	Violet flame	
	Rubidium	Red violet flame	
	Caesium	Blue flame	
Group 1 metals burn in sufficient oxygen form various oxides	Sodium	Sodium starts burning immediately in air when heated or put under a flame. Sodium burns with yellow flame and forms a mixture of sodium oxide and sodium peroxide.	$4\text{Na}_{(s)} + \text{O}_{2(g)} \longrightarrow 2\text{Na}_2\text{O}_{(s)}$ <small>sodium oxide</small> $2\text{Na}_{(s)} + \text{O}_{2(g)} \longrightarrow \text{Na}_2\text{O}_{2(s)}$ <small>sodium peroxide</small>
	Potassium	Sodium burns vigorously with lilac / violet flame and forms a mixture of potassium superoxide and potassium peroxide.	$\text{K}_{(s)} + \text{O}_{2(g)} \longrightarrow \text{KO}_2(s)$ <small>potassium superoxide</small> $2\text{K}_{(s)} + \text{O}_{2(g)} \longrightarrow \text{K}_2\text{O}_2(s)$ <small>potassium peroxide</small>
Group 1 metals react with halogen to form white metal halides	Lithium	1. All of the alkali metals react vigorously with chlorine gas.	$2\text{Li}_{(s)} + \text{Cl}_{2(g)} \longrightarrow 2\text{LiCl}_{(s)}$
	Sodium	2. Each reaction produces a white crystalline salt.	$2\text{Na}_{(s)} + \text{Cl}_{2(g)} \longrightarrow 2\text{NaCl}_{(s)}$
	Potassium	3. The reaction gets more violent as you move down Group 1.	$2\text{K}_{(s)} + \text{Cl}_{2(g)} \longrightarrow 2\text{KCl}_{(s)}$

## Group 17 (Group VIIA, halogens)

- Group 17 elements: Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I), Astatine (At)
- The halogens molecules exist as **diatomic molecules**:  $\text{F}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$ ,  $\text{I}_2$ ,  $\text{At}_2$

## Physical Properties

- All Group 17 elements are non-metals. They exist as diatomic molecules.
- They're **non-conductors** of heat and electricity.
- Low density.
- Low melting and boiling point**, because the halogen molecules are attracted by **weak van der Waals force**.

- van der Waals force  $\propto$  molecule size
- $\text{I}_{2(s)} > \text{Br}_{2(l)} > \text{Cl}_{2(g)} > \text{F}_{2(g)}$  [at 25°C]

### Trend in physical properties when going down the group

Characteristic	Trend	Reason
Atomic radius	Increases	The number of electron shells increases when descending the group.
Density	Increases	Mass increases faster than the increase in radius.
Melting point / Boiling point	Increases	<ul style="list-style-type: none"> <li>- Molecular size increases.</li> <li>- Causes van der Waals force between the molecules become stronger.</li> <li>- More heat is required to overcome the attractive force.</li> </ul>
Electronegativity	Decreases	<ul style="list-style-type: none"> <li>- Atomic radius increases.</li> <li>- Causes the attractive force between the nucleus and the valence electrons becomes weaker.</li> <li>- Hence, the ability of atoms to attract electrons decreases.</li> </ul>

### \*Electronegativity

H 1		Electronegativity																He 2									
Li 3		Be 4		The electronegativity of an atom is how strongly it attracts electrons towards itself. It depends on the atomic radius and the atomic number of the element. Electronegativity is most commonly measured on the Pauling scale. Values are shown relative to fluorine, the element with the highest electronegativity.														B 5	C 6	N 7	O 8	F 9	Ne 10				
Na 11		Mg 12																Al 13	Si 14	P 15	S 16	Cl 17	Ar 18				
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36										
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54										
Cs 55	Ba 56	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86										
Fr 87	Ra 88	Ac 89	Rf 104	Db 105	Sg 106	Bh 107	Hs 108	Mt 109	Ds 110	Rg 111	Cn 112	Nh 113	Fl 114	Mc 115	Lv 116	Ts 117	Og 118										
Ce 58		Pr 59		Nd 60		Pm 61		Sm 62		Eu 63		Gd 64		Tb 65		Dy 66		Ho 67		Er 68		Tm 69		Yb 70		Lu 71	
Th 90		Pa 91		U 92		Np 93		Pu 94		Am 95		Cm 96		Bk 97		Cf 98		Es 99		Fm 100		Md 101		No 102		Lr 103	

- A measure of the **tendency of an atom to attract a bonding pair of electrons**.
- When two atoms with different electronegativity bond covalently, the electron pair will be pulled more to the atom with higher electronegativity.
- The Pauling scale is most commonly used. Fluorine (the most electronegative element) is assigned a value of 4.0, and values range down to francium which is the least electronegative at 0.7.
- **F > O > Cl > N > Br > I > S > C > H > metals**
- No electronegativity difference between two atoms leads to a pure **non-polar covalent bond**.
- A small electronegativity difference leads to a **polar covalent bond**.
- A large electronegativity difference leads to an **ionic bond**.

## Chemical Properties

- Each of Group 17 elements have **similar chemical properties** because all atoms have 7 valence electrons.
- All group 17 elements are very reactive non-metals. An atom of group 17 elements will **accept an valence electron** to form 1- ion to **achieve octet electron arrangement**.
  - $X_2 + 2e^- \rightarrow 2X^-$  (X = F, Cl, Br, I)
- The reactivity of Group 17 elements depends on its ability to gain an electron. **The reactivity decreases when going down the group**, because:
  - The number of shells increases so the atomic radius of group 17 elements increases down the group.*
  - The valence electron in the outermost shell gets further away from the nucleus, thus causes attractive force of proton to valence electron becomes weaker.*
  - The ability of a halogen atom to attract electron decreases from fluorine to astatine.*

## Chemical Reactions

### (i) Group 17 elements react with water to produce acid

Halogen	Observation: Solubility	Observation: Effect on Litmus Paper	Equation
Chlorine	Dissolves rapidly; light yellow solution	Blue turns red, then decolorized	$Cl_{2(g)} + H_2O_{(l)} \rightleftharpoons HCl_{(aq)} + HOCl_{(aq)}$
Bromine	Dissolves slowly; reddish brown solution	Blue turns red, then decolorized after a longer time	$Br_{2(g)} + H_2O_{(l)} \rightleftharpoons HBr_{(aq)} + HOBr_{(aq)}$
Iodine	Dissolves slightly; pale brown solution	Blue turns red, is not decolorized	$I_{2(g)} + H_2O_{(l)} \rightleftharpoons HI_{(aq)} + HOI_{(aq)}$

### (ii) Group 17 elements react with sodium hydroxide to produce water and salt

- $Cl_{2(g)} + 2NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NaClO_{(aq)} + H_2O_{(l)}$
- $Br_{2(g)} + 2NaOH_{(aq)} \rightarrow NaBr_{(aq)} + NaBrO_{(aq)} + H_2O_{(l)}$
- $I_{2(g)} + 2NaOH_{(aq)} \rightarrow NaI_{(aq)} + NaIO_{(aq)} + H_2O_{(l)}$

## Group 18 (Group VIIIA, noble gases / inert gases)

- Group 18 elements: Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe), Radon (Rn)

## Physical Properties

- All noble gases are **insoluble** in water.
- They're **non-conductors** of heat and electricity.
- Very **low density**. Due to the atoms are very far apart.
- The **density** of noble gases **increases** when descending the group, due to the atomic mass increases faster than the increase in radius.
- Noble gases have **very low melting and boiling point**. This is because noble gases exist as **monoatomic molecules** which are attracted by **weak van der Waals forces**, less energy is required to overcome the forces.

6. However, the **melting and boiling point increases** when going down the group, because atomic size increases, causing van der Waals forces stronger and more energy is required to overcome the forces.

## Chemical Properties

1. They exist as **monoatomic molecules** (single atoms) because the atom does **not lose, gain or share electrons**.
2. Noble gases are **chemically inert** (unreactive) because the atoms have achieved **duplet electron arrangement** for helium and **octet electron arrangement** for others.

## Pattern in the Periodic Table - Period

### Period 3

- The elements in period 3: Sodium (Na), Magnesium (Mg), Aluminum (Al), Silicon (Si), Phosphorus (P), Sulphur (S), Chlorine (Cl), Argon (Ar).

### Physical and Chemical Properties across Period 3

Characteristic	Trend	Reason
1. Atomic radius	Decreases	<ul style="list-style-type: none"> <li>- All atoms of the elements have 3 shells occupied with electrons.</li> <li>- <b>Proton number increases</b> one by one from Na to Cl</li> <li>- It causes the number of positive charge in the nucleus increases.</li> <li>- The strength of <b>attraction</b> between proton in nucleus and electrons in the shells <b>increases</b>.</li> <li>- All electrons in the shells are pulled closer toward nucleus</li> <li>- The <b>size of atom decreases</b> across the period.</li> </ul>
2. Electronegativity	Increases	<ul style="list-style-type: none"> <li>- Electronegativity: a measurement of <b>tendency of an element to attract protons</b>.</li> <li>- The increase in number of protons and decrease in atomic radius cause <b>increase in attractive force between nucleus and outermost shell electron</b>.</li> <li>- The <b>electronegativity increases</b> when moving across period.</li> </ul>
3. Melting point / Boiling Point	a. Increases (from left to the middle of period)	<ul style="list-style-type: none"> <li>- <b>Na, Mg and Al</b> are metals with strong metallic bond between atoms.</li> <li>- The strength of <b>metallic bond increases</b> from Na to Al with <b>increase</b> in number of protons and <b>valence electrons</b>.</li> <li>- Hence, they have high melting and boiling point</li> </ul>
	b. Very high melting and boiling point (at the middle of the period)	<ul style="list-style-type: none"> <li>- <b>Si</b> has very high melting and boiling point.</li> <li>- It forms a <b>giant covalent network structure</b>, atoms are held by <b>strong covalent bond</b>.</li> <li>- A large amount of energy is required to break to strong covalent bond.</li> </ul>

Characteristic	Trend	Reason
	c. Decreases (from the middle to right of period)	- <b>P, S, Cl and Ar</b> are non-metals with <b>weak van der Waals force</b> between molecules. - They have low melting and boiling point. - m.p. <b>S<sub>8</sub> &gt; P<sub>4</sub> &gt; Cl<sub>2</sub> &gt; Ar</b>

Characteristic	Trend
4. Metallic property	<b>Metals</b> [Na, Mg, Al] —→ <b>Metalloid</b> [Si] —→ <b>Non-metal</b> [P, S, Cl, Ar]
5. Nature of oxides	<b>Basic</b> oxides [Na <sub>2</sub> O, MgO] —→ <b>Amphoteric</b> oxide [Al <sub>2</sub> O <sub>3</sub> ] —→ <b>Acidic</b> oxides [SiO <sub>2</sub> , P <sub>4</sub> O <sub>10</sub> , SO <sub>2</sub> , Cl <sub>2</sub> O <sub>7</sub> ]

## Transition Elements

- located between group IIA and IIIA (group 2 and group 13)
- shows metal properties: shiny conduct heat / electricity, malleable, high tensile strength, high melting and boiling point and high density

### Special Characteristic

- Have **more than one oxidation state**

Transition Metals	Oxidation States
Fe	+2, +3
Cr	+3, +6
Cu	+1, +2
Ni	+2, +3

- Exception: Zinc, only form Zn<sup>2+</sup>

- Form **colored compounds**

Transition Ions	Color
Fe <sup>2+</sup>	green
Fe <sup>3+</sup>	brown
Cu <sup>2+</sup>	blue
Co <sup>2+</sup>	pink

- Exception: Zn<sup>2+</sup> & Sc<sup>3+</sup>

- From **complex ions**

- e.g. [Cu(NH<sub>3</sub>)<sub>4</sub>]<sup>2+</sup>

4. Used as **catalyst**

Transition metals / compounds	Process	Usage
$V_2O_5$	Contact process	produce sulfuric acid
Pt	Ostwald process	produce nitric acid
Fe	Haber process	produce ammonia