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Wang Hai-Jun

ABSTRACT

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I. INTRODUCTION

Ion potential is an important physical quantity used to measure the ability of ions or charged atoms to attract (or repel) electrons (clouds). It can qualitatively or semi-quantitatively explain some chemical facts and is widely used in the field of chemistry.

1.1 The meaning of ionic potential

1.1.1 Definition of G.H.carteledge [1]

The ratio of ionic charge to ionic radius, i. e.

$$\varphi=Z/r$$

In the upper formula, Z is charged with ions and r is ionic radius (pm).

1.1.2 Definition of structural chemistry

$$\varphi=Z^*/r^*$$

In the formula above, Z^* is the effective charge [2] (net charge or formal charge) of an ion or atom, and r^* is the effective ion radius or the effective covalent radius of an atom. The effective charge Z^* can be determined by Slater approximation rule [2] 65 according to the principle of drilling through effect and shielding effect. The meaning and numerical value of r^* can be found in reference [2] These definitions are applicable to simple ions and charged atoms (such as B^{+3}, S^{+6}, Cl^{+7}, etc.), and can also be regarded as positive ions. They are more accurate, scientific, factual and widely used. For example:

$$S^{+6}: Z^*=+6,\quad r^*=30pm,\quad \varphi(S^{+6})=0.200$$

$$Mn^{+7}: Z^*=+7,\quad r^*=25pm,\quad \varphi(Mn^{+7})=0.280$$

$$S^{2-}: Z^*=-2,\quad r^*=184pm,\quad \varphi(S^{2-})=-0.019$$

The physical meaning of ion potential can be simply understood as: ion potential is a physical quantity used to measure the ability of an ion or charged atom to attract or repulse electrons (clouds).

1. $\varphi>0$, Positive ion $\varphi>0$, Negative ion $\varphi<0$; φ(Positive ion)》φ(Negative ion) ;
2. Positive ions play an important role in attracting electrons (clouds).φThe greater the φ , the stronger the ability to attract electrons (clouds) ;
3. The negative ions are mainly repulsive to electrons (clouds). The smaller the φ, the stronger the ability to reject electrons (clouds).

II. MAIN FACTORS AFFECTING ION POTENTIAL OF POSITIVE IONS

2.1 Z^* Effective charge Z^* of ions

hen r^* is close, the larger the Z^*, the larger the φ .

2.2 Effective radius of ions r^*

When Z^* is the same, the smaller the r^*, the larger the φ.

2.3 Electron layer structure of ions

The positive ions, which are similar to Z^* and r^*, are mainly determined by the valence electron structure of ions. The relationship between ionic potential and valence electron layer structure (configuration) is roughly:

$$\varphi(8e^-)<\varphi(9e^-—17e^-)<\varphi(18e^-,(18+2)e^-)$$

For example, Zn^{2+} and Ag^+ are $18e^-$ positive ions with larger ionic potential. While Na^+ and K^+ are the positive ions of the $8e^-$ electron configuration, the ionic potential is smaller. This is mainly due to the fact that the valence electron layer of $9e^-$-$17e^-$or $18e^-$or $(18+2)$ e^-positive ions contains $\boldsymbol{d^x}$ electrons, while that of $8e^-$positive ions has no ***d*** electrons. Because d electron cloud has smaller penetration effect, weaker nuclear control, farther extension and smaller shielding effect on the nucleus, its effective nuclear charge Z^* increases with the increase of x, which leads to the increase of ion potential φ.

[Explanation] The factors affecting ion potential are complex and multi-factors cooperate; ion potential is dynamic and changeable, i.e. ion potential corresponds to the state of ion one by one; ion potential has clear physical meaning, but it has not been given dimension at present.

G. H. carteledge's definition of ionic potential: $\varphi=Z/r$ is based on the ionic bond model, so it can only meet some facts. Some people also use ionic function such as $\varphi=Z/r^2$ or $\varphi=Z^2/r$[3] to express ionic potential to match some other facts. It should be said that the definition of $\varphi=Z^*/r^*$ is more reasonable. At present, ionic potential is still a developing concept.

III. LAWS OF IONIC POTENTIAL

3.1 The same period of periodic table of elements

From left to right: the positive ions distributed outside the core are reduced with the increase of Z^*, and r^* decreases. Such as:

$$\varphi(Al^{3+})>\varphi(Mg^{2+})>\varphi(Na^+),\varphi(Tl^{3+})>\varphi(Hg^{2+})>\varphi(Au^+)$$

3.2 The periodic table of elements is the same as the main group (elements s and p).

From top to bottom: positive ion: Z^* is the same, r^* increases, therefore, with the increase of atomic number, φ decreases; negative ion: $Z^* < 0$, Z^* is the same, r^* increases, therefore, with the increase of atomic number, φ increases.

[Explanation] When discussing the properties of the same period or group elements (e.g. metallicity), it is necessary to discuss them according to the actual size of the ion potential. At this time, Z^* and r^* determine the ion potential together.

3.3 Positive ions of different valence states of the same elements

The larger the Z^*, the smaller the r^* and the larger the pH. Such as:

$$\varphi(Cr^{6+})>\varphi(Cr^{4+})>\varphi(Cr^{3+})>\varphi(Cr^{2+})$$

3.4 Rare earth element

Because of the shrinkage of lanthanide, the ion potential of rare earth elements is very close. Therefore, the properties of rare-earth elements are very close, coexistence and symbiosis, and separation is very difficult. Our country adopts Xu Guangxian's innovative grading extraction technology to separate the technology, leading the world.

3.5 The elements with similar ionic potentials are similar in nature.

Diagonal Similarity Principle: The adjacent elements (Li and Mg; Be and Al; B and Si, etc.) in the diagonal periodic table are similar in nature, but far from other elements of the family. This is mainly due to the fact that the values of the ionic potentials of the two elements are similar, while the values of the ionic potentials of the other elements of the family are quite different.

3.6 The relationship between ionic potential and electronegativity

According to the definition of electronegativity by Allred-Rochow [2]:

$$X_{AR}=3590Z^*/r^2+0.744=3590\varphi/r+0.744$$

It can be seen that the larger the ion potential, the larger the electronegativity X_{AR}, the positive correlation between the two, with essential consistency.

3.7 Relationship between ionic potential and ionic deformability

3.7.1 The meaning of ionic deformability

Under the electric field of other ions around the specified ion, the polarization deformation of the electron cloud around the ion occurs. Usually, the polarizability of the ion is described quantitatively by the ionic polarizability μ[3]. The larger the μ, the stronger the ion deformability.

3.7.2 Relationship between ionic potential and ionic polarizability

Ⅰ. Positive ions: the larger the ionic potential, the smaller the polarizability is, for example:

$$\varphi(Fe^{3+})>\varphi(Fe^{2+}),\mu(Fe^{3+})<\mu(Fe^{2+})$$

$$\varphi(B^{3+})>\varphi(Al^{3+})>\varphi(Sc^{3+})>\varphi(Y^{3+})>\varphi(La^{3+})$$

$$\mu(B^{3+})<\mu(Al^{3+})<\mu(Sc^{3+})<\mu(Y^{3+})<\mu(La^{3+})$$

Ⅱ. Negative ions: Because of its $Z^*<0$, therefore, the greater the ion potential, the greater the polarizability, such as:

$$\varphi(O^{2-})<\varphi(S^{2-})<\varphi(Se^{2-})<\varphi(Te^{2-})$$

$$\mu(O^{2-})<\mu(S^{2-})<\mu(Se^{2-})<\mu(Te^{2-})$$

Ⅲ. The difference between Z^* and r^* is different. The larger the r^*, the greater the polarizability.

$$r^*(I^-)>r^*(Br^-)>r^*(S^{2-})>r^*(Cl^-)>r^*(O^{2-})>r^*(F^-)[3]$$

$$\mu(I^-)>\mu(Br^-)>\mu(S^{2-})>\mu(Cl^-)>\mu(O^{2-})>\mu(F^-)[3]$$

Ⅳ. Positive ions with d^x electrons: Because of the large deformability of d orbital, the deformability of positive ions with d^x electrons is generally strong. For example, Ag^+ has larger φ and μ.

[Conclusion]] Positive ions: The larger the ion potential is, the stronger the polarization ability is, and the weaker the deformability is, that is, the smaller the polarizability is. Negative ions: The bigger the ion potential is, the bigger the polarizability is, the bigger the deformability of the ion itself is. The ionic potential generally corresponds to the polarization ability of ions, and the rate of polarizability generally corresponds to the deformability of ions.

[Note] I^-, Br^-, $S^{2-}$are strongly deformable negative ions, while O^{2-}, $F^-$are weakly deformable negative ions; Ag^+, Zn^{2+}, Cd^{2+}, Hg^{2+} are strongly deformable positive ions, while Li^+, Na^+, Be^{2+}, B^{3+} are weakly deformable positive ions.

3.7.3 The matching principle between the ionic potential and the polarizability

1. The larger the ion potential is, the larger the polarizability is. The polarization deformation of the negative ion is mainly caused by the positive ion. For example, $\varphi(Fe^{3+})>\varphi(Fe^{2+})$, $\mu(I^-)>\mu(S^{2-})$, both FeI_2 and FeS can exist stably; Fe_2S_3 exists, but its stability is poor; FeI_3 does not exist. How to explain? The answer is in this article!
2. The interactions between the two ions with large ionic potentials (or small polarizabilities) are relatively small.
3. There is a strong interaction between positive ions with larger ionic potential and polarizability (the valence layer contains d^x electrons) and negative ions with larger polarizability.

[Explanation] Experience shows that the ionic potential is determined by Z^* and r^* together, while the polarizability is mainly determined by r^* [3]; the ionic potential is generally used to describe positive ions, and the polarizability is generally used to describe negative ions. Therefore, it can be said that the surface of ionic potential and polarizability are opposite and essentially unified, "i.e., enemies and narrow paths"; the two models are different, the angles of dealing with the problem are different, the paths are the same, and can not be neglected.

IV. SIMPLE APPLICATION OF IONIC POTENTIAL

4.1 Influence of ionic potential on the type of chemical bond

The enhanced polarization between positive and negative ions bound by chemical bonds can lead to changes in the types of chemical bonds.

Generally speaking, when positive ions with large polarization ability (big φ) and strong deformability (big μ) are combined with negative ions with strong deformability (big μ), the distribution of charges (electronic clouds) of each other is greatly changed by strengthening the interaction of polarization, thus shortening the distance between ions, leading to orbital overlap (overlap, reorganization), and chemical bonds. The ionic component decreases while the covalent component of the chemical bond increases. The transition from the ionic bond to the covalent bond leads to the transformation of the chemical bond type and even the corresponding crystal type, which causes the change of some properties of the substance.

Particularly noteworthy is that positive ions (such as Ag^+, Zn^{2+}, Cd^{2+}, Hg^{2+}) with strong polarization ability and deformability of the electronic layer structure containing d^{10} and negative ions (such as I^-, $S^{2-}$and so on) with strong deformability will deform the electron cloud around the ion due to the polarization interaction (additional polarization). Both the type of chemical bond and the corresponding crystal type are transformed. For AgCl, if the distance $d(d=r_+ + r_-)$ is calculated according to the ionic bond model, it should be 294 *pm*, whereas the actual measurement is 287 *pm*. Obviously, due to the interaction of polarization, the distance between nuclei is shortened by 17 *pm*, the covalent bond composition in AgCl increases significantly, and AgCl crystals become "mixed or transitional crystals" rather than pure ionic crystals. Based on the stronger interaction of polarization, the distance between nuclei of AgI is even shortened by 52 *pm*, so that the chemical bonds in AgI are mainly covalent bonds, and AgI crystals become ZnS-type molecular crystals rather than NaCl-type ionic crystals.

Another example is that Li^+ has a small radius (r_{Li}+=68pm, r_{Na}+=98pm) and a large ionic potential. Therefore, Li^+ has a strong polarization ability in alkali metals, and has the greatest tendency to form covalent bonds. It can form covalent compounds, while other alkali metals basically form ionic compounds. Based on this, compared with other alkali metal ions of the same family, Li^+ exhibits many different special properties, for example, LiOH is moderately strong alkali. This is mainly due to the obvious covalence of Li-O bond due to the polarization of Li^+ and the change of the dissociation mode of LiOH (partially acid dissociation). Other alkali metal hydroxides (MOH) are strongly alkaline, mainly ionic compounds.

He radius of Be^{2+} ion in the main group of IIA is smaller than that of $Li^+$and the charge is twice that of Li^+so that the charge density of Be^{2+} is higher, the ionic potential is larger and the polarization ability is stronger. It can even cause polarization deformation of $F^-$and $O^{2-}$with very low polarizability, which results in the covalent bonding of Be with almost all other elements. There is no simple ion of Be^{2+} in the solution, only in the form of hydrated ion $Be(H_2O)_4^{2+}$ in strong acidic solution (strong polarization of O in water molecule), and even the Be-F bond in BeF_2 formed by the most electronegative fluorine element has obvious covalence. BeF_2 molecule is a few metals that can not be completely ionized in solution. One of fluoride.

The valence electron configuration of boron (B) atoms in the main group of III A in the periodic table of elements is $2s^2 2p^1$, the radius of boron atom is small (85~90pm), and the ionization energy is high (I_1=8.296ev, I_2 and I_3 are much larger than I_1). This shows that the effective nuclear charge Z^* of B^{+3} is large, the attraction of valence layer electrons is great, the ionic potential is large, and boron is almost all bound by covalent bonds with other elements. On the other hand, the polarization ability of B^{3+} ions is very strong. In the solution, there is no simple ion of B^{3+}, but it acts with H_2O molecule to form $B(OH)_3$.

4.2 Influence of ionic potential on solubility of compounds

The solubility of compounds is related to many factors, and the dissolution process is a very complex physical and chemical process. Various theories, hypotheses and models [4] 550 are still immature, and most of them are only approximate and empirical laws. For example, the principle of similarity and compatibility, which is an empirical rule summarized from a large number of experimental facts, can explain some facts. According to this principle, ionic compounds and strong polar covalent compounds are easy to dissolve in polar solvents (such as H_2O (l), NH_3 (l), etc.), while non-polar or weak polar covalent compounds are easy to dissolve in non-polar solvents (such as benzene, carbon tetrachloride, etc.). Strictly speaking, compounds with similar dipole moment are easy to dissolve each other.

Because of the polarization between ions, the chemical bond type in some compounds transits from ionic bond to covalent bond, resulting in the decrease of dipole moment and polarity of the compounds, and the corresponding decrease of solubility of the compounds in strong polar solvent water. For example, the radius of Ag^+ is between Na^+ and K^+(r_{Ag+}=126pm, r_{Na+}=95pm, r_{K+}=133pm), but the solubility of AgCl(s), NaCl(s), KCl(s) is quite different. The change of Gibbs free energy ($\Delta_{sol}G$), which indicates the spontaneous dissolution trend, is respectively: +55.6$kJ \cdot mol^{-1}$, -92$kJ \cdot mol^{-1}$, -6.7 $kJ \cdot mol^{-1}$, so it can be inferred that their solubilities are quite different. The same is true: AgCl(s) is insoluble in water (p^{Ksp} $_{(AgCl)}$=9.75), NaCl(s) and KCl(s) are soluble in water. The solubility of AgCl(s) is obviously related to the valence electron structure of Ag^+. On the one hand, because the valence electron configuration of $Ag^+$is $4d^{10}5s^0$(9 e^--17e^- configuration, while Na^+ and K^+ are 8e^-configuration), the effective nuclear charge Z^* of $Ag^+$is larger and the ionic potential is larger; on the other hand, the polarizability of Ag^+ is mu because the D electron cloud is easily polarized and deformed. It is also larger. In this way, the Ag-Cl bond is strengthened due to the interaction of polarization, the ionic component of the bond decreases while the covalent component increases, resulting in the decrease of dipole moment and the decrease of solubility in polar solvent water. It can be expected that the solubility of NaCl (s) and KCl (s) is also quite different (why? Please seek an explanation in this article).

Compared with silver halide AgX:AgF(s), the F^-ion potential of fluoride ion in AgX:AgF(s) is smaller, the polarizability is smaller and the deformability is weaker, so the interaction between $Ag^+$and $F^-$is very small. The ionic composition of Ag-F chemical bond is larger. AgF(s) is still an ionic crystal, and the dipole moment of AgF is larger, so it is soluble in polar solvent water. For AgCl, AgBr and AgI, due to the strong mutual polarization between Ag^+ and X^-ions, the covalent component of Ag-X chemical bond is increased, the dipole moment of AgX is reduced and the polarity is weakened, so that they are insoluble in water (p^{Ksp} (AgBr)=12.3, p^{Ksp} (AgI)=16.08), and with the mutual polarization between $Ag^+$and X^-gradually. The solubility of $AgCl_{(s)}$ to $AgBr_{(s)}$ to $AgI_{(s)}$ decreases gradually.

The solubility of $Li_3PO_{4(s)}$ and $Li_2CO_{3\,(s)}$ in water is much lower than that of salts corresponding to other alkali metals in the periodic table of elements.

4.3 Influence of ionic potential on the color of compounds

Basic Principle of Matter's Color Presentation: The interaction between light and matter causes the transition of electrons in matter between different energy levels, leading to the selective absorption and emission of light by matter. The physical processes of reflection, transmission, deflection and scattering can occur when light of different wavelengths is transmitted in matter. Human's perception of material color is based on the comprehensive effect of the above process on human's visual perception. All of these are related to material structure.

In addition to changing the type of chemical bond, positive ions with high ionic potential can also change colorless ions into colored ones. These changes are related to the internal structure of molecules or ions that constitute substances. In the formation of compounds, the polarization of the electron cloud of the ligand (atom, molecule or ion which distributes around the positive ion and binds directly with it) can be deformed, so that the energy levels of the electrons in the system can be recombined, and the energy level difference $\Delta E=h\nu$ (ν where_is the frequency of light, the same below) can be changed. The change results in the change of the frequency_and wavelength lambda of the absorbed light and the change of the frequency_and wavelength lambda of the emitted light when the electron undergoes the energy level transition, so that the original colorless ions become colored after the formation of compounds.

Typical ionic compounds generally do not absorb light in the visible region of the spectrum, so they appear colorless or white under visible light (white light), for example, KCl is a colorless solid. The basic principle is that in typical ionic compounds, if it takes a great deal of energy to make electrons transit from negative ions to positive ions, the ultraviolet light of high frequency and short wave should be absorbed from$\Delta E=h\nu$ and $\lambda=c/\nu$ (c is the speed of light). When a typical ionic compound absorbs light from the ultraviolet region, it generates charge transfer (charge transfer transition, see below) absorption spectra. However, when the light of the ultraviolet region is absorbed, the band is colorless, that is, the substance is colorless.

By strengthening the polarization of the ligand, the central ions with high ionic potential promote the electron cloud to migrate gradually from the ligand to the central ion, which decreases the density of the electron cloud around the ligand and increases the energy level of the atomic orbital or molecular orbital, which leads to the increase of the energy of the electrons in the original ground state of the ligand. The energy level difference between the highest occupied orbital level of the ligand and the lowest empty orbital of the central ion decreases, and the transition (transfer) of electrons between the ligand and the central ion becomes easier. From $\Delta E=h\nu$ and $\lambda=c/\nu$, we know that the absorbed light will move in the direction of low frequency and long wave. Therefore, only absorbing a certain wavelength of light in the visible region can cause the excited state electrons in the highest occupied orbit of the ligand to transit to the low energy space orbit adjacent to the central ion, thus showing the color in the visible region, which is called color rendering.

The transition of electrons between central ions and ligands is usually called charge transfer transition, which corresponds to the electron transfer occurring inside the system and is a form of electron transition.

The necessary condition for charge transfer transition is that the energy of the lowest empty orbit of the ligand is higher than that of the lowest empty orbit of the central ion, so as to ensure that the excited state electrons in the highest occupied orbit of the ligand can transit to the lowest empty orbit of the central ion more easily without transiting to the ligand itself. The lowest orbit on the body. This self consistent condition can be satisfied by polarization.

The wavelength range of the light absorbed by the charge-transfer transition depends on the energy difference between the lowest space orbit of the central ion and the highest occupied orbit energy level of the ligand. Thus, it is not difficult to understand: (1) For compounds with the same central ion but different ligands, the more deformable the ligand (the greater the polarizability mu) is, the less energy is needed for the charge transfer transition, and the more color the material presents moves to the direction of long wave, i.e. the color deepens. For example, the color of AgCl to AgBr to AgI gradually deepens. (2) For compounds with the same ligand but different central ions, if the central ions belong to the isomorphic ions, the energy required for charge transfer transition generally decreases with the increase of the ion potential of the central ions, and the color of the substance gradually moves towards the long wave direction, that is, the color gradually deepens. For example, $ZnI_{2(s)})$ is colorless, $CdI_{2(s)}$ is yellow, $HgI_{2(s)}$ is red, and color gradually deepens.

Some covalent halides or sulphides with d^{10} electronic structure of central ions show darker colors. For example, AgBr is yellowish, AgI, CdS, PbI_2 are yellowish, HgS is red and so on. The color deepens gradually, indicating that the charge transfer transition becomes easier. The main reasons for the above changes are as follows: ionic potential $\varphi(Hg^{2+}) > \varphi(Pb^{2+}) > \varphi(Cd^{2+}) > \varphi(Ag^{+})$, and the polarizability of I^- and S^{2-} are larger, and the interaction between positive and negative ions is stronger. In addition, the central metal ions in these compounds have d^{10} electronic structure, and the d electron layer is full, that is, the low-level orbits are full. In order to have their own electronic transitions (i.e., *d*-*d* transitions), it is necessary to jump to the higher-level orbits. The energy required is much larger than that of visible light. If the electron of the ligand moves to its own high-level orbit, it also needs higher energy, but the charge transfer transition between the ligand and the central ion requires less energy, which can be achieved by absorbing some wavelength of light in the visible region. Apparently, the Colouring Mechanism of these compounds is the result of charge transfer transition, i.e. electron transfer in the system caused by strong polarization between ligand and central ion, which is the evidence of charge transfer transition.

It should be pointed out that the ionic potential can only affect the color of the compound, but it can not determine the color of the compound (the intensity of the color). The color of the compound depends on the probability of electron transition. The greater the probability of transition (allowing transition), the darker the color; the smaller the probability of transition (forbidden transition), the lighter the color. Ion potential only affects the wavelength and frequency of the light absorbed during the transition. If the light absorbed in the ultraviolet or infrared region is colorless. Therefore, the larger the ionic potential is, the more complex the color of the compound is.

4.4 Influence of ionic potential on acid-base ionization

Oxyacids and bases are essentially isomorphic substances, that is, they are compounds containing hydroxyl groups (OH). It can be simply expressed by **R-O-H**. According to structural analysis, R-O-H may have two ways of breaking chemical bonds: 1. breaking R-O bond; 2. breaking O-H bond. If the R-O bond is broken mainly by alkali ionization, if the O-H bond is broken mainly by acid ionization, if the two bond breaking abilities are equal, it is "amphoteric" ionization. The way of ionization depends on the polarization of the central ion (atom) and the polarity induction effect of the water molecule. In the final analysis, it is equivalent to the "contest" for the central O atom of R-O-H. A model of action is proposed here: $H_2O \ldots R\text{-}O\text{-}H \ldots OH_2$, H-O-H molecules at both ends (R-end and H-end) induce electron clouds in the system in opposite directions and with different intensities. If the value of the ionic potential of the central ion is larger, the polarization effect is stronger, and the power of R^{+n} attracting oxygen atom electron cloud is greater. Through the interaction of R^{+n}

polarization and the polarity induction effect of the left-end hydrogen peroxide, the density of oxygen atom O electron cloud is reduced, the R-O bond is shortened, the O-H bond is lengthened, and the bond force of O-H bond is weakened. The O-H bond is covalent bond. The trend of conversion to ion bond increases, the polarity of O-H increases, and the dipole moment of the bond increases, then acid ionization occurs mainly by breaking the O-H bond, resulting in H^+. On the contrary, if the value of the ionic potential of the central ion is smaller, the polarity induction of the right-end hydrogen oxide molecule will be greater than that of the central ion (atom) R^{+n} and the left-end hydrogen oxide molecule pole. The sum of sexual induction makes the R-O bond lengthen and weakens the bonding force of the R-O bond. The weaker the covalent component of the R-O bond, the larger the ionic component of the bond, the greater the polarity of the R-O bond and the increase of the bond dipole moment, the basic ionization of the R-O bond mainly breaks and produces OH^-. The acidity and alkalinity of ROH are mainly determined by the synergistic effect of the ion potential of R^{+n} and the polarity induction effect of water molecule. A semi-quantitative relationship between acid and alkali of ROH-type compounds related to the ionic potential of R^{+n} has been proposed.

1. $\varphi<0.048$, ROH The main manifestation is basic ionization.
2. $0.048< \varphi<0.102$, ROH Two types of ionization coexist, that is, acid type and basic ionization.
3. $\varphi>0.102$, ROH Acid ionization is the main feature.

When the electronic layer structure of the central atom R^{+n} of ROH is the same, the greater the ionic potential of the central ion (atom), the stronger the acidity, the smaller the ionic potential, and the stronger the alkalinity. For example, in NaOH, $\varphi(Na^+) = 0.001 < 0.048$, NaOH mainly occurs alkaline ionization, showing strong alkalinity; in $Al(OH)_3$, $0.048 < \varphi(Al^{3+}) = 0.059 < 0.102$, $Al(OH)_3$ occurs two kinds of properties of ionization coexist, and alkaline ionization is stronger than acidic ionization, showing very weak alkalinity; in H_2SO_4, $\varphi(S^{6+}) = 0.200 > 0.102$, H_2SO_4 mainly occurs acidic ionization, showing strong alkalinity. Strong acid.

[question] by contrast, why is "gender" ionization very weak? Please seek an explanation from this article.

4.5 Effect of Ion Potential on Other Properties of Compounds

4.5.1 Effect of Ion Potential on the Reaction Performance of Salts with Water

When salts dissolve in water, the positive and negative ions released by salts hydrate with water molecules and coordinate with a certain number of water molecules around the positive and negative ions, which are then strengthened by ion polarization. If the ion potential of metal positive ions is large and the polarization ability is strong, the bond force of O-H chemical bond in water molecules can be greatly weakened by polarization. At the same time, under the constant collision of water molecules or other particles in water, some additional energy can be obtained. At this time, the O-H bond in more water molecules may break, resulting in $OH^-$and H^+. Then the positive ions of salt can capture $OH^-$and then OH^-can be obtained. If the H^+ released by water itself can polarize the negative ions of salt and penetrate into the electron cloud of the negative ions, it will combine with the negative ions directly, which is equivalent to the catching of H^+ by the negative ions and releasing OH^-, the above two processes can also be carried out simultaneously. As a result, the ionization equilibrium of water is destroyed, resulting in the continuous ionization of $OH^-$and H^+. When the equilibrium is re-established, the relative concentration of $OH^-$and H^+($c(H^+)/c(OH^-)$) in the solution changes, making the solution acidic or alkaline, which is the essence of saline hydrolysis.

[Conclusion] The hydrolysis reaction of anion is complex, and its hydrolysis ability is inversely proportional to the acid strength of its conjugated acid (Ktheta). The weaker the conjugated acid is, the stronger the hydrolysis reaction of its acid radical anion (i.e. its conjugated base), such as CO_3^{2-}, SiO_3^{2-}, etc. The higher the ion potential of positive ions, the stronger the hydrolysis reaction ability.

4.5.2 Effect of Ion Potential on Crystal Lattice Type

Because of the strong polarization, the type of chemical bond changes from ion bond to covalent bond. This transition not only shortens the distance between ions, but also tends to reduce the coordination number of crystals, leading to the change of crystal type. For example, according to the ionic bond model, the radius ratio of positive and negative ions of CdS is calculated to be $r(+)/r(-)=0.53$. Therefore, CdS should belong to NaCl-type crystal with coordination number 6, whereas CdS belongs to ZnS-type crystal with coordination number 4. The reason is that Cd^{2+} has larger ionic potential and stronger polarization ability, while S^{2-}has strong deformability. Cd^{2+} can partially penetrate into the electronic cloud of S^{2-}as if it reduced the ionic radius of Cd^{2+} to 0.44 instead of 0.53, thus changing the lattice type.

4.5.3 Effect of Ion Potential on Conductivity and Metallicity of Compounds

Because of the polarization, the electrons in negative ions tend to be separated from the negative ions, and the fluidity of these electrons is enhanced, which is equivalent to the free electrons in metals. Thus, the lattice of the crystal can be transformed from the lattice of the ionic crystal to the lattice of the metallic crystal, and the conductivity increases, and the metallicity increases accordingly. The opacity and metallic luster of sulfides are related to this, such as FeS, CoS, NiS and other chemical powders, especially their ores, all have metallic luster.

4.5.4 Effect of Ion Potential on Thermal Stability of Compounds

The polarization of ions with high ionic potential is stronger. Because of the polarization of positive ions in the compound, the thermal stability of the compound is weakened and the decomposition reaction is easy to take place. For example: $\varphi(Li^+)>\varphi(Cu^{2+})>\varphi(Ag^+)>\varphi(Na^+)$, LiOH red heat is decomposition, $Cu(OH)_2$ heating decomposition, NaOH does not decompose; $AgNO_3$ heating decomposition, $NaNO_3$ strong thermal decomposition.

4.5.5 Effect of Ion Potential on Thermal Stability of Metal Hydride

In alkali metal and alkaline earth metal binary hydrides, because H^-has a larger radius (208*pm*), it is easy to deform and polarize, so when it is combined with positive ions with larger ion potential, the type of chemical bond will transit from ion bond to covalent bond, and the strength of chemical bond will also increase, thus enhancing its stability. For example: LiH heating does not decompose, while NaH heating is easy to decompose into metal sodium and hydrogen, because: $\varphi(Li^+)>\varphi(Na^+)$, Li^+ polarizes more strongly than Na^+ and LiH is more stable than NaH.

4.5.6 Influence of ionic potential on other properties of compounds

Many elements, especially transition metal elements, often show different oxidation states in different compounds, and the existence of various stable oxidation states is closely related to ion potential. Ions with high ionic potential generally can not form simple ions, they all exist in the form of complex ions; ions with low ionic potential can generally form simple ions. For example, Mn^{2+} ion can exist stably in aqueous solution, while Mn^{+7} ion can only exist in the form of MnO_4^- in aqueous solution; S^{+6} ion can not exist in the form of S^{6+} simple ion, but only in the form of SO_4^{2-} hydrated ion.

[Summary] Thinking model: ionic potential, deformability→polarization, mutual polarization→ polarity change of bond→transition from ionic bond to covalent bond (bond type change, crystal type change)→variety of properties change.

It should be pointed out that at present the ion potential can only explain qualitatively or semi-quantitatively some properties of elements or compounds, and the properties of elements or compounds are often the synergistic representation of many factors. For example, the strength of oxyacid depends not only on the ionic potential of the central ion, but also on the number of non-hydroxyl oxygen atoms around the central ion and other factors [1]94; the solubility of ionic compounds is also related to their lattice energy, and so on. Therefore, when using ion potential to explain the problem, we should pay attention to its scope of application and role, and consider the problem comprehensively.

In this paper, the principle and application of ionic potential are simply summarized and sorted out, focusing on throwing bricks and attracting jade. The shortcomings are unavoidable, such as inadequate structural analysis, no thermodynamic analysis and so on. Thank you for your comments and comments.

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