CHEM 101

Chapter 8 Problems

8.25 A M^{2+} ion derived from a metal in the first transition metal series has four electrons in the 3*d* subshell.
What element might M be?

Transition metals lose electrons from the *ns* valence subshell before they are lost from the (*n* - 1)*d* subshell. Since this metal ion has a +2 charge and the 4*s* subshell is less stable than the 3*d*, two electrons have been removed from the 4*s*. For the neutral atom there are only six valence electrons. Therefore, the electron configuration of the neutral atom is $[Ar]4s^2 3d^4$ The element can be identified as **Cr** (**chromium**)

8.27 Write the ground-state electron configurations of the following ions: (a) Li^+, (b) H^-, (c) N^{3-}, (d) F^-, (e) S^{2-}, (f) Al^{3+}, (g) Se^{2-}, (h) Br^-, (i) Rb^+, (j) Sr^{2+}, (k) Sn^{2+}, (l) Te^{2-}, (m) Ba^{2+}, (n) Pb^{2+}, (o) In^{3+}, (p) Tl^+, (q) Tl^{3+}.

(a) $1s^2$

(b) $1s^2$

(c) $1s^2 2s^2 2p^6$

(d) $1s^2 2s^2 2p^6$

(e) $[Ne]3s^2 3p^6$

(f) $[Ne]$

(g) $[Ar]4s^2 3d^{10} 4p^6$

(h) $[Ar]4s^2 3d^{10} 4p^6$

(i) $[Kr]$

(j) $[Kr]$

(k) $[Kr]5s^2 4d^{10}$

(l) $[Kr]5s^2 4d^{10} 5p^6$

(m) $[Xe]$

(n) $[Xe]6s^2 4f^{14} 5d^{10}$

(o) $[Kr]5d^{10}$

(p) $[Xe]6s^2 4f^{14} 5d^{10}$

(q) $[Xe]4f^{14} 5d^{10}$

8.30 Name the ions with +3 charges that have the following electron configurations: (a) $[Ar]3d^3$, (b) $[Ar]$, (c) $[Kr]4d^6$, (d) $[Xe]4f^{14}5d^6$.

(a) Cr^{3+} **(b)** Sc^{3+} **(c)** Rh^{3+} **(d)** Ir^{3+}

8.31 Which of the following species are isoelectronic with each other? C, Cl^-, Mn^{2+}, B^-, Ar, Zn, Fe^{3+}, Ge^{2+}.

Two species are isoelectronic if they have the same number of electrons. Can two neutral atoms of different elements be isoelectronic?

(a) C and B^- are isoelectronic. **(b)** Mn^{2+} and Fe^{3+} are isoelectronic.

(c) Ar and Cl^- are isoelectronic. **(d)** Zn and Ge^{2+} are isoelectronic.

8.37 On the basis of their positions in the periodic table, select the atom with the larger atomic radius in each of the following pairs: (a) Na, Cs; (b) Be, Ba; (c) N, Sb; (d) F, Br; (e) Ne, Xe.

(a) Cs is larger. It is below Na in Group 1A.

(b) Ba is larger. It is below Be in Group 2A.

(c) Sb is larger. It is below N in Group 5A.

(d) Br is larger. It is below F in Group 7A.

(e) Xe is larger. It is below Ne in Group 8A.

8.38 Arrange the following atoms in order of decreasing atomic radius: Na, Al, P, Cl, Mg.

Na > Mg > Al > P > Cl

8.39 Which is the largest atom in Group 4A?

Pb

8.40 Which is the smallest atom in Group 7A?

F

8.41 Why is the radius of the Li considerably larger than the radius of the H?

The electron configuration of lithium is $1s^2 2s^1$. The two $1s$ electrons shield the $2s$ electron effectively from the nucleus. Consequently, the lithium atom is considerably larger than the hydrogen atom.

- **8.51** Arrange the following in order of increasing first ionization energy Na, Cl, Al, S, and Cs

$$\mathbf{Cs < Na < Al < S < Cl}$$

- **8.52** Arrange the following in order of increasing first ionization energy F, K, P, Ca, and Ne.

$$\mathbf{K < Ca < P < F < Ne}$$

8.54 In general, ionization energy increases from left to right across a given period. Al, however, has a lower ionization energy than Mg. Explain.

The Group 3A elements (such as Al) all have a single electron in the outermost *p* subshell, which is well shielded from the nuclear charge by the inner electrons and the ns^2 electrons. Therefore, less energy is needed to remove a single *p* electron than to remove a paired *s* electron from the same principal energy level (such as for Mg).

8.55 The first and second ionization energies of K are 419 KJ/mol and 3052 kJ/mol, and those of Ca are 590 kJ/mol and 1145 kJ/mol, respectively. Compare their values and comments on differences

To form the +2 ion of Ca, it is only necessary to remove two valence electrons. For K, however, the second electron must come from the atom's noble gas core which accounts for the much higher second ionization energy.

8.56 Two atoms have the electron configurations $1s^22s^22p^6$ and $1s^22s^22p^63s^1$. The first ionization energy of one is 2080 kJ/mol, and that of the other is 496 kJ/mol.
Match each ionization energy with one of the given electron configurations. Justify your choice.

Solution: The lone electron in the 3*s* orbital will be much easier to remove. This lone electron is shielded from the nuclear charge by the filled inner shell. Therefore, the ionization energy of 496 kJ/mol is paired with the electron configuration $1s^22s^22p^63s^1$.

A noble gas electron configuration, such as $1s^22s^22p^6$, is a very stable configuration, making it extremely difficult to remove an electron. The 2*p* electron is not as effectively shielded by electrons in the same energy level. The high ionization energy of 2080 kJ/mol would be associated with the element having this noble gas electron configuration.

8.61 Arrange the elements in each of the following groups in increasing order of the most positive electron affinity:

(a) Li, Na, K; (b) F, Cl, Br, I; (c) O, Si, P, Ca, Ba.

(a) $K < Na < Li$ **(b)** $I < Br < F < Cl$ **(c)** $Ca < Ba < P < Si < O$

8.64 Explain why alkali metals (1A) have a greater affinity for electrons than alkaline earth metals (2A).

Alkali metals have a valence electron configuration of ns^1 so they can accept another electron in the ns orbital. On the other hand, alkaline earth metals have a valence electron configuration of ns^2. Alkaline earth metals have little tendency to accept another electron, as it would have to go into a higher energy *p* orbital

8.77 Write equations representing the following processes:

(a) The electron affinity of S^-.

(b) The third ionization energy of Ti.

(c) The electron affinity of Mg^{+2}.

(d) The ionization energy of O^{-2}.

(a) $S^- + e^- \rightarrow S^{2-}$

(b) $Ti^{2+} \rightarrow Ti^{3+} + e^-$

(c) $Mg^{2+} + e^- \rightarrow Mg^+$

(d) $O^{2-} \rightarrow O^- + e^-$

8.90 The H^- ion and the He atom have two 1*s* electrons each. Which of the two species is larger? Explain

H^- and He are isoelectronic species with two electrons. Since H^- has only one proton compared to two protons for He, the nucleus of H^- will attract the two electrons less strongly compared to He. Therefore, $\mathbf{H^-}$ is larger.