THE PERIODIC TABLE

On February 17, 1869, Dmitri Mendeleev came up with an arrangement of the elements in the order of increasing atomic mass in a single day. Because atomic mass "happens" to increase with atomic number (=number of protons) and hence the number of electrons).

- **Groups** (vertical column): group number follows the number of valence electrons (except for He);
  members have the same valence electronic configuration.

- **Periods** (horizontal row): period number follows the principal quantum number (n) or the shell number.

- **Blocks**:
  s and p blocks (n shell): main groups;
  d blocks (n-1 shell): transition elements;
  f blocks (n-2 shell): inner-transition elements - Lanthanides (4f), Actinides (5f)

- **Metals vs non-metals vs metalloids** (REF: Table 8.4, p.329).
  metals are luster;
  conduct heat or electricity;
  are malleable and ductile;
  lose electrons easily.

The periodic table summarizes the periodic recurrence of analogous ground-state electron configuration as the atomic number increases, which accounts for the periodic trend of the properties of atoms.

Atomic radius is half the distance between the centers of neighboring atoms (covalent bond-length); (REF Fig. 8.10, p. 322)

increases from top to bottom (with Z) in a group;
decreases from left to right (with Z) across a period.

Ionic radius is the contribution to the distance between neighboring ions in a solid ionic compound; (REF: Fig. 8.9, p. 323) e.g. $r(Mg^{2+})$ = 65 pm and $r(O^{2-})$ = 140 pm.

monatomic cations are smaller than parent atoms;
monatomic anions are larger than parent atoms.

Ionization energy is the energy required to remove an electron from a gaseous ground-state atom or ion [NB: 2nd ionization energy is more than 1st];

$$A\ (g) \longrightarrow A^{+}\ (g) + e^{-}\ (g) \qquad \Delta H_{ion}$$

decreases from top to bottom (with Z) in a group;
increases from left to right (with Z) across a period.

Note that this trend is opposite to the atomic radius.

Electron affinity is the enthalpy change involved in the addition of an electron to a ground-state atom or ion;

$$A\ (g) + e^{-}\ (g) \longrightarrow A^{-}\ (g) \qquad \Delta H_{EA}$$

increases toward the upper right corner.

EG Arrange (a) Mg^{2+} and Ca^{2+} and (b) O^{2-} and F^{-} in order of increasing ionic radius.

We note that monatomic ions follow the same patterns that neutral atoms follow. This means that the smaller member of the pair will be an ion of an element that lies towards the top right corner of the periodic table (i.e. further to the right in a period or higher in a group).

(a) Mg is above Ca in group II, so we have $r(Mg^{2+}) < r(Ca^{2+})$.
65 pm 99 pm

(b) F is to the right of O, so $r(F^{-}) < r(O^{2-})$.
136 pm 140 pm

EG Suggest a reason for the small decrease of ionization energy between nitrogen (1400 kJ/mol) and oxygen (1310 kJ/mol).

Here, we expect O to have a higher ionization energy than N because the outermost electron is closer to the nucleus in O (than in N) so the O nucleus is more strongly charged (higher Z_{eff}). When we have experimental result (given) disagrees with our expectation, we should examine our electronic configuration more closely. For the valence orbital diagram, we have:

	2s	$2p_x$	$2p_y$	$2p_z$
N (Z=7): [He] $2s^2 2p^3$	↑↓	↑	↑	↑
O (Z=8): [He] $2s^2 2p^4$	↑↓	↑↓	↑	↑

Here, we note that one of the 2p orbitals in O is full. The electrons in this orbital repel each other and therefore less energy is required to remove one of them. So, the ionization energy of O is less than N.