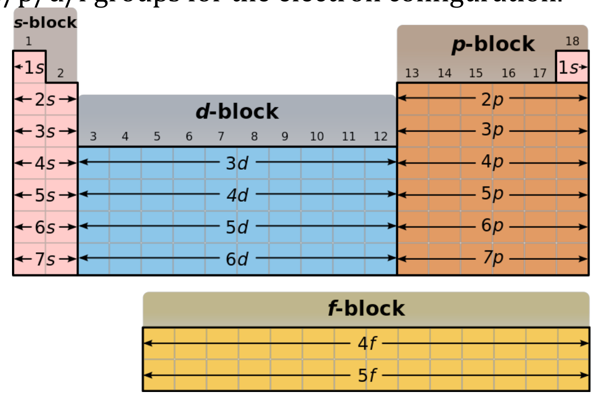
# TOPIC: 1.7 PERIODIC TRENDS

| ENDURING UNDERSTANDING: | |
|---|---|
| SAP-2 | The periodic table shows patterns in electronic structure and trends in atomic properties. |
| **LEARNING OBJECTIVE:** | |
| SAP-2.A | Explain the relationship between trends in atomic properties of elements and electronic structure and periodicity. |
| **ESSENTIAL KNOWLEDGE:** | |
| SAP-2.A.1 | The organization of the periodic table is based on the recurring properties of the elements and explained by the pattern of electron configurations and the presence of completely or partially filled shells (and subshells) of electrons in atoms. |
| | WRITING THE ELECTRON CONFIGURATION OF ELEMENTS THAT ARE EXCEPTIONS TO THE AUFBAU PRINCIPLE WILL NOT BE ASSESSED ON THE AP EXAM. *Rationale: The mere rote recall of the exceptions does not match the goals of the curriculum revision.* |
| SAP-2.A.2 | Trends in atomic properties within the periodic table (periodicity) can be qualitatively understood through the position of the element in the periodic table, Coulomb's law, the shell model, and the concept of shielding/effective nuclear charge. These properties include: a. Ionization energy b. Atomic and ionic radii c. Electron affinity d. Electronegativity. |
| SAP-2.A.3 | The periodicity (in SAP-2.A.2) is useful to predict /estimate values of properties in the absence of data. |
| **EQUATION(S):** | |
| | N/A |

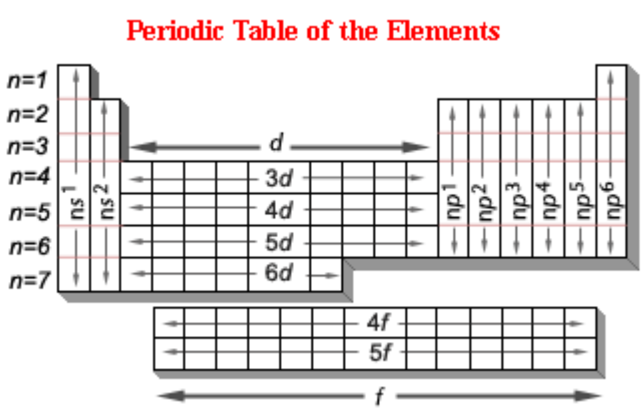
## NOTES:

The periodic table is arranged in order from lowest atomic number (# of protons) to highest. The blocks of the periodic table correspond to the s/p/d/f groups for the electron configuration.



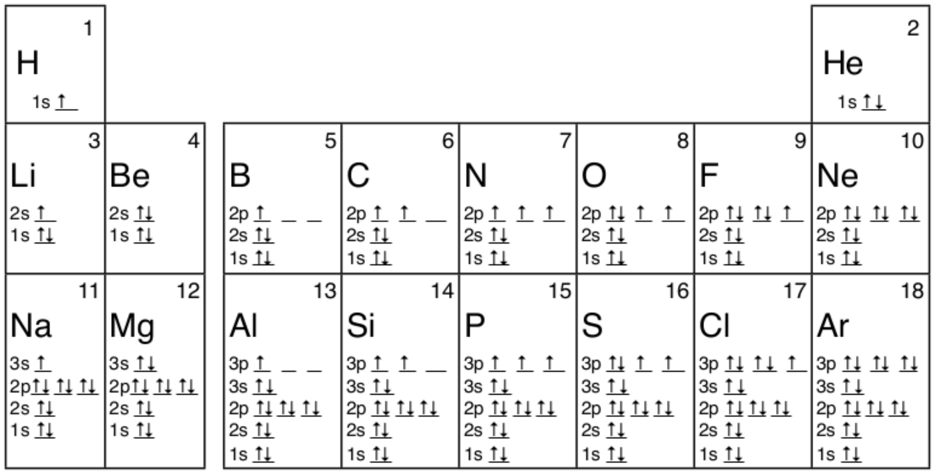

https://socratic.org/questions/what-is-the-electron-configuration-for-francium

Elements that have the same valence electron configuration tend to have similar chemical properties.




http://nobel.scas.bcit.ca/wiki/index.php/File:Ptable_econfig.gif#filelinks

Most, if not all, periodic trends can be explained by the arrangement of the electrons and the number of protons in the atoms.




https://chem.libretexts.org/Under_Construction/Purgatory/Essential_Chemistry_(Curriki)/Unit_1%3A_Atomic_and_Molecular_Structure/1.4%3A_Electron_Configuration_and_Orbital_Diagrams

REMEMBER: Stating a trend is not EXPLAINING a trend. Explanations of trends should never be in terms of the location of the periodic table.

Coulombic Attraction is the electrostatic attraction between two charged particles. Often when discussing periodic trends the charged particles are the nucleus (specifically the total number of protons) and the electrons. Often we are referring to the outermost electrons, the valence electrons.

Coulomb's law states that the attraction between two charged particles is proportional to the magnitude of the charge and inversely proportional to the distance between them. To make this simpler, the larger the charge, the more attractive forces between the particles. The further away the particles are from each other, the weaker the attraction.

# PERIODIC TRENDS

Key Terms:

### *Coulombic Attraction/ Electrostatic Interactions*

The positive-negative attraction which takes place when you have two charged particles in close proximity.

- Increases with increase in charge
- Increases with decrease in distance between particles

### *Effective Nuclear Charge and Electron Shielding*

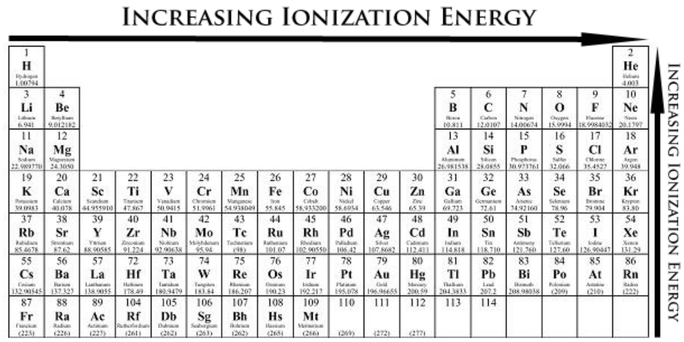
The **effective nuclear charge** is the net positive **charge** experienced by valence electrons. It can be approximated by the equation: $Z_{eff} = Z - S$, where Z is the atomic number and S is the number of electrons in orbitals that are closer to the nucleus.

## A) First Ionization Energy

The energy required to remove the outermost (highest energy) electron from the gas from of a neutral atom in its ground state.

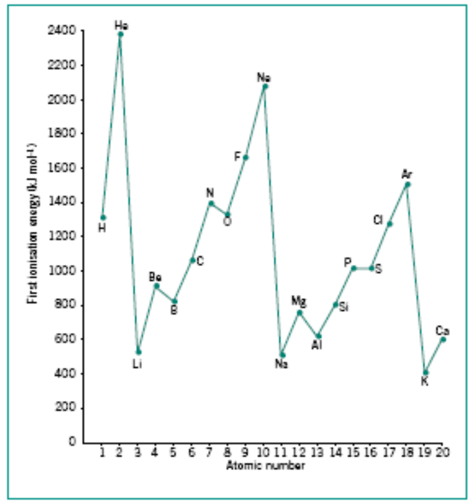
First Ionization energy ***decreases*** as you move down a group. Electrons are further from the nucleus and therefore have a lower Coulombic attraction. Additionally, the inner shells of electrons ***shield*** or block the protons force of attraction, so that outermost electrons do not feel as much of the nuclear force. This results in the outer electrons being even easier to remove.

First Ionization energy ***increases*** as you move across a period on the periodic table, from left to right. As you move across the period the atomic radius is smaller and there is an increase in protons in the nucleus. Both factors result in greater Coulombic attraction, which in turn means that it will require more energy to remove the first electron.




https://wps.pearsoned.com.au/ibcsl/89/22896/5861561.cw/content/index.html

There are a few places where the ionization doesn't appear to follow a trend. You can see this on the graph between Be and B or between N and O. These are actually for two slightly different reasons. https://useruploads.socratic.org/N5qKJ5fTLiJK3MXQAifQ_Ionization_Energy_Trend_IK.png




*Be and B exception (s² to s²p¹)*

Be = $1s^2\ 2s^2$

B = $1s^2\ 2s^2\ 2p^1$

When the first electron is removed from the boron, B, atom, the electron is being removed from the 2p orbital. Since the 2p orbital is further away from the nucleus it takes less energy to remove it even though there are more protons in the atom.

*N and O exception($s^2p^3$ to $s^2p^4$)*

N = $1s^2\ 2s^2\ 2p^3$

O = $1s^2\ 2s^2\ 2p^4$

When the first electron is removed from oxygen it takes less energy (despite the increase in protons) than from nitrogen because the electrons in oxygen are sharing the $2p_x$ orbital and therefore have greater electron-electron repulsions making it easier to remove one electron.

The second ionization energy is the energy to remove a second electron from the atom and so on for each successive electron.

By examining the successive ionization energies for an element we can determine how many valence electrons there are in that element. When all of the valence electrons have been removed, you will see a large "jump" in the ionization energy values. This "jump" is due to the fact that the core electrons are closer to and less shielded from the nucleus and therefore it requires more energy to remove them.

| Ionization Energy Number | Enthalpy (kJ/mole) |
|---|---|
| 1st | 738 |
| 2nd | 1451 |
| 3rd | 7733 |
| 4th | 10543 |
| 5th | 13636 |
| 6th | 18020 |
| 7th | 21711 |
| 8th | 25658 |
| 9th | 31646 |
| 10th | 35457 |
| 11th | 169988 |

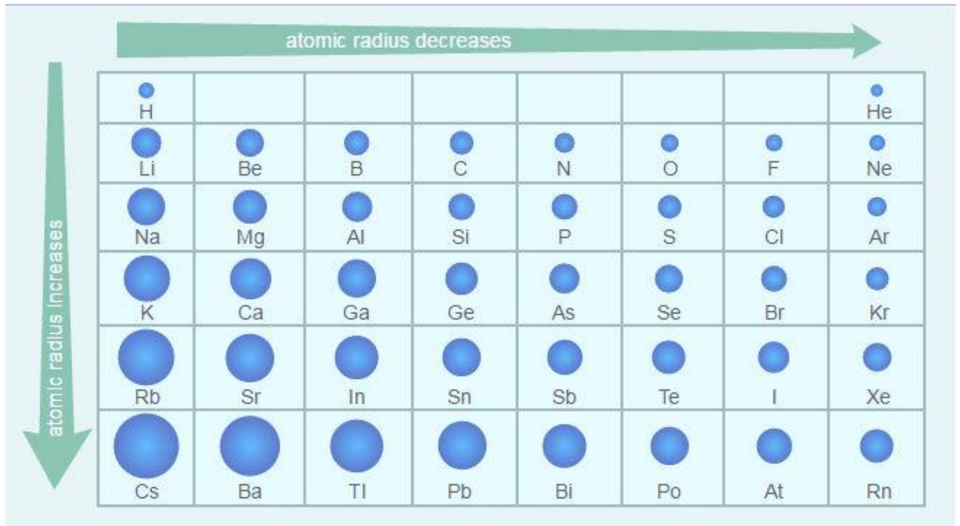
For example:

Consider magnesium, Mg, the electron configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2$ and we can see that it has 2 valence electrons. https://www.webelements.com/magnesium/atoms.html

You can see that there is a big jump between the 2nd and 3rd ionization energies and again between the 10th and 11th ionization energies. This shows when electrons are being removed from a shell that is closer to the nucleus.

## B) Atomic Radius

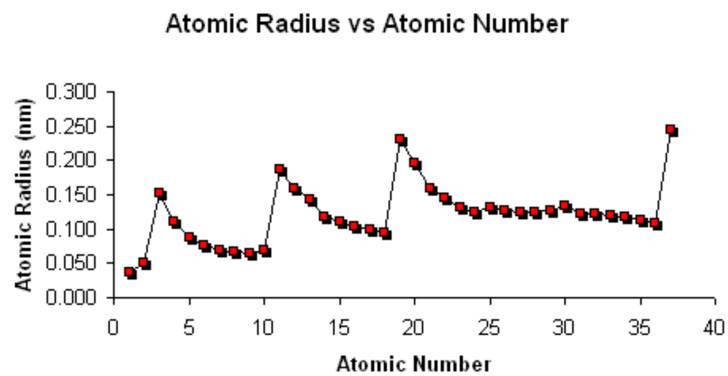
The **atomic radius** of a chemical element is a measure of the size of its **atoms**, usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding cloud of electrons.




Atomic Radii ***increases*** as you move down a column as there are more electron shells.

https://byjus.com/chemistry/atomic-radius-in-periodic-table-in-basic-chemistry/

Atomic Radii ***decreases*** as you move across a period on the periodic table, from left to right. Electrons are being added to the same energy level. At the same time, protons are being added to the nucleus. Increasing the number of protons gives a ***higher effective nuclear charge***. In other words, there is a stronger force of attraction pulling the electrons closer to the nucleus. This results in a smaller atomic radius, as with greater numbers of protons there is more pull on the electrons.
https://www.geocities.ws/junebug_sophia/atmRad.gif

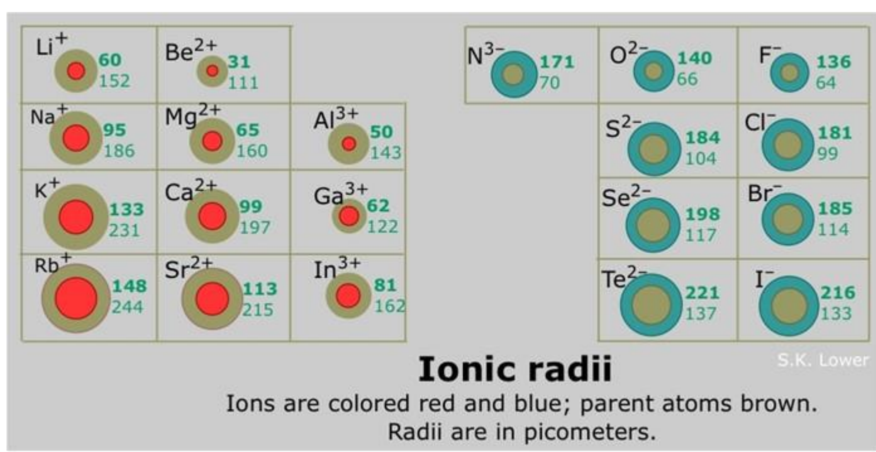



## Ionic Radius

The trends for ionic radii are similar to those of atomic radii, except that cations and anions are different from each other.

Cations are always smaller than the parent atoms, because they have lost their valence shell. This causes them to be smaller. They also decrease in size because the nuclear attraction is now acting on fewer electrons so they are drawn in toward the nucleus due to the greater attraction. Additionally there are fewer electron-electron repulsions.

Anions, on the other hand, are always larger than the parent atom. Electrons are added to the same valence shell; however, there are greater electron-electron repulsions so the ion increases in size.

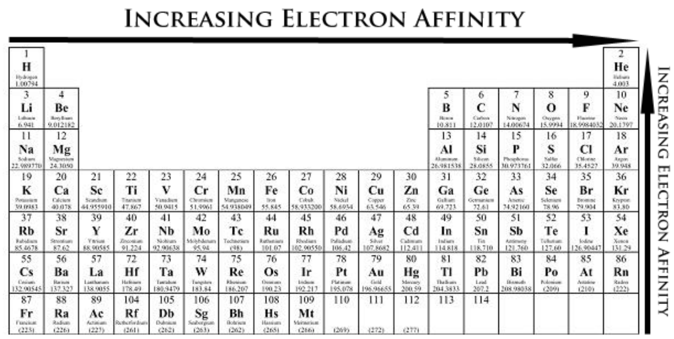



https://slideplayer.com/slide/8861824/

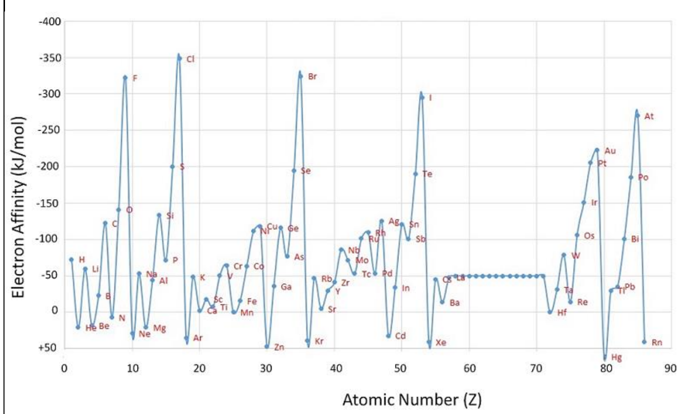
## C) ELECTRON AFFINITY

Electron affinity is the amount of energy involved when an electron is accepted by a gaseous atom to form a negative ion. In other words, the neutral atom's likelihood of gaining an electron. The values tend to be negative to show the energy is released as electrons are added to the atoms.

In general, the electron affinity increases from left to right on the periodic table. This is caused by the filling of the valence shell of the atom; a Group 17atom releases more energy than a Group 1 atom when it gains an electron, this indicates that it is more stable.




https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Physical_Properties_of_Matter/Atomic_and_Molecular_Properties/Electron_Affinity




A trend of decreasing electron affinity when moving down the groups in the periodic table might be expected. The additional electron will be entering an orbital farther away from the nucleus. Since this electron is farther from the nucleus it is less attracted to the nucleus and would release less energy when added. However, a clear counterexample to this trend can be found in Group 2, and inspecting the entire periodic table, it turns out that the proposed trend only applies to Group 1 atoms.

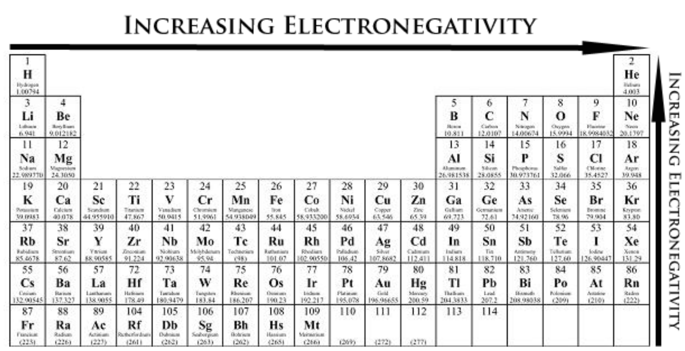
## D) ELECTRONEGATIVITY

Electronegativity is a measure of the ability of an atom (or group of atoms) to attract shared electrons.

Electronegativity **decreases** as you move down a column as there is a greater distance from the nucleus and because there is also more electron shielding.

Electronegativity **increases** as you move across a period on the periodic table, from left to right. This is because the atomic radius is decreasing while the number of protons (and effective nuclear charge) is increasing.

Fluorine is the most electronegative element.




| Example | Difference in electronegativity | Type of Bond |
|---|---|---|
| H-H | No difference – Electrons are shared equally | Nonpolar covalent bond |
| H-Br | Slight difference in values – Electrons are shared unequally | Polar covalent bond |
| NaCl | Large difference in values – Electrons are not shared, they are transferred | Ionic Bond |

## I DO:

For each of the following pairs of elements

Na or Li F and O

Choose the atom with:

a) Higher first ionization energy

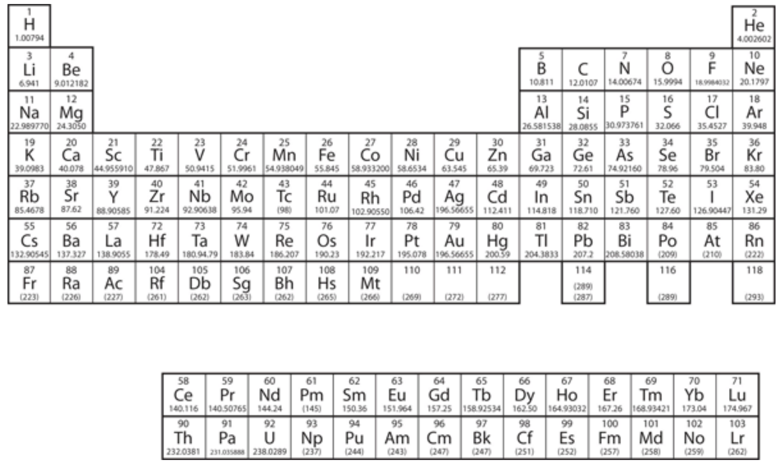
Li , F

b) Larger atomic radius

Na , O

c) Higher electronegativity

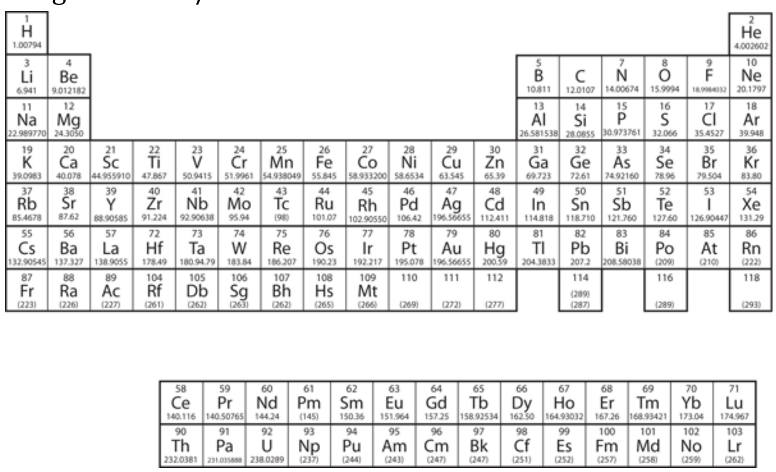
Li , F



## WE DO:

Rank the following from smallest to largest atomic/ionic radius.

a) $Na^+$, Na, $Na^-$
b) C, N, O
c) Cl, Ar, K
d) Be, Mg, Ca



SCAN ME

## YOU DO:

1) On the basis of their position on the periodic table determine which element in the pair would have a larger atomic radius
   a) P or S
   b) Cl or Br
   c) Sr or Sc

2) Based on the successive ionization energies for the following element "X", predict the formula that would be formed when "X" reacts with chlorine, Cl.

| Ionization Energy Number | Enthalpy (kJ/mole) |
|---|---|
| 1st | 577 |
| 2nd | 1820 |
| 3rd | 2740 |
| 4th | 11600 |
| 5th | 14841 |

3) The first ionization energy for potassium, K, is 419 kJ/mol and the second ionization energy for calcium, Ca, is 1145 kJ/mol. Using concepts from this unit explain why they are different even though they are isoelectric (have the same number of electrons).

$K_{(g)} \rightarrow K^+_{(g)} + e^-$

$Ca^+_{(g)} \rightarrow Ca^{2+}_{(g)} + e^-$

4) Element X has an electron configuration of $1s^22s^22p^63s^1$, while element Z has an electron configuration of $1s^22s^22p^5$.

a) Which element would have greater first ionization energy?

b) Which element would have a larger radius?

c) Which element would have higher electronegativity?

d) Which element would form an ion that has a larger radius?

e) Which element would release more energy when it gains an electron?

5) Predict two elements that would have properties similar to:

a) Chlorine

b) Sodium

c) Calcium

6) Nitrogen is in column 5A of the periodic table, which is called the pnictogens. When nitrogen reacts with iodine it forms nitrogen triiodide, $NI_3$, which is a contact explosive that explodes with a snap releasing clouds of purple iodine vapor. Select another pnictogen and predict the formula of the compound that would be formed with a reaction with bromine.

7) Based on the given electron configurations, group together the elements that would have similar chemical properties.

a) $1s^22s^22p^63s^1$
b) $1s^22s^22p^63s^23p^64s^2$
c) $1s^22s^22p^5$
d) $1s^22s^22p^63s^23p^64s^24p^65s^2$
e) $1s^22s^1$
f) $1s^22s^22p^63s^23p^5$