

1. Fill in the blanks:

Atom	A	B	C
Number of valence electrons	3	1	4
Number of core electrons	2	10	2
Neutral element	B	Na	C

2. Write the electron configurations for the following chemical species:

P: $[\text{Ne}] 3s^2 3p^3$

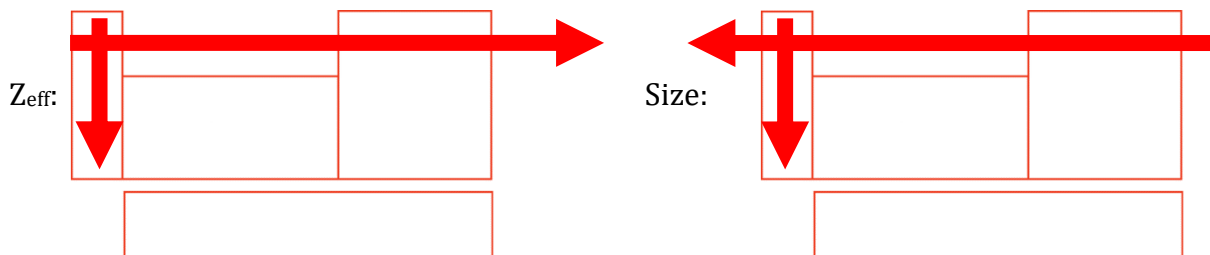
S^{2-} : $[\text{Ne}] 3s^2 3p^6$ or $[\text{Ar}]$

Se: $[\text{Ar}] 4s^2 3d^{10} 4p^4$

Mg^+ : $[\text{Ne}] 3s$

In: $[\text{Kr}] 5s^2 4d^{10} 5p$

3. Draw arrows to represent the trends of Z_{eff} and size on the periodic tables below.



Z_{eff} and size increase in the direction of the arrows.

4. Calculate the atomic radius, in Angstroms, of silicon (one of the most important elements in the microelectronics industry), given the following information:

- the H-H bond distance in H_2 is 0.74 \AA
- the Si-H bond distance in SiH_4 is 1.46 \AA

(If needed, $1 \text{ \AA} = 1 \times 10^{-10} \text{ m}$) Bond distance is sum of atomic radii of atoms in bond.

$$\text{Atomic radius of H} = \frac{0.74 \text{ \AA}}{2} = 0.37 \text{ \AA}$$

$$\text{Atomic radius of Si} = 1.46 \text{ \AA} - 0.37 \text{ \AA} = 1.09 \text{ \AA}$$

5. Rationalize the difference in the atomic and ionic radii of the following species:

a) Se (117 pm) and Se^{2-} (198 pm)

Se^{2-} has the same Z as Se, however, Se^{2-} has two more electrons in the valence shell. The electrons in the valence shell will experience greater electron-electron repulsion causing the Selenium anion (Se^{2-}) to have a larger ionic radius than neutral Selenium.

b) K (231 pm) and K^+ (133 pm)

K^+ has lost an electron compared to neutral K. The ionic radius of K^+ will be determined by the size of the 3rd shell. The atomic radius of neutral K will be determined by the size of the fourth shell. Because the third shell has a smaller radius than the fourth shell, K^+ is smaller than K.

6. Using Al as an example, write equations for the processes corresponding to the ionization energy and the electron affinity, and in each case indicate whether energy is absorbed or released.

Ionization energy: $\text{Al(g)} \rightarrow \text{Al}^+(\text{g}) + \text{e}^-$ **Energy is absorbed**

Electron affinity: $\text{Al(g)} + \text{e}^- \rightarrow \text{Al}^-(\text{g})$ **Energy is released**

7. **Good Question.** Explain why elements with very large ionization energies also tend to have highly exothermic electron affinities.

This trend is not general since the noble gases have the largest ionization energies but have no electron affinities. Elements with both high IE and EA are clustered in the top right hand corner of the periodic table, and have empty or partially empty p sub-shells that accommodate an extra electron. Such elements have relatively large nuclear charges acting on orbitals with small n (thus being smaller and closer to the nucleus). This creates a strong attraction for electrons that causes IE to be large. These same attractions though favour addition of electrons to the element, provided they are added to the same sub-shell (for example, the same valence p orbital). This in turn is energetically favourable, producing highly energetic EA values.

8. It is sometimes said, “A noble gas configuration is especially stable so atoms or ions with such a configuration are energetically favourable.” This would imply, for example, that Na^+ is more stable than Na, or that N^{3-} is more stable than N. Comment on the validity of this statement using arguments based upon content in this Unit.

All neutral atoms have a positive ionization energy which means it is never energetically favourable to remove an electron. So, Na^+ is not lower in energy than Na even though it has a noble gas electron configuration. Similarly, for some elements, it IS energetically favourable to gain an electron. For example, Cl^- has a lower energy than Cl (since Cl has a measured EA value) and in this particular case Cl^- does adopt a noble gas configuration. However, N^{3-} is highly energetically unfavourable even though it has a noble gas configuration. In fact, N^- alone is energetically unfavourable since N does not have a measured EA value. Because of electron repulsion, atoms usually do not adopt negative charge states above -1. This shows the statement above must be interpreted carefully and one should not conflate “stable”, “chemically inert”, and “energetically favourable”.

9. **Good Question.** Arrange the following species in order of increasing electron density on each oxygen atom. O_2 , CaO , OF_2 , CO_2 . (*Hint: think about what type of bonds are present*).



10. Rationalize the trend in lattice energies for the compounds given in the table.

Compound	Lattice Energy (kJ/mol)
NaCl	787
NaF	923
CaCl_2	2528
MgO	3791
CaO	3401

In general, predicting lattice energies is difficult because compounds can adopt different crystal structures. However, electrostatic interactions usually are the driving force which, according to Coulomb's Law, make the lattice energy greater as the distance between ions decreases, or the product of the ion charges, q_1q_2 , increases. So, ionic compounds formed from singly-charged ions (NaCl, NaF) should have the smallest lattice energies with $\text{NaF} > \text{NaCl}$ because F^- is smaller than Cl^- (charges are closer together in NaF than NaCl). Compounds with both ions doubly-charged (MgO, CaO) should have large lattice energies with $\text{MgO} > \text{CaO}$ because Mg^{2+} is smaller than Ca^{2+} . Finally, CaCl_2 has doubly-charged cations and singly-charged anions. It is between the two cases discussed above, as is its lattice energy. Even though there are twice as many Cl^- ions, their total contribution to the electrostatic energy is not as great as having fewer ions with greater charge (i.e. q_1q_2 is smaller).