In liquid pentanol which intermolecular forces are present

Show transcribed image text 100% (2 ratings)Previous question Next questionPart A In liquid pentanol, CH3CH2CH2CH2CH2OH which intermolecular forces are present? Only ion-dipole forces are present. Dispersion, hydrogen bonding and dipole-dipole forces are present. Only dipole-dipole and ion-dipole forces are present. Only dispersion and hydrogen bonding forces are present. Recall the two conceptual steps necessary to dissolve a solute and form a solution There are two conceptual steps to form a solution, each corresponding to one of the two opposing forces that dictate solubility. The first conceptual step is dissolution, which corresponds to the force of the solvent-solvent and solute-solute intermolecular attractions that needs to be broken down. The second conceptual step is solvation, which corresponds to the force of the solute-solvent intermolecular attraction that needs to be formed in order to form a solution. Many intermolecular forces can contribute to solvation, including hydrogen bonding, dipole-dipole forces, Van Der Waals forces, and ion-dipole interactions. The strength of the intermolecular forces between solutes and solvents determines the solubility of a given solute in a given solvent. In order to form a solution, the solute must be surrounded, or solvated, by the solvent. Solutes successfully dissolve into solvents when solute-solvent bonds are stronger than either solute-solute bonds or solvent-solvent bonds. Qualitatively, one can determine the solubility of a solute in a solvent by using the rule "like dissolves like". In general, solutes whose polarity matches that of the solvent will generally be soluble. For example, table salt (NaCl) dissolves easily into water (H2O) because both molecules are polar. Intermolecular Forces and Their Importance in Solution Formation There are two conceptual steps to form a solution, each corresponding to one of the two opposing forces that dictate solubility. If the solute is a solid or liquid, it must first be dispersed — that is, its molecular units must be pulled apart. This requires energy, and so this step always works against solution formation (always endothermic, or requires that energy be put into the system). Step 1 of dissolutionMolecules going from an ordered state, such as a solid, to a disordered state require an input of energy.The nature of the solute (X) and solvent (Y) determines whether dissolution is energetically favorable or unfavorable. If the solute binds to other solute (X-X bond) more strongly than the solute binds to the solvent (X-Y bond), then the dissolution is not energetically favorable. Step 1: Dissolving Solutes The nature of the solute (X) and solvent (Y) determines whether dissolution is energetically favorable or unfavorable. If the solute binds to other solute (X-X bond) more strongly than the solute binds to the solvent (X-Y bond), then the dissolution is not energetically favorable. On the other hand, dissolution is favorable when solute-solvent bonds (X-Y) are stronger than X-X or Y-Y bonds. In this case, the potential energy is lower when the solute and solvent can form bonds. If the X-Y attractions are stronger than the X-X or Y-Y attractions, the dissolution reaction is exothermic and releases energy when the solute and solvent are combined. Since the dissolution of the solvent (X-X) and solute (Y-Y) is always positive, the determining factor for solution formation is the value of X-Y. Remember that the interactions between X and Y, represented above as X-Y, are classified as intermolecular forces, which are not covalent (bonding) interactions. Step 2: Forming a Solution After dissolution occurs, solvation follows. If solvation releases more energy than is consumed during dissolute, then solution formation is favored and the solute is soluble in the solvent. Many intermolecular forces can contribute to solvation, including hydrogen bonding, dipole-dipole forces, and Van Der Waals forces. Ion-Dipole Interactions Another common example of these forces at work is an ion-dipole interaction, which arises when water solvates ions in solution. This interaction arises most prevalently when strong or weak electrolytes are place in water. Consider the dissolution of table salt (sodium chloride) in water: [latex]NaCl (s) \rightarrow Na^+(aq)+Cl^-(aq)[/latex] The water molecules form a solvent cage around each Na+ or Cl– ion, as implied by the aqueous state symbol (aq) following each of the products. The positive ion, Na+, is surrounded by water molecules that have the negative dipoles of the water, or the oxygen, pointing towards the cation. Solvation of a cation by water.Water molecules (gray/green is hydrogen, orange is oxygen) surround a sodium cation in a solution. Notice the negative dipole or the oxygen molecules are 'facing' the Na+. In this case, the anion Cl– is solvated by the positive dipoles of water, which are represented by hyrogen atoms. In general, when ions are present in water, each cation and anion is surrounded by a 'cage' of partial negative or partial positive charge, respectively. These interactions explain why most ionic compounds are considered soluble in water, unless specifically labeled otherwise. Which will have the strongest dipole (a) F2 or HF, (b) CH3Cl or CH3Br? Answer (a) HF (b) CH3Cl Which is more polarizable? (a) Cl2 or I2, (b) C2H6 or C10H22? Answer (a) I2 (b) C10H22 Which of the following materials is likely to have (a) no dipole-dipole forces, but the largest London dispersion forces, (b) the largest dipole-dipole intermolecular forces: I2, He, H2S, H2Te. Answer (a) I2, it is the largest nonpolar molecule. (b) H2S, S is more electronegative and will make the molecule more polar. What kind of attractive forces must be overcome to (a) sublime At2; (b) vaporize C2H5F; (c) boil hydrogen fluoride, HF; (d) melt LiBr? Explain. Answer (a) Nonpolar - London Dispersion Forces. (b) Polar - Dipole-dipole attractions (and dispersion forces). (c) Very polar - Hydrogen bonds between the molecules. Hydrogen bonds form when hydrogen is covalently bonded to N, O, or F. (d) No molecules, so there are no intermolecular forces - Ionic bonds. Which of the following materials is likely to have (a) no dipole-dipole forces, but the largest London dispersion forces, (b) the largest dipole-dipole intermolecular forces: I2, He, H2S, H2Te. Answer (a) I2, it is the largest nonpolar molecule. (b) H2S, S is more electronegative and will make the molecule more polar. Identify the following crystals (A-E) as ionic, polar-molecular, nonpolar-molecular, covalent (network), or metallic. Explain. There could be more than one or none of the crystal types in the chart. CrystalMelting Point (°C)Boiling Point (°C)Electrical Conductivity Solid Liquid A-83.119.54NoNo B-259.14-252.5NoNo C15353000YesYes D6861330NoYes E-56.6-78.5NoNo F-182.48-164.8NoNo G35504827NoNo Answer A ⇒ good separation between melting and boiling points, relatively high m.p. for molecular, no conductivity - polar-molecular. (HF - polar molecule) B ⇒ close and low melting and boiling points, no condutivity - nonpolar-molecular. (H2 - nonpolar molecule) C ⇒ high mp and bp, good conductor - metallic. (Fe - metal) D ⇒ high melting and boiling points, conducts as a liquid - ionic. (KI - ionic compound) E ⇒ low and close together mp and bp, nonconductor - nonpolar-molecular. (CO2 - nonpolar molecule) F ⇒ low and close together mp and bp, nonconductor - nonpolar-molecular. (CH4 - nonpolar molecule) G ⇒ very high mp and bp, nonconductor - covalent (network). (Diamond - covalent network) Indicate the type of crystal (ionic, polar-molecular, nonpolar-moleculas, covalent, or metallic) and predict some of the properties (difference in melting and boiling points, conductivity, etc.) for each of the following substances upon solidification: HF, H2, Fe, KI, CO2, CH4, and diamond. Answer HF - polar molecule, good separation between melting and boiling points, relatively high m.p. for molecular, no conductivity - polar-molecular. H2 - nonpolar molecule, close and low melting and boiling points, no condutivity - nonpolar-molecular. Fe - metal, high or low mp and bp, good conductor - metallic. KI - ionic compound, high melting and boiling points, conducts as a liquid - ionic. CO2 - nonpolar molecule, low and close together mp and bp, nonconductor - nonpolar-molecular. CH4 - nonpolar molecule, low and close together mp and bp, nonconductor - nonpolar-molecular. Diamond - covalent network, very high mp and bp, nonconductor - covalent (network) CrystalMelting Point (°C)Boiling Point (°C)Electrical Conductivity Solid Liquid HF-83.119.54NoNo H2-259.14-252.5NoNo Fe15353000YesYes KI6861330NoYes CO2-56.6-78.5NoNo CH4-182.48-164.8NoNo Diamond35504827NoNo List all the intermolecular interactions that take place in each of the follow kings of molecules: \(CCl_3F\), \(CCl_2F_2\), \(CClF_3\), and \(CF_4\). Determine what type of intermolecular forces exist in the following molecules: LiF, MgF2, H2O, and HF. S13.1b H2O: London Force, Dipole-Dipole interaction, Hydrogen bonds. HF: Dipole-Dipole intermolecular forces, Hydrogen bonds. MgF2 and LiF: strong ionic attraction. Arrange the follow species in order of decreasing melting points: CsBr, KI, KCL, MgF2. Which has the highest boiling point I2, Br2, and Cl2. Explain why? The atomic weigh of Iodine = 127, Bromine = 80, and Chlorine = 35.5. The weigh is proportion to the London dispersion force, and the higher molecular weigh, the larger the force. Thus, I2 has a highest boiling point. 1-Propanol C3H7OH and methoxyethane CH3O C2H5 have the same molecular weigh. Which has the higher boiling point? The 1-Propanol can form London Force, Dipole- Dipole, and H- bonding due to the H bonded to O atom of OH group, whereas the methoxyethane can not form the H-bonding. Therefore, the 1-Propanol has higher intermolecular attractive force and thus a higher boiling point. Why do the lightest compounds such as NH3, H2O, and HF have the highest boiling points? What kind of attractive interaction exists between atoms and between nonpolar molecules? Why nature gas CH4 is a good choice to storage tank in winter? It's about boiling point. The methane has the boiling point at -161 °C, making it to be a good choice for winter season. Explain why methane (CH_4\) is used as the primary heating gas in Alaska during wintertime instead of the more commonly used butant or propane gases use in the lower 48 states. Methane (\(CH_4\)) remains gas because its boiling point is about -160°C. Other gases, such as propane or butane, would liquefy under freezing condition. Therefore, methane is more likely to be used during wintertime at Alaska Define types of intermolecular forces and give example for each. There are 3 types of intermolecular force: London Dispersion, Dipole-Dipole (Example: Two \(NaCl\)) and Ion-Dipole (Example: \(Mg^+\) and \(HCl\)) Dipole- Dipole occurs between polar molecules Ion- Dipole occurs between an ion and polar molecules London Dispersion occurs between the nonpolar molecules. How does the intermolecular determine the boiling point? The weakest intermolecular, the lowest boiling point. What is their dipole-dipole interaction of wo HCl molecules are co-linear head-to tail. Given: The dipole moment of HF is 1.86 D. The dipole moment of HCl is 1.05 D. The distance between the two is 1.78 H+-F- - - - - H+-Cl- \[V=-\dfrac{2\mu _{A}\mu _{B}}{4\pi \varepsilon _{0}r^{3}}\] \[V=-\dfrac{2(1.05)(1.86)}{4\pi (8.854187817\cdot 10^{-12})(1.78)^{3}}\] Calculate an ion-ion interaction energy between \(K^+\) and \(Cl^-\) at a distance of 600 pm. The ion-ion interaction energy is given by Coulomb's law. \[V = \dfrac{q_1q_2}{4 \pi \epsilon_o r}\] \[V = \dfrac{- (1.602 \times 10^{-19}\;\cancel{C})(1.602 \times 10^{-19} \cancel{C})}{4 \pi (8.853 \times 10^{-12} \cancel{C^2} \cdot N^{−1} \cdot m \cancel{^{−2}})(6 \times 10^{-10}\; \cancel{m})} = -3.84 \times 10^{-19} \; J\] Calculate an ion-dipole interaction energy between \(K^+\) and \(HCl\) at a distance of 600 pm. \(HCl\) has a dipole moment of \(1.08\;D\). \[\mu = 1.08 \cancel{D} \times \dfrac{3.3356 \times 10^{30} \; C \cdot m}{1\;\cancel{D}} = 3.6 \times 10^{-30}\; C \cdot m\] \[V = \dfrac{-q\;\mu}{4 \pi \epsilon_o r^2} = \dfrac{- (1.602 \times 10^{-19}\;\cancel{C})(3.6 \times 10^{-30} \cancel{C} \cdot \cancel{m})}{4 \pi (8.853 \times 10^{-12} \cancel{C^2} \cdot N^{−1} \cdot m \cancel{^{−2}})(6 \times 10^{-10}\; \cancel{m})^2} = -1.44 \times 10^{-20} \; J\] As expected this is appreciably smaller in energy than covalent bonds (e..g, \(HCl\) has a bond enthalpy of \(7.0 \times 10^{-19}\;J\)). Rank the interactions from weakest to strongest: H2O - - OH2 Li+ - - F- Li+ - - OH2 S13.8a ion-ion interaction: Li+ - - F- ion-dipole interaction: Li+ - - OH2 dipole-dipole interaction: H2O - - OH2 A low concentration electrolytic solution behaves non-ideally while a high concentration of the same solution behaves ideally. Explain this phenomenon in terms of forces, noting that Coulomb forces depend on 1/r2 while van der Waals forces depend on 1/r7. Which of these forces are low concentration electrolytic solutions likely to follow? High concentration? The interatomic distances in a low concentration electrolytic solution are greater than those in a high concentration solution. They follow van der Waals forces and thus behave less ideally. High concentration electrolytic solutions follow Coulomb forces. Calculate the ion-dipole interaction between H2O and Li+. You are given the dipole moment of H2O is 1.82 D. The distance between these two is 2 Å. \[V=-\frac{q\mu }{4\pi \varepsilon _{0}r^{2}}\] \[=\frac{1.82D\cdot(\frac{3.3356\cdot 10^{-30}Cm)}{1D}}{4\pi (8.85\cdot 10^{-12})(2\cdot 10^{-10}m)^{^{2}}} =1.36\; kJ/mol\] Calculate the potential energy of interaction between a Cl- ion situated 120 pm away from an \(H_2O\) molecule with a dipole moment of 1.85 D. \[\mu = 1.85 \cancel{D} \times \dfrac{3.3356 \times 10^{30} \; C \cdot m}{1\;\cancel{D}} = 6.18 \times 10^{-30}\; C \cdot m\] \[r = 1.2 \times 10^{-10}\; m\] \[V = \dfrac{q\mu}{4\pi \varepsilon _{o}r^{2}} = \dfrac{(-1.602\times10^{-19}\;C)(6.18 \times 10^{-30}\; C \cdot m)}{4\pi (8.851 \times 10^{-12}\; C^{-2}N^{-1}m^{-2})(1.2 \times 10^{-10} \; m) ^2}\] Do you expect a greater dipole-dipole interaction between two molecules that are antiparallel or between two molecules that are co-linear head-to-tail? You expect a stronger interaction when the two are co-linear head-to tail. This can be seen by looking at the formula or in the images of the two. Express the equilibrium distance re in term ð and show V = - € Q13.12 Determine Vander Waals radius of Argon. Use this radius to find fraction of volume by 2 mole of argon at room temperature at 1 atm. a. r= σ/2 = 3.40 A0 /2 = 1.70 A0 b. Volume of 2 mole of Ar 4/3 πr^3 ((6.022 x 10^23)/(2 mol))= 4/3 π (1.70 x 10^(-10) m)^3 ((6.022 x 10^23)/(2 mol)) = 6.19 x 10-6 L mol-1 V/n=RT/P= ((0.08206 L atm K^(-1) mol^(-1) (298.2 K))/1atm = 24.5 L mol-1 The fraction of this volume occupied by 2 mole of Ar (1.239 x 10^-2 L mol-1)/ 24.47 L mol-1 = 2.5 x10 -7 a) What is the original of polarity in a molecule? b) Is CO2 polar? Explain. Of the following compound, which one(s) is/are soluble? What makes a compound soluble in water? Explain using examples. Explain why does water have a high specific heat. The energy of a hydrogen bond for each base pair in DNA is 15 kJ/mol. Two complimentary strands has 50 base pairs each. What is the ratio of the 2 different strands to hydrogen double helix in a solution given a temperature of 300 K. First calculate the ratio of the two different strands for just one pair. \[e^{\Delta E/RT}=exp[(15 \times 10^3\; J/mol)/(8.314\; J/K*mol)(300\; K) = 2.4 \times 10^{-3}\] Since there are 50 base pairs, we need to multiply by 50 to account for all the base pairs. exp[100X(15X103 J/mol)/(8.314 J/K*mol)(300K) = 0 Consider two pure liquids. One has strong intermolecular interactions, and the other has relatively weak intermolecular interactions. For the following properties, indicate which of the liquids you would expect to have a higher value (answer with "strong" or "weak"). viscosity vapor pressure freezing point surface tension S13.24 Strong. Higher viscosity results from stronger interactions between the liquid molecules. Weak. The liquid with weaker bonds takes less energy to turn into vapor, so it will exert a higher vapor pressure. Strong. The freezing point is the same as the melting point; it takes more energy to melt a solid with stronger intermolecular interactions. Strong. Surface tension is a result of intermolecular interactions. The stronger these interactions, the greater the surface tension. Fun fact: if the DNA in a single human cell were stretched out (but still in its familiar double helix conformation), it would be approximately 2 meters long. The distance, along the helix, between nucleotides is 3.4 Å. Estimate the number of basepairs in the haploid human genome, from the 2 meter fun fact. The human body contains about 100 trillion cells. About a quarter of these are erythrocytes (red blood cells) and contain no genomic DNA. Use the average molar mass for a basepair, 650 grams per mole, to estimate how much of a human's mass is human genomic DNA. At its closest, Pluto is 4.28 billion km from Earth. Do you have enough DNA to reach Pluto? Hint: Humans are diploid. (a) $$ 2\ m/cell \times \dfrac{bp}{3.4\ Å} \times \dfrac{10^{10}\ Å}{m} \times \dfrac{cell}{2\ haploid\ genomes} = 3 \times 10^{9} \dfrac{bp}{haploid\ genome} \] Three billion basepairs. (b) 75 trillion of the human cells in your body have genomic DNA. $$ 75 \times 10^{20}\ cells \times \dfrac{haploid\ genomes}{cell} \times \dfrac{3 \times 10^9\ bp}{haploid\ genome} \times \dfrac{mol}{6.022 \times 10^{23}} \times 650 \dfrac{g}{mol\ bp} = 200\ g \] That's about half a pound. (c) $$ \dfrac{2\ m}{cell} \times 75 \times 10^{12}\ cells \times \dfrac{km}{1000\ m} = 2 \times 10^{11}\ km \] Yes, you have way more DNA than you need to stretch it from Earth to Pluto. Page 2This page explains the origin of hydrogen bonding - a relatively strong form of intermolecular attraction. Many elements form compounds with hydrogen. If you plot the boiling points of the compounds of the Group 4 elements with hydrogen, you find that the boiling points increase as you go down the group. The increase in boiling point happens because the molecules are getting larger with more electrons, and so van der Waals dispersion forces become greater. If you repeat this exercise with the compounds of the elements in Groups 5, 6 and 7 with hydrogen, something odd happens. Although for the most part the trend is exactly the same as in group 4 (for exactly the same reasons), the boiling point of the compound of hydrogen with the first element in each group is abnormally high. In the cases of NH3, H2O and HF there must be some additional intermolecular forces of attraction, requiring significantly more heat energy to break. These relatively powerful intermolecular forces are described as hydrogen bonds. The molecules which have this extra bonding are: The solid line represents a bond in the plane of the screen or paper. Dotted bonds are going back into the screen or paper away from you, and wedge-shaped ones are coming out towards you. Notice that in each of these molecules: The hydrogen is attached directly to one of the most electronegative elements, causing the hydrogen to acquire a significant amount of positive charge. Each of the elements to which the hydrogen is attached is not only significantly negative, but also has at least one "active" lone pair. Lone pairs at the 2-level have the electrons contained in a relatively small volume of space which therefore has a high density of negative charge. Lone pairs at higher levels are more diffuse and not so attractive to positive things. Consider two water molecules coming close together. The + hydrogen is so strongly attracted to the lone pair that it is almost as if you were beginning to form a co-ordinate (dative covalent) bond. It doesn't go that far, but the attraction is significantly stronger than an ordinary dipole-dipole interaction. Hydrogen bonds have about a tenth of the strength of an average covalent bond, and are being constantly broken and reformed in liquid water. If you liken the covalent bond between the oxygen and hydrogen to a stable marriage, the hydrogen bond has "just good friends" status. Notice that each water molecule can potentially form four hydrogen bonds with surrounding water molecules. There are exactly the right numbers of + hydrogens and lone pairs so that every one of them can be involved in hydrogen bonding. This is why the boiling point of water is higher than that of ammonia or hydrogen fluoride. In the case of ammonia, the amount of hydrogen bonding is limited by the fact that each nitrogen only has one lone pair. In a group of ammonia molecules, there aren't enough lone pairs to go around to satisfy all the hydrogens. In hydrogen fluoride, the problem is a shortage of hydrogens. In water, there are exactly the right number of each. Water could be considered as the "perfect" hydrogen bonded system. The diagram shows the potential hydrogen bonds formed to a chloride ion, Cl-. Although the lone pairs in the chloride ion are at the 3-level and wouldn't normally be active enough to form hydrogen bonds, in this case they are made more attractive by the full negative charge on the chlorine. However complicated the negative ion, there will always be lone pairs that the hydrogen atoms from the water molecules can hydrogen bond to. An alcohol is an organic molecule containing an -O-H group. Any molecule which has a hydrogen atom attached directly to an oxygen or a nitrogen is capable of hydrogen bonding. Such molecules will always have higher boiling points than similarly sized molecules which don't have an -O-H or an -N-H group. The hydrogen bonding makes the molecules "stickier", and more heat is necessary to separate them. Ethanol, CH3CH2-O-H, and methoxymethane, CH3-O-CH3, both have the same molecular formula, C2H6O. They have the same number of electrons, and a similar length to the molecule. The van der Waals attractions (both dispersion forces and dipole-dipole attractions) in each will be much the same. However, ethanol has a hydrogen atom attached directly to an oxygen - and that oxygen still has exactly the same two lone pairs as in a water molecule. Hydrogen bonding can occur between ethanol molecules, although not as effectively as in water. The hydrogen bonding is limited by the fact that there is only one hydrogen in each ethanol molecule with sufficient + charge. In methoxymethane, the lone pairs on the oxygen are still there, but the hydrogens aren't sufficiently + for hydrogen bonds to form. Except in some rather unusual cases, the hydrogen atom has to be attached directly to the very electronegative element for hydrogen bonding to occur. The boiling points of ethanol and methoxymethane show the dramatic effect that the hydrogen bonding has on the stickiness of the ethanol molecules: ethanol (with hydrogen bonding) 78.5°C methoxymethane (without hydrogen bonding) -24.8°C The hydrogen bonding in the ethanol has lifted its boiling point about 100°C. It is important to realize that hydrogen bonding exists in addition to van der Waals attractions. For example, all the following molecules contain the same number of electrons, and the first two are much the same length. The higher boiling point of the butan-1-ol is due to the additional hydrogen bonding. Comparing the two alcohols (containing -OH groups), both boiling points are high because of the additional hydrogen bonding due to the hydrogen attached directly to the oxygen - but they are not the same. The boiling point of the 2-methylpropan-1-ol isn't as high as the butan-1-ol because the branching in the molecule makes the van der Waals attractions less effective than in the longer butan-1-ol. Hydrogen bonding also occurs in organic molecules containing N-H groups - in the same sort of way that it occurs in ammonia. Examples range from simple molecules like CH3NH2 (methylamine) to large molecules like proteins and DNA. The two strands of the famous double helix in DNA are held together by hydrogen bonds between hydrogen atoms attached to nitrogen on one strand, and lone pairs on another nitrogen or an oxygen on the other one. Contributors and Attributions Jim Clark (Chemguide.co.uk)