Electron Configuration

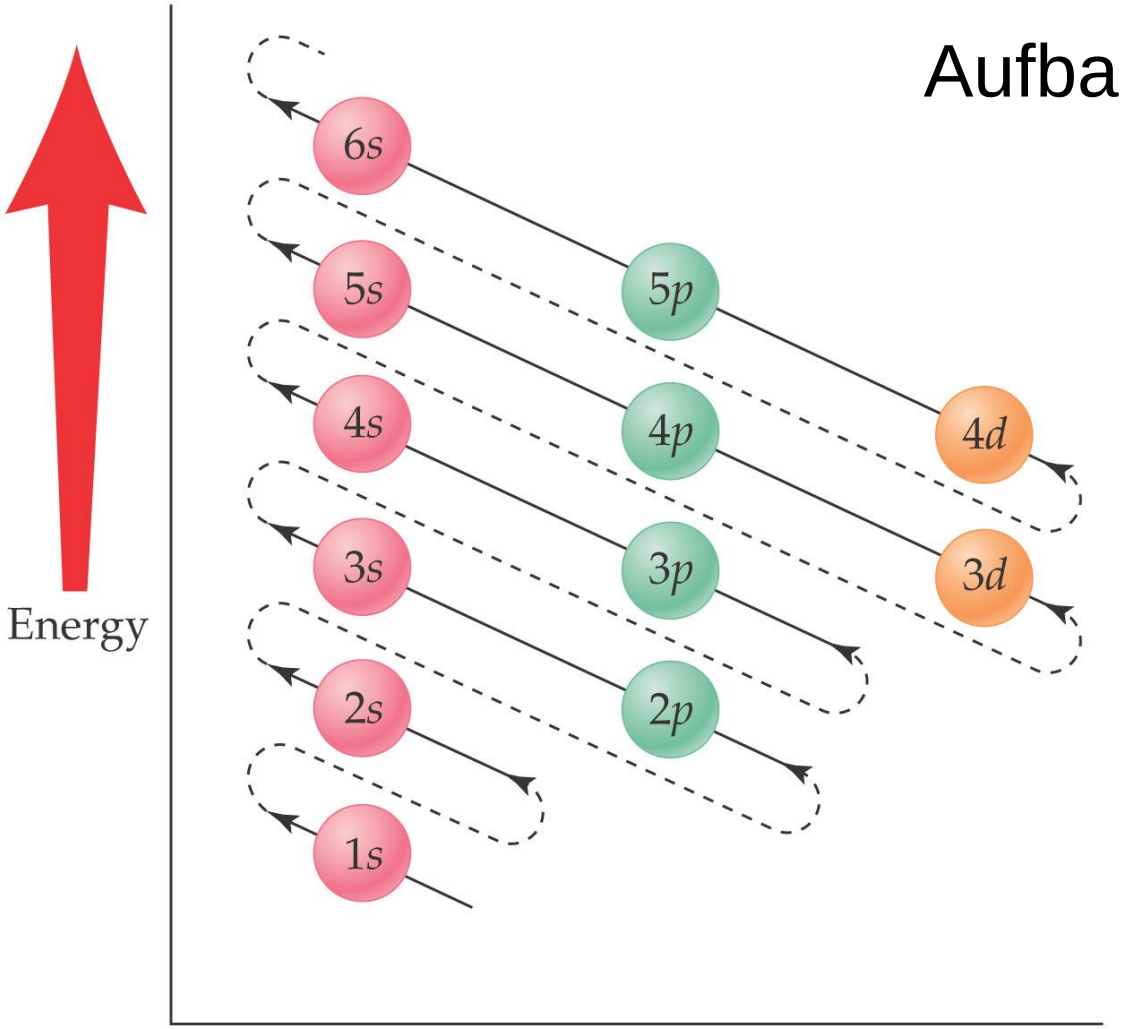
Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Na: [Ne] $3s^1$

Electron Configurations

- Electron configurations tells us in which orbitals the electrons for an element are located.
- Three rules:
  - **electrons fill orbitals starting with lowest n and moving upwards;**
  - **no two electrons can fill one orbital with the same spin (Pauli);**
  - **for degenerate orbitals, electrons fill each orbital singly before any orbital gets a second electron (Hund's rule).**

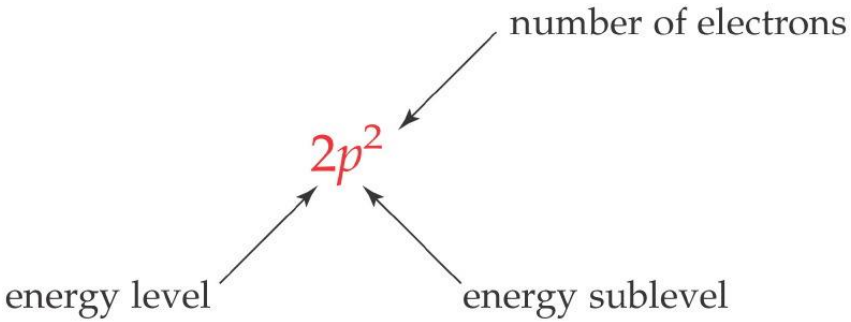
Filling Diagram for Sublevels



Aufbau Principle

Electron Configurations

- The ***electron configuration*** of an atom is a shorthand method of writing the location of electrons by sublevel.
- The sublevel is written followed by a superscript with the number of electrons in the sublevel.
  - If the $2p$ sublevel contains 2 electrons, it is written $2p^2$



Writing Electron Configurations

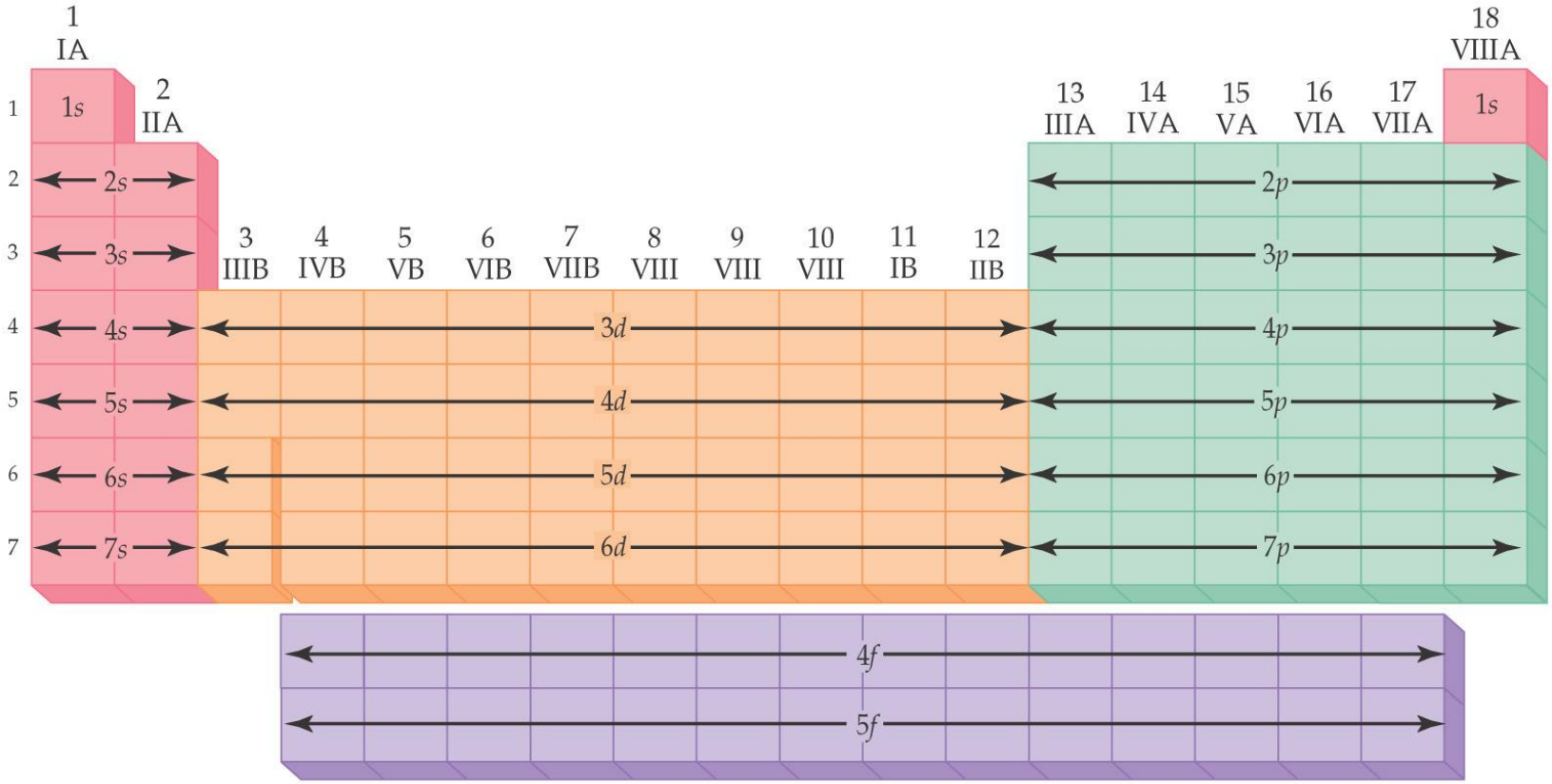
- First, determine how many electrons are in the atom. Iron has 26 electrons.
- Arrange the energy sublevels according to increasing energy:
  - $1s \;\; 2s \;\; 2p \;\; 3s \;\; 3p \;\; 4s \;\; 3d \ldots$
- Fill each sublevel with electrons until you have used all the electrons in the atom:
  - Fe: $1s^2 \;\; 2s^2 \;\; 2p^6 \;\; 3s^2 \;\; 3p^6 \;\; 4s^2 \;\; 3d^6$
- The sum of the superscripts equals the atomic number of iron (26)

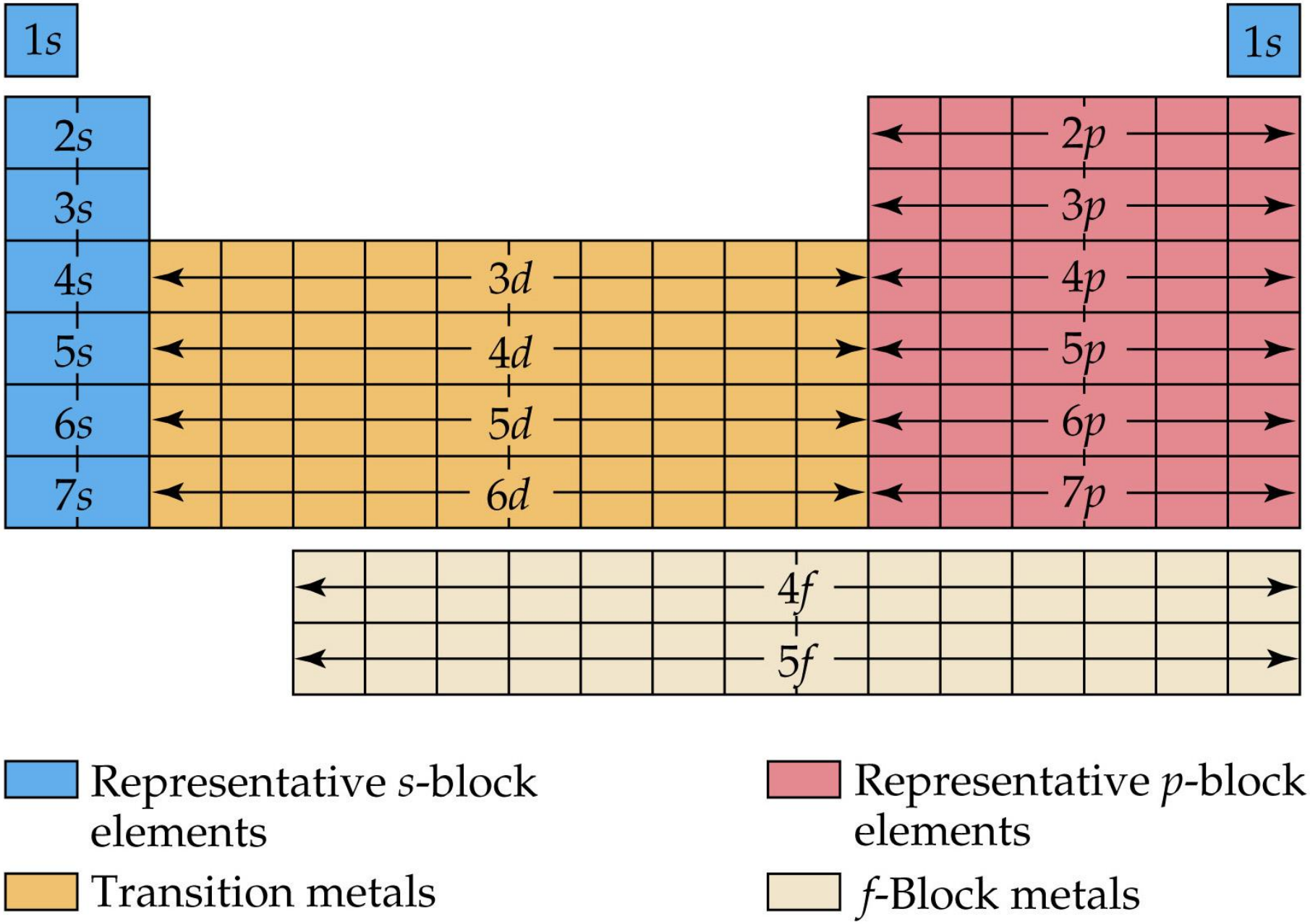
Electron Configurations and the Periodic Table

- The periodic table can be used as a guide for electron configurations.
- The period number is the value of *n*.
- Groups 1A and 2A have the *s*-orbital filled.
- Groups 3A - 8A have the *p*-orbital filled.
- Groups 3B - 2B have the *d*-orbital filled.
- The lanthanides and actinides have the *f*-orbital filled.

Blocks and Sublevels

- We can use the periodic table to predict which sublevel is being filled by a particular element.




1s
1s
2s
3s
4s
5s
6s
7s
3d
4d
5d
6d
2p
3p
4p
5p
6p
7p
4f
5f
Representative s-block elements
Representative p-block elements
Transition metals
f-Block metals

Noble Gas Core Electron Configurations

- Recall, the electron configuration for Na is:

  Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

- We can abbreviate the electron configuration by indicating the innermost electrons with the symbol of the preceding noble gas.

- The preceding noble gas with an atomic number less than sodium is neon, Ne. We rewrite the electron configuration:

  Na: [Ne] $3s^1$

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
Core	1 **H** $1s^1$																	2 **He** $1s^2$
[He]	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^2 2p^1$	6 **C** $2s^2 2p^2$	7 **N** $2s^2 2p^3$	8 **O** $2s^2 2p^4$	9 **F** $2s^2 2p^5$	10 **Ne** $2s^2 2p^6$
[Ne]	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^2 3p^1$	14 **Si** $3s^2 3p^2$	15 **P** $3s^2 3p^3$	16 **S** $3s^2 3p^4$	17 **Cl** $3s^2 3p^5$	18 **Ar** $3s^2 3p^6$
[Ar]	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $3d^1 4s^2$	22 **Ti** $3d^2 4s^2$	23 **V** $3d^3 4s^2$	24 **Cr** $3d^5 4s^1$	25 **Mn** $3d^5 4s^2$	26 **Fe** $3d^6 4s^2$	27 **Co** $3d^7 4s^2$	28 **Ni** $3d^8 4s^2$	29 **Cu** $3d^{10} 4s^1$	30 **Zn** $3d^{10} 4s^2$	31 **Ga** $3d^{10} 4s^2 4p^1$	32 **Ge** $3d^{10} 4s^2 4p^2$	33 **As** $3d^{10} 4s^2 4p^3$	34 **Se** $3d^{10} 4s^2 4p^4$	35 **Br** $3d^{10} 4s^2 4p^5$	36 **Kr** $3d^{10} 4s^2 4p^6$
[Kr]	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $4d^1 5s^2$	40 **Zr** $4d^2 5s^2$	41 **Nb** $4d^3 5s^2$	42 **Mo** $4d^5 5s^1$	43 **Tc** $4d^5 5s^2$	44 **Ru** $4d^7 5s^1$	45 **Rh** $4d^8 5s^1$	46 **Pd** $4d^{10}$	47 **Ag** $4d^{10} 5s^1$	48 **Cd** $4d^{10} 5s^2$	49 **In** $4d^{10} 5s^2 5p^1$	50 **Sn** $4d^{10} 5s^2 5p^2$	51 **Sb** $4d^{10} 5s^2 5p^3$	52 **Te** $4d^{10} 5s^2 5p^4$	53 **I** $4d^{10} 5s^2 5p^5$	54 **Xe** $4d^{10} 5s^2 5p^6$
[Xe]	55 **Cs** $6s^1$	56 **Ba** $6s^2$	71 **Lu** $4f^{14} 5d^1 6s^2$	72 **Hf** $4f^{14} 5d^2 6s^2$	73 **Ta** $4f^{14} 5d^3 6s^2$	74 **W** $4f^{14} 5d^4 6s^2$	75 **Re** $4f^{14} 5d^5 6s^2$	76 **Os** $4f^{14} 5d^6 6s^2$	77 **Ir** $4f^{14} 5d^7 6s^2$	78 **Pt** $4f^{14} 5d^9 6s^1$	79 **Au** $4f^{14} 5d^{10} 6s^1$	80 **Hg** $4f^{14} 5d^{10} 6s^2$	81 **Tl** $4f^{14} 5d^{10} 6s^2 6p^1$	82 **Pb** $4f^{14} 5d^{10} 6s^2 6p^2$	83 **Bi** $4f^{14} 5d^{10} 6s^2 6p^3$	84 **Po** $4f^{14} 5d^{10} 6s^2 6p^4$	85 **At** $4f^{14} 5d^{10} 6s^2 6p^5$	86 **Rn** $4f^{14} 5d^{10} 6s^2 6p^6$
[Rn]	87 **Fr** $7s^1$	88 **Ra** $7s^2$	103 **Lr** $5f^{14} 6d^1 7s^2$	104 **Rf** $5f^{14} 6d^2 7s^2$	105 **Db** $5f^{14} 6d^3 7s^2$	106 **Sg** $5f^{14} 6d^4 7s^2$	107 **Bh** $5f^{14} 6d^5 7s^2$	108 **Hs** $5f^{14} 6d^6 7s^2$	109 **Mt** $5f^{14} 6d^7 7s^2$	110	111	112		114		116		

Lanthanide series [Xe]	57 **La** $5d^1 6s^2$	58 **Ce** $4f^1 5d^1 6s^2$	59 **Pr** $4f^3 6s^2$	60 **Nd** $4f^7 6s^2$	61 **Pm** $4f^5 6s^2$	62 **Sm** $4f^6 6s^2$	63 **Eu** $4f^7 6s^2$	64 **Gd** $4f^7 5d^1 6s^2$	65 **Tb** $4f^9 6s^2$	66 **Dy** $4f^{10} 6s^2$	67 **Ho** $4f^{11} 6s^2$	68 **Er** $4f^{12} 6s^2$	69 **Tm** $4f^{13} 6s^2$	70 **Yb** $4f^{14} 6s^2$
Actinide series [Rn]	89 **Ac** $6d^1 7s^2$	90 **Th** $6d^2 7s^2$	91 **Pa** $5f^2 6d^1 7s^2$	92 **U** $5f^3 6d^1 7s^2$	93 **Np** $5f^4 6d^1 7s^2$	94 **Pu** $5f^6 7s^2$	95 **Am** $5f^7 7s^2$	96 **Cm** $5f^7 6d^1 7s^2$	97 **Bk** $5f^9 7s^2$	98 **Cf** $5f^{10} 7s^2$	99 **Es** $5f^{11} 7s^2$	100 **Fm** $5f^{12} 7s^2$	101 **Md** $5f^{13} 7s^2$	102 **No** $5f^{14} 7s^2$

Metals Metalloids Nonmetals

Electron Configurations

Condensed Electron Configurations

- Neon completes the 2*p* subshell.
- Sodium marks the beginning of a new row.
- So, we write the condensed electron configuration for sodium as

$$\text{Na: [Ne] } 3s^1$$

- [Ne] represents the electron configuration of neon.
- **Core electrons:** electrons in [Noble Gas].
- **Valence electrons:** electrons outside of [Noble Gas].

Valence Electrons

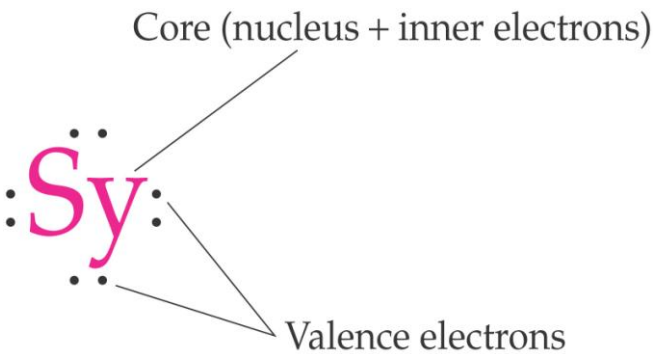
- When an atom undergoes a chemical reaction, only the outermost electrons are involved.
- These electrons are of the highest energy and are furthest away from the nucleus. These are the ***valence electrons***.
- The valence electrons are the *s* and *p* electrons beyond the noble gas core.

Predicting Valence Electrons

- The Roman numeral in the American convention indicates the number of valence electrons.
  - Group IA elements have 1 valence electron
  - Group VA elements have 5 valence electrons
- When using the IUPAC designations for group numbers, the last digit indicates the number of valence electrons.
  - Group 14 elements have 4 valence electrons
  - Group 2 elements have 2 valence electrons

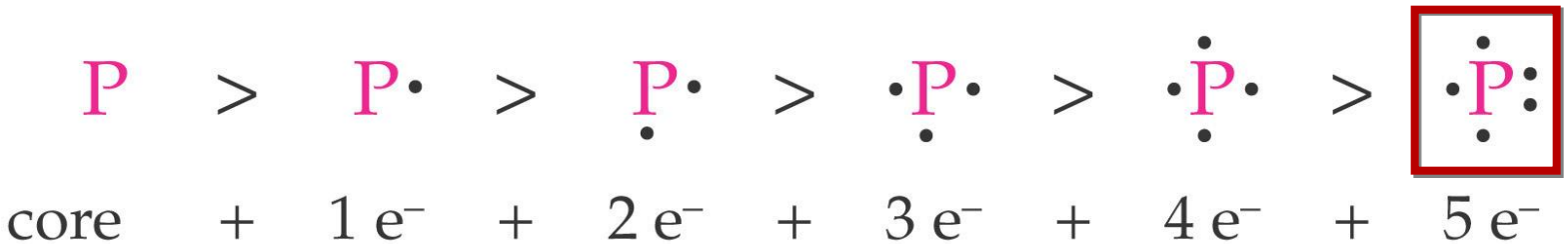
Electron Dot Formulas

- An electron dot formula of an elements shows the symbol of the element surrounded by its valence electrons.



- We use one dot for each valence electron.

- Consider phosphorous, P, which has 5 valence electrons. Here is the method for writing the electron dot formula.



Ionic Charge

- Recall, that atoms *lose or gain electrons* to form ions.
- The charge of an ion is related to the number of valence electrons on the atom.
- Group IA/1 metals lose their one valence electron to form 1+ ions.
  - $Na \rightarrow Na^{+} + e^{-}$
- Metals lose their valence electrons to form ions.

Predicting Ionic Charge

- Group IA/1 metals form 1+ ions, group IIA/2 metals form 2+ ions, group IIIA/13 metals form 3+ ions, and group IVA/14 metals from 4+ ions.
- By losing their valence electrons, they achieve a noble gas configuration.
- Similarly, nonmetals can gain electrons to achieve a noble gas configuration.
- Group VA/15 elements form -3 ions, group VIA/16 elements form -2 ions, and group VIIA/17 elements form -1 ions.

Ion Electron Configurations

- When we write the electron configuration of a positive ion, we remove one electron for each positive charge:

$$\mathrm{Na} \rightarrow \mathrm{Na}^{+}$$
$$1s^2\,2s^2\,2p^6\,3s^1 \rightarrow 1s^2\,2s^2\,2p^6$$

- When we write the electron configuration of a negative ion, we add one electron for each negative charge:

$$\mathrm{O} \rightarrow \mathrm{O}^{2-}$$
$$1s^2\,2s^2\,2p^4 \rightarrow 1s^2\,2s^2\,2p^6$$

Conclusions Continued

- We can Write the electron configuration of an element based on its position on the periodic table.
- Valence electrons are the outermost electrons and are involved in chemical reactions.
- We can write electron dot formulas for elements which indicate the number of valence electrons.

Conclusions Continued

- We can predict the charge on the ion of an element from its position on the periodic table.