

# An Overview of The Properties of Elements in The Periodic Table

## The Key Atomic Properties

Atomic properties that are critical to the behavior of elements are *electron configuration*, *atomic size*, *ionization energy*, *electron affinity*, and *electronegativity*.

The electron configuration ( $nl^{\#}$ : *spdf* notation) gives the distribution of electrons in *principal energy levels* and *sublevels* of an atom.

- The principal quantum number  $n$  determines the *energy* and *relative distance* of orbitals from the nucleus;
- The quantum number  $l$  yields *s*, *p*, *d*, and *f* notations, which describes the shape of orbitals in the sublevel;
- The *superscript* ( $\#$ ) tells the number of electrons in the sublevels.

Electrons in inner levels or *shells* tend to *shield* outer electrons from the full nuclear charge, which is reduced to *effective nuclear charge* ( $Z_{\text{eff}}$ ). Electrons that have a greater *penetration* shield others more effectively. For example, electrons in level  $n = 1$  shield those in level  $n = 2$  very effectively, and those in  $n = 1$  and  $n = 2$  shield electrons in level  $n = 3$ . Electrons at the same level, but in different sublevels, also *shield* other electrons to some extent. The extent of *penetration* and *shielding effect* is in the order:  $s > p > d > f$ .

The *effective nuclear charge* ( $Z_{\text{eff}}$ ) greatly influences atomic properties. In general,

- $Z_{\text{eff}}$  increases significantly across a period (left-to-right)
- $Z_{\text{eff}}$  decreases slightly down a group.

In the periodic table, elements are divided into:

- the *s-block* (contains reactive metals of Group 1A (1) and 2A (2)),
- the *p-block* (contains *metals* and *nonmetals* of Group 3A (13) through 8A (18)),
- the *d-block* (contains transition metals (Group 3B (3) through Group 2B (12)), and
- the *f-block* (contains *lanthanide* and *actinide series* or *inner transition* metals).

Elements within the same *group* have the same number of electrons in their *valence* (outermost) shells, and they have *similar electron configurations*. They exhibit *similar chemical properties*. Elements within the same *period* have different number of electrons in their *valence shells* (the number is *increasing from left to right*) and *different valence shell electron configuration*. Therefore, elements in the same period are chemically different.

Atoms do not have specific outer boundary; their **atomic size** is based on atomic radius, which is *one-half the internuclear distance of two bonded identical atoms*. Two factors that greatly influenced atomic size are *principal quantum number* ( $n$ ) of valence shell and the *effective nuclear charge* ( $Z_{\text{eff}}$ ). Since principal quantum number  $n$  increases down a group and  $Z_{\text{eff}}$  increase significantly across a period, the following general trends in *atomic size* are observed:

- *Atomic size increases down a group* because the principal quantum number increases;
- *Atomic size generally decreases across a period* because  $Z_{\text{eff}}$  increases.

**Ionization energy (IE)** is the energy needed to remove a valence electron from 1 mol of gaseous atoms. The relative size of the IE influences the types of ions and bonds an atom forms: an element with a low IE is more likely to lose electrons and form cations, and one with a high IE is more likely to gain electrons and form anions or share electrons in a covalent bond (exclude noble gases). Since the  $Z_{eff}$  of the atom influences the IE of elements, the following trends are observed in the periodic table:

- IE generally *increases across a period*:  $Z_{eff}$  increases and valence shell electrons are held more firmly;
- IE *decreases down a group*: valence electrons are farther away from the nucleus and held less firmly.

The trends in IE are opposite those in atomic size: it is easier to remove an electron (lower IE) from a large atom than from a small one.

**Electron affinity (EA)** is the energy change when electrons are added to 1 mol of gaseous atoms. Two factors influenced EA, the  $Z_{eff}$  of the atom and electron-electron repulsions. The following trends of EA are observed in the periodic table:

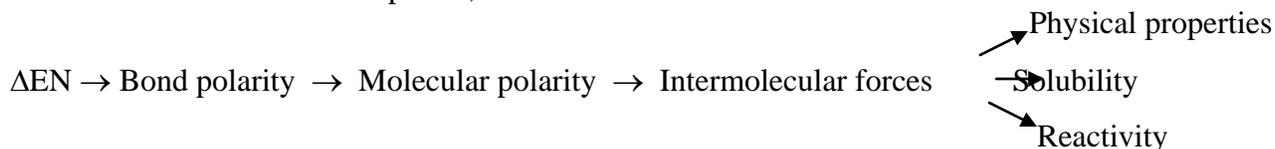
- EA generally *becomes more negative* (increases) *across a period* (exclude noble gases): as  $Z_{eff}$  increases the added electron comes into a stronger nuclear attraction.
- EA *becomes less negative* (decreases) *down a group*: nuclear attraction is less effective when an electron is added to the valence shell of a large atom.

However, if the atomic size is very small electron-electron repulsions become quite large. Thus, EA of elements in the second period ( $n = 2$ ) are generally less negative (lower) than the same group elements in the third period ( $n = 3$ ). Nonmetals (except the noble gases), especially those in Group 7A (17) and 6A (16), generally have a high IE (more positive) and EA (more negative) and they gain electron quite readily to become anions.

**Electronegativity (EN)** is the *relative ability of a covalently bonded atom to attract shared electrons*. EN follows a trend that is opposite to that in atomic size and the same as that in IE.

- EN generally *increases across a period*: higher  $Z_{eff}$  and shorter distance from the nucleus strengthen the attraction for the shared electron pair(s);
- EN generally *decreases down a group*: greater distance from the nucleus weakens the attraction for the shared electron pair(s).

The difference in electronegativity ( $\Delta EN$ ) between bonded atoms influences the physical and chemical behavior of the compound, as shown below:



### Characteristics of Chemical Bond

Chemical bonds are forces that hold atoms (or ions) together in an element or compound. The type of bonds, bond properties, bond strength, and the number of bonds determine the chemical behavior of the compound.

**Types of Bonding:** *ionic, covalent, and metallic.*

- **Ionic bond** is the electrostatic attraction between cations and anions. The strength of ionic bond *increases as the ion's charge becomes larger and ion's size decreases*. This type of bonding leads to *crystalline solids* with ions packed tightly in regular *lattice structures*.
- **Covalent bond** results from the attraction between two nuclei and a localized electron pair. The bond arises through electron sharing between nonmetal atoms or between a semimetal and a nonmetal atoms, which leads to discrete molecules with specific shapes or to extended networks.

*Covalent bonds* are characterized in terms of:

- **bond length** - the distance between the nuclei of bonded atoms;
- **bond energy** - (a measure of *bond strength*) is the enthalpy change when 1 mol of bonds in gaseous molecules are broken.
- **bond order** - the number of electron pairs shared between bonded atoms (bond order of 1 is single bond, 2 is double bond, and 3 is triple bond).

Bond length and bond energy are related to extent of overlap of atomic orbitals. Among similar compounds, they relate to reactivity. For example, C—X bond length in CX<sub>4</sub> molecules increases in the order: C—F < C—Cl < C—Br < C—I.

The bond energy decreases in the order: C—F > C—Cl > C—Br > C—I.

The reactivity of the compounds increases in the order: CF<sub>4</sub> < CCl<sub>4</sub> < CBr<sub>4</sub> < CI<sub>4</sub>. Thus,

- *As bond length increases, bond energy decreases*: shorter bonds are stronger;
- *As bond energy decreases, reactivity increases*: easier to break weaker bonds.

In a covalent bond, the shared electrons reside in the entire region of the overlapping orbitals of bonded atoms. Orbitals overlap in two ways, which leads to two types of bonds:

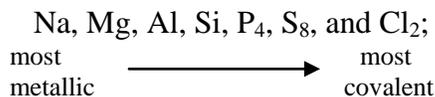
- **End-to-end overlap** (involving *s*, *p*, and *hybrid atomic orbitals*) leads to a sigma ( $\sigma$ ) bonds, which has a symmetrical electron density distribution along the bond axis. A single bond is a  $\sigma$ -bond.
- **Side-to-side overlap** (involving two parallel *p* orbitals) leads to a pi ( $\pi$ ) bond, which has electron density distribution above and below the bond axis. A double bond consists of one  $\sigma$ -bond and one  $\pi$ -bond. A  $\pi$ -bond restricts rotation around the bond axis, leading to different spatial arrangements of the atoms in such molecules. In organic chemistry,  $\pi$ -bonds are often sites of reactivity.

**Metallic bond** results from the attraction between the cores of metal atoms (nuclei plus inner-shell electrons) and their delocalized valence electrons. This bonding, which is called the “sea of electrons” model, leads to crystalline solids (similar to ionic crystals) and accounts for the special characteristics of metals, such as their *lustrous appearance, conductivity, malleability, and ductility*.

Most substances exhibit bonding characteristics that lie between these distinct models. For example, the electron density depictions show that electron clouds overlap slightly even in ionic bonding (NaCl, KBr, etc.). They overlap significantly in polar covalent bonding (HCl, H<sub>2</sub>O, etc.) and completely in nonpolar covalent bonding (H<sub>2</sub>, Cl<sub>2</sub>, etc.).

### *Some generalization:*

For the following series of elements:



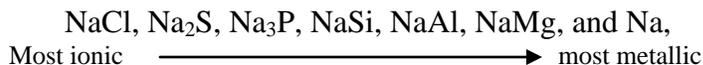
Metallic character increases from right to left; covalent character increases from left to right.

For the following series of compounds:



Ionic character increases from right to left; covalent character increases from left to right.

For the following series of substances:



Ionic character increases from right to left; metallic character increases from left to right.

The ionic bond character increases as  $\Delta EN$  becomes greater. The polarity of covalent bonds decreases as  $\Delta EN$  becomes smaller. Thus, compounds formed between metals of Group 1A(1) and nonmetals of Group 7A(17) exhibit the most ionic character. Those made up of elements (nonmetals) in Groups 6A(16) and 7A(17) are the most covalent and are nonpolar or least polar. Substances made up of Groups 1A and 2A elements are mostly metallic.

### **Physical States of Elements**

The physical state and heat of fusion or vaporization of elements reflects the relative strength of bonding and/or interparticle forces within the element.

- *Metals are solids* (except mercury): strong metallic bonding holds atoms in their crystal structures.
- *Metalloids and carbon are hard solids* because of strong covalent bonds holding atoms together in extensive covalent networks;
- Nonmetals such as sulfur, phosphorus, and iodine are “soft” solids because of relatively *strong dispersion forces* holding their molecules together. Bromine is a liquid.
- Other nonmetals are gases because dispersion forces between small molecules are weak.
- *The hardness and melting points of metals increase as the number of valence electrons increases from left to right across periods;*
- In Groups 1A(1) and 2A(2), *melting point, boiling point,  $\Delta H_{\text{fus}}$ , and  $\Delta H_{\text{vap}}$*  decrease from top to bottom down the groups;
- In Groups 7A(17) and 8A(18), *melting point, boiling point,  $\Delta H_{\text{fus}}$ , and  $\Delta H_{\text{vap}}$*  increase from top to bottom down the groups.
- Among transition metals, the hardest metals are found approximately in the middle of the series. Thus, Cr and Mn are the hardest metals in the first transition series.

## Molecular Shapes

- The shape of a molecule is defined by the relative positions of the nuclei of the bonded atoms.
- According to the VSEPR theory for “simple” molecules, the number of bond pairs and lone pairs and their geometrical orientation around the central atom determines molecular shapes. The geometrical orientations of bond pairs around the central atom are as follows:

| No.of e-pairs | El. Pr. Geometry   |
|---------------|--------------------|
| 2             | linear             |
| 3             | trigonal planar    |
| 4             | tetrahedral        |
| 5             | trigonal bipyramid |
| 6             | octahedral         |

If a molecule contains no lone pair, the molecular shape will be the same as the electron pair geometry. If one or more lone pairs are present, the molecule may adopt shapes such as, a *V-shape*, *trigonal pyramid*, *seesaw*, *T-shape*, *square planar*, or *square planar*.

For generalization,

- *Elements in Period 2 cannot form more than four bonds* because they can only accommodate 4 pairs of electrons in the valence shell (which contains *s* and *p* orbitals).
- *Elements in period 3 or higher can form more than four bonds*, which involves the *d* orbitals.

## Metallic & Nonmetallic Characteristics

General Properties of Elements in the Periodic Table:

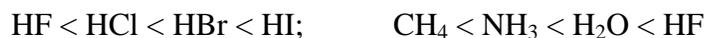
|                            | Metals  | Nonmetals   |
|----------------------------|---|---|
| <b>Atomic Properties</b>   | Have fewer valence electrons<br>Larger atomic size<br>Lower ionization energy<br>Lower electronegativity                                | Have more valence electrons<br>Smaller atomic size<br>Higher ionization energy<br>Higher electronegativity  |
| <b>Physical Properties</b> | Solid at room temperature<br>(except Hg)<br>Good conductors of electricity<br>and heat<br>Lustrous, malleable & ductile                 | Mostly gases, Br <sub>2</sub> is a liquid,<br>others are solids<br>Generally nonconductors<br>(except graphite)<br>None of these                    |
| <b>Chemical Properties</b> | Lose electrons to become cations<br>Reacts with nonmetals to form<br>ionic compounds (salts)<br>Metals do not react with<br>one another | Gain electrons to become anion<br>Reacts with metals to form<br>ionic compounds<br>Nonmetals (except noble gases)<br>react to form covalent compds. |

### **General trends of metallic and nonmetallic behaviors of elements in the periodic table:**

- Metallic properties increase and nonmetallic properties decrease top to bottom within the group;
- Metallic properties decrease and nonmetallic increase left to right across periods;
- Reactivity of metals in the main groups increases from top to bottom and decrease from left to right across periods;
- Reactivity of nonmetals (noble gases excluded) decreases top down within groups and increases from left to right across period.

### **Acid-Base Properties of the Element Oxides**

- An acid reacts with water to increase the  $\text{H}_3\text{O}^+$  concentration;
- A base reacts with water to increase the  $\text{OH}^-$  concentration;
- Acids and bases react to form salts (ionic compounds)
- Oxides of metals are generally basic; the basicity increases from top to bottom within groups and decreases left to right across periods; example:  
Reaction with water:  $\text{Li}_2\text{O}(\text{s}) + \text{H}_2\text{O} \rightarrow 2\text{Li}^+(\text{aq}) + 2\text{OH}^-(\text{aq})$ ;  
Reaction with acid:  $\text{Na}_2\text{O}(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{H}_2\text{O}$ ;
- Oxides of nonmetals (noble gases excluded) are acidic; the acidity decreases from top to bottom within groups and increases left to right across periods.  
Reaction with water:  $\text{SO}_2(\text{g}) + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{SO}_3(\text{aq}) \leftrightarrow \text{H}^+(\text{aq}) + \text{HSO}_3^-(\text{aq})$ ;  
Reaction with base:  $\text{SO}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}$
- Some oxides of metals located in the middle of periods are *amphoteric* - behave like acids or bases, depending on the condition. For example:  
Reaction with acid:  $\text{Al}_2\text{O}_3(\text{s}) + 6\text{H}_3\text{O}^+(\text{aq}) \rightarrow 2\text{Al}^{3+}(\text{aq}) + 9\text{H}_2\text{O}$ ;  
Reaction with base:  $\text{Al}_2\text{O}_3(\text{s}) + 2\text{OH}^-(\text{aq}) + 3\text{H}_2\text{O} \rightarrow 2\text{Al}(\text{OH})_4^-(\text{aq})$
- For binary acids, HX (where X is a simple anion), acid strength increases from top to bottom within groups and left to right across periods. The order of acid strength is:



### **Oxidation-Reduction Properties**

The relative ability of an element to lose or gain electrons when reacting with other elements defines its oxidation-reduction (or redox) property. A redox reaction occurs if the *oxidation number* (o.n.) of any atoms in the reactants is different from those in the products.

- In a redox reaction, the *reactant that gives up electrons is the reducing agents and those that gain electrons are the oxidizing agents*;
- In redox reactions, *reducing agents got oxidized and oxidizing agents got reduced*;
- *Metals are generally reducing agents and nonmetals are oxidizing agents*;
- *Reducing strength of metals increases from top to bottom down groups and decreases left to right across periods*;
- *Oxidizing strength of nonmetals (noble gases excluded) decreases from top to bottom down groups and increases from left to right across periods*;

### ***Oxidation Number (or State):***

- *The oxidation number (or state) of atoms in free elements is zero;*
- *In compounds, the oxidation state of fluorine is -1; the oxidation state of a Group 1A (1) metal is +1, and the oxidation state of a Group 2A (2) metal is +2;*
- *In compounds, the oxidation state of oxygen is -2; except in peroxides which is -1;*
- *The oxidation state of hydrogen is +1 in compounds with nonmetals and -1 in hydrides of metals;*
- *When forming compounds with metals, the oxidation state of Cl, Br, and I is -1;*
- *The highest oxidation state in a group equals the A-group number; for nonmetals the most negative oxidation state equals to (A-group minus 8).*
- *Nonmetals have more oxidation states than metals in the same group or period;*
- *Generally, elements in odd-numbered groups have odd-numbered oxidation states and those in even-numbered groups have even-numbered oxidation states that are stable;*
- *Metals with more than one oxidation states exhibit more metallic behavior in their lower oxidation states - they acts as a reducing agent;*
- *An element with more than one oxidation state is a strong oxidizing agent when in its highest oxidation state.*

### ***Trends of Atomic Properties Across a Period for Elements of the Second***

The second period elements possess certain chemical characteristics that are not shared by the rest of the elements in their respective groups. General trends of physical and chemical properties across (left-to-right) the period 2 elements: Li, Be, B, C, N, O, F, and Ne.

- Atomic number changes from 3 to 10;
- Electron configuration of valence shell changes from  $2s^1$  through  $2s^2 2p^6$ ; number of electrons in the *valence shell* increases from 1 (in Group 1A) to 8 (in Group 8A).
- Effective nuclear charge increases across the period;
- Atomic size generally decreases with increasing effective nuclear charge; anomalous behavior is observed at Be and B ( $Be < B$ ), and at N and O ( $N < O$ )
- Ionization energy generally increases; anomalous behavior is observed at Be and B ( $Be > B$ ), and at N and O ( $N > O$ )
- Electronegativity increases gradually from left to right across period (exclude noble gas);
- Metallic character decreases with increasing nuclear charge;
- Reactivity is highest on the left end (for metal) and right end (for nonmetal) of the period, and decreasing towards the middle;
- Chemical bonding between atoms of an element changes from metallic (Li & Be) to covalent in network structure (B & C) to covalent in individual molecules ( $N_2$ ,  $O_2$ , &  $F_2$ ) to none in Ne.
- Bonding between each element and an active nonmetal (such as  $F_2$ ) changes from ionic to polar covalent, to nonpolar covalent.
- The acid-base characteristics of the common oxides changes from most basic to mild basic to weakly acidic to strongly acidic;
- The reducing strength decreases through the metals; oxidizing strength increases through the nonmetal (exclude noble gases).
- The trend in physical properties is such that it goes from soft and light metal with low melting point (for Li) followed by a hard metal with high mp as occurs in beryllium to very hard metalloid with extremely high mp that occurs in boron. Then the nonmetals start with Carbon (middle of the period), the only nonmetal that is a solid. Important natural allotropes of carbon are *graphite* and *diamond*, which are physically very different. Graphite is opaque, soft and flaky, while diamond is very hard and shiny. The other nonmetals ( $N_2$ ,  $O_2$ ,  $F_2$ , and Ne) are all gases.

## Various Uses of the Second Period Elements and Their Compounds

- Lithium is now extensively used to make batteries for laptop computers, cellphones, hybrid and electric cars, cameras and pacemakers. Among lithium compounds, LiCl is used as dehumidifier in air-conditioners and  $\text{Li}_2\text{CO}_3$  is used to treat treating patients suffering from manic-depression.
  - Beryllium is used in X-ray tube windows; to make alloys for springs and gears (especially those used in watches), and spark free tools.
  - The compounds of boron are used as cleaning agent (such as borax), eyewash antiseptic (boric acid), borosilicate glass;  $\text{B}_4\text{C}$  is used in armor; regulator in nuclear plants; composite for making tennis rackets.
  - Carbon is one of the primary elements in the molecules of life; it is the second most abundant element in the human body. It forms the skeletons of all organic compounds; Graphite is used as lubricant, to make strong structural fiber, in steel making, as pencil lead, manufacture of plastics. Diamond is used in jewelry, cutting tools, protective films.
  - Nitrogen is one of the primary element in the molecules of life, especially proteins and nucleic acids; used in the manufacture of ammonia, fertilizers, explosive, rocket fuel, nitric acid; nitric oxide in atmosphere is responsible for acid rain and smog; nitrogen gas is used as inert blanket for electronic equipments; liquid nitrogen is used as refrigerant and in cryogenic surgery.
  - Oxygen is the most abundant element in the Earth's crust; oxygen gas is essential for life; it is needed for combustion; used in water purification and in steel production; water is 88.9% oxygen (by mass) and it is essential for life.
  - Fluorine compounds are used as additives in toothpaste; for making Teflon; in uranium enrichment, and CFCs is used as refrigerant, but causes the depletion of ozone layer.
  - Neon is used in elemental form primarily in neon lighting for advertising signs, in fog lamps, TV and monitor tubes, laser, and voltage detectors.
-