Review Packet – Independent Practice

Check your answers at http://shakeribchem.weebly.com

1. The shielding effect increases with increasing atomic number within a ___.
   a. period b. group c. both d. neither

2. ~~In~~ any ___, the number of electrons between the nucleus and the outer energy level is the same.
   a. period b. group c. both d. neither

3. Within a ____, the nucleus has a stronger ability to pull on the outermost (valence) electrons in elements of high atomic number.
   a. period b. group c. both d. neither

4. In a ____, electron affinity values become more negative as atomic number increases.
   a. period b. group c. both d. neither

5. The halogens are considered a ____.
   a. period b. group c. both d. neither

6. Which atom has the greater nuclear charge? ______
   a. Na b. Al c. P d. Ar

7. Which atom demonstrates the greatest shielding effect? _______
   a. Na Period 3 b. Al c. P d. Ar same shielding for all

8. The atoms Na, Al, P, and Ar all have the same __________
   a. number of valence electrons b. size atomic radius c. number of kernel electrons
   inner (nonvalence)

9. Which element on the periodic table has
   a. lowest ionization energy H (Cs by trend)
   b. highest second ionization energy He
   c. highest electronegativity F
   d. highest ionization energy He
   e. largest atomic radius Fr

10. Explain the relationship between the relative size of an cation and anion (ionic radius) to its atom (atomic radius).

smaller – loses outer E level

larger – gains e⁻; gains repulsion

11. Explain why noble gases are inert and do not form ions.

full outer shell – stable as is

12. Will the shielding effect be more noticeable in metals or non metals? Explain your answer.

Important in both, but across the same period it will be more noticeable in metals. The nonmetals within the same period will have a larger effective nuclear charge due to higher # of protons with no change in shielding. Valence are pulled more tightly.

13. Why do elements in the same family generally have similar properties? Choose one as an example to support your reasoning.

Same number of valence – Eg: alkali metals are all reactive because 1 valence e^- & low ↓IE

14. Arrange each of the following in order of increasing ionization energy and explain your reasoning: Calcium, iron, copper, bromine and krypton.

Ca Fe Cu Br Kr

increasing →

Ca, Cu, Fe, Br, Kr

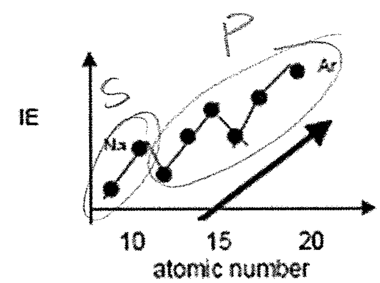
increases across a period (exc. in some transition metals)

15. Factors affecting ionization energy include effective nuclear charge, the shielding effect, the atomic radius and the electron arrangement in a sublevel. Use the appropriate factors to explain the overall trend indicated by the dark line and the exceptions to it.

1) Effective Nuclear energy: Increase, Decrease or Constant

2) Shielding: Increase, Decrease or Constant

3) Atomic Radius: Increase, Decrease or Constant



4) Explain exceptions to the overall trend based on electron configuration. Exceptions @ sublevels (eg. Mg to Al)

16. What element am I?. (Brief periodic table location description for each clue)

Clue #1)

I have a high electron affinity, (highly negative value), and my atomic number is X.

Clue #2)

The element with atomic number X-1 has a lower ionization energy and a lower electron affinity.

← period

Clue #3)

The element with atomic number X+1 has a higher ionization energy and basically no electron affinity (positive value).

d) Within my group, I have the second highest ionization energy.

What element am I? Cl

17. What do transition metals have in common with respect to their electron configurations? And WHY is Zinc or Scandium not considered a transition metal?

incomplete d sublevels or ability to have incomplete d sublevel when ion

never has an incomplete d sublevel

18. Consider the table of the first four ionization energies for an element we will call A.

Ionization Energy (kJ/mole)	1st	2nd	3rd	4th
	576	1817	2745	11580

a. In which group does A appear on the periodic table?

group 13 probably

b. What is the most likely oxidation number for element A?

+3 (loses 3e⁻ from outer shell easily)

c. What is the minimum number of electrons that A must have?

must have @ least 4

d. Write the valence electron sublevel configuration for this element.
(Hint: use "n" as the energy level)

np^3

19. Can anions (negative ions) of two different elements have the same valence electron arrangement? If so, give examples and discuss. If no, explain why not.

Yes — examples: O^{2-} $1s^2 2s^2 2p^6$

Cl^{1-} $1s^2 2s^2 2p^6$

Because although they have gained diff # of e⁻, they are both striving for octet

20. The first ionization energy of beryllium is 9.322 eV, the second ionization energy is 18.211 eV, and the third ionization energy is 153.893 eV. Explain why the third ionization energy of beryllium so much higher than the first two.

1st $Be_{(g)} \rightarrow Be^{1+}_{(g)} + e^-$ 9.322eV

2nd $Be^{1+}_{(g)} \rightarrow Be^{2+}_{(g)} + e^-$ 18.211eV

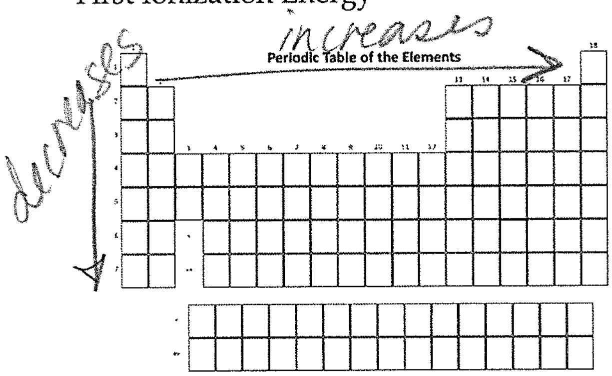
Lower because valence — easier to lose, require less E

3rd $Be^{2+}_{(g)} \rightarrow Be^{3+}_{(g)} + e^-$ 153.893eV

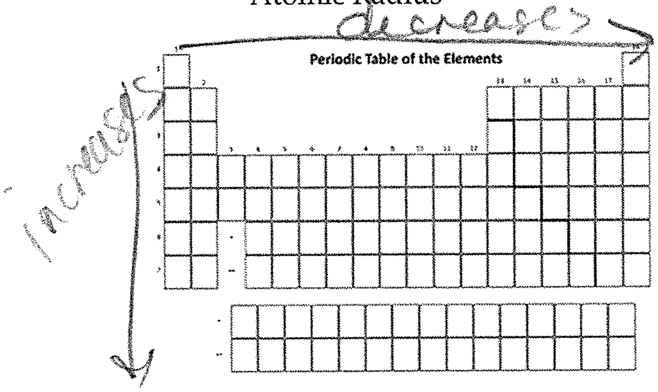
nonvalence e⁻ much harder to lose, more tightly held by nucleus

Quiz yourself- Draw in the general trends on the following periodic tables:
Next to the periodic table, you should be able to define the term, and describe why the trend occurs the way it does. You should also be able to discuss any inconsistencies or exceptions and explain why they occur.

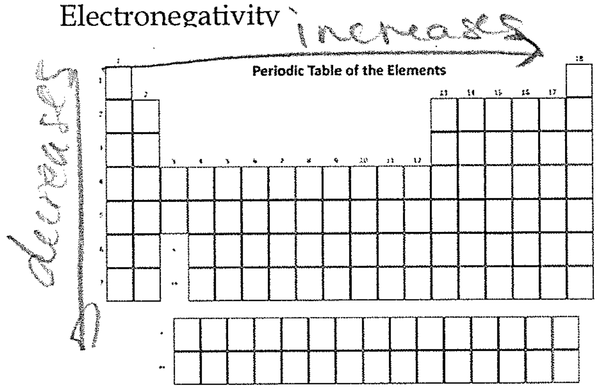
First Ionization Energy



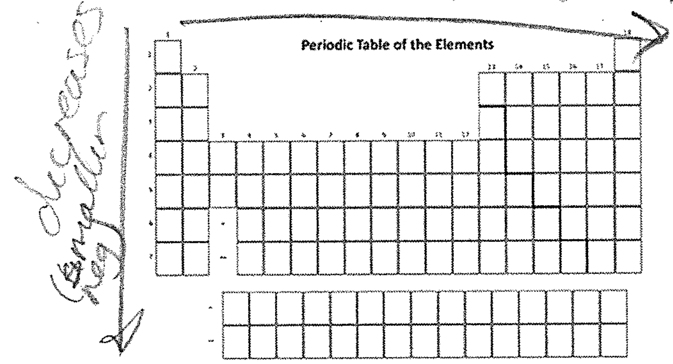
Atomic Radius



Electronegativity

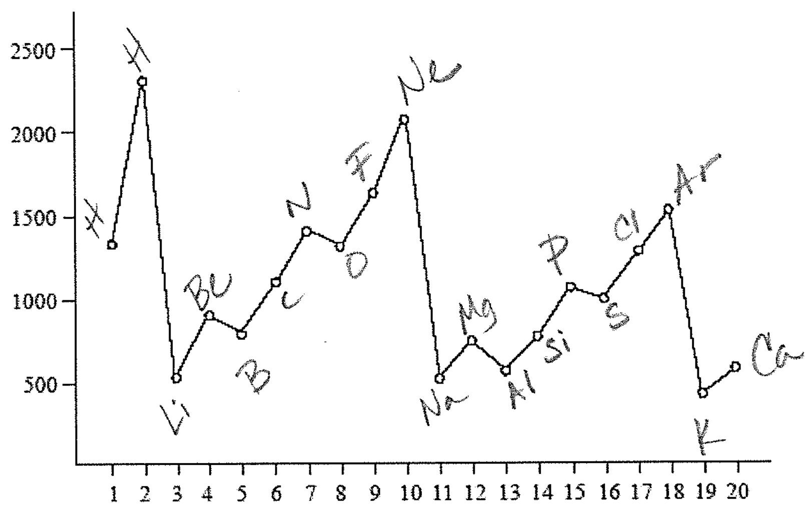


Electron Affinity increases (larger negative)



19. The graph of the first ionization energy plotted against atomic number for the first twenty elements shows periodicity.

First ionization energy / kJ mol^{-1}



Atomic number

periodic recurrance of pattern/trend

(i) Define the term first ionization energy and state what is meant by the term periodicity.

Energy required to remove 1 mole e⁻ from 1 mol of gaseous atoms

(ii) State the electron arrangement of argon and explain why the noble gases, helium, neon and argon show the highest first ionization energies for their respective periods.

$1s^2 2s^2 2p^6 3s^2 3p^6$ full valence octet

Very high IE because effective nuclear charge is the highest in the period due to no increase in shielding but greater atomic number & therefor nuclear charge. Valence shell is held tightly and electron arrangement is stable.

(iii) A graph of atomic radius plotted against atomic number shows that the atomic radius decreases across a period. Explain why chlorine has a smaller atomic radius than sodium.

Chlorine (17) has a greater nuclear charge than Sodium (11) with no additional shielding, so therefore it has a greater effective nuclear charge. The larger positive charge attracts the principal & valence electrons with greater force, pulling tighter and causing the atom to have a smaller radius.

(iv) Explain why a sulfide ion, S2–, is larger than a chloride ion, Cl–.

S^{2-} has gained $2e^-$, causing greater repulsion in the valence shell and therefore a larger radius.

1. Which compound contains both covalent and ionic bonds?
   - A. Sodium carbonate, Na2CO3
   - B. Magnesium bromide, MgBr2
   - C. Dichloromethane, CH2Cl2
   - D. Ethanoic acid, CH3COOH

2. Which pair of elements is most likely to form a covalent bond?
   - A. Li and Cl
   - B. P and O
   - C. Ca and S
   - D. Zn and Br

3. What is the formula of the compound formed when aluminum reacts with oxygen?
   - A. Al3O2
   - B. Al2O3
   - C. AlO2
   - D. AlO3

Al^{3+} O^{2-}

Al_2O_3

4. What happens when potassium and oxygen combine together?
   - A. Each potassium atom gains one electron.
   - B. Each potassium atom loses one electron.
   - C. Each oxygen atom gains one electron.
   - D. Each oxygen atom loses one electron.

will gain 2