

# CHAPTER 6: THE PERIODIC TABLE

Problems: 5,9,11,13,17,27,29,31,33,35,41,43,47,49,59,61,63,65,67,69, 75,77,79

## 6.1 CLASSIFICATION OF THE ELEMENTS

Dimitri Mendeleev (1869) arranged elements in a table in order of increasing atomic mass.

⇒ Elements with similar properties were grouped together in the same column or row.

⇒ Mendeleev even left some holes in table for elements that had not yet been discovered.

## 6.2 THE PERIODIC LAW CONCEPT

H. G.J. Moseley arranged elements in order of **increasing atomic number**,  $Z$ . (recall that  $Z$  = # of protons!)

**Periodic Law** – properties occur in repeating pattern when elements are arranged according to increasing atomic number.

- *Atomic mass usually increases when the number of protons increases* **except for Ni & Co, Ar & K, Te & I.**

Neils Bohr's introduction of electron energy levels altered the **shape of the Periodic Table**.

- The modern Periodic Table indicates filling of s, p, d and f sublevels.

## 6.3 GROUPS AND PERIODS OF ELEMENTS

A **horizontal row** is called a period or series. (There are 7 rows = 7 periods.)

A **vertical column** is called a group or family ⇒ elements in the same group exhibit similar properties.

### Representative Elements: Group A Elements (elements in Groups IA-VIIIA)

Group IA: alkali metals

Group IIA: alkaline earth metals

Group VIIA: halogens

Group VIIIA: noble gases

### Transition Metals: Group B elements (in middle of the Periodic Table)

⇒ The behavior and properties of transition metals is not very predictable.

### Inner Transition Elements (beneath the main body of Periodic Table)

**lanthanide series:** Ce-Lu, also called rare earth metals

**actinide series:** Th-Lr, radioactive elements that exist for only very short periods of time before decaying to other elements

**transuranium** – all elements with  $Z > 93$  are man-made in particle accelerators.

Example: Give the element symbol that matches each description below:

a. The noble gas in the third period: \_\_\_\_\_

b. The halogen in the fourth period: \_\_\_\_\_

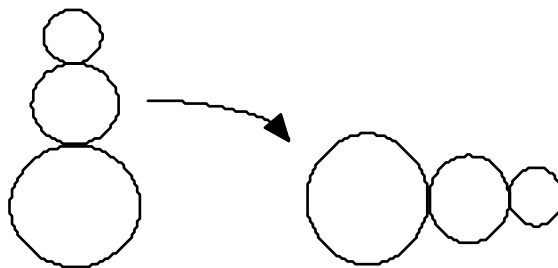
## 6.4 PERIODIC TRENDS

**Atomic Size:** distance from nucleus to outermost electrons

- ◇ **Increases down a group**  
⇒ As we go ↓, there are more energy levels → bigger radius
- ◇ **Decreases across a period (from left to right)**  
⇒ As we go →, the nuclear charge increases since there are more protons in the nucleus.  
As nuclear charge increases, the electrons are pulled closer, so the atom is smaller.

Trend from top to bottom → like a snowman

Trend from left to right → like a snowman  
that fell to the right



**Metallic Character:**

- ◇ **Decreases across a period (from left to right)**  
⇒ metals are concentrated on left-hand side of Periodic Table
- ◇ **Increases down a group**  
⇒ For groups IVA and VA, we go from nonmetals (C & N) at the top to metals (Sn & Bi) at the bottom.

★ **Same snowman trends as for atomic radius!**

## 6.5 PROPERTIES OF ELEMENTS

One can predict physical properties such as the density, MP, atomic mass, and the atomic radius for an element in a group from examining trends for other elements in the group.

Elements in the same group typically have similar chemical behavior and form similar compounds.

⇒ For example, Na reacts with  $\text{Cl}_2$  to form NaCl and K reacts with  $\text{Cl}_2$  to form KCl.

Ex. 1: Sodium reacts violently with water. Which element below is also likely to react violently with water?

K      Mg      Al      Si      P      Li      Cl      Ar

Ex. 2: Nitrogen reacts with hydrogen to form  $\text{NH}_3$ . Give the formula for the compound that forms when phosphorus reacts with hydrogen.

## 6.7 VALENCE ELECTRONS

**core electrons:** innermost electrons belonging to completely *filled electron shells*

**valence electrons:** s and p electrons in the *outermost shell*

⇒ valence electrons form chemical bonds between atoms

⇒ valence electrons dictate the properties and chemical behavior of an element

For Group A elements: **Group # = # of valence electrons**

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
# valence e <sup>-</sup>	1	2	3	4	5	6	7	8

Example. Indicate the number of valence electrons for each of the following:

B \_\_\_\_ H \_\_\_\_ O \_\_\_\_ N \_\_\_\_ C \_\_\_\_ F \_\_\_\_

## 6.8 ELECTRON DOT FORMULAS

**Electron Dot Formula:** shows the symbol of an element surrounded by valence electrons.

- ⇒ **Element symbol** represents the nucleus and inner core electrons
- ⇒ **Dots** represent the valence e<sup>-</sup>

### Rules for writing Electron Dot Formulas

1. Write down the element symbol
2. Determine the number of valence electrons from the group number
3. Assume the atom has four sides, and distribute electrons *with one electron per side before pairing electrons*. **There are a maximum of 2 e<sup>-</sup> on each side!**

Example. Write the electron dot formula for each of the following:

boron	nitrogen	silicon	oxygen
fluorine	sodium	calcium	argon

## 6.9 IONIZATION ENERGY

**Ionization energy:** amount of energy required to remove an electron from a neutral atom.

### Ionization energy trends

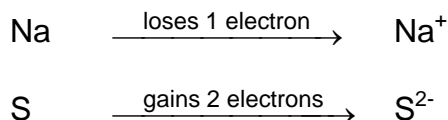
- ◇ **IE decreases down a group**
  - ⇒ easier to remove electrons from a larger atom (e<sup>-</sup>'s are further from the nucleus)
- ◇ **IE increases across a period (from left to right)**
  - ⇒ Elements with fewer (1–3) valence electrons have lower IE's because they want to **give up** electrons to attain the noble gas configuration (greater stability).
  - ⇒ Nonmetal elements on the right side of the P.T. have higher IE's because they don't want to give up electrons. Group VA-VIIA elements need to **gain** electrons to attain the noble gas configuration.

**\*These trends are the opposite of the snowman trends!**

## 6.10 IONIC CHARGES

Representative (Group A) elements gain or lose electrons and form ions to achieve a noble gas electron configuration, so they can be stable.

Example: Indicate the number of protons and electrons for the following:



**Metals lose electrons** from their valence shell

⇒ **positively charged ions = cations**

Group	Group IA metals	Group IIA metals	Group IIIA metals
Charge	+1	+2	+3
Example	Li <sup>+</sup>	Mg <sup>2+</sup>	Al <sup>3+</sup>

**Nonmetals gain electrons**, adding electrons to their valence shell

⇒ **negatively charged ions = anions**

Group	Group VA nonmetals	Group VIA nonmetals	Group VIIA nonmetals
Charge	-3	-2	-1
Example	N <sup>3-</sup>	O <sup>2-</sup>	F <sup>-</sup>

- ◆ Charges on ions can be shown with the sign before or after the number—e.g. magnesium ion can be shown as Mg<sup>+2</sup> or Mg<sup>2+</sup>.
- ◆ The 1 is omitted for ions with a +1 or -1 charge

Example: Give the **formula for the ion** formed by each of the following:

chlorine: \_\_\_\_\_ calcium: \_\_\_\_\_ phosphorus: \_\_\_\_\_

sodium: \_\_\_\_\_ oxygen: \_\_\_\_\_ aluminum: \_\_\_\_\_

**isoelectronic:** two or more species that have the same number of electrons and the same configuration.

### Electron Configuration of ions

+ charge means electrons have been lost, so we subtract the number of electrons given by its ionic charge

⇒ move to the left on the periodic table for electrons lost

-charge means electrons have been gained, so we add the number of electrons given by its ionic charge.

⇒ Move to the right on the periodic table for electrons gained

Use the P.T. to find the isoelectronic element (usually a noble gas), then give the electron configuration for that element.

**Note: Some ions are not isoelectronic with a noble gas (e.g. Sn or Pb).**

Example 1: Fill in the blanks for the following ions:

Ion	Isoelectronic with which Noble gas?	Electron Configuration (using Noble gas core notation)
Na <sup>+</sup>		
P <sup>3-</sup>		
Ti <sup>4+</sup>		
Ba <sup>2+</sup>		
Al <sup>3+</sup>		
Br <sup>-</sup>		

**Note that the Noble gas the ion is isoelectronic with is the same Noble gas used in the core notation!**

Example 2: Circle all the ions below that are isoelectronic with argon:

