Grade 12 Chemistry

Structure and Properties of Matter

Class 6

Condensed Notation

- For elements with many electrons, use a shorthand notation

1. Locate the closest noble gas in the previous row and write it in square brackets.
2. Write the rest of the electron configuration from that noble gas.

Ex: Mg = $1s^2 2s^2 2p^6 3s^2 \rightarrow [Ne]3s^2$

Ne = $1s^2 2s^2 2p^6$

Checkpoint

- Write the electron configuration and orbital filling diagram of:
  - Radium (Z=88)
  - Mercury (Z=80)

Anomalous Electron Configurations

- Consider the expected electron configurations of Chromium and Copper
  - Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$
  - Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
- Actual Configuration
  - Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
  - Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
- The reason: half-filled or completely filled orbitals are particularly stable

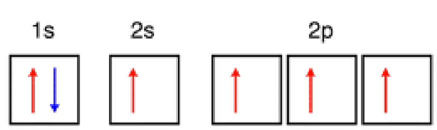
*These special cases occur with the transition metals

Electron Configurations of Ions

- Ion – atom that has an electric charge due to the presence or absence of electrons
  - Na^+ = Sodium with 1 less electron
  - O^{2-} = Oxygen with 2 more electrons
- To write the electronic configurations, add/subtract electrons from the exponent
  - Ex: F = $1s^2 2s^2 2p^5$ F^- = $1s^2 2s^2 2p^6$
- Since F^- and Ne have the same electronic configurations, they are called isoelectronic

Electron Configuration of Excited Electrons

- When electrons become excited, the electrons will move to higher orbitals

Excited Carbon: 

$1s^2 2s^1 2p^3$

- Collect the electrons and rewrite the electron configuration to find the element's ground state configuration

Checkpoint

- Write the electron configuration of:

  P^{3-} Sr^{+}

- Find the isoelectronic element of the above elements

- Write this element's ground state electron configuration:

  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^2$

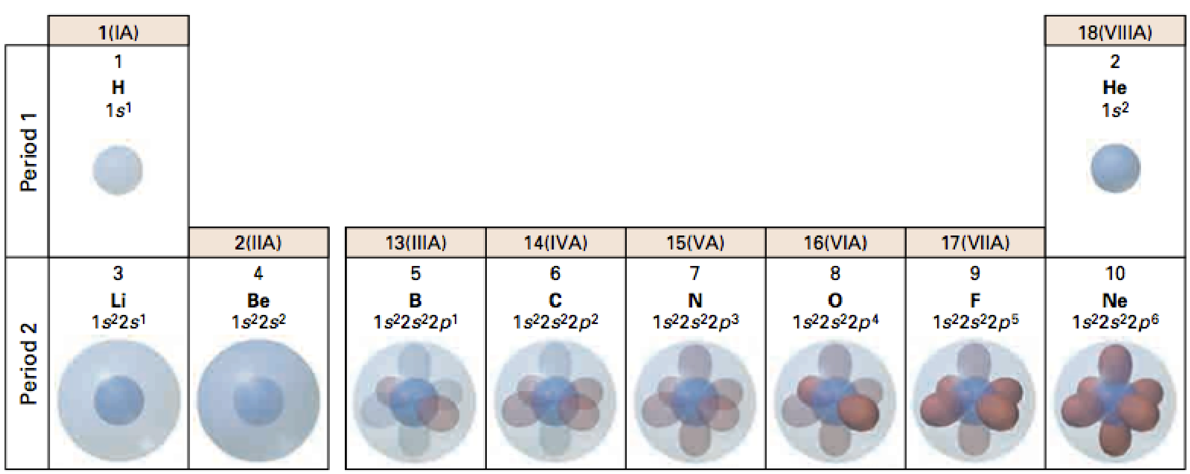
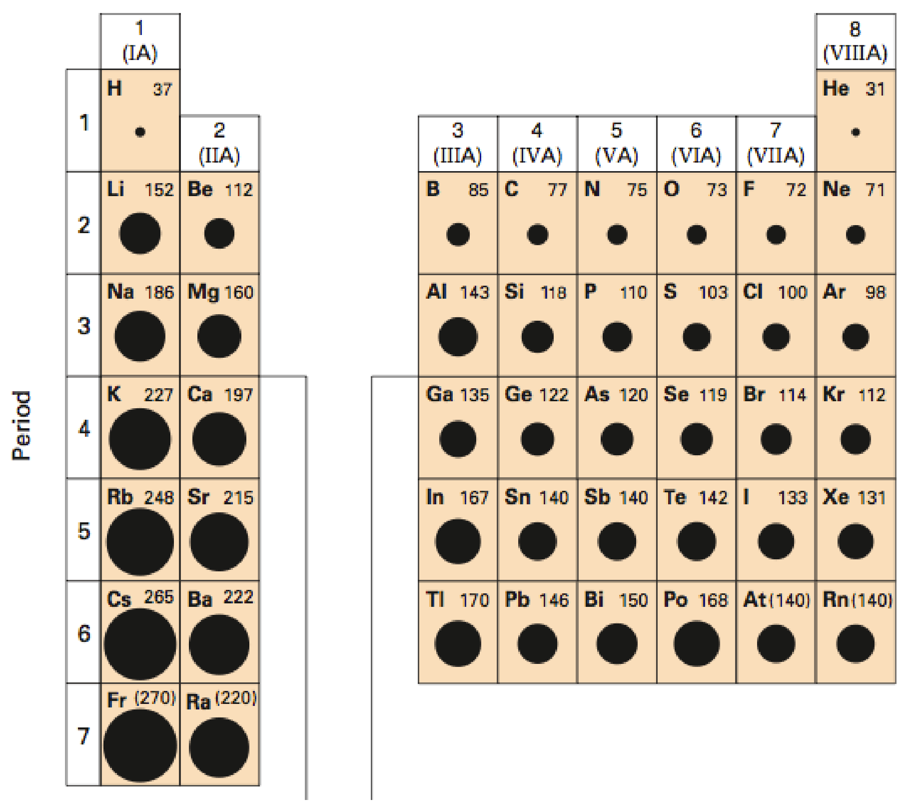
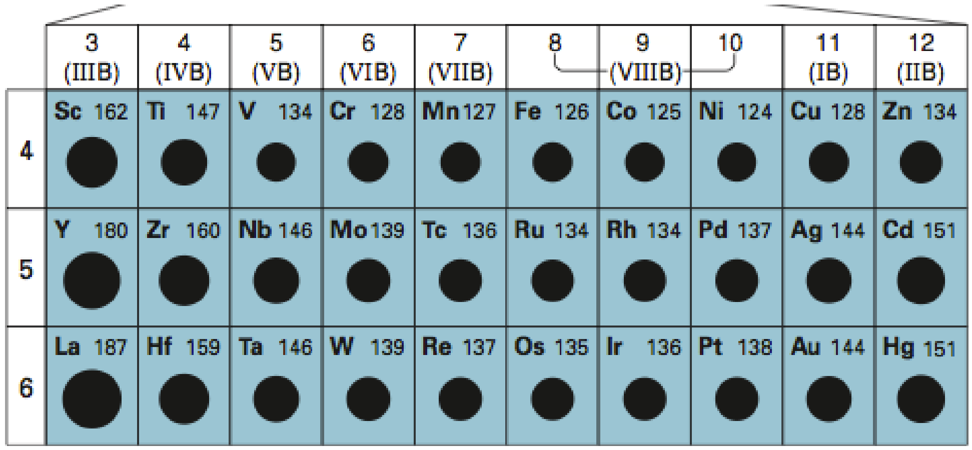


Figure 3.20 How orbitals are filled for atoms of the first 10 elements of the periodic table

Periodic Trends Explained

- **Atomic Radius**
  – Increases down a group
  – Decreases across a period
- Factors:
  – Changing *n*, as *n* increases, there is a higher probability of finding electrons farther from the nucleus
  – Changing the Effective Nuclear Charge (Z_{eff}) – net force of attraction between electrons and the nucleus; offset by some electron repulsion





- Atomic radius trend does not apply to the transition metals because electrons are added to the d-orbitals rather than the outer energy levels; Z_{eff} changes relatively little

- **Ionization Energy** – the energy required to remove one electron from its ground state
  - 1st ionization energy = least energy required
  - Successive ionization energies are always greater than the 1st ionization energy because electron must be removed from a positively charged ion
- Down a group: IE decreases
- Across a period: IE increases
- Variations occur between Group 2A and Group 3A; Group 5A and Group 6A, which can be explained with electron configuration

Checkpoint

Draw the orbital filling diagrams for each of the following and explain why they do not follow the trend of increasing ionization energy:

a) Be and B

a) P and S

- **Electron Affinity** – change in energy that occurs when an electron is added to a gaseous atom
  - 1^{st} electron affinity results in the formation of an anion with a charge of -1

1	2 (IIA)	13 (IIIA)	14 (IVA)	15 (VA)	16 (VIA)	17 (VIIA)	18 (VIIIA)
H −72.8							He (+21)
Li −59.6	Be (+241)	B −26.7	C −122	N 0	O −141	F −328	Ne (+29)
Na −52.9	Mg (+230)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+34)
K −48.4	Ca (+156)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+167)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+40)
Cs −45.5	Ba (+52)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

High negative numbers = High EA

Low negative numbers/positive numbers = Low EA

- Down a group: EA decreases
- Across a group: EA increases

- Noble Gases (Group 18) have very high ionization energies and very low electron affinities
- Variations occur between Group 1A and Group 2A; Group 4A and Group 5A

Checkpoint

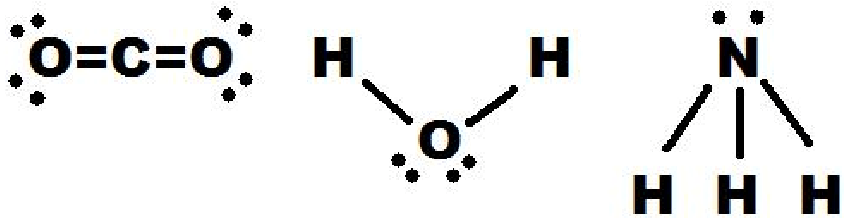
Draw the orbital filling diagrams for each of the following and explain why they do not follow the trend of increasing electron affinity:

a) Na and Mg

a) Ge and As

Chemical Bonds

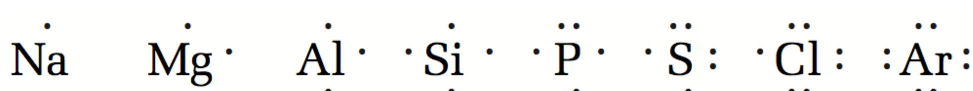
- Why do chemical bonds form?
  - Bonded atoms have lower energy than single, uncombined atoms



  - Bonds formed from the interaction of the valence electrons

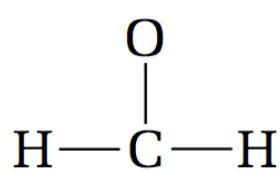
Lewis Structures

- Show the interaction of the valence electrons
  - Write the element symbol
  - Add dots around the element symbol to symbolize the atom's valence electrons
  - To find valence, look at the group's Roman Numeral

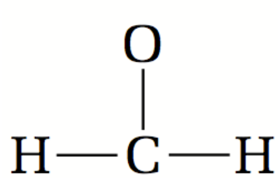


- How do you draw the Lewis structure for molecules and polyatomic ions?
  1. Position the least electronegative atom in the centre of the molecule or polyatomic ion. H and F are usually on the outside
  2. Write the other atoms around this central atom attached by a single bond.

Draw the Lewis Structure for Methanal (CH_2O)



3. Determine the total number of valence electrons in the molecule or ion.
   - Subtract valence electrons for cations
   - Add valence electrons for anions



(1 C atom × 4 e^-/C atom) + (1 O atom × 6 e^-/O atom)
+ (2 H atoms × 1 e^-/H atom)
= 4 e^- + 6 e^- + 2 e^-
= 12 e^-

4. Determine the total number of electrons needed for each atom to achieve a noble gas configuration

C = 8 e^- O = 8 e^- H = 2 e^- (x2) = 20 e^-

5. Subtract the valence electrons from the total number of electrons to find the number of shared electrons

   $20\ e^- - 12e^- = 8\ e^-$

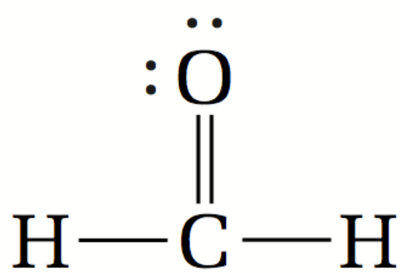
6. Divide the number of shared electrons by 2 to give the number of bonds

   $8\ e^- \div 2 = 4$ covalent bonds

– Double or triple bonds may be needed to account for the number of bonds
  - Double Bonds = 2 Bonds
  - Triple Bonds = 3 Bonds

7. Determine the number of non-bonding electrons by subtracting the number of shared electrons from the total number of valence electrons.

$12\ e^- - 8\ e^- = 4$ non-bonding electrons (2 lone-pairs)



Checkpoint

Draw a Lewis-structure for:

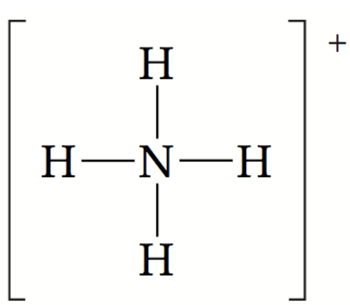
a) ClNO

b) H_2S

c) BrO^-

Coordinate Covalent Bonds

- Some atoms contribute both of their electrons – the bond formed is called a coordinate covalent bond
- Ex: NH_4^+

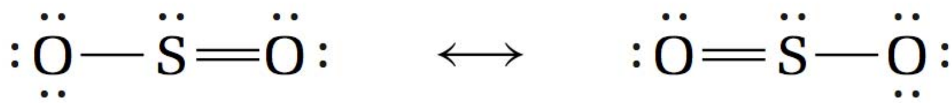


Lewis structure does not indicate which atom contributes both electrons.

Quantum model $[He]2s^22p^3$ shows that each nitrogen has three unpaired 2p electrons for bonding. Since there are four covalent bonds, electrons in one of the bonds must have come from the filled orbitals of nitrogen.

Resonance

- If you draw the Lewis structure of SO_2, there are two possible structures



- The actual structure is an average of the different structures – **resonance hybrid**
- Draw a double-ended arrow to indicate resonance

Checkpoint

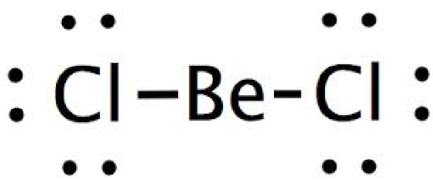
Draw a Lewis-structure for:

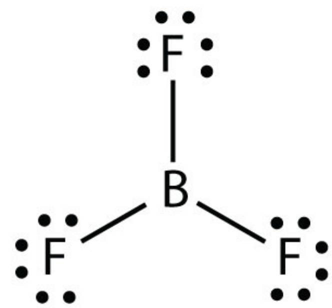
a) CO_3^{2-}

b) SO_3

Incomplete Octet

- Octet rules mainly applies to the second-period elements
- Some compounds are stable as incomplete octets





Checkpoint

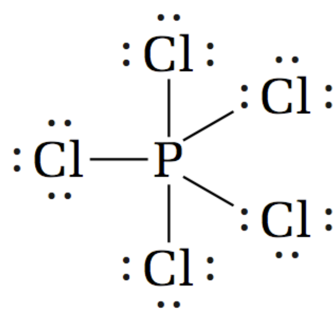
Draw a Lewis-structure for:

a) AlI_3

b) BeF_2

Expanded Valence Levels

- Octet rule allows four bonds to form around an atom but central atoms can bind with more than eight electrons
- Ex: PCl_5



Applies to atoms in the third period and beyond.

Larger atoms can accommodate additional valence electrons because of their size.

Checkpoint

Draw a Lewis-structure for:

a) SF_6

b) BrF_5

c) XeF_4