

Chapter 24

CONFIGURATIONS AND
CONSEQUENCES, Part 2

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Right now we need to come back to ions. 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 all 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. Virtually all monatomic ions of *s*- and *p*-block elements will possess a charge which gives them one of the more favorable electron configurations.

I limited that statement 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 one of the more favorable configurations 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 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, 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 the first 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 [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. In outer 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 move 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 [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 [Ne]. Don't bother trying for Mg^{3+} ; you're not going to pull out core 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 cations but Al does. (Boron only does covalents.) Neutral Al has the configuration [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 [Ne]. That's it. You get Al^{3+} and that's where it ends.

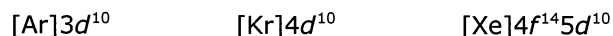
Let me do some anions. Let's jump over to Group 17 and check out a halogen. We'll do Cl again. Neutral chlorine has the configuration [Ne] $3s^23p^5$. Cl has a very high EA, so it accepts another electron very readily to form Cl^- . This has the noble gas configuration [Ar]. Chloride ion has no desire whatsoever to exceed noble gas configuration by taking on a second electron, and it cannot be forced to do so. Thus, it 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 [Ar] $4s^23d^{10}4p^4$. It has a reasonably high EA for one more electron. It does not have a favorable energy 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 [Kr], and this is where it stops. It has no desire whatsoever to take on a third electron and it cannot be forced to do so. 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 [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 [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 tough to do and it's pretty much limited to the highest members in the Group.

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 one of the other "more favorable" electron configurations.

Our second of the "more favorable" electron configurations is called pseudo-noble gas configuration. It's called "pseudo-" because it's not a true noble gas configuration: it's got a full *d*-subshell out beyond the noble gas part. Some cases will also have a full *f*-subshell but this will only apply for elements in the Sixth Period. There are only three pseudo-noble gas configurations.



(A fourth, [Rn] $5f^{14}6d^{10}$, is also possible for some Period 7 ions but not much is known about those cases yet.) Let's illustrate this with gallium, since we looked at Ga^+ and Ga^{3+} ions in the last Chapter. We only

need Ga^{3+} here. Neutral Ga is $[\text{Ar}]4s^23d^{10}4p^1$. The Ga^{3+} ion is $[\text{Ar}]3d^{10}$; that's pseudo-noble gas. The lower members of Group 13, In and Tl, also form 3+ cations with pseudo-noble gas configurations. Tl^{3+} has the f part included: $[\text{Xe}]4f^{14}5d^{10}$. Tin and lead in Group 14 will form 4+ ions of pseudo-noble gas configurations. As a matter of fact, Pb^{4+} has the same configuration as Tl^{3+} . Go ahead, check it out yourself.

With noble gas configuration and pseudo-noble gas configuration, we are now able to explain some of the guidelines from Chapter 3 regarding 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. ”

Now we can see why: these charges correspond to favorable configurations. 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 Group 15 forms anions of 3- charge Group 16 forms anions of 2- charge Group 17 forms anions of 1- charge	and	Ga, In, Tl form cations of 3+ charge
noble gas configuration		pseudo-noble gas 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. The cation of 3+ charge is explained as noble gas or pseudo-noble gas configuration, as shown above. On the other hand, some elements in Group 13 can also form a 1+ cation. How so? This refinement brings us to our third "more favorable" type of configuration. This one has no special name. I'll just say "another full subshell configuration". The key adjective is full; a partly filled subshell doesn't count here. For ions with full subshell configurations, every electron is paired.

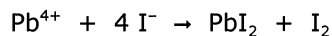
Things with full subshells are extremely common and these dominate Main Group compounds. As a matter of fact, both of the above configuration types, noble gas and pseudo-noble gas, are also full-subshell configurations. For this third type of configuration, when I say "another full subshell configuration", I mean other than noble gas or pseudo-noble gas.

This third type covers a number of other cations and it also covers several "uncommon" anions. I've actually mentioned one cation example already, Ga^+ which has the configuration $[\text{Ar}]4s^23d^{10}$. It's not noble gas and it's not pseudo-noble gas (since it's got the $4s^2$ part), but it's still a full subshell configuration. The other members of Group 13 which can form a 1+ cation also achieve this. For example, Tl^+ has the configuration $[\text{Xe}]6s^24f^{14}5d^{10}$.

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 have 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+} because its configuration is not favorable.

In Chapter 3, we left tin and lead as variable charge cations. The charges are 2+ or 4+ and nothing else. Sn^{4+} and Pb^{4+} have pseudo-noble gas configurations; Sn^{2+} and Pb^{2+} have another full subshell configuration. 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 decent EA and that I^- is a good 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 very badly. As a result, Pb^{4+} oxidizes some of the I^- ; you get Pb^{2+} , corresponding to a full subshell configuration. On the other hand, let's compare that to PbF_4 . Although Pb^{4+} still wants some electrons back very 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^+ , simply because those are not favored configurations.

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 cases here.

I mentioned that several "uncommon" anions also have another full subshell configuration. As I said in Chapter 23, all common anions come from the p -block and all of those anions have a noble gas configuration. Anions from other blocks of the Periodic Table are not commonly encountered, but they do exist, such as the gold anion, Au^- , as noted in Chapter 23. It's called "auride"; the aur- prefix is from the symbol's Latin name, aurum. 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 $[\text{Xe}]6s^24f^{14}5d^{10}$. Notice that this is a full subshell configuration.

Let's recap.

Of the many different electron configurations, some are more favorable than others. These include noble gas configuration, pseudo-noble gas configuration, and any other full subshell configuration. These three types account for virtually all cations and anions from the Main Groups. These also account for some ions near the ends of the d - and f -blocks, but not from the middle of those blocks. (Additional factors operate within the d - and f -blocks which we aren't covering here.)

Let's put our "more favorable" configurations to some use with one more example: predict the charge for a cadmium cation. Cadmium is just barely in the d -block 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 it forms, start taking out electrons and look for a favorable configuration. Cd^+ would have the configuration $[\text{Kr}]5s^14d^{10}$, which is not particularly favorable. Cd^{2+} would have the configuration $[\text{Kr}]5d^{10}$, which is pseudo-noble gas 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 stop at 2+ for its 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

Notice the big drop in size when emptying the largest shell.

Anions are of the opposite effect: anions are always larger than the neutral atom. The reasons can be anticipated. More electrons gives more repulsions and everybody spreads out. These effects are drastic and the expansion of anions is very large, even for the same shell. For example, a neutral Cl atom is 99 pm in radius, while the radius of Cl^- is 181 pm.

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.

O^{2-} F^- Na^+ Mg^{2+} Al^{3+}

All of these have ten electrons and they all have the [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

It would have been nice to include neutral Ne in our comparison here, but that would not be a fair comparison. As noted in Chapter 23, there are different ways of measuring sizes. The atomic radii for neutrals given in Chapter 23 were derived from covalent compounds. The ionic radii given here are derived from ionic compounds. Neon doesn't do either type, so it does not have a value which can fit into our comparison here.

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, magnetism is 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 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.

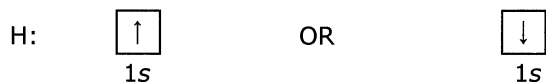
These common examples are bulk magnetic properties. They apply to a sample as a whole. I would like to get down to individual chemical units, either monatomic or polyatomic. Some chemical units have all of their electrons paired and others have some electrons that are unpaired. This can have important consequences. Those with one or more unpaired electrons are attracted to a magnetic field. This leads to many forms 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 O_2 in the air has two unpaired electrons in the molecule. You can make another form of O_2 with all of its electrons paired. These two react very differently. The version of O_2 with all electrons paired is even far more deadly than the normal O_2 . 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

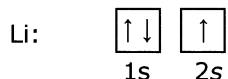
ALL electrons paired. Paramagnetic can have some paired but it's got to have one or more unpaired. It's the number unpaired that is important.

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.

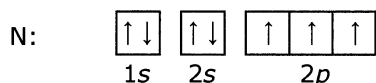
Let's return to some of the box configurations in Chapter 22. Here's H.



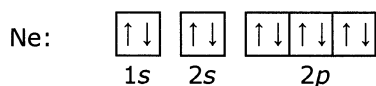
We can now say that a hydrogen atom is paramagnetic. It has one unpaired electron. Let's do lithium again.



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.



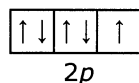
Neon is diamagnetic.

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, everything with an odd number of electrons **MUST** be paramagnetic. This is straightforward: you can't get pairs out of an odd number of anything. 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 exactly 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. ”

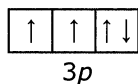
I had pointed this out in Chapter 22 and it's very important now for determining diamagnetism and paramagnetism. Keep these things in mind. Look for full versus partly filled subshells. Let's run through a few more examples of dia/paramagnetic monatomics.

► 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.

► S^{2-}

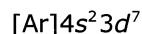
Eighteen electrons, even number, could be dia or para. We need the configuration.



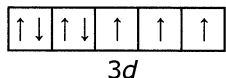
No question here: all subshells are filled. That means everybody is paired. Diamagnetic.

► 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.

► Co^{3+}

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.

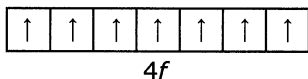
That ends our examples.

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

“ Gadolinium (Gd) compounds are routinely used in medicine for MRI (Magnetic Resonance Imaging). 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 the water molecules in your cells. MRI uses radiofrequency EM radiation: water molecules in different cells 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 Gd^{3+} . What's so great about Gd^{3+} ? Configuration. The configuration of Gd^{3+} is $[\text{Xe}] 4f^7$. Check out the $4f$ boxes.



Now, that's a whopper of a paramagnet: seven unpaired electrons. That's the most unpaired electrons which you can have for any monatomic cation.

Problems

- True or false.
 - No element in Period 3 can have a pseudo-noble gas configuration.
 - All neutral atoms in Group 2 have a full subshell configuration.
 - A cation is larger than its neutral atom.
 - Sulfide anion and potassium cation are isoelectronic.
 - All atoms with an even number of electrons are diamagnetic.
- True or false.
 - All Group 1 metals form cations which have a noble gas configuration.
 - F^- is larger than Na^+ .
 - Br^- and Rb^+ are isoelectronic.
 - V^{3+} is bigger than V^{2+} .
 - All neutral atoms in Group 14 are diamagnetic.
- Which of the following have a pseudo-noble gas configuration?
 In^{3+} Zn Hg^{2+} Sn^{4+} Bi^{3+} Al^{3+}
- Which of the following have a full subshell configuration?
 Ca Ca^{2+} Y^{2+} Mg^{2+} Cu^{2+} Cd^{2+}
- Which of the following are diamagnetic?
 Al^{3+} Ni^{2+} P Zn^{2+} V^{3+} H
- How many unpaired electrons are in each of the following?
 - Co^{3+}
 - Sn
 - Ru^{3+}
 - F
- How many unpaired electrons are in each of the following?
 - W^{2+}
 - O
 - Ba
 - Nd
- 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 to Al^{3+} ?
- Which one of the following is smallest?
 Br Br^- Rb Se^{2-} Te I
- Which one of the following is largest?
 Te^{2-} I^- Br Kr Rb^+ Cs^+