**Keeping Up With the Trends - In-class assignment – Periodic trends – Answer Key**

When Mendeleev proposed the Periodic Table of the Elements, he based the positions of different elements primarily on shared chemical properties. Although little was known of the atom at the time, it is now widely known that periodic law actually follows elements as they increase in atomic number. Additionally, modern atomic theory provides a framework for explaining observed phenomena such as trends in:

* Atomic size
* Ionic size
* Ionization energy
* Electron affinity

In this exercise, you will use the various statements provided to correctly explain observed periodic trends. When you have constructed an argument to explain the observed trend, call your teacher over to see if you got it right.

1. Complete the following table to help you explain the trend that atomic size increases down a group.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Element**  | **Atomic radius (pm)**  | **Ionization energy** | **Electron configuration. Write the last orbitals representing the largest shell in a different colour** | **# protons** | **# electrons** | **# core electrons** | **# valence electrons** |
| H | 37 | 1312.0 | 1s1 | 1 | 1 | 0 | 1 |
| Li | 152 | 520.2 | 1s22s1 | 3 | 3 | 2 | 1 |
| Na | 186 | 495.8 | 1s22s22p63s1 | 11 | 11 | 10 | 1 |
| K | 231 | 418.8 | 1s22s22p63s23p64s1 | 19 | 19 | 18 | 1 |

Choose at least 4 of the cards provided to help you construct an explanation for the trend in increasing atomic size. Your explanation must invoke Coulomb’s law and relate the size of the atom to the attraction of electrons within an atom to its nucleus.

According to *Coulomb’s law,* electrons located closer to the nucleus experience a *greater attraction to the nucleus*. All of these elements have their electrons in the same type of orbital (s) but the *shell size increases* down a group. This means the *outermost electron is further* away from the nucleus making the atom larger down a group.

2. Complete the following table to help you explain the trend that atomic size decreases across a period. Use the same statements from Q1 to help you construct an explanation for why this trend occurs.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Element** | **Atomic radius (pm)** | **Ionization energy (kJ/mol)** | **Electron configuration. Write the last orbitals representing the largest shell in a different colour** | **# protons** | **# electrons** | **# core electrons** | **# valence electrons** |
| Li | 152 | 520.2 | 1s22s1 | 3 | 3 | 2 | 1 |
| Be | 111 | 899.5 | 1s22s2 | 4 | 4 | 2 | 2 |
| B | 88 | 800.6 | 1s22s22p1 | 5 | 5 | 2 | 3 |
| C | 77 | 1086.5 | 1s22s22p2 | 6 | 6 | 2 | 4 |

According to *Coulomb’s law,* electrons located closer to the nucleus experience a *greater attraction to the nucleus*. All of these elements have their outermost electrons in the *same size orbital* (2) and thus they experience the *same shielding (can invoke Zeff here instead).* But there are *more protons* across a period. This means the outermost electrons feel a greater attraction to the nucleus and the orbitals shrink as a result (making the atom smaller).

3. Refer back to the table in Q1. Use the cards to develop a statement that explains the periodic trend of ionization energy decreasing down a group (HINT: think of what orbital the electron is being removed from).

According to *Coulomb’s law,* electrons located closer to the nucleus experience a *greater attraction to the nucleus*. All of these elements have their electrons in the same type of orbital (s) but the *shell size increases* down a group. This means the *outermost electron is further* away from the nucleus and it feels *less attraction to the nucleus.* Thus less energy is required to remove the electron in bigger elements.

4. Refer back to the table in Q2. Use the cards to develop a statement that explains the periodic trend of ionization energy increasing across a period.

According to *Coulomb’s law,* electrons located closer to the nucleus experience a *greater attraction to the nucleus*. All of these elements have their outermost electrons in the *same size orbital* (2) and thus they experience the *same shielding (can invoke Zeff here instead).* But there are *more protons* across a period. This means the outermost electrons feel a greater attraction to their nuclei across the period and the energy required to remove the electron is greater for these elements.

5. When atoms turn into ions, they gain a charge but they also change in size. Fill out the 1st four rows of the following table then use the information therein to explain why *cations are smaller than parent atoms but anions are larger.*

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Atom and element** | **Radius (pm)** | **Electron configuration. Write the last orbitals representing the largest shell in a different colour** | **# protons** | **# electrons** | **# core electrons** | **# valence electrons** |
| Li | 152 | 1s22s1 | 3 | 3 | 2 | 1 |
| Li+ | 60 | 1s2 | 3 | 2 | 0 | 2 |
| F | 57 | 1s22s22p5 | 9 | 9 | 2 | 7 |
| F- | 133 | 1s22s22p6 | 9 | 10 | 2 | 8 |
| Na+ | 102 | 1s22s22p6 | 11 | 10 | 2 | 8 |
| Be2+ | 45 | 1s2 | 4 | 2 | 0 | 2 |

For Li and Li+, both elements share the *same number of protons* but the Li+ ion has one *less valence electron*. This means there are *less electron-electron repulsions* and so the electrons are more closely linked to their nuclei causing the atom to shrink. Also, in this case, the loss of the 2s electron in Li removes and entire shell, greatly shrinking the size of the ion. For F, it is the opposite. Adding an electron to F means *more valence electrons* for the *same number of protons*. There are *more electron-electron repulsions* and so these must spread out *further away from the nucleus* to minimize repulsions.

6. Complete the 5th row of the table in Q5 and compare it to Li+. Explain why the sodium ion is larger than the lithium ion. Use the cards to help you construct an argument.

Both Li+ and Na+ share the *same number of valence electrons* but the last electron in sodium ion is in a *larger shell*. This means the last electron is *further from the nucleus* making this ion larger.

7. Complete the 6th row of the table in Q5 and compare it to Li+. Explain why the beryllium ion is smaller than the lithium ion. Use the cards to help you construct an argument.

Both ions are *isoelectronic* (*same number of valence* electrons) but the Be2+ ion has *more protons* in the nucleus. This means a *greater attraction to the nucleus* for the last electron in Be2+ and so it is smaller.

8. Periodic properties can also be used to infer other phenomenon, for example, the preferred charge states of different elements. Consider the ionization energies associated with aluminum:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Species and change in radius | Al à Al+118 pm à 98 pm | Al+ à Al2+98 pm à 80 pm | Al2+ à Al3+80 pm à 50 pm | Al3+ à Al4+50 pm à slightly smaller |
| Ionization energy (kJ/mol) | 1st577 | 2nd1817 | 3rd2745 | 4th11577 |

Create your own table, similar to what is in Q5, then use the cards to help you develop an argument for a) why the ionization energy increases with the removal of sequential electrons and b) why the aluminum ion’s preferred charge state is 3+.

As electrons are removed from aluminum, the same number of protons are present but the number of electrons keeps decreasing. This means the remaining electrons fell and increased attraction to the nucleus making the next ionization energy bigger. Al3+ is the preferred charge state because after the 3rd electron is removed, the energetic cost of removing the last one is from a much smaller orbital which feels much greater attraction to the nucleus.

9. Design your own table using the elements N, O and F to explain the following trend: electron affinity becomes more exothermic across a period.

The electrons added to N, O and F are similar sized shells but the number of protons is greater in F, then O and last N. The greater nuclear charge means greater attraction of the electron being added to the F atom and thus the electron affinity is more exothermic (more negative).

Table should have electron configurations highlighting the valence shell, as well as the number of protons.

10. Design your own table using the elements F, Cl and Br to explain the following trend: electron affinity becomes less exothermic down a group.

Electrons when added to larger orbitals are further away from the nucleus where they feel less attraction to the nucleus. Electron affinity is thus less exothermic for larger elements.