**CHEM131 – General Chemistry I**

**Chemical Periodicity Worksheet**

* Knowledge of periodicity is valuable in understanding bonding in simple compounds
* Variations useful in predicting chemical behaviour
* Changes in properties depend on:
  + electron configurations, especially configuration in outmost occupied shell
  + How far away that shell is from the nucleus

**Questions**

1. (a) Why is the effective nuclear charge, Zeff, experienced by an electron in an outer shell is less than the actual nuclear charge, Z?

**This is because the inner electrons block/ screen/shield the nuclear charge’s effect on the outer electrons.**

**The concept of *shielding* or *screening* helps us to understand many periodic trends in atomic properties.**

(b) Within a family (group) of representative elements, atomic radii increase from the top to bottom of the periodic table. Use lithium and sodium as examples to explain this trend.

**As we move down a family (group) electrons are added to shells further from the nucleus.**

**E.g. 3Li has a 1s2 2s1 configuration.**

* + **The outermost 2s1 electron is not as effectively shielded as an electron in a shell further from nucleus**

**E.g. 11Na has 10 inner e-s 1s2 2s2 2p6 and one in an outer shell, 3s1**

* + **The 10 inner e-s shield the outer-shell electron from most of the +11 nuclear charge**

(c) Atomic radii decrease going from left to right *across* the periodic table. Use lithium and beryllium as examples to explain this trend.

**Moving from left to right across a period a proton is added to the nucleus and an electron is added to a particular shell.**

* + **Moving across a period, each element has an increased nuclear charge and the electrons are going into the same shell (2s and 2p or 3s and 3p, etc.).**
    - **Consequently, the outer electrons feel a stronger effective nuclear charge.**
    - **For Li, Zeff ~ +1 For Be, Zeff ~ +2**

(d) Arrange the following elements in order of increasing atomic radii: Se, S, O, Te

* **O < S < Se < Te**

2. This question is about ionization energy.

(a) What do you understand by the term first ionization energy (IE1)?

* **First ionization energy (IE1)** 
  + **The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a 1+ ion.**
* **Symbolically: Atom(g) + energy → ion+(g) + e-**
* **Example: Mg(g) + 738kJ/mol** → **Mg+ + e-**

(b) What do you understand by the term second ionization energy (IE2)?

* **Second ionization energy (IE2)**
  + **The amount of energy required to remove the second electron from a gaseous 1+ ion.**
  + **Symbolically: ion+ + energy  ion2+ + e-**

**Example: Mg+ + 1451 kJ/mol  Mg2+ + e-**

* **Atoms can have 3rd  (IE3), 4th (IE4), etc. ionization energies.**

(c) Summarize the periodic trends for ionization energy.

* **Periodic trends for Ionization Energy:**
  1. **IE2 > IE1**

**It always takes moreenergy to remove a second electron from an ion than from a neutral atom.**

* 1. **IE1 generally increases moving from IA elements to VIIIA elements.**

**Important exceptions at Be & Mg, N & P, etc. due to filled and half-filled subshells.**

* 1. **IE1 generally decreases moving down a family.**

**IE1 for Li > IE1 for Na, etc.**

(d) Arrange the following elements based on their first ionization energies: Sr, Be, Ca, Mg

**Sr < Ca < Mg < Be**

3. This question is about electron affinity.

(a) What do you understand by the term electron affinity?

* **Electron affinity (EA) is the amount of energy *absorbed* when an electron is added to an isolated gaseous atom to form an ion with a 1- charge.**

(b) What are the sign conventions for electron affinity?

* **Sign conventions for electron affinity.**
  + **If electron affinity > 0 energy is absorbed.**
  + **If electron affinity < 0 energy is released.**
* **Electron affinity is a measure of an atom’s ability to form negative ions.**
* **Elements with very negative electron affinities gain electrons easily to form negative ions (anions)**
* **Symbolically: atom(g) + e- + EA  ion-(g)**

(c) What are the general periodic trends for electron affinity?

* **General periodic trend for electron affinity is**
  + **the values become more negative from left to right across a period on the periodic chart.**
  + **the values become more negative from bottom to top up a row on the periodic chart.**

(d) Arrange the following elements based on their electron affinities: Al, Mg, Si, Na

**Si < Al < Na < Mg**

**N.B. Noteworthy exceptions:**

**Group 2A – very difficult to add an e- because these elements have their outer *s* subshell filled**

**Group 5A – an additional e- would have to be added to a half-filled set of *np* orbitals**

4. This question is about ionic radii.

(a) Cations (+ve ions) are always ***smaller*** than their respective neutral atoms. Explain why.

**- Due to the increase in effective nuclear charge.**

(b) What is the trend in cation radii vary across a period?

* **Cation (positive ions) radii decrease from left to right across a period.**
  + **Increasing nuclear charge attracts the electrons and decreases the radius.**

(c) Anions (negative ions) are always ***larger*** than their neutral atoms. Explain why.

**- Due to the increase in effective nuclear charge.**

(d) What is the trend in anion radii across a period?

* **Anion (negative ions) radii decrease from left to right across a period.**
  + **Increasing electron numbers in highly charged ions cause the electrons to repel and increase the ionic radius.**
  + **Example: O2- is larger than the *isoelectric* F- because the oxide ion contains 10 e-s held by a nuclear charge of 8+, whereas the F- ion has 10 e-s held by a nuclear charge of 9+**

(e) What is the trend in ionic radii down a group?

* **Both cation and anion sizes increase going down a group**



(f) Arrange the following elements based on their ionic radii: Ga, K, Ca

**K1+ > Ca2+ > Ga3+**

5. This question is about electronegativity.

(a) What do you understand by the term electronegativity?

* **Electronegativity is a measure of the relative tendency of an atom to attract electrons to itself when *chemically combined with another element*.**
  + **Electronegativity is measured on the Pauling scale.**
  + **Fluorine is the most electronegative element.**
    - **E.g. EN value for F is 4.0 🡪 when F is chemically bonded to other elements, it has a greater tendency to attract electron density to itself than any other element**
  + **Cesium and francium are the least electronegative elements.**
  + **Elements with high electronegativities (nonmetals) often gain electrons to form anions.**
  + **Elements with low electronegativities (metals) often lose electrons to form cations.**

(b) What is the trend in electronegativity across a period?

* **For the representative elements, electronegativities usually increase from left to right across periods.**

(c) What is the trend in electronegativity down a group?

* **For the representative elements, electronegativities usually decrease from top to bottom within groups.**

(d) Arrange the following elements based on their electronegativity: Se, Ge, Br, As

**Ge < As < Se < Br**

6. Write short notes on the following:

(a) the chemical periodicity in the reactions of hydrogen

(b) the chemical periodicity in the reactions of oxygen

**See PowerPoint for Answers**