

CHM101 SUMMARY WITH PAST QUESTIONS

Development of the periodic law and the periodic table of the elements is one such attempt. This has brought order in the study of the vast chemistry of more than a hundred elements known now. In this unit, you will be starting from the very beginning, that is, with the very first attempt made at classification of the elements. By the mid-19th century, more than 60 elements were known and many more were being discovered.

During this period, it was also realised that certain groups of elements exhibited similar physical and chemical properties. List accurately at least two scientists who attempted to classify the elements into periods write with at least 70% accuracy, brief accounts of the attempts made by the two scientists and the result of these attempts state Mendeleev's periodic law state the property used by Mendeleev to classify the elements in his periodic table demonstrate an understanding of Mendeleev's law by applying it to predict properties of undiscovered elements.

Attempts Made By John Newlands

« He suggested that if the elements were arranged in order of increasing atomic weight, every eighth element would have properties similar to the first element. For example, he arranged the elements in the following manner. He called it the »Law of octaves« because he says the cycle of repetition shown by the elements is like that shown by octaves of music where every eight-note resembles the first in octaves of music. Firstly, it did not hold good for elements heavier than Ca.

The Work of Lothar Meyer

In 1869, Lothar Meyer reported that when physical properties like atomic volume, boiling point etc. were plotted against atomic weight, a periodically repeating curve was obtained in each case. Figure 1.1 is a graph showing the variation in atomic volume with atomic number. The atomic volume behaviour is periodic. Each of the cycles is called a period.

The location of an element on the peak or in the troughs has an important correlation with their chemical reactivity. The elements of the peaks are the most reactive. 1.1 Periodic Dependence of Atomic Volume on Atomic Number.

Mendeleev's Periodic Law

In contrast to Lothar Meyer, Mendeleev used chemical properties like valence and formulae of hydrides, chloride, and oxides of the elements to illustrate his periodic law. According to Mendeleev's periodic law, if the elements are arranged sequentially in the order of increasing atomic weight, a periodic repetition, that is, periodicity in properties is observed. Mendeleev arranged elements in horizontal rows and vertical columns in order of increasing atomic weight so that the elements having similar properties were kept in the same vertical column. He included along with his table, a detailed analysis of the properties of all known elements and correlated a broad range of physical and chemical properties with atomic weights.

He kept his primary goal of arranging similar elements in the same group quite clear. Therefore he was bold enough in reversing the order of certain elements. For example, iodine with lower atomic weight than that of tellurium was placed in group VII along with fluorine, chlorine and bromine because of similarities in properties. He also corrected the atomic weight of certain elements to include them in proper groups.

For example, he corrected the atomic weight of beryllium and indium without doing any actual measurement. Keeping to his primary goal of arranging similar elements in the same vertical column, he realised that some of the elements were still undiscovered and therefore left their places vacant in the table and predicted their properties. He predicted the existence in nature of over ten new elements and predicted properties of three of them, example eka-boron, Eka aluminium and Eka silicon from the properties of known elements surrounding them. When these elements were eventually discovered, Mendeleev prediction proved to be amazingly accurate.

The validity of Mendeleev periodic law was dramatically and conclusively proven by the discovery of three out of the more than ten elements predicted by Mendeleev. Both elements were found to have properties corresponding to those of earlier predicted for them by Mendeleev. On this basis, he predicted the properties of undiscovered elements representing the gaps in his table.

The periodic table of today has many similarities with that formed by

In some cases, such a practice has led to disputes between scientists who have discovered the same elements working independently in different parts of the world.

I on the periodic table state the relationship between the atomic number and the periodic classification of elements apply IUPAC nomenclature rules in naming new elements having

You must have noticed that there were anomalies in Mendeleev's original periodic table. He could not predict the existence of noble gases, nor could he properly place hydrogen. Mendeleev's periodic law was therefore accordingly modified. Arrangement of the elements in order of their increasing atomic number removes most of the anomalies of Mendeleev's periodic table.

The positions of K and Ar, Co and Ni, Te and I do not remain anomalous any longer since atomic number, not atomic weight is used in arranging the elements. As isotopes of an element have the same atomic number, they can all be placed at one and the same place in the periodic table. We know that the atomic number cannot be fractional. It increases by the integer from one element to the next.

It has thus placed a limit on the number of elements. Today, 109 elements have been discovered and any more elements that may be discovered in future will be beyond 109. You have now seen that in the modern form of Mendeleev periodic table, elements are arranged in seven horizontal rows and eight vertical columns.

Long Form of the Periodic Table

To avoid this controversy, International Union of Pure and Applied Chemistry has adopted Arabic numerals 1, 2, 3, 18 as the new group designation in the form of the periodic table. Sc to Zn families become groups 3, 4, 5, 12 and finally, the P block elements become groups 13, 14, 15, 16, 17, 18 of the table.

Nomenclature of Elements Having $Z > 100$

It has been a historical practice to allow the discoverer of the elements to assign the element's name. In recent times, this has led to some controversy because elements with very high atomic number are so unstable that only minute quantities of them, sometimes only one or two atoms are prepared before scientists claim credit for their discovery. The names are derived directly from the atomic number of the element using the following numerical root.

The root be put together in the order of the digit which make up the atomic number and be terminated by »mum and ending occurring in the names of the metallic elements as these are the final «n» of enn be dropped when it occurs before 'nil' and '1' of

The symbol of the element is composed of the initial letters of the numerical roots which make up the names.

The periodic law first proposed by Dmitri Mendeleev had to be modified the modification came as a result of the recovery of the concept of the 'atomic number' in 1913 by Hendry Moseley the introduction of the concept of atomic number further clarified the arrangement of the elements in the periodic table and removal any ambiguities which were observed the atomic number rather than the atomic weight is the most important determinant of the properties of an element as a result of a consensus reached by IUPAC the recommendation of any newly discovered element must follow IUPAC nomenclature rules.

In the next unit, you will be studying the distributions of the electrons within the atom, and how they govern the properties of the elements. The electronic configurations of isolated atoms of elements are usually verified experimentally by a detailed analysis of atomic spectra. State the principle involved in determining which electron goes into which atomic orbital fill out correctly electrons in a given atom once given the number list the four blocks of elements on the periodic table and determine to which block an element belongs if given the electronic configuration.

The electronic configuration of atoms can be predicted with the help of

In the Aufbau process, it is assumed that there exist a set of empty hydrogen-like orbital around the nucleus of an atom. The electronic configuration of the atom in the ground state is then derived by adding electrons one at a time to the orbitals of the lowest energy in the sequence shown by arrows in Table 3.1. This rule reminds us that the energy of subshells of multi-electron atoms depends upon the value of both the quantum numbers n and l , but mainly on the value of n . Filling of electrons in orbitals is also governed by Pauli's Exclusion Principle and Hund's rule.

According to the Pauli Exclusion Principle, no two electrons in the same atom can have the same value of n , l and m_l , they will differ in their m_s values. In other words, an orbital can have at most, two electrons of opposite spin. Since there is only one s orbital for any given value of n , it can contain only two electrons. However, the three p orbitals for any given value of n can contain six electrons, the five d orbitals, for any given value of n can hold a total of ten electrons and the seven f orbitals can have fourteen electrons.

Permitted combinations of all the four quantum numbers for the electrons in different orbitals are given below in Table 3.2. Hund's rule of maximum multiplicity states that as far as possible in a given atom in the ground state, electrons in the same subshell will occupy different orbitals and will have parallel spins. So far you have studied the rules governing the filling of electrons in the orbitals of atoms. The electronic configuration of hydrogen and helium are $1s^1$ and $1s^2$ respectively.

The $1s^2$ configuration of helium is usually represented by. This period contains elements from lithium to neon. Next six elements from boron to neon, the $2p$ orbitals are filled. Neon thus has the electronic configuration of $2s^2 2p^6$ which as was done in the case of He, is represented by.

An electronic configuration of means $1s^2 2s^2 2p^6$. Although the third principal shell can accommodate 10 more electrons in $3d$ orbitals filling of $4s$ orbital takes place first because of its lower energy. This period contains 18 elements from potassium to krypton. In K and Ca, the first two elements of this period, the successive electrons go into the $4s$ orbitals giving them the configuration $4s^1$ and $4s^2$ respectively.

Then in the following 10 elements filling of hitherto unoccupied $3d$ orbitals takes place. Thus the electronic configuration of zinc becomes $3d^{10} 4s^2$ and $3d^9 4s^2$ respectively.

The next 18 elements from rubidium to Xenon belong to this period. In building up of the atoms of these elements, $5s$ and $4d$

After that in elements from Y to Cd filling of $4d$ orbitals takes place. You can see from table 3.1 that once again there are minor irregularities in the distribution of electron between $4d$ and $5s$ orbitals.

$4d^5 5s^1$ and $4d^{10} 5s^1$ configurations similar to those of Cr and

Now in the next six elements, that is I, Sn, Sb, Te, I and Xe filling of sp orbitals take place and thus Xe $Z = 54$ attains $4d^{10} 5s^2 5p^6$ configuration. This period contains 32 elements from caesium to radon in which the $6s$, $4f$, $5d$ and $6p$ orbitals are filled. The first two elements of this period have configurations analogous to those of corresponding member of the lower periods, thus caesium and barium have $6s^1$ and $6s^2$ configuration respectively.

In the next 14 elements from cerium to lutecium, the $4f$ orbital is successively filled pertaining to $4f^1 5d^1 6s^2$ and

4f¹⁵5d¹ 6s² configuration, respectively, but you should remember, it is only Ce, Gd and Lu that 5d orbitals have one electron while in all the remaining Lanthanides the 5d orbitals remain vacant. After Lutetium, successive electrons occupy 5d orbitals and the electronic configuration builds up from 4f¹⁴ 5d²6s² to for hafnium to 4f¹⁴ 5d¹⁰ 6s² for mercury the homologue of zinc and cadmium. Again a minor departure from a steady increase in the number of d electrons occurs. This period is still incomplete and contains 23 elements from francium to ununennium.

In these elements, electrons are filled in 7s, 5f and 6d orbitals. In the rest of the elements, the 6d orbitals remain vacant, thus the electronic configuration of Lr is 5f¹⁴ 6d² 7s². Having examined the electronic configuration of elements in the periodic table, you can see from Table 3.4 that the elements occupying the same group of the periodic table have the same valence-shell electronic configuration. In other words, the elements having the same valence shell electronic configuration recur periodically, that is after intervals of 2, 8, 8, 18, 18 and 32 in their atomic number.

Therefore periodicity in the properties of elements can easily be understood. So far, you have studied the electronic configuration of neutral atoms of elements. I am sure you will be interested in knowing the electronic configuration of ions that are obtained by removal of electrons from the elements. When the gaseous iron atom having 3d⁶ 4s² ground state electronic configuration loses an electron, the Fe⁺ ion is formed.

This ion has its minimum energy in the configuration 3d⁷, although the isoelectronic manganese atom has the configuration 3d⁵ 4s² in the ground state.

3d⁶ and 3d⁵ respectively rather than 3d⁵ 4s¹ and

Evidently, the differences in nuclear charge between Fe⁺ and Mn, Fe²⁺ and Cr and Fe³⁺ and V are important in determining the orbital to be occupied by the electrons. Since the chemistry of elements is essentially that of their free ions, the regularities in configuration of ions are much more important than the irregularities in the electronic configuration of the neutral atoms.

Electronic Configuration and Division of Elements into

Elements of the periodic table have been divided into four blocks s, p, d and f depending upon the nature of the atomic orbitals into which the differentiating or the last electron enters. As you know the valence shell electronic configuration of these groups are ns¹ and ns² respectively. We also know that each period of the periodic table begins with alkali metals. All the elements in this block are metals.

These elements are placed in the middle of the periodic table between the s and p-block elements. The electronic configuration of the atoms of the elements of this block can be represented by d¹⁻¹⁰ ns¹⁻². The elements from Sc to Zn, Y to Cd, La and from Hf to Hg are the members of 3d, 4d, and 5d series respectively. The atoms of these elements have the general configuration f⁰⁻¹⁴ d⁰⁻¹ ns².

These elements belong to two series depending upon the filling of 4f and 5f orbitals. Elements from Ce to Lu are the members of the 4f series, while those from Th to Lr belong to the 5f series. Elements of 4f series which follow lanthanum in the periodic table are known as Lanthanides whereas those of 5f series following actinium are called Actinides.

We learned about the properties of elements being periodic function of their atomic numbers. Arrangements that give rise to similarities and differences in the properties of elements whose valence electrons appear in the same group and those whose valence electrons are in different groups respectively.

Measurement of Atomic Radii

Atomic radii are the measure of the size of the atom. Atomic radii are important because other atomic properties like ionisation energy, electron affinity and electronegativity are related to them. The wave mechanical picture of an atom depicts an atom as composed of a compact nucleus surrounded by an electron cloud. This electron cloud does not have a definite boundary surface like that of a ball.

There is a definite but very small probability of finding an electron at an infinite distance from the nucleus of the atom. Now that we have defined the size of an atom, we have to tackle the problem of measuring that size. Consequently, we cannot have one set of atomic radii applicable under all conditions.

Covalent radius can be defined as one half of the distance between the nuclei of two like atoms bonded together by a single covalent bond. If in a homonuclear diatomic molecule of A_2 type r_{A-A} is bond length or inter nucleus distance and r_A is the covalent radius of the atom A, then $r_A = 1/2 r_{A-A}$. In the heteronuclear, diatomic molecule of AB type, if the bonding is purely covalent, then the bond length r_{A-B} is equal to the sum of covalent radii of A and B that is $r_{A-B} = r_A + r_B$. Thus covalent radii are additive. It is possible to calculate the radius of one of the atoms in a heteronuclear diatomic molecule of AB type.

If we know the internuclear distance r_{A-B} and radius of the other atom. As stated earlier, the above relation holds good only if the bond between the atoms A and B is purely covalent. If there is a difference in the electronegativities of the bonded atoms, it causes shortening of the bonds. Usually, a double bond is about 0.86 times and a triple bond about 0.78 times the single bond length for the second-period elements.

Van Der Waal's Radius

In the solid-state, non-metallic elements usually exist as aggregates of molecules. The bonding within a non-metal molecule is largely covalent. Table 4.1 lists the values of Van der Waals radii of some elements. Figure 4.1 illustrates the difference between the covalent and van der Waals radii of chlorine.

Metallic or Crystal Radius

In both these structure, a given metal atom has twelve nearest neighbours. However, a significant number of metals adopt a body-centred cubic lattice in which the number of nearest neighbours is eight. The number of nearest neighbours of a metal atom in a lattice is known as the coordination number of the metal. Experimental studies on a number of metals having more than one crystal lattice have shown that the radius of a metal in an eight coordinate lattice is about 0.97 of the radius of the same metal in a twelve coordinate environment.

Table 4.2 gives a set of twelve coordinate radii for metal atoms. The metallic radii are generally larger than the corresponding covalent radii. Although both involve a sharing of electrons this is because the average bond order of an individual metal-metal bond is considerably less than one and therefore the individual bond is weaker and longer than the covalent. This does not mean that the overall bonding is weak as there are a large number of these bonds, eight or twelve per metal atom.

Ionic radius is defined as the distance between the nucleus of an ion and the point up to which the nucleus has an influence on the electron cloud. This distance is the sum of radii of Na^+ and Cl^- ions. From the electron density maps obtained by x-ray analysis, it has become possible, in some cases, to apportion the internuclear distance into the radius of cation and anion. A small number of ionic crystals has thus been studied and the ionic radii of some of the elements have been determined.

These radii have become the basis for assigning the ionic radii of most of the other elements. Ionic radii are of two types, cation radii and anion radii. This is not too surprising since not only is there a loss of electron from a partially filled outer shell on cation formations, but there is also an increase in the overall positive charge on the ion. Conversely, in anion formation, the addition of an electron to an atom increases the size due to an increase in inter-electronic repulsion in the valence-shell and decrease in effective nuclear charge.

In general, there is a decrease in size of anions to covalent radii of corresponding atoms to cations thus in the series of isoelectronic species.

Factors Affecting the Atomic Radii

So far, we have defined and explained types of atomic radii. Electrons in any group higher in this sequence than the electron under consideration contribute nothing to s .

B

Electronic configuration of K = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ Grouping of orbitals will be $4s + 3d = 16$.

Electronic configuration of Zn = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

You can see from the table that there is a steady increase in Slater's Z^* across rows of the periodic table. Effective nuclear charge felt by electrons also depends on the oxidation state of an

atom in a compound. The higher the oxidation state of the atom, the higher will be the effective nuclear charge felt by the electrons and therefore, smaller will be the atomic radius. Now that we know the various types of atomic radii and the factors that affect them, we will consider the periodicity in them.

Before doing that, however, we would like to emphasize that trends observed in one type of radii are generally found in the other type of radii also. Two general periodic trends are found for all types of atomic radii. These are the atomic radii decreases along a period and generally, increase down a group in the long form of the periodic table. These changes in the atomic radii can be related to the changes in effective nuclear charge and the principal quantum number in the periodic table.

If you examine Table 4.3 you will find out that there is a steady increase in the value of Z^* from alkali metals to halogens for the elements of period 2 and 3, but there is no change in the value of n because the electrons fill the same principal shell. In comparison to the above, the decrease in covalent radii across the transition series is much smaller. As you know, electrons are successfully filled in the d orbitals across a transition series and hence screen the size determining ns electrons from the nuclear charge more effectively. Therefore across a transition series, there is only a small increase in effective nuclear charge, therefore only a small increase in effective nuclear charge decrease in atomic radius from one element to another takes place.

In 3d series, covalent radius decreases from 144 pm for Sc to 115 pm for

In the Lanthanide elements, filling off orbitals take place, while simultaneously the nuclear charge increases. The electrons in the f orbital shield the ns electrons, almost completely. As a result of this, there is only a small decrease in the atomic radius from one element to another. But there are 14 elements in the series. This is known as lanthanide contraction, because of which the atoms of elements following Lu are usually smaller than they would be if the lanthanide had not been built up before them.

Lanthanide contraction almost exactly cancel out the effect of the last shell added in the sixth period and therefore, the transition elements of

On descending any group of the periodic table, the number of electron in the valence shell remains constant but the number of shells around nucleus increases monotonically so that the effective nuclear charge felt by valence electrons stays nearly the same. Similarly, because of the inclusion of Lanthanide elements in period 6, atoms of the transition elements of this period are almost of the same size as atoms above than in period 5. After that, only a small increase in the size of elements of period 6 as compared to the size of elements above them in period 5 is observed.

Metals Downward the Group

Having gone through this unit, we can conclude that knowledge of the size of an atom is indeed very essential. It is through the knowledge of the atomic radii that we can predict accurately the

reaction of the atom. Periodicity in atomic radii. Arrange the following iso-electronic species in order of decreasing atomic radius.

When elements react, they do so by gaining, losing or sharing of electrons. The energy required to remove the least strongly bound electron from an isolated gaseous atom in its ground state is known as the ionisation energy. Since more than one electron may be removed from an atom, the energy required for the above process is called the first ionisation energy. The second ionisation energy is the energy required to remove an electron from a univalent cation that process is represented by this reaction.

This is because in this case an electron is being removed from a positively charged cation.

Periodicity in Ionisation Energy across Periods

Figure 5.1 shows the plot of first ionisation energies of the elements of the first six periods against their atomic numbers. Figure, the first ionisation energy generally increases from alkali metals to noble gases across any row of the periodic table. 5.1 The Plot of First Ionisation Energies of the Elements of the.

First Six Periods against their Atomic Numbers

You have seen from an earlier section that across any row of the periodic table, the effective nuclear charge steadily increases and the atomic radii decrease. The 2p subshell is higher in energy than the 2s, so the 2p electron of boron is more easily removed than the 2s electron of beryllium. When we come to nitrogen, we will find out that we have a half-filled 2p subshell while in oxygen the 2p subshell is occupied by four electrons. The fourth electron in this 2p subshell is in an orbital already occupied by another electron, so it experiences considerable repulsion.

As a result, this electron is more easily removed than one of the electrons from a singly occupied orbital in nitrogen atom. Thus the ionisation energy of oxygen becomes less than that of nitrogen. Similar anomalies are observed in elements of Period 3 where the first ionisation energies of magnesium and phosphorous are higher than those of aluminium and sulphur respectively. We have earlier seen that across the transition series, the increase in effective nuclear charge and consequent decrease in atomic radius is small.

This again is due to the removal of an electron from a singly occupied np orbitals which are of relatively higher energy than the ns orbital of Zn, Cd and Hg.

Trends in Successive Ionisation Energies

We have already defined successive ionisation energies, that is second, third etc. Values of eight successive ionisation energies of the first twenty elements are listed in Table 5.2. It is evident from the values in the table that the successive ionisation energies of an element inevitably

become larger because the removal of successive electron leaves a higher charge on the nucleus to hold the remaining electrons. It is also clear from the table that the difference between successive ionisation energies of the same element is not constant. Big jumps occur whenever an electron from a subshell of lower principal quantum number is removed for the first time.

For example, for alkali metals, the second ionisation energies are much higher than the first. For alkaline earth metal, the third ionisation energies are much larger than the second and for the halogen, the eighth ionisation energies are much greater than the seventh. So far in our study, you have seen that ionisation energy generally increases across a period and decreases down the group in the periodic table. We can conclude this unit by observing that chemical reaction inevitably involves either again, a loss or sharing of electrons by atoms.

It follows therefore that, whether a reaction takes place or not, the ease with which it can happen and how regularly it happens, depends very much on the factors which affect the formation of ions by the element.

Factors Affecting Electron Affinity

The smaller the size of the atom, the greater will be the force of attraction of the nucleus for the extra action added. More energy will be released in picking up an electron. We have so far defined electron affinity and considered the factors that affect it. In this subsection, we will discuss how the electron affinity varies in the provided descriptions.

Trend across groups

We know from the previous section that on moving down the group of s- and p- block elements in the periodic table, the effective nuclear charge remains almost steady, but there is a general increase in atomic radius due to increase in the value of the principal quantum number n . As a result, the electron affinity generally decreases down any group in the periodic table. This is evident from the values given listed in Table values of electron affinities of second-row non-metals that is B, C, N, O, F are however against the general trend, being smaller than those of corresponding elements that is Al, Si, P, S, Cl of Period 3. This is apparently an indirect result of the small size of the atoms of these elements that is B, C, N, O, F. In conclusion, we can say the electron affinity of an element determines how easy or difficult it is for that element to participate in a chemical reaction.

In this unit, you are going to learn about the electronegativity of an element. The electronegativity of an element is a measure of the power of an atom in a molecule to attract shared electrons to itself. Unlike ionisation and electron affinity, it is not a directly measurable physical quantity but rather, a theoretical concept for which several numerical scales have been developed. Define the concept of electronegativity calculate, given all the necessary impact, the electronegativity of an element, using Pauling, Mulliken-Taffe and Alfred-Rochow electronegativity scales discuss with at least eighty per cent accuracy periodicity in electronegativity.

Pauling Electronegativity Scale

As you know in homonuclear diatomic molecule like A_2 and B_2 , the electron pair is equally shared between the atoms bonded together. Atom A slowly start stripping off its electrons thereby becoming a partially cationic species. As the positively develops on A, its tendency to attract electron increases. Meanwhile, a similar process also takes place on atom B, thus both atoms A and B start attracting the electron pair towards themselves.

They continue doing so until the tendencies of both the atoms in the bonded state to attract the electron pair towards themselves balance. If one of the atoms say B, has higher tendency to attract the electron pair toward itself compared to that the other atom say A, the electron pair will spend more time on B than A. He defined it on the basis of the patterns desirable in the single bond energies of elements which were derived from the thermochemical data. He realized that bond energy, E_{A-B} between two unlike ion like atom A and B is greater than $\sqrt{E_{A-A} E_{B-B}}$ where E_{A-A} and E_{B-B} are bond energies of A-A and B-B homonuclear bonds. He assigned the course of this excess bond energy $E_{A-B} - \sqrt{E_{A-A} E_{B-B}}$ to the electrostatic attraction between partially charged atoms are separated due to difference in electron attracting tendencies of A & B that is the difference in the electronegativity of A and B. Pauling assigned arbitrarily a whole number value 4.0 for fluorine so that values of electronegativity of all elements remain positive.

Table 2.1 shows the values of electronegativity of different elements as calculated by Pauling using his Formulae.

Jaffe Electronegativity Scale

So it will be difficult to remove its electrons. It also has a very high electron affinity. Hence, a very stable species results when electrons are added. On the other hand, an element of low electronegativity will have low ionisation energy and low electron affinity.

So it loses electrons readily and has little tendency to pick up electron. It is very difficult to measure electron affinity for all elements.

Electronegativity Scale

Where Z^* is Slater's effective nuclear charge e the electronic charge and r - the covalent radius. The ability of an atom to attract electrons depends upon the charge on the atom and the hybridisation of the atom. An atom which has acquired a positive charge will tend to attract electrons to it more readily than will a neutral atom. In turn, a negatively charged atom will be less attractive to electron than a neutral atom.

Hybridisation also affects electronegativity because of lower energy and hence, greater electron attracting power of s-orbital. Thus hybrid orbitals having greater s-character possess higher electronegativity. sp^2 hybridised state which will, in turn, be more electronegative than the same atom in sp^3 hybridised state. Thus the carbon atom in CH_4 , C_2H_4 and C_2H_2 has different values of electronegativity.

The larger the difference between the electronegativities of the two elements, the more polar will be the bond between these elements. On the contrary, a difference in the range of 0.4 to 1.7 results in a covalent bond's partial ionic characters or polar covalent bond. Electronegativity values of elements show a fairly discernible periodic trend throughout the periodic table. The trend is similar to that of ionisation energies.

Thus, as expected, the electronegativity of elements increases sharply across a row of S and P-block element. This is as a result of the sharp increase in effective nuclear charge of these elements example from lithium to fluorine. However, across a series of transition elements, the increase in electronegativity is much smaller. This is because the additional electron is being added to an inner shell which provides relatively good shielding for the outer electron from the nucleus.

On moving down a group of representative elements, for example in the lithium group, there is a general decrease in electronegativity. The decrease is relatively small except between the first two elements. The much greater electronegativity of lithium row elements correlates well with their small size. As expected, the elements of Period 4 from gallium onwards that is Ga, Ge, As, Se and Br have greater electronegativities than would be expected by extrapolation from values for the first two elements in the respective groups.

This is due to the insertion of transition elements because of which the effective nuclear charge of these elements is greater than that if the transition elements were not there. Similarly, the presence of the lanthanide elements is responsible for greater electronegativity of the elements of 5d series than would be expected by extrapolation from values of the elements of 3d and 4d series. Let us conclude this unit by pointing out that, studies of ionisation energy and electron affinity show us the tendency of elements to lose or gain electron while going into a relationship. In this unit, on the other hand, we studied the power of these elements while in a relationship with each other to attract to it the electrons in the bond between them.

Explain why the electronegativity values of noble gases are zero while those of halogens are the highest in each period.

Deuterium compounds

In the Units 1 and 2, you studied the development of the periodic table and periodicity in the properties of elements. I am sure that you noticed the fact that the very first element in the periodic table is hydrogen. The hydrogen atom consists of only one proton and one electron. In spite of that, hydrogen forms more compounds than any other element.

In addition, it is the most abundant of all the elements in the universe. In this unit, you will be studying some important aspects of the chemistry of hydrogen. Justify the position of hydrogen in the periodic table describe isotopes of hydrogen differentiate between Ortho and Para forms of hydrogen. The position of hydrogen in the periodic table is of particular interest.

Hydrogen is the first element of the periodic table, with an electronic configuration of $1s^1$. Like alkali metals, hydrogen forms halides, oxides and sulphides. With high ionisation energy, hydrogen resembles halogens. Due to its high ionisation energy, hydrogen forms large numbers of covalent compounds by sharing a pair of electrons.

Hydrogen like halogens forms a diatomic molecule by sharing a pair of electrons between the two atoms. By picking up an electron, hydrogen forms the hydride ion, just like the halogens from the halide ion. From the previous discussion, it is clear that hydrogen resembles both the alkali metals as well as the halogens. So hydrogen can be placed with either of them in the periodic table.

Naturally occurring hydrogen contains 0.0156% deuterium. Like water, which is the oxide of hydrogen, deuterium also forms an oxide, D_2O , which is known as Heavy Water. Deuterium oxide is used as a moderator in nuclear reactions since it slows down neutrons quickly.

We can also employ exchange reactions like those given below for the preparation of deuterium compounds

Tritium differs from the other two isotopes of hydrogen in being radioactive. Naturally occurring hydrogen contains nearly 10-15% tritium.

The following reaction occurs in nature

Tritium can be easily incorporated into biological molecules because it behaves chemically, just like ordinary hydrogen. When two hydrogen atoms combine to form a molecule, there are two possibilities. The two nuclei will either spin in the same direction to give the form called Ortho Hydrogen, or they would spin in opposite directions to give Para Hydrogen. Para hydrogen has a lower internal energy than that of ortho hydrogen.

Hydrogen gas is an equilibrium mixture of Ortho and Para hydrogen. The ratio of Ortho to Para hydrogen varies with temperature as shown in Figure 3.3. The percentage of hydrogen at 300K and above is 75%. This means it is not possible to get 100% ortho hydrogen at any temperature.

The equilibrium mixture of Para and Ortho hydrogen changes to almost

100% Para hydrogen when cooled to nearly absolute zero. Ortho hydrogen, deuterium and tritium also exhibit spin isomerism and exist in Ortho and Para forms. However, the ratio of Ortho to Para forms in deuterium and tritium is different from that in hydrogen. We can conclude this unit by observing that hydrogen holds a unique position in the periodic table.

It is the first element in the periodic table and also exhibits properties of the alkali metals as well as that of the hydrogen. Hydrogen also exists in different forms thus explaining some of its properties. Tritium 3H hydrogen has in addition to H, D and T two other forms, Ortho and Para Hydrogen these two forms arise as a result of differences in spins of the two molecules that make up the H_2 molecule. At room percentage of Ortho hydrogen in the mixture is 70%.

MANUFACTURE OF HYDROGEN

In the last unit, you studied the unique position of hydrogen in the periodic table, you also studied the various isotopes of hydrogen and also the two forms of the hydrogen molecule. In this unit, you shall be studying the ways hydrogen is manufactured, its properties and uses. Water is a naturally abundant source for the manufacture of hydrogen. Water can be reduced to hydrogen either chemically or electrically.

Manufacturing by electrolysis

On a large scale, very pure hydrogen can be obtained from the electrolysis of an aqueous solution of barium hydroxide between nickel electrodes. Hydrogen obtained by electrolysis of water is relatively expensive because of the cost of electrical energy. Hydrogen can, however, be obtained economically as a byproduct in the electrolysis of brine during the manufacture of sodium hydroxide.

The anode and the cathode reactions are combined to give the reactions thus

The reaction can be explosively violent with alkali metals convenient laboratory methods employ sodium amalgam or calcium with water or zinc and tin with hydrochloric acid. Hydrogen is the lightest element known. The hydrogen molecule is.

Thermally stable and has little tendency to dissociate at normal temperatures, the reaction $H_2(g)$

The atomic hydrogen produced, exists for less than half a second after which it recombines to give molecular hydrogen and liberates a large amount of energy, in form of heat. Most of the transition metals catalyse the combination reaction of hydrogen. Atomic hydrogen is a powerful reducing agent and reduces copper, silver and mercury salts to the metallic state of.

H_2O_2

Atomic hydrogen is produced by passing ordinary hydrogen through electric arc maintained between two electrodes. The atoms set free are carried away by a stream of incoming hydrogen gas. This principle is utilised in the making of the 'atomic hydrogen welding torch' It provides an opportunity of welding at a very high temperature but in a reducing atmosphere. Despite the fairly high bond dissociation energy of the hydrogen molecules it is moderately reactive and forms strong bonds with many other elements.

It reacts with almost all elements except the noble gases. 4.1 Atomic Hydrogen Welding Torch.

Hydrogen is easily oxidised to water and; therefore it acts as a very good reducing agent in a variety of situations

Hydrogen is used in metallurgy to reduce metal oxides to metals in cases where carbon cannot be used because the metal can form carbide. Such metals include Mo and W. Hydrogen adds on the

multiple bonds in organic compounds. Hydrogen reacts with carbon monoxide in the presence of catalysts to form methanol.

The half cell reactions are given below

We can see from the equations above that the electromotive force of the cell remains the same whether we use alkaline or acidic electrolyte. Fuel cells have several advantages over other sources of energy. Firstly in a fuel cell unlike in the dry cell or storage battery, the cathode and anode reactants are continuously supplied so that energy can be indefinitely withdrawn from it. Secondly, in a fuel cell energy is extracted from the reactants under almost ideal conditions.

Combustion of hydrogen is a highly exothermic reaction and produces no pollutants

With these facts in mind, there is now an active search for alternative source of energy. In addition to solar power, hydrogen is being considered a potential fuel for the future. Hydrogen as a fuel has many advantages over the conventional fossil fuels and electric power. It releases greater energy per unit weight of fuel in comparison to gasoline and other fuels.

Hydrogen can be transported as a gas in high-pressure pipelines, as a liquid in tankers and even as a solid in form of metal hydrides. Unlike electricity, hydrogen can be stored and used when needed. Hydrogen like electricity is a secondary source of energy because it is produced using energy from a primary source such as coal, nuclear fission or sun. Preparation of hydrogen through electrolysis is not economical at present, in fact, more energy has to be spent in electrolysis of water than what can be liberated by burning hydrogen as a fuel.

Decomposition of water by solar energy in the presence of catalysis is known as photochemical decomposition of water. If this process can be made industrial, a convenient method of converting solar energy directly to a useful form of stored chemical energy will be available. We can conclude this unit by observing that the major material for the manufacture of hydrogen is water. Since water is abundant, the production of hydrogen should be simple.

However, it is not efficient to produce hydrogen by electrolysis. Manufacture by electrolysis even though manufacturing by electrolysis would have been preferred it is too expensive, because of the cost of the energy needed for electrolysis the largest use of hydrogen is in the manufacture of ammonia, which is used to manufacture nitric acid and nitrogenous fertilizers hydrogen is made use of in the production of fuel cells for space crafts hydrogen is also used in extraction of metal and hydrogenation of vegetable oils. Complete the following chemical reactions which take place during the formation of hydrogen.

Covalent Hydrides

In the previous units, you became aware of the fact that hydrogen forms more compounds than any other element. You also saw that despite the fairly high bond dissociation energy of the hydrogen molecule, it is moderately reactive and forms strong bonds with many other elements. You also learnt that it forms ionic and covalent hydrides with metals and non-metals

respectively. As already pointed out at the beginning of this unit, hydrogen combines with a number of elements to form hydrides.

Alkali and Alkaline-earth metal of Groups 1 and 2 are sufficiently electropositive and force the hydrogen atom to accept an electron to form the hydride ion, if eg Lithium hydride LiH - and calcium hydride

Ionic hydrides are white crystalline solids. They have high melting points and conduct electricity in liquid state, liberating hydrogen at the anode. The complex hydrides are frequently used in the reduction of aldehydes, ketones, acids and their derivatives to give alcohols.

The covalent hydrides can be prepared either by direct reaction of nonmetals with hydrogen under suitable conditions or by the reaction of

These hydrides have molecular lattice made up of individual saturated covalent molecules, with only weak van der Waals forces and in some cases along with hydrogen bonds. Some covalent hydrides are unstable in the presence of air, e. Some covalent hydride hydrides of Groups 2 and 13 are electron deficient. These have structures between ionic and covalent hydrides.

When heated, hydrogen reacts with many transition metals to form metallic hydrides. They are almost always non-stoichiometric, being deficient in hydrogen. Most of these hydrides are stable to water up to 375K but are quantitatively decomposed by acids and show some reducing properties. Formerly, these hydrides were formed as interstitial compounds in which hydrogen was through to be accommodated in the interstices in the metal lattice producing distortion but no change in its type.

But recent studies have shown that except for hydrides of nickel palladium, cerium and actinium, other hydrides of this class have a lattice of a type different from that of the parent metal. These characteristics suggest that hydrogen is present in the metal lattice as hydrogen atoms rather than as hydrogen molecules. The single electron of hydrogen is paired with an electron of the metal, thereby reducing the extent of metallic bonding. Breaking of the H-H bond is in agreement with the fact that their metals catalyse reactions of hydrogen.

Effects of Hydrogen Bonding

Hydrogen bonding plays a very significant role in determining the properties of compounds.

If you examine the values for the melting and boiling points shown in

Figure 1.1, you will see that the melting and boiling points of the hydrides of Group 14 elements i. This anomaly is explained on the basis of hydrogen bond formation. In compounds where the molecules are linked by hydrogen bonds, some extra energy is required to break the intermolecular hydrogen bond and this is responsible for their higher boiling and melting points. Intra-molecular hydrogen bond, however, has the opposite effects.

For example, in ortho-nitrophenol, the groups present in ortho position are involved in intramolecular hydrogen bonding thus preventing the intermolecular hydrogen bond formation, i. Solubility of a substance increases markedly when hydrogen bonding is possible between the solvent and the solute molecules. For example, lower alcohols like, methanol, ethanol etc are highly miscible with water due to the hydrogen bonding with water molecules. As the hydrogen cation, i.

We can conclude this unit by stating that hydrogen bond formation is a very important phenomenon and it helps in explaining the behaviour of some compounds. Variations in boiling and melting points and solubility of substances are explainable by the existence of hydrogen bonds.

METALS

You will no doubt recall that at the beginning of our studies we learnt how the efforts of leading scientists resulted in the formation of the periodic table as we know it today. You recall that the end product of that effort is the arrangement of the elements according to their atomic numbers. You also recall how in the last few units, we discussed that the properties of the elements in a periodic table are indeed a periodic function of their atomic numbers as stated by the periodic law. In earlier units, you studied hydrogen and learnt about its unique position.

1 elements in some respects and the Group 17 elements in others. In this unit, you will be studying the elements of Group 1, their occurrence extraction and uses. The elements of Group 1 and 2 are called the S-block elements because the outermost electron in these elements occupy the S-orbital. 1 elements consist of Li, Na, K, Rb, C and Fr.

Metals because they foam hydroxides which are strong Alkalis. List for each member number of the alkali group where they can be land give for each member at least one use discuss, the atomic size, density, melting and boiling points, ionisation energy and electronegativity of members of the group. Alkali metals are useful as metals as well as in the form of their compounds. The alkali metals are highly reactive so they do not occur in the free state in nature.

Francium being a radioactive element with a very short half-life period occurs in very minute's traces in nature. The alkali metals are very good conductors of heat and electricity. Because of the high specific heat and thermal conductivity, liquid sodium is used as a coolant in nuclear reactors. Caesium has the distinction of being the metal from which electrons are ejected most easily on exposure to light.

Other every day uses of alkali metal compounds include the following

The hydrides of lithium and sodium viz LiAlH_4 and NaBH_4 are used as reducing agents in synthetic organic chemistry. Most of the physical properties of the alkali metals are directly relocated to atomic properties of elements. Variation of physical properties from one element to the other in a group is governed by the trends of the various atomic properties earlier discussed in Module 2.

Atomic size

The size of the atom or its ion increases on descending the group. This is due to the addition of an extra shell of electrons as we move down the group from one element to the next. The addition of the extra shell of electrons outweighs the effects of increased nuclear charge and thus there is an increase in size from Li to Cs. 2.2 Trend in Ionic Radii of Group 1 Elements.

Density

Related to atomic size is the density of the elements. There are two general trends observed in the densities of the elements in the periodic table. Since the elemental structuring are often the same within any group, the factors which determine the density are atomic mass and volume. This means that the increase in atomic weight from one element to the next in passing down the group overweighs the effect of increase in the size of the atom.

There are, however, some exceptions to this general trend and in this particular group of alkali metals, you can see from Table 2.1 that the density of potassium is less than that of sodium.

Melting Points and Boiling Points

These metals are soft and can be cut with a knife. These metals have low melting and boiling points which also reflect the low values of cohesive energy between the atoms. 2.3 Trend in the Melting Point of Group 1 Elements.

Thermal and Electrical Conductivity

In alkali metals, electrons of the noble gas core efficiently shield the lone valence shell electron from the nuclear charge. This makes the alkali metals good conductors of heat and electricity. This loosely bound electron is also responsive for the silvery lustre of the alkali metals when freshly cut.

Ionisation energy

By losing the loosely bond solitary outermost electron, these elements can acquire the electronic configuration of the preceding noble gas elements. Figure 2.4 shows the trend in the first ionisation energies of the alkali metals down the group. On account of their low ionisation energies, these elements have a high tendency to form cations. Because alkali metals have a tendency to lose electron easily rather than to gain, values of electronegativity of these elements are very small.

In fact, alkali metals are the least electronegative elements in the periodic table.

Ionic Character of Compounds

Because of the small size of Li^+ , it has more polarising power and therefore favours covalent bonding. Alkali metal salts like halides, oxides, hydroxides, carbonates, sulphates etc exhibit

some interesting trends in their solubility in water. First, let us remind ourselves that lattice energy is the driving force for the formation of an ionic compound and its stability. Lattice energy is directly proportional to the charge on the ions and inversely proportional to the distance between the cation.

This distance is taken as the sum of radii of cation and anion. Let us again remind ourselves that lattice energy is the enthalpy change when one mole of crystal lattice is formed from the isolated gaseous ions and hydration energy is the enthalpy change when one mole of solute is dissolved in water. The higher the charge and the smaller the size of the ion, the more is the hydration energy. In a group, lattice energy and hydration energy decrease as we move down.

While the decrease in lattice energy favours the solubility, the decrease in hydration energy makes the compound insoluble. For the salts of small anions the lattice energy which is inversely proportional to $r_c + r_a$, is very sensitive to the change in the size of the cation, anion being very small in size has little contribution in the total and decreases sharply as we move down the group. Thus, in these salts, the decrease in lattice energy is greater than the decrease in hydration energy and therefore, the solubility of these salts increases as we go down the group. For example, in the case of alkali metal fluorides, the solubility increases in the order $\text{LiF} < \text{NaF} < \text{KF}$, the radii of the cation has little contribution in the total and decreases sharply as we move down the group.

For example, in the case of alkali metal carbonates, lithium carbonate is highly soluble while the solubility of calcium carbonate is very little.

Solutions of Alkali Metals in Liquid Ammonia

All the alkali metals are highly soluble in liquid ammonia giving a deep blue colour. The dissolution of the alkali metal is accompanied by its dissociation into the metal ions and the electrons. The metal ion and the electrons then get associated with ammonia solvent molecules. Solvated Electrons.

The alkali metal solutions in liquid ammonia are highly conducting and behave almost as metals. solvated electron, in all the cases, is the same. The solution of alkali metals in liquid ammonia is blue in colour due to the presence of solvated electrons, and therefore the solutions are also paramagnetic. With increasing concentration, there is a decrease in Paramagnetism suggesting that the electron can get associated to form diamagnetic electron pairs although there may be other equilibria also.

Apart from lithium, other alkali metals can be recovered unchanged from solution. The blue solutions of alkali metals are moderately stable at temperatures where ammonia is still a liquid. The alkali metal solutions in liquid ammonia are powerful reducing agents and are used for this purpose in inorganic and organic reactions.

Complexation Behaviour of Alkali Metals

A complex compound can be defined as a compound with a central atom or ion surrounded by a group of ions or molecules called «ligands». This is because they have a strong tendency of interacting with electron clouds of other anionic or neutral electron-rich species. According to the model above, a very weak coordinating ability is expected in the group 1 metals because of their large size and low charge of the cations M^+ .

Alkali metals form few complexes mostly chelates with the ligands like

Lithium, being the most polarising cation of all the alkali metals forms tetrahedral complexes with ligands like NH_3 , LiH_5N etc. With ammonia, it forms the complex of the type Li , whereas with pyridine a complex of the type formed.

Anomalous Natures of Lithium

On descending any group of S or P block elements of the periodic table, one notices that there is a general decrease in electronegativity or increase in electronegativity. The difference in electronegativity between the first and second elements of each group is much greater than that between any two successive elements. This is reflected in the properties of the elements. Thus not only is the element more electronegative than the other elements of the group, but it is much more electronegative than expected by simple extrapolations.

This trend of bigger than expected differences in properties between the first member of a group and the rest of the elements is shown in group 1 where Li elements markedly different from the rest of the members of the group. Due to the very small size of lithium, the metallic bonding between the atoms in the metallic lattice is very strong giving rise to strong cohesive forces. This is shown in its relatively higher melting point, and boiling point, hardness and homonuclear bond energy. The relatively higher attraction of lithium for its outer electron results in its relatively higher electronegativity, ionisation energy, hydration energy, electron affinity and of course smaller atomic radii relative to the other homologues.

Similar anomalies are also found in the chemical properties, but the differences appear greater as we shall see in the following accounts. Lithium salts of large polarizable anions are thermally less stable than those of other alkali metals. Lithium salts of anions of high charge density are less soluble than those of other alkali metals. We can conclude this unit by stating that the alkali metal ions are very soluble in both water and liquid behaviour.

Also, that lithium shows the anomalous ammonia observed for first members of any group of the elements in the periodic table. Li^+ ion is the least conducting in solution when compared to the ions of the other members of the alkali metal ions in solution. The stability of alkali metal complexes is low the stability of the complexes decrease as you go down the group.

E the Alkali metals, and their compounds. Groups 1 and 2 elements belong to the s-block of the periodic table. Their electronic configurations show an outer shell of ns^1 and ns^2 for alkali and

alkaline earth metals respectively. s-block elements are known to be very reactive metals and generally form ionic compounds.

In this unit, you will be studying the elements of Group 2 consisting of beryllium, magnesium, calcium, strontium, barium and radium. Elements Ca, Sr, Ba and Ra are called alkaline earth metals because their earths are alkaline in nature. Beryllium is not counted as an alkaline earth metal since its oxides is not alkaline. We shall start our study of the alkaline earth metals by studying their occurrence, extraction uses and physical properties.

List at least four places where the alkaline earth metals can be found described at least two methods used in the extraction of the list at least three uses of members of the group compare the physical properties of members of the group with each other. Beryllium, the first member of the group is found in small quantities in the silicate mineral, phenacite, Be_2SiO_4 , and beryl, 3BeO . Magnesium is the second most abundant metallic element next only sodium in seawater. These metals are well known because they occur as concentrated ores and are easy to extract.

Radium is extremely scarce and it is a radioactive element. These metals are extracted by electrolysis of their fused chlorides, though magnesium has been manufactured by the carbon reduction of its oxide. Beryllium is obtained by the electrolysis of molten beryllium chloride.

Sodium chloride must be added to the melt as an electrolyte because

Beryllium alloys are not sparking, therefore, they are used in making hand tools for use in the petroleum industry. The alkaline earth metals are quick soft metals but are harder than the corresponding Group 1 elements. This is because of their two valence electrons which participate in metallic bonding. The atoms of the alkaline earth metals are smaller than those of the corresponding Group 1 elements.

This is because of the increasing atomic number. Because of the resulting increase in effective nuclear charge, valence shell electrons are pulled in more firmly by the nucleus, thereby reducing the size of the atom. Similarly, their ionic radii are also smaller than those of Group 1 elements, because the removal of two orbital electrons increases the effective nuclear charge even further. The elements are denser than Group 1 metals because they have two valence electrons per atom for bonding the atoms into a metallic lattice and as a result, more mass can be packed into a smaller volume.

The atomic/ionic radii increase from Be to Ra due to the effect of extra shells of electrons added. This outweighs the effect of the increased nuclear charge. Group 2 metals have higher melting points when compared to group 1 metals. The reason being the +2 charge on the cations in the metallic lattice causing them to be more strongly attracted to the 'Sea of electrons and making it difficult to pull them apart.

The first ionisation energy of alkaline earth metals is more than that of corresponding alkali metals. This is because the alkaline earth metals have higher nuclear charge and are smaller in size. The electrons are therefore more tightly held to the nucleus. The second ionization energy of these elements is almost twice their first ionization energy.

This is because once one electron has been removed, the effective nuclear charge felt by the orbital electrons is increased so that the remaining electrons are more lightly held and hence much more energy is needed to remove the second electron. However, their second ionisation energy is less than that of the corresponding alkali metals because of the stability of a closed shell configuration of the univalent cations that are formed in the cases of the alkali earth metals. The ionization energy of alkaline earth metals decreases on moving down the group. The metals of this group form ionic compounds.

This is because the assembly of positive and negative ions into a symmetrical crystal lattice results in the liberation of large amounts of energy. Since alkaline earth metals lose electrons easily, they form divalent anions which have noble gas structure with no unpaired electrons. The solubility of alkaline earth metal compounds shows some interesting trends. Decreasing lattice energy favours increased solubility, whilst decreasing hydration energy favours decreased solubility.

If on moving down the group the hydration energy decreases more rapidly than the lattice energy, the compound becomes less soluble. Due to their small ionic radii, Be^{2+} and Mg^{2+} have high hydration energies. In the case of fluorides and hydroxides, the lattice energy decreases more rapidly than the hydration energy. In conclusion, we can say alkaline earth metals, like metals, are only found combined in the form of their salts, because of their reactivity.

Alkaline earth metals are usually extracted by electrolysis. Beryllium is used as fuel containers.

This shows the anomaly in the behaviours of beryllium, in having amphoteric properties

All the alkaline earth metals burn in oxygen to form oxides, MO . With the exception of Beryllium oxide which is covalent, all other oxides are ionic in nature.

Barium peroxide, BaO_2 is formed by passing air overheated

Calcium peroxide, CaO_2 is obtained as a hydrate by treating Ca_2 with hydrogen peroxide H_2O_2 and then dehydrating the product. Magnesium peroxide, MgO_2 is obtained only in the crude form by using hydrogen peroxide but no peroxide of beryllium is known. Alkaline earth metals react less readily with water than alkali metals to give hydrogen and metal hydroxides. Beryllium does not react with water or steam even at red heat.

Metal halides are obtained by direct combination with halogens as well as by the action of halogen on metals. Beryllium halides are covalent and other metal halides are ionic. Beryllium halides are hygroscopic and fume in air due to hydrolysis. Anhydrous beryllium halides are polymeric.

Beryllium chloride vapours contain BeCl_2 and.

Polymer

All the Group 2 elements except beryllium form hydrides, MH_2 , by direct combination with hydrogen. Beryllium hydride can be formed by reducing beryllium chloride with lithium aluminium hydride $LiAlH_4$. All these hydrides are reducing agents which react with water to liberate hydrogen. Calcium strontium and barium hydrides are ionic and contain the hydride ion H^- .

Beryllium and magnesium hydrides are covalent and polymeric. n has an interesting structure. Be has two valence electrons and H only one, it means that there are not enough electrons to form the usual type of bonds in which two electrons are shared between two atoms. The monomeric molecule BeH_2 if formed with normal bonds would have only four electrons in the outer shell of the beryllium atom and would be electron deficient. All the metals in the Mg - Ba series or their oxides react directly with carbon to give the carbides, MC_2 .

BeC₂ with acetylene

Alkaline earth metals burn in nitrogen to form nitrides, M_3N_2 it requires a lot of energy to convert. The stable N_2 molecule into nitride ion, N^{3-} , and this is recovered from the very high lattice energies of the alkaline earth metal nitrides. Beryllium compound is rather volatile while others are not. All Group 2 elements form oxy salts.

The carbonates of alkaline earth metals are stable at room temperature.

The order of decomposition temperature of the sulphate is

Some of the reactions of the alkali earth metals include their action on acid to liberate hydrogen and their reaction with oxygen to form oxides which with the exception of that of beryllium are ionic in nature. That alkaline earth metals are less reactive than alkali metals that the alkaline earth metals react with acids to liberate hydrogen that all alkaline earth metals burn in oxygen to form oxides that the structure of BeO is the wurtzite type, and that of other oxides is the NaCl of structure that the alkaline earth metals react less readily with water that the result of the reaction between alkaline earth metals and water is the formation of hydrogen and metal hydroxides that alkaline earth metals combine directly with halogens to form metal halides which are hygroscopic in nature and some e.

Complexation Behaviour of Alkaline Earth Metals

In Unit 1, you studied the chemical properties of alkaline earth metals. You learnt about the reactions of the alkaline earth metals with oxygen, water and acids. List at least one complex that is formed by each member of the describe the structure of the complexes formed by members of the group list the properties in which beryllium differs from other members of the alkaline earth metals.

Complex-Forming Behaviour of Alkaline Earth Metals

An interesting property of the alkaline earth metals is their complex behaviour. As we have seen earlier, complex formation is favoured by small, highly charged cations with suitable empty

orbitals of approximately the right energy with which the Ligand orbitals can combine. Alkaline earth metal form more complexes as compared to alkaline metals. The tendency to form complex decreases with increasing atomic number.

Beryllium having the smallest ion in the group tends to form complex most readily. It mostly forms complexes with tetrahedral arrangement because of the available orbitals as shown in Figure 2.1. In the hydrated salt, example $\text{BeCO}_3 \cdot 4\text{H}_2\text{O}$ and $\text{BeSO}_4 \cdot 4\text{H}_2\text{O}$ beryllium ions exist in the form $2+$ where they show a coordination number of four.

Magnesium in chlorophyll is coordinated by four nitrogen atoms in the heterocyclic porphyrin ring system

Beryllium, the first member of the alkaline earth metal group, differs from the members of the group just as lithium differs from three members of its group. In fact, beryllium differs more from magnesium than lithium does from sodium. Let us now look at these properties one at a time. The cohesive properties of beryllium are much greater than those of magnesium and other elements in the group.

It has smaller atomic radii higher electron affinity and higher ionization energy. Its higher polarizing power leads to all its compounds being largely covalent with lower melting and boiling points, and enthalpies of formation. In conclusion, you have seen in this unit that alkaline earth metals form more complex than the alkali metals of Group 1. Beryllium shows behaviours which differ from that expected from a member of Group 2.

Alkaline earth metals form complexes beryllium being smaller in size than the other member of the group forms more complexes magnesium forms important complexes that occur in nature. An example of such a complex is the one found in chlorophyll, the green pigment found in plant beryllium the first member of the alkaline earth metal shows anomalous behaviour that is, it behaves differently than it is expected to do as a member of the group this anomalous behaviour is manifested in it having much higher melting and boiling points. It also has smaller atomic radii, higher electron affinity and higher ionisation energy. Explain why noble metals like Au, Pt, Rh and Ir are not attacked by nitric acid, but Au and Pt are dissolved by Aqua regia.

FBQ1: When two hydrogen atoms combine to form a molecule and the two nuclei spin in the same direction, parallel spins, to give the form known as _____

Answer: Ortho Hydrogen

FBQ2: The two hydrogen nuclei spin in opposite directions to give _____

Answer: Para Hydrogen

FBQ3: Hydrogen can be produced by the reaction of methane with _____ in the presence of nickel catalyst

Answer: Steam

FBQ4: Metals are very powerful _____ agent

Answer: Reducing

FBQ5: Sulphur reacts with _____ to form hydrogen sulphide

Answer: Hydrogen

FBQ6: Carbon monoxide is reduced strongly by hydrogen to yield a product appropriately described as _____

Answer: Formaldehyde

FBQ7: In a fuel cell, electrical energy is generated by the reaction of hydrogen and oxygen, a process which is called _____

Answer: Cold combustion

FBQ8: Fuel cells have efficiencies approaching 75% whereas power plants that burn fuels have efficiencies of only about _____

Answer: 40%

FBQ9: _____ are soft have low melting points and are poor conductors of electricity.

Answer: covalent hydrides

FBQ10: _____ hydrogen bond is formed between two atoms of the same molecule

Answer: Intramolecular

FBQ11: Half-life period of tritium is _____ years

Answer: 12.3

FBQ12: Melting point and boiling point increases in group 14 elements because of _____

Answer: Hydrogen bonding

FBQ13: Solubility of a substance increases markedly when _____ is possible between the solvent and the solute molecules

Answer: Hydrogen bonding

FBQ14: How many types of hydrogen bonding do we have?

Answer: Two

FBQ15: What percentage of rubidium chloride is contained in Carmallite?

Answer: 0.94%

FBQ16: The hydrides of lithium and sodium are used as _____ agents in synthetic organic chemistry

Answer: Reducing

FBQ17: the alkali metals are very useful, some of their uses include being used as _____ conductors.

Answer: Electrical

FBQ18: Normal oxide and peroxide of alkaline metals are colourless and _____

Answer: Diamagnetic

FBQ19: Super oxides of alkaline metals are usually Coloured and _____ in nature

Answer: Paramagnetic

FBQ20: All the Group 1 metal oxides are strongly _____ and react vigorously to give hydroxide

Answer: Basic

FBQ21: polysulphides of sodium have a _____ chain structure

Answer: zig-zag

FBQ22: The density of gallium was ____ kg as predicted by Mendeleev

Answer: 5.8×10^3

FBQ23: The works of Lars Fredrick Nilson led to discovered one very important element known as _____

Answer: Scandium

FBQ24: _____ is the scientist who first discovered germanium as an element

Answer: Winkler

FBQ25: The maximum number of electrons that can be contained by d-orbital in opposite spin is _____

Answer: Ten

FBQ26: F orbital can hold maximally _____ number of electrons

Answer: Fourteen

FBQ27: The maximum number of electrons that can be contained by p-orbital in opposite spin is _____

Answer: Six

FBQ28: How many sub orbitals has f-orbital?

Answer: Seven

FBQ29: The electronic configuration of _____ is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

Answer: Zinc

FBQ30: Period 1 of the modern periodic table consists of how many elements _____

Answer: Two

FBQ31: The metallic radius depends to some extent on _____ structure of the metal.

Answer: Crystal

FBQ32: Hydrogen may not be advantageous as a fuel because it is a secondary _____

Answer: Source of energy

FBQ33: Elements of the periodic table have been divided into how many blocks?

Answer: Four

FBQ34: Lanthanides and actinides are collectively known as _____ element.

Answer: F-block

FBQ35: A periodic element that behaves both as a metal in group 1A and also as a halogen is most likely to be

Answer: Hydrogen

MCQ1: Second ionisation energy is greater than the first because

Answer: Electron is removed from the positively charged cation that is held firmly due to nuclear forces.

MCQ2: The valence shell electron of _____ are more stable and requires very great energy to remove them

Answer: Noble gases

MCQ3: Ionisation energy of an element depends on all of the following except;

Answer: Crystal lattices of the atom of element

MCQ4: Which of the following element has the lowest ionisation energies?

Answer: Ne

MCQ5: The energy required to remove the least strongly bond electrons from an isolated gaseous atom on ground state is appropriately described as

Answer: Ionisation energy

MCQ6: The energy released or absorbed when an electron is added to the gaseous atom in its ground state is described as

Answer: Electron affinity

MCQ7: Which of these elements shows the highest reluctance to form an anion?

Answer: K

MCQ8: Which of these factors would not affect electron affinity?

Answer: Steric effect

MCQ9: _____ is the tendency of an atom to attract toward itself the shared electron pair of a bond in which it is involved

Answer: Electronegativity

MCQ10: One of these is not an isotope of hydrogen

Answer: Polonium

MCQ11: Which of these isotopes of hydrogen is radioactive?

Answer: Tritium

MCQ12: The tendency of an atom to attract toward itself the shared electron pair of a bond in which it is involved can be measured by all these scale except;

Answer: Lothar Meyer electronegativity scale

MCQ13: How would you effectively separate a mixture of carbon (iv) oxide and Hydrogen gas?

Answer: Pass the mixture through water which absorbs CO₂ and hydrogen gas remains insoluble

MCQ14: Mixture of CO and H₂ is known as

Answer: water gas

MCQ15: It is not advisable to prepared hydrogen gas by one of these methods

Answer: Reaction of potassium metal with warm water

MCQ16: Which of these is used in metallurgy to reduce metal oxides to metals in cases where carbon cannot be used because the metal can form carbide?

Answer: Hydrogen

MCQ17: Stability of the alkali metal complexes decrease as you go down the group one of these order

Answer: Li > Na > K > Rb > Cs

MCQ18: The following metallic elements are most likely to be extracted by electrolysis of their fused chloride except

Answer: Magnesium

MCQ19: One of these is used as a window materials in x-ray apparatus and also in making atomic fuel containers

Answer: Beryllium

MCQ20: The process by which metal ion is surrounded by solvent molecules is appropriately described as

Answer: Solvation

MCQ21: All are the three types of oxides which are formed by the alkali metals except

Answer: Amphoteric oxide

MCQ22: One of these elements in its stearate form may be used as grease

Answer: Li

MCQ23: One of these metal element is obtained by the reduction of its chloride with sodium vapour

Answer: Potassium

MCQ24: Sodium occur naturally in combine state due to its reactivity. Sodium does not occur in one of these ores

Answer: Kainite

MCQ25: One of these is not a hydride

Answer: Stoichiometric hydrides.

MCQ26: The attractive force which binds hydrogen atom of one molecule with electronegative atom of another molecule, generally of the same compound is known as

Answer: Hydrogen bonding

MCQ27: One of these is not a property of metals

Answer: They form acidic oxides

MCQ28: The concept of atomic number was essentially discovered in 1913 by one of these scientists

Answer: Henry Moseley

MCQ29: _____ states that as far as possible in a given atom in the ground state, electrons in the same sub shell will occupy different orbitals and will have parallel spins

Answer: Hund's rule

MCQ30: There exist a set of empty hydrogen like orbitals into which electrons can be added.

This assumption was made by _____

Answer: Aufbau principle

MCQ31: Which of these is the electronic configuration of scandium?

Answer: [Ar] 3d14s2

MCQ32: The electronic configuration of nickel is _____

Answer: [Ar] 3d74s2

MCQ33: Metal halides may be obtained by the direct combination of a metal and _____

Answer: Halogen

MCQ34: One of these is the electronic configuration of copper?

Answer: 1s2 2s2 2p6 3s2 3p6 3d10 4s1

MCQ35: Identify the element that has this electronic configuration: 1s2 2s2 2p6 3s2 3p6 4s2.

Answer: Ca