

Chapter 24

CONFIGURATIONS AND CONSEQUENCES, Part 2

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Right now we need to come back to ions as found in typical ionic compounds. When we first set up ions in Chapter 3, we distinguished two categories, "constant charge" and "variable charge". Let's return to several points from then.

“ Some elements can form different ions with different charges. Iron is good at this: some iron ions are Fe^{2+} ($24 e^-$) and some are Fe^{3+} ($23 e^-$). Other ions, however, are locked into the same charge all the time. Calcium ions will only be Ca^{2+} . We distinguish these cases as "variable charge ions" and "constant charge ions". "Variable charge" applies for elements (such as Fe) which can form ions of different charges in their different, common ionic compounds. "Constant charge" involves those elements (such as Ca) which will always form the same ion of the same charge in their common ionic compounds. Ca^{2+} and numerous examples above (Na^+ , F^- , O^{2-} , Mg^{2+} , Al^{3+} and N^{3-}) are constant charge cases. You won't find an ionic compound with a Na^{2+} cation. You won't find an ionic compound with a F^{2-} anion. Et cetera. ”

“ Why do ions of some elements form the same charge all the time and others can form different charges? Why are some ions 1+ while others are 2- while others are 3+ or whatever? It's in the electron arrangements. It depends on how many electrons are in the atom, where they are in the atom and what are their energies. Again, greater details and explanation can be found in later Chapters. We don't need that right now because we are more concerned with what are the charges. We can do why later. We already have a tool to help us with what are the charges: it's the Periodic Table. Remember in the last Chapter that I said the Periodic Table is related to electron arrangements. I just now said that charges are also related to electron arrangements. Put the two together and you realize that charges and the Periodic Table are related. You can use the Periodic Table as a guide for what charges a particular element typically will adopt in its monatomic ions. These are guidelines, not rules. They're good but not perfect. There will be exceptions. ”

I said all that in Chapter 3. We've come a long way, haven't we? Here we are in Chapter 24 with much of the explanation at hand: configurations. We can now consider some of the details of why some ions are 1+ while others are 2- while others are 3+, etc. With the details, we can understand the guidelines better and we can see how exceptions happen.

24.1 Ion charge

There are two Points to note for the general relationship between configuration and ion charge.

POINT 1. Monatomic ions of *s*- and *p*-block elements will possess a charge which gives them a more favorable electron configuration.

Note that I limited Point 1 to Main Group cases. Monatomic ions of elements within the *d*- and *f*-blocks are subject to more subtleties. We will mention some cases but only to the extent that they adopt a more favorable configuration as covered here.

POINT 2. The ability to achieve a more favorable electron configuration will be tempered by energy constraints.

The energy constraints are very important and they will affect cations and anions differently. For cations, all ionizations have a positive ΔE which means that, somewhere along the way, you have to pay for each and every electron removed to form that cation. Actually, it's the reaction that pays, either in bond energies or the energies of phases or whatever. CATIONS CAN FORM FROM A NEUTRAL ATOM WHEN THE REACTION PROVIDES ENOUGH KJ'S TO COVER THE ENERGIES OF IONIZATION. When you hit a step whose IE is too expensive, then further ionization is blocked. This leaves you at some ion charge and at some configuration, which typically corresponds to one of the more favorable configurations. In the case of anions, ELEMENTS TEND TO FORM ANIONS WHEN THEY HAVE A REASONABLE ELECTRON AFFINITY. In most of the common cases of elements, gaining one electron has a negative ΔE so this much is already favorable. Gaining two or more electrons is not favorable, however, and we are again faced with the fact that the reaction must adequately cover the cost. For the common anions, this is indeed the case.

So what are these "more favorable" electron configurations which I keep mentioning? We've already covered one: the noble gas configuration. Many Main Group elements will form ions of some charge which leaves them with a noble gas configuration. For example, Group 1 metals have a small IE for the first

ionization. Let's use potassium as an example. K has the configuration $[\text{Ar}]4s^1$. That lone electron in $4s$ is not held very tightly and its loss is easily paid for by the energies of a typical chemical reaction. By losing the $4s$ electron, we now have K^+ with the configuration $[\text{Ar}]$. Now, look at what you've got! You've got a noble gas configuration of just core electrons. You would have to pay BIG time to pull out a second electron. It's not going to happen, not with the energy available. So K will form K^+ , but it will not form K^{2+} .

Notice that I refer to Group 1 metals, and I'm leaving out H. H does not form compounds with monatomic H^+ , remember? Also, as we go along here, keep in mind that the energy constraints refer to the energies of typical chemical reactions. With some instruments, you can pull core electrons out of monatomic ions. Out in space, with the astronomical energies available out there, you can find all sorts of bizarre ions such as O^{7+} , C^{6+} , Si^{9+} , etc. which are present in the solar wind. All of that is well beyond common reaction energies and we are not dealing with those cases here.

The other Group 1 metals are similar to K. The neutral atoms have some noble gas configuration followed by ns^1 , which I will depict as $[]ns^1$. They can lose one electron fairly easily but that's all. No more. You end up with a $1+$ cation with a noble gas configuration.

Let's mosey to Group 2. Let's do Mg, since we looked at the actual numbers for its first three ionization steps in the last Chapter. Neutral Mg is $[\text{Ne}]3s^2$. It's somewhat easy to remove the first electron; it's harder to remove the second but it's still in the range of chemical reaction energies. This puts us at Mg^{2+} , which has the noble gas configuration $[\text{Ne}]$. Don't bother trying for Mg^{3+} ; you're not going to pull out core electrons with reaction energies. It's too expensive. We can extend this result to all of Group 2. Their neutral atoms are $[]ns^2$. They form $2+$ ions and this corresponds to a noble gas configuration.

In Group 13, boron doesn't do monatomic cations but Al does. (Boron does covalent bonds.) Neutral Al has the configuration $[\text{Ne}]3s^23p^1$. It can ionize over three steps although this is now getting quite costly. Nevertheless, we are still in the range of reaction energies to do this. This leaves Al^{3+} with the noble gas configuration $[\text{Ne}]$. That's it. You get Al^{3+} and that's where it ends.

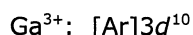
Let me do some anions. Let's jump over to Group 17 and check out a halogen. We'll do Cl again. Neutral chlorine has the configuration $[\text{Ne}]3s^23p^5$. Cl has a very high EA, so it accepts another electron very readily to form Cl^- . This has the noble gas configuration $[\text{Ar}]$. Chloride ion has no desire whatsoever to exceed noble gas configuration by taking on a second electron; that would give the $[\text{Ar}]4s^1$ configuration, and that electron in the bigger $4s$ orbital cannot hold on. Thus, adding an electron stops at $1-$. All other halogens are similar. They will all readily add one electron (and only one electron) to form a $1-$ anion with a noble gas configuration.

Go to Group 16. Let's pick selenium. Neutral Se has the configuration $[\text{Ar}]4s^23d^{10}4p^4$. It has a reasonably high EA for one more electron. It does not have a favorable affinity to take a second electron on its own, but this energy can be provided by a typical chemical reaction. This brings it to Se^{2-} with the noble gas configuration $[\text{Kr}]$, and this is where it stops. Adding another electron to give $[\text{Kr}]5s^1$ won't happen. Other members of Group 16 behave similarly, forming $2-$ anions with a noble gas configuration.

We can continue this into Group 15 but it's getting a bit tough to do. Nitrogen is the best of the Group to form an anion. A neutral nitrogen atom has the configuration $[\text{He}]2s^22p^3$. It is not willing, but it is able under some conditions, to take on three more electrons; this forms the nitride ion with the configuration $[\text{Ne}]$. You can forget about forcing it to take a fourth electron. It won't happen. Other members of Group 15 would behave similarly, forming a $3-$ anion with a noble gas configuration but, in all honesty, this is getting difficult and it's pretty much limited to N^{3-} and a few compounds of P^{3-} .

All of the above examples so far adopt a charge corresponding to a noble gas configuration. This works very well for all common anions and for many cations, but many other cations cannot do this. For example, gallium is right below Al in Group 13 but there's no way it could hit a noble gas configuration for its cation; it would have to lose thirteen electrons. Forget it! Don't even try. So what's a gallium to do? Simple: it goes for another of the "more favorable" electron configurations.

The "more favorable" electron configurations are all characterized by having full subshells. As such, instead of calling them "more favorable" configurations, they are called full subshell configurations. In a full subshell configuration, every subshell with electrons is completely filled. All of the ions above have noble gas configurations and all of those are full subshell configurations. But there are other configurations also with full subshells. For example, we discussed Ga^+ and Ga^{3+} in the last Chapter.



These are not noble gas configurations but they are still full subshell configurations because all occupied subshells are completely full. Compare those to neutral Ga, $[\text{Ar}]4s^23d^{10}4p^1$, which is not a full subshell configuration because the $4p$ is occupied but it is not full. Thallium, Tl, two Periods below Ga in Group 13, is similar to Ga but Tl now involves an f portion.



Continuing into Group 14, tin and lead commonly form ions of 2+ and 4+ charge, all of which are full subshell configurations. In fact, the configurations of Pb^{2+} and Pb^{4+} are the same as for Tl^+ and Tl^{3+} .

With full subshell configurations, we are now able to explain some of the guidelines from Chapter 3 regarding monatomic ion charges. Here are two of the guidelines (Categories 1 and 3) from Chapter 3.

“ • CATEGORY 1. CONSTANT CHARGE CATIONS

A. Metal elements of Groups 1, 2 and 13 form cations of charge 1+, 2+, and 3+, respectively. ”

“ • CATEGORY 3. CONSTANT CHARGE ANIONS

The nonmetal and metalloid elements commonly form anions in which the charge is equal to the Group number minus eighteen. ”

All of these are full subshell cases. Let's summarize this.

Group 1 forms cations of 1+ charge

Group 2 forms cations of 2+ charge

Group 13: Al forms a cation of 3+ charge and

Ga, In, Tl form cations of 3+ charge

Group 15 forms anions of 3- charge

Group 16 forms anions of 2- charge

Group 17 forms anions of 1- charge

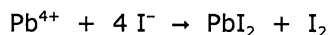
noble gas configuration

other full subshell configuration

I said in Chapter 3 that those guidelines were just guidelines. We are now in a position to refine them. I've already made reference to one refinement: some members of Group 13 can form variable charge cations, 3+ or 1+. All are full subshell configuration. Monatomic ions with full subshells are extremely common, and these dominate Main Group ionic compounds.

We are now in a position to explain another very important notion that goes back to Chapter 3. Why are there variable charge cations? Look at thallium as an example. Ionic compounds of thallium are either Tl^+ or Tl^{3+} . Why is it sometimes 1+ and sometimes 3+? This has to do with the energies involved in making the compounds. When making a Tl^+ compound, the total chemical energies of the process must pay for ending up with a thallium cation of 1+ charge. This is easy to do. When making a Tl^{3+} compound, the total chemical energies of the process must pay for ending up with a thallium cation of 3+ charge. Not all chemical reactions for thallium compounds can cover the expense of Tl^{3+} but usually there's enough energy for Tl^+ . So, depending on the reaction, you get a compound with Tl^+ cations or with Tl^{3+} cations. On the other hand, you will NOT get a compound with monatomic Tl^{2+} .

Here's another example. In Chapter 3, we left tin and lead as variable charge cations. The charges are 2+ or 4+, as noted above, and nothing else. In all honesty, however, it gets very difficult to cover the expense of 4+ cations and this is especially true for lead. Most (not all) ionic compounds of lead are 2+. For example, if you try to make PbI_4 , you end up with PbI_2 . You could look at this in terms of redox.



Despite the fact that iodine has a considerable EA and that I^- is a very common anion, there's not enough energy in the total picture to keep Pb^{4+} as the 4+ ion. The third and fourth IEs for Pb^{4+} are just too expensive and Pb wants those electrons back badly. As a result, Pb^{4+} oxidizes some of the I^- ions and you get a product of Pb^{2+} . Now, let's compare that to PbF_4 . Although Pb^{4+} still wants some electrons back badly, the much greater EA of fluorine to form F^- helps to oppose that, and this compound is stable. These are the various considerations which can go into the variable charge cases. As far as configuration goes, notice that you do NOT get Pb^{3+} and you do NOT get Pb^+ .

The above consideration nicely explains the variable charge cations of the Main Groups. Actually, however, most variable charge cations reside among the transition elements and they have other factors operating. I'm not covering those other factors here.

Let me add a brief note about anions here. This is what I said in Chapter 23 when it was first noted that the common anions had noble gas configurations.

“ Notice that I said this applies for the common anions. Except for H^- , the common anions are all in the p -block. There are uncommon anions from other blocks which don't follow a noble gas configuration, like Au^- . Be aware that this can happen. ”

Au^- ? Anions from other blocks of the Periodic Table are not commonly encountered but they do exist, such as the gold anion, Au^- . It's called "auride"; the aur- prefix is from the symbol's Latin name, aurum, and the suffix -ide is the standard ending for an anion. Although I didn't cover EA for d - and f -block elements, gold has a very high electron affinity. In fact, gold's EA is greater than the EA of all elements except for the halogens. The configuration of auride is $[\text{Xe}]6s^24f^{14}5d^{10}$. This is yet another example to highlight the importance of full subshell configurations.

Although I've been emphasizing the importance of full subshell configurations for the monatomic ions of the Main Groups, we can extend this notion to the left and right ends of the d - and f -blocks. Let's illustrate this with one more example: predict the charge for a cadmium cation. Cadmium (Cd) is just barely in the d -block (Group 12) and its ion follows our coverage here. In Chapter 23, we did the configuration of a neutral cadmium atom: $[\text{Kr}]5s^24d^{10}$. To predict what cation Cd forms, start taking out electrons and look for a full subshell configuration. Cd^+ would have the configuration $[\text{Kr}]5s^14d^{10}$, which is not particularly favorable. Cd^{2+} would have the configuration $[\text{Kr}]4d^{10}$, which is a full subshell configuration. Taking out another electron and another and another does not get you anything better; it just gets more and more expensive. So cadmium is expected to form a $2+$ cation, and that is what it does.

24.2 Sizing up an ion

We talked about sizes of neutral atoms earlier. Now we talk about ion sizes. In order to talk about ion sizes, we have to discuss cations separately from anions.

The size of any cation is always smaller than its neutral atom. The explanation is related to various aspects which we've covered already. Cations have fewer electrons than their neutral atoms. Fewer electrons mean fewer electron-electron repulsions. Fewer repulsions mean the remaining electrons can come in closer to the nucleus. This gives an overall size contraction. For each electron lost, the size decreases. This means that a higher cation charge gives a smaller cation. For example, Fe^{2+} is 78 pm in radius, while Fe^{3+} is 64 pm.

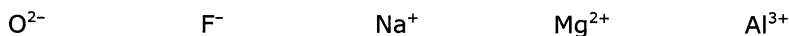
There is an added effect to this contraction when the valence shell empties and you are left with only core electrons. Why? Remember that the valence shell is the largest n -shell and this is what determines size. When you empty the largest shell, the size is determined by a smaller shell. Here's a comparison to illustrate this point.

	Na	versus	Na^+
configuration	$1s^22s^22p^63s^1$		$1s^22s^22p^6$
size (radius)	154 pm		102 pm

Both effects are operating here: the loss of one electron and the loss of the highest n -shell. Unfortunately, this comparison is somewhat approximate because we are mixing the types of bonds which are used in these measurements. As noted in Chapter 23, our atomic radii are derived from bonds in covalent compounds; the ionic radii here are derived from ionic compounds. So, although the difference in the numbers can be regarded as approximate, the comparison is valid qualitatively.

Anions are of the opposite effect: anions are always larger than the neutral atom. The reasons can be anticipated. More electrons give more repulsions and everybody spreads out. For example, a neutral Cl atom is 99 pm in radius, while the radius of Cl^- is 181 pm. Again and for the same reasons, the difference in the numbers can be regarded as approximate, but the comparison is valid qualitatively.

There is one more illustrative comparison to make for ions, and it deals with "isoelectronic" ions. The prefix "iso-" is used in the normal sense, meaning equal or same: "isoelectronic" means having the same number of electrons. Let's compare sizes for the following isoelectronic series of ions.



All of these have ten electrons and they all have the $[\text{Ne}]$ configuration. The BIG difference between them is that they all differ in their number of protons: nuclear charge increases left-to-right in the line. A higher nuclear charge draws in all electrons more, and this is reflected in the smaller sizes.

	O ²⁻	F ⁻	Na ⁺	Mg ²⁺	Al ³⁺
Z	8	9	11	12	13
pm	140.	133	102	72	54

(This is now a comparison of ionic sizes only, so this is a valid comparison quantitatively.) The difference in the sizes is quite substantial.

24.3 Taking a spin on magnetism

Now we come back to the notion of magnetism. I'm going to tie a few things together from this Chapter and from Chapter 22, and I'll even go back to a question I posed in Chapter 2.

As previously noted in the very beginning of Chapter 22, the electron has a magnetism as a consequence of m_s . This can lead to magnetic properties for the item as a whole. The "item" here can be an atom or a molecule or a polyatomic ion or even a chunk of metal. If all electrons in the item are paired, then half of the electrons have m_s of $+1/2$ and the other half have m_s of $-1/2$. Under these conditions, all magnetic effects of the electrons are cancelled and the item cannot be magnetic overall. If the item has one or more unpaired electrons, then some magnetic properties can occur.

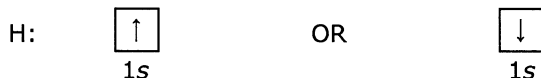
I keep saying "magnetic properties". There are actually several kinds of magnetic properties. Permanent magnets are one kind and you are familiar with these. Another familiar kind of magnetic property is exhibited by some metals that are not permanently magnetic but which can be attracted to a magnet. Many everyday things which are made of iron are in this category: nails, paper clips, etc. These things are not magnetic by themselves but you can pick them up with a magnet. You can't do this to aluminum foil. It won't work. You can't do this with brass or some forms of stainless steel. It won't work. By the way, if someone ever tries to sell you something which is supposed to be all brass, put a magnet to it. If it's attracted to the magnet, then it may not be brass or it may only be brass-plated. I keep this in mind even when I'm just buying brass nuts, bolts, fixtures or whatever: I like to check them with a magnet.

These common examples are bulk magnetic properties. They apply to a sample as a whole. I would like to get down to individual chemical units, either monatomic or polyatomic. Some chemical units have all of their electrons paired and others have some electrons that are unpaired. This can have important consequences. Those with one or more unpaired electrons are attracted to a magnetic field. This leads to various types of instruments which are used to measure unpaired electrons. Some of these instruments can tell you how many unpaired electrons are in the chemical unit; other instruments can also tell you which atoms in a polyatomic unit hold the unpaired electrons. There's also an extremely important difference in the way chemical units react, depending on whether they have all of their electrons paired or not. Perhaps the most obvious example is the very oxygen which you need to live. Normal O₂ in the air has two unpaired electrons in the molecule. You can make another form of O₂ with all of its electrons paired. These two react very differently. The version of O₂ with all electrons paired is even far more deadly than the normal O₂. That's Nature at work again. Subtle differences, dire consequences.

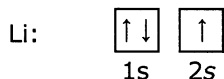
We need to introduce some new terminology here. A "diamagnetic" chemical unit has all of its electrons paired. A "paramagnetic" chemical unit has one or more electrons unpaired. These words also come as nouns: diamagnetism and paramagnetism. Be careful with these definitions. Diamagnetic means ALL electrons are paired. Paramagnetic can have some paired but it's got to have one or more unpaired. It's the number unpaired that is important.

For the rest of this discussion in this Chapter, I will limit to monatomics. In Chapter 26 and later Chapters, we will see diamagnetism and paramagnetism for polyatomics.

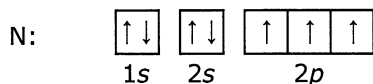
Let's return to some of the box configurations in Chapter 22. Here's H.



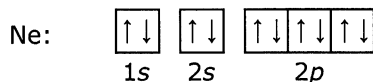
We can now say that a hydrogen atom is paramagnetic. It has one unpaired electron. Let's do lithium again.



A lithium atom is paramagnetic. It has one unpaired electron. Let's go to nitrogen.



A nitrogen atom is paramagnetic. It has three unpaired electrons. Let's go to neon.



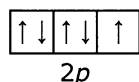
Neon is diamagnetic.

To help you with the question of dia/paramagnetism, here are some clues. These are handy now for monatomics and they will also be related to polyatomics in Chapter 32. Flag this part: put a blue star in the margin here, so I can refer you back to this later.

Firstly, any chemical unit with an odd number of electrons **MUST** be paramagnetic. This is straightforward: you can't pair everything when the total is an odd number. Secondly, if the unit contains an even number of electrons, then it may be diamagnetic or it may be paramagnetic and you need to look more closely at it. Remember what I said in Section 22.3.

- Whenever subshells are completely filled, all electrons in them will be paired.
- Whenever a subshell is only partly filled, one or more electrons will be unpaired.

These statements are always true. As I mentioned earlier, paired and unpaired electrons are more easily seen in box notation than in written notation. For example, can you tell how many electrons are unpaired in the configuration $1s^2 2s^2 2p^5$? The number of unpaired electrons has important consequences, but it's not so straightforward to see from the written configuration alone without boxes. It will help if you apply these two points. For this configuration, 1s and 2s are filled so their electrons are all paired up; don't bother to look for unpaired electrons there. 2p is only partly filled, so that guarantees that one or more electrons are unpaired. Once you've reached that conclusion, it's easier if you visualize the box notation for that portion. How would five electrons in a 2p subshell look? Picture it.

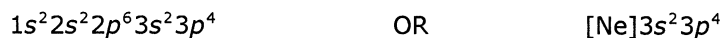


This clearly shows that there is one unpaired electron. ”

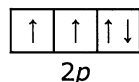
I had pointed this out in Chapter 22 and it's very important now for determining diamagnetism and paramagnetism. Keep these things in mind. Look for full versus partly filled subshells. Let's run through a few more examples of dia/paramagnetic monatomics.

► S

A sulfur atom has sixteen electrons. This is an even number, so it could be either dia or para. Check the configuration.



The 3p subshell is only partly filled; therefore, S must be paramagnetic. To find out how many electrons are unpaired, look only at the partly-filled subshell. Do not even bother with full subshells, since all electrons in full subshells must be paired. This also means you can skip everything within the noble gas part, []. For S, only 3p is partly filled. Consider its box diagram.



You can see it has two unpaired electrons.

► S²⁻

Eighteen electrons, even number, could be dia or para. We need the configuration.



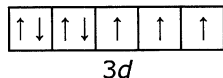
No question here: all subshells are filled. That means everybody is paired. Diamagnetic.

► Co

27 electrons: odd number, must be paramagnetic. How many electrons are unpaired? Configuration, please.



Only the $3d$ subshell is partly filled. That's the only place we have to look.



There are three unpaired electrons.

► Co³⁺

You do this one by yourself.

I will give you a clue: compared to neutral Co, the ion has fewer total electrons but it has more unpaired electrons. Go ahead and work it out. But be careful! A common error is messing up the configuration for the cation. You need to remember the steps for doing cations.

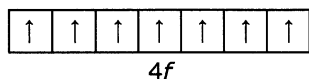
That ends our examples.

OK, I told you above that I would return you to a question I posed in Chapter 2. Here's that question.

“ I will also note that gadolinium (Gd) has a very important medicinal role: gadolinium compounds have been used for years in MRI (Magnetic Resonance Imaging), although side effects are known. If you've ever had MRI, you may have had a gadolinium compound injected into you. Why gadolinium? Ask me in Chapter 24. ”

Well, here we are in Chapter 24. Time to answer the question.

In MRI, the instrument is detecting the hydrogen atoms in water and in fats within your body. MRI uses radiofrequency EM radiation: water and fat molecules in different tissues and organs respond differently to this radiation. The MRI instrument detects this difference, and this allows an image of the tissues and organs to be obtained. To get better pictures with better contrast, they often inject you with a solution of a paramagnetic compound. For this to work best, they want as many unpaired electrons as are possible in that compound. This is where gadolinium fits in. Gd is smack dab in the middle of the f -block and the compound they use involves Gd³⁺. What's so great about Gd³⁺? Configuration. The configuration of Gd³⁺ is [Xe]4f⁷. Check out the $4f$ boxes.



Now, that's a whopper of a paramagnet: seven unpaired electrons.

Although that is indeed a lot of unpaired electrons, it's actually not the maximum for monatomic ions in general. For example, europium (Eu) can form a Eu⁺ cation in the gas phase which has [Xe]6s¹4f⁷ configuration with eight unpaired electrons. But that's gas phase; Eu⁺ does not readily form stable compounds, and Eu³⁺ is the typical cation for europium.

Problems

1. True or false.
 - a. All neutral atoms in Group 2 have a full subshell configuration.
 - b. All neutral atoms in Group 13 are paramagnetic.
 - c. A cation is larger than its neutral atom.
 - d. Sulfide anion and potassium cation are isoelectronic.

- e. All atoms with an even number of electrons are diamagnetic.
2. True or false.
- All Group 1 metal elements form cations which have a noble gas configuration.
 - Na^+ is larger than F^- .
 - Br^- and Rb^+ are isoelectronic.
 - V^{3+} is larger than V^{2+} .
 - All neutral atoms in Group 14 are diamagnetic.
3. Which of the following have a full subshell configuration?
- Ca Ca^{2+} Y^{2+} Mg^{2+} Cu^{2+} U^{4+}
4. Which of the following are diamagnetic?
- Al^{3+} Ni^{2+} P Zn^{2+} V^{3+} H
5. How many unpaired electrons are in each of the following?
- Cr^{3+}
 - Sn
 - Ru^{3+}
 - F
 - Fl^{2+}
6. How many unpaired electrons are in each of the following?
- W^{2+}
 - O
 - Ba
 - Nd^{3+}
 - Rb
- 7.
- What elements in Period 5 form a 2+ cation with four unpaired electrons?
 - What element forms a 2+ cation which is isoelectronic with Xe?
 - What element forms a 2+ cation which is isoelectronic with Al^{3+} ?
- 8.
- How many elements in the whole Periodic Table can form a 2+ cation with noble gas configuration?
 - What is the most common cation for Bi? (It's not Bi^{5+} .)
 - Which elements in the Fourth Period have two unpaired electrons in their neutral atoms?
9. Which one of the following is smallest?
- Br Br^- Rb Se^{2-} Te I
10. Which one of the following is largest?
- Te^{2-} I^- Br Kr Rb^+ Cs^+