Name KEY Date Period

Answer the following in the space provided.

1. Use your knowledge of the periodic table of the elements to explain the following trends. Be sure to state the trend **and** give reasons for the trend.

A. The trend in electronegativity from P to S to Cl

EN increases L to R across a period because the # of protons increases, increasing the effective nuclear charge. ↑ in protons w/o an ↑ in ELs.

B. The trend in electronegativity from Cl to Br to I

EN decreases going down a group because of shielding effect ↳ electrons in higher ELs are shielded from the pull of the nucleus by the inner ELs

The shared electrons are farther from the nucleus due to the addition of ELs.

(The increase in the # of protons as you go down a group cannot overcome the shielding effect.)

C. The trend in atomic radius from Li to Na to K

Atomic radius increases going down a group because EL are being added. The more ELs between the nucleus and the valence EL, the greater the shielding effect.

D. The trend in atomic radius from Al to Mg to Na

Atomic radius increases as you move R to L across a period because the # of protons decreases (decrease in effective nuclear charge).

2. Explain the reason(s) for each of the following:

A. The first ionization energy for magnesium is greater than the first ionization energy for calcium.

IE ↓

Mg
Ca

The valence e⁻ in Mg is closer to the nucleus than the valence e⁻ in Ca (less shielding effect). Therefore Mg's valence e⁻ is held more tightly by the nucleus and requires more Energy to remove. Ca's valence e⁻ is in an EL farther from the nucleus (more shielding) and thus requires less Energy to remove.

B. The first and second ionization energies for calcium are comparable, but the third ionization energy is much greater.

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

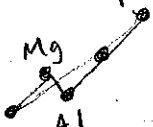
1° The first and second outermost electrons are in the 4s orbital, so they are farther from the nucleus and more shielded from the nucleus than the third outermost electron which is in a 3p orbital.

2° The 3rd electron is being taken from a full p sublevel which is very stable.

C. The second ionization energy for sodium is much greater than the first ionization energy, but the second ionization energy for magnesium is comparable to the first ionization energy.

The outermost electron in sodium is easier to remove because it is in the 3s orbital, whereas the second outermost electron is in the 2nd EL which is closer to the nucleus and less shielded making it harder to remove. Magnesium's two outermost electrons are both in the 3s orbital, so the energies required to remove them are comparable.

D. The first ionization energy of aluminum is lower than the first ionization energy for magnesium. → Does not follow trend!



Mg $1s^2 2s^2 2p^6 3s^2$

Al $1s^2 2s^2 2p^6 3s^2 3p^1$

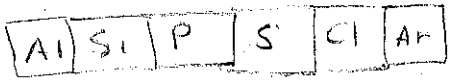
1° The outermost electron in Al is in a 3p orbital making it easier to remove than the outermost electron in magnesium, which is in a 3s orbital. P electrons are slightly shielded by S electrons.

2° The outermost electron in Mg is in a full s orbital which is a stable configuration.

3. As shown on the table below, the first ionization energies of Si, P, and Cl show a trend.

Element	Ionization Energy
Si	786
P	1012
Cl	1251

Explain the reason for the trend.



IE increases →

As you move L to R across the period the # of protons increases so the pull on the electrons increases making it harder to remove an electron from chlorine than from phosphorous or silicon.

4. The calcium atom is much larger than the calcium ion, while the fluorine atom is much smaller than the fluorine ion. Explain this natural occurrence.

Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Ca^{2+} $1s^2 2s^2 2p^6 3s^2 3p^6$

F $1s^2 2s^2 2p^5$

F^- $1s^2 2s^2 2p^6$

Calcium ion is smaller than calcium atom because EL 4 is lost and because the protons outnumber the electrons → effective nuclear charge increases in the ion.

The fluoride ion is bigger than the fluoride atom because when the electrons outnumber the protons, each electron feels less pull from the nucleus (decrease in effective nuclear charge.) Also ↑ $e^- - e^-$ repulsion

5. The first three ionization energies (I_1, I_2, I_3) for beryllium and neon are given in the following table:

kJ/mole	I_1	I_2	I_3
Be	900	1757	14840
Ne	2080	3963	6276

A. Write the complete electron configuration for Ne.

Ne $1s^2 2s^2 2p^6$

Be $1s^2 2s^2$

B. Explain any trends or significant discrepancies found in the ionization energies for Be and Ne.

* Beryllium's 1st and 2nd IE show a small increase because its two outermost electrons are in the 2s orbital, while its 3rd outermost electron is in the full 1st EL with no shielding effect making the 3rd IE much higher.
* Neon's 1st three IE show a small increase because the first three electrons are from the 2p orbital. The IE's increase because after one electron is removed, the protons' pull on the electrons increases.

C. If chlorine gas is passed into separate containers of heated Be and heated Ne, explain what compounds, if any, might be formed, and explain your answer in terms of the electron configurations of these two elements.

Be has a greater chance of bonding with Cl than does Ne

Be has an incomplete second EL so has the room to accomodate more electrons. Ne has an octet (full s & p) so is stable and unlikely to bond.

:Cl:Be:Cl:

big increase

D. An unknown element, X, has the following three ionization energies:

kJ/mole	I_1	I_2	I_3
X	419	3069	4600

On the basis of the ionization energies given, what is most likely to be the compound produced when chlorine reacts with element X?

+1 -1
X Cl

Changed EL between IE_1 and IE_2

X can be any metal from group 1

6. $3Br_2(aq) + 2Al(s) \rightarrow 2Al^{+3}(aq) + 6Br^-(aq)$

The reaction between aluminum metal and bromine in aqueous solution occurs according to the reaction above.

A. Give the ground state electron configurations of Al and Al^{+3}.

Al $1s^2 2s^2 2p^6 3s^2 3p^1$

Al^{3+} $1s^2 2s^2 2p^6$

B. Give the ground state electron configurations of Br and Br^-.

Br $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Br^- $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

C. Are Al^{3+} and Br^- isoelectronic? Explain.

No, because they have different electron configurations

higher

Kr and Br are isoelectronic

3 more

3 more

less

p^+