



NATIONAL MATH + SCIENCE INITIATIVE

AP CHEMISTRY

IMFs and Condensed States of Matter

2016 EDITION

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https://www.surveymonkey.com/r/S_SSS



Periodic Table of the Elements

1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.012															9 F 19.00	10 Ne 20.179
11 Na 22.99	12 Mg 24.30															17 Cl 35.453	18 Ar 39.948
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.938	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 93.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.02	89 †Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 § (269)	111 § (272)	112 § (277)	§Not yet named					

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

*Lanthanide Series:

†Actinide Series:

AP Chemistry Equations & Constants

Throughout the test the following symbols have the definitions specified unless otherwise noted.

L, mL = liter(s), milliliter(s)

g = gram(s)

nm = nanometer(s)

atm = atmosphere(s)

mm Hg = millimeters of mercury

J, kJ = joule(s), kilojoule(s)

V = volt(s)

mol = mole(s)

ATOMIC STRUCTURE

$$E = h\nu$$

$$c = \lambda\nu$$

E = energy

ν = frequency

λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^8$ m s⁻¹

Avogadro's number = 6.022×10^{23} mol⁻¹

Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}, \text{ where } a A + b B \rightleftharpoons c C + d D$$

$$K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$= K_a \times K_b$$

$$\text{pH} = -\log [H^+], \text{ pOH} = -\log [OH^-]$$

$$14 = \text{pH} + \text{pOH}$$

$$\text{pH} = \text{p}K_a + \log \frac{[A^-]}{[HA]}$$

$$\text{p}K_a = -\log K_a, \text{ p}K_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)

K_p (gas pressures)

K_a (weak acid)

K_b (weak base)

K_w (water)

KINETICS

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant

t = time

$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{\text{total}} = P_A + P_B + P_C + \dots$$

$$n = \frac{m}{M}$$

$$K = ^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure

V = volume

T = temperature

n = number of moles

m = mass

M = molar mass

D = density

KE = kinetic energy

v = velocity

A = absorbance

a = molar absorptivity

b = path length

c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

$$= 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$$

$$= 62.36 \text{ L torr mol}^{-1} \text{ K}^{-1}$$

1 atm = 760 mm Hg

$$= 760 \text{ torr}$$

STP = 0.00°C and 1.000 atm

THERMOCHEMISTRY/ ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat

m = mass

c = specific heat capacity

T = temperature

S° = standard entropy

H° = standard enthalpy

G° = standard free energy

n = number of moles

E° = standard reduction potential

I = current (amperes)

q = charge (coulombs)

t = time (seconds)

Faraday's constant, $F = 96,485$ coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

Intermolecular Forces

Solids, Liquids, Solutions, and Phase Changes

What I Absolutely Have to Know to Survive the AP* Exam

The following might indicate that the question deals with intermolecular forces:
Boiling points; vapor pressure; melting points; network solid; crystalline solids; metallic solids; sea of electrons; delocalized electrons; triple point; sublimation; deposition; condensation; boiling; melting; freezing; intermolecular forces; vapor pressure; Coulombic interactions, etc...

Coulombic Interactions – *It's All About the Attractions*

The strength and nature of the electrostatic forces or Coulombic interactions that exist *between* particles explain the properties or behaviors of substances.

- When a substance undergoes a physical change the forces being overcome (the intermolecular interactions) are Coulombic in nature.
- When answering questions about the physical changes and properties of ionic solids always consider Coulomb's law – it is all about the charge and the distance the charges are apart.

$$U_E \propto \frac{q^+q^-}{d}$$

Properties of Matter – *Let's Get Physical*

The different properties can be explained by differences in their structures at the particulate level. These properties reflect:

- Relative orderliness of the arrangement of their particles; i.e. the spacing between their (atoms, molecules, ions)
- Their relative freedom of motion
- *Nature and strength of the interactions between them*

SOLIDS	<p>Can be crystalline, where the particles are arranged in a regular 3-D structure, or they can be amorphous, where the particles do not have a regular, orderly arrangement.</p> <ul style="list-style-type: none"> • Motion of the individual particles is limited • They do not undergo any overall translation with respect to each other. • Inter-particle interactions provide the main criteria for the structures of solids
LIQUIDS	<p>Particles in liquids are very close to each other, and they are continually moving and colliding.</p> <ul style="list-style-type: none"> • Their particles are able to undergo translation with respect to each other and their arrangement • Movement is influenced by the nature and strength of the inter-particle interactions that are present
NOTE	<p>The solid and liquid phases for a particular substance generally have relatively small differences in molar volume. In both cases their particles are very close to each other at all times. This not true for the gas phase – where we assume the particles move independently having neither a definite shape or volume – as their particles are not “bound together” by inter-particle attractions</p>

So.... What are these so called inter-particle attractions..?

Interparticle Attractions – *They're Sticky!*

Intermolecular interactions are based on the structural features of their particles. Although there are some trends in the relative strengths of these interactions, the specific structure and size of the particles involved can play a very important role in determining the overall strength of a particular intermolecular interaction.

- The properties of condensed phases and of many crucial biological structures are determined by the nature and strength of these interactions.
- Deviation from ideal gas behavior is generally a reflection of the presence of intermolecular interactions between gas particles (which we completely ignore when we consider gases as being *ideal*).

NOTE

An *interparticle attraction* or **intermolecular force (IMF)** is an attraction that occurs *BETWEEN* particles.

They are NOT bonds; bonds are known as *INTRA*particle attractions or *INTRA*molecular forces.

Without these forces there would be no liquids or solids – everything would be in the gas phase, behaving *ideally*

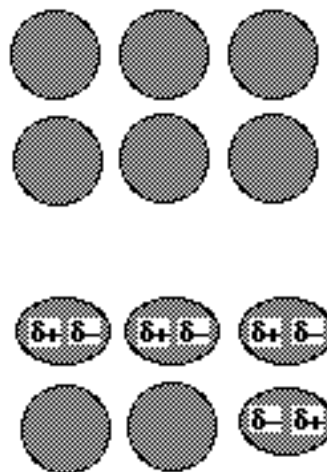
NOTE

On the AP exam, if the questions is about phase changes you need to explain your response in terms of IMFs

- Be on the look out for questions asking about melting points, vapor pressure, boiling points, etc...
- These particles have to overcome whatever stickiness is holding them together – i.e. there must be enough energy added to accomplish this

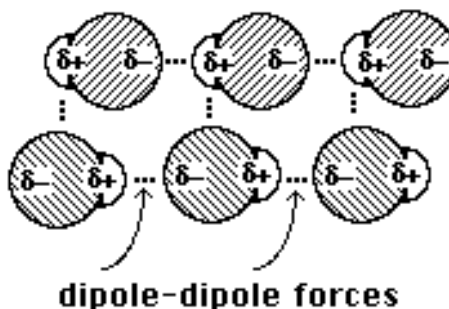
London dispersion forces

- Exist between all atoms and molecules
- Often are the strongest net IMF between large molecules
- ONLY IMF that for non-polar molecules
- In the example on the right six particles are shown in a form depicting the symmetric distribution of electron density.
- In the second group of six, some of the particles have formed instantaneous dipoles.
- The instantaneous dipoles result from an unequal distribution of electrons.
- One particle with an instantaneous dipole will affect other particles adjacent to it producing a short-range attractive interaction. The larger the particle, the more electrons, the more polarizable its electron cloud, the stronger the force of attraction, the stronger the **London Dispersion forces**.

London dispersion forces

Interparticle Attractions – *They're Sticky!* con't.**Dipole-dipole interactions**

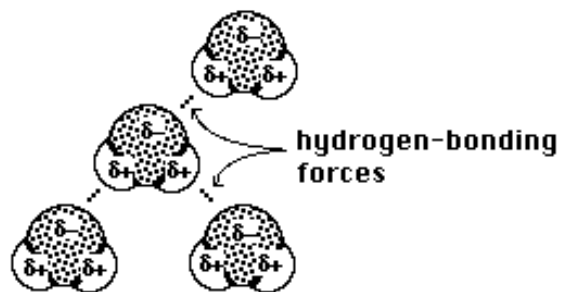
- Dipole forces result from the attraction among the positive ends and negative ends of polar molecules.
- The molecules align themselves such that the opposite poles align – they maximize the attraction and minimize the repulsion
- These dipoles result from the unequal distribution of electron density in the molecule
- Intermolecular dipole-dipole forces are weaker than ionic forces or covalent bonds.
- The larger the dipoles, the stronger the force of attraction between the two molecules, the stronger the ***Dipole-Dipole force***

**Dipole-induced dipole interactions**

- Dipole-induced dipole interactions are present between a polar and nonpolar molecule.
- The strength of these forces increases with the magnitude of the dipole of the polar molecule and with the polarizability of the nonpolar molecule

Hydrogen bonding forces (*hydrogen bonding*)

- Type of dipole force
- Hydrogen bonding is a relatively strong type of intermolecular interaction
- Exists when hydrogen atoms that are covalently bonded to the highly electronegative atoms (N, O, and F) are also attracted to the negative end of a dipole formed by the electronegative atom (N, O, and F) in a ***different*** molecule, or a ***different part*** of the same molecule (think about the complex structures of a large biomolecule – like *DNA*)
- When hydrogen bonding is present, even small molecules may have strong intermolecular attractions. Think H₂O



NOTE: Ionic interactions with dipoles are important in the solubility of ionic compounds in polar solvents. The positive and negative ions interact with the positive and negative ends of polar molecules, creating strong interactions – helping explain why NaCl is soluble in water.

WATCH OUT! When comparing ***similar sized particles***, hydrogen bonding forces > dipole forces > London dispersion forces. However, do not assume this is always the case. In substances with only London dispersion forces that have a considerably larger (thus very polarizable) electron cloud than the polar molecules, the London dispersion forces can be quite substantial and can be stronger than both Hydrogen bonding forces or dipole-dipole forces...

Solids – <i>It's All About their Types</i>	
Molecular Solids	<p>Composed of distinct, individual units of covalently bonded molecules attracted to each other through relatively weak intermolecular forces. Molecular solids:</p> <ul style="list-style-type: none"> • Are not expected to conduct electricity because their electrons are tightly held within the covalent bonds of each constituent molecule. • Generally have a low melting point because of the relatively weak intermolecular forces present between the molecules. • Sometimes composed of very large molecules, or polymers, with important commercial and biological applications. • The solid contains arrangements of atoms or molecules that are organized in an orderly three-dimensional pattern
Covalent Network Solids	<p>Consist of atoms that are covalently bonded together into a 2-dimensional or 3-dimensional network.</p> <ul style="list-style-type: none"> • Are only formed from nonmetals: elemental (diamond, graphite, silicon) or two nonmetals (silicon dioxide and silicon carbide). • The properties of covalent network solids are a reflection of their structure. • Covalent network solids have high melting points because all of the atoms are covalently bonded. • Three-dimensional covalent networks tend to be rigid and hard because the covalent bond angles are fixed. • Generally, covalent network solids form in the carbon group because of their ability to form four covalent bonds. <ul style="list-style-type: none"> • Carbon forms diamond and graphite • Silicon forms a three-dimensional network similar in geometry to a diamond. • Silicon's conductivity increases as temperature increases.
Metallic Solids	<p>Are good conductors of heat and electricity, have a wide range of melting points, and are shiny, malleable, ductile, and readily alloyed. Represented as positive kernels (or cores) consisting of the nucleus and inner electrons of each atom surrounded by a sea of mobile valence electrons.</p> <ul style="list-style-type: none"> • Metals are good conductors because the electrons are delocalized and relatively free to move. • Metals are malleable and ductile because deforming the solid does not change the environment immediately surrounding each metal core. • Metallic solids are often pure substances, but may also be mixtures called alloys.

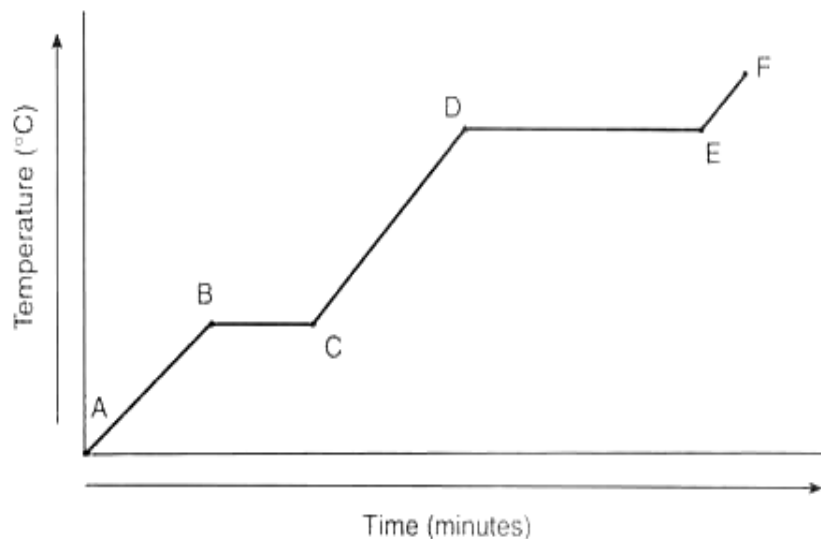
Solids – It's All About their Types <i>con't</i>	
Ionic Solids	<p>Properties are related to the strong Coulombic interactions of positive and negative ions arranged in a regular three- dimensional array.</p> <ul style="list-style-type: none"> • Generally have low vapor pressure • Tend to be brittle due to the repulsion of like charges caused when one layer slides across another layer. • Their solids do not conduct electricity. However, when ionic solids are melted, they do conduct electricity because the ions are free to move. • When ionic solids are dissolved in water, the separated ions are free to move; therefore, these solutions will conduct electricity. • The attractive force between any two ions is governed by Coulomb's law: $U_E \propto \frac{q^+q^-}{d}$ <ul style="list-style-type: none"> • For ions of a given charge, the smaller the ions, and thus the smaller the distance between ion centers, the stronger the Coulombic force of attraction, and the higher the melting point. • Ions with higher charges lead to higher Coulombic forces, and therefore higher melting points.

Liquids - All About their Properties	
Surface Tension	Molecules in the interior of a liquid are attracted by surrounding molecules in all 3– dimensions. A molecule at the surface of a liquid is attracted only by the molecules below it and on each side of it. This leads to surface tension. High surface tension indicates strong IMFs.
Capillary Action	Described by spontaneous rising of a liquid in a narrow tube. Adhesive forces between the molecules and the glass overcome the forces (IMFs) between molecules themselves. Water has a higher attraction for glass than for itself so its meniscus is inverted or concave, while Hg has a higher attraction for other Hg atoms, thus its meniscus is convex.
Boiling Point	Boiling point is the temperature at which the vapor pressure of a liquid equals the atmospheric pressure; at this point the molecules have enough energy to overcome their IMFs and enter the vapor phase. At this temperature the substance exists in equilibrium as liquid and gas particles.
Vapor Pressure	<p>Vapor pressure is the pressure resulting from particles of a substance, which exist in the gas phase above the liquid (or solid). The weaker the IMFs, the higher the vapor pressure. <i>Why?</i></p> <ul style="list-style-type: none"> • The substance can more easily overcome those IMFs and break away into the gas phase, increasing the number of molecules in the gas phase (at that temperature), thus increasing the pressure above the liquid. • Increasing the temperature on the substance will increase the vapor pressure; again due to more molecules being able to overcome their IMFs and move into the gas phase.

Heating and Cooling Curves – All About the Change

Graphically represents the relationship of a pure substance in terms of how the temperature changes over time...

Different substances have different phase change points (melting/freezing and vaporization/condensation) but the shapes of their heating and/or cooling curves are very similar.



A to B represents the substance as a solid ($q = mC\Delta T$)

B to C represents the process of melting/freezing: at this point the energy is overcoming the IMFs between the particles – thus the temperature doesn't change ($q = \Delta H_{\text{fusion}}$)

C to D represents the substance as a liquid ($q = mC\Delta T$)

D to E represents the process of vaporization/condensation: at this point the energy is further overcoming the IMFs between the particles – thus the temperature doesn't change ($q = \Delta H_{\text{vap}}$)

E to F (and beyond) represents the substance in the vapor/gas phase ($q = mC\Delta T$)

Remember on the slopes the substance is all in the same phase... $q = mC\Delta T$

During the phase changes (flat lines) the energy is being used to overcome IMFs (increasing potential energy), not to increase the KE of the particles...

Solution Formation – All About those IMFs

Solutions are homogenous mixtures in which the physical properties are dependent on the concentration of the solute and the strengths of all interactions among the particles of the solutes and solvent.

Substances with similar intermolecular interactions tend to be miscible or soluble in one another.

Solutions come in the form of solids, liquids, and gases

- For liquid solutions, the solute may be a gas, a liquid, or a solid.

Liquid solutions exhibit several general properties:

- The components cannot be separated using filter paper.
- There are no components large enough to scatter visible light.
- The components can be separated using processes that are a result of the intermolecular interactions between and among the components.
 - Chromatography (paper and column) separates chemical species by taking advantage of the differential strength of intermolecular interactions between and among the components.
 - Distillation is used to separate chemical species by taking advantage of the differential strength of intermolecular interactions between and among the components and the effects these interactions have on the vapor pressures of the components in the mixture.
- The formation of a solution may be an exothermic or endothermic process, depending on the relative strengths of intermolecular/inter-particle interactions before and after the dissolution process.

CAUTION Never use “like dissolves like” on the AP exam. EXPLAIN in terms of IMFs and energy....

Solution Formation:

Particles exhibit a number of inter-particle interactions (London dispersion, hydrogen bonding, dipole-dipole, ion-dipole, etc...) and it is these interactions that either PROMOTE or PREVENT solution formation.

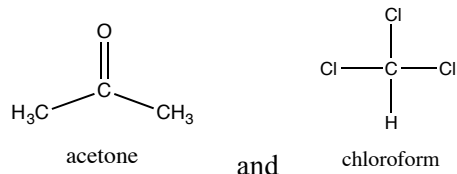
In order to dissolve a substance must...

Overcome (requires energy) *both* solute-solute and solvent-solvent interactions

Form solute-solvent interactions upon mixing (releases energy)

A solution WILL form when the energy released due to the solute-solvent interactions is the same as (or more than) the energy needed overcome the solute-solute and the solvent-solvent interactions.

Example: dissolving acetone in chloroform

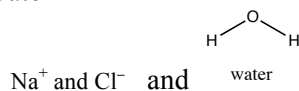


Solute-Solute: acetone forms dipole-dipole interactions between their particles

Solvent-Solvent: chloroform forms dipole-dipole interactions between their particles

Solute-Solvent: when mixed together dipole-dipole interactions form between the acetone and chloroform particles

Example: dissolving sodium chloride in water



Solute-Solute: positive sodium and negative chloride ions form ionic interactions between their particles

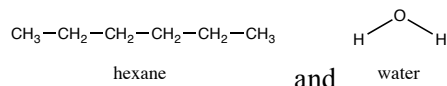
Solvent-Solvent: water forms hydrogen bonding interactions between their particles

Solute-Solvent: when mixed together ion-dipole interactions form between the positive and negative ions of the salt and the polar water particles

Solution Formation – All About those IMFs con't

A solution MAY or MAY NOT form when the energy released due to the solute-solvent interactions is less than the energy needed overcome the solute-solute and the solvent-solvent interactions. Whether the solution forms or not depends on the disparity between the two.

Example: dissolving hexane in water



- Solute-Solute: hexane molecules form London dispersion interactions between their particles
- Solvent-Solvent: water forms hydrogen bonding interactions between their particles
- Solute-Solvent: when mixed together polar water molecules can only interact with the non-polar hexane molecules through London dispersion forces

Thus a solution does NOT form.

Although the tendency to mix is strong (increased entropy) the large energy disparity between overcoming the hydrogen bonding forces in water and the relatively small energy gain from forming weak interactions between water and hexane is too large.

IMPORTANT

Melting points, vapor pressure, boiling points (really all phase change processes) are all about the strength of the intermolecular attractive forces. The physical change that accompanies any of these processes will require particles to overcome the attractive forces holding them together.

REMEMBER: If you are asked to compare 2 substances (higher boiling point, lower vapor pressure, which melts first, etc...) it's Coulombic! Write about Coulomb's Law for ionic substances and IMFs for molecular compounds

Connections to Other Chapters

Periodicity

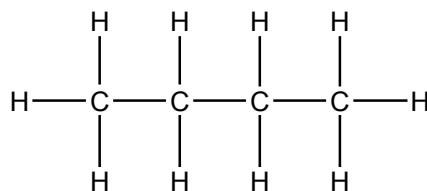
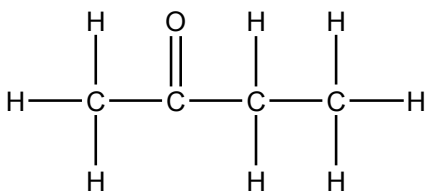
Bonding

Lewis Structures and Geometry

NMSI SUPER PROBLEM

Answer the following questions using your knowledge of intermolecular forces and molecular structure. Your response must include specific information about all substances in each question.

- A. The structures for butanone, $\text{CH}_3\text{COCH}_2\text{CH}_3$, and *n*-butane, $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$, are shown below.



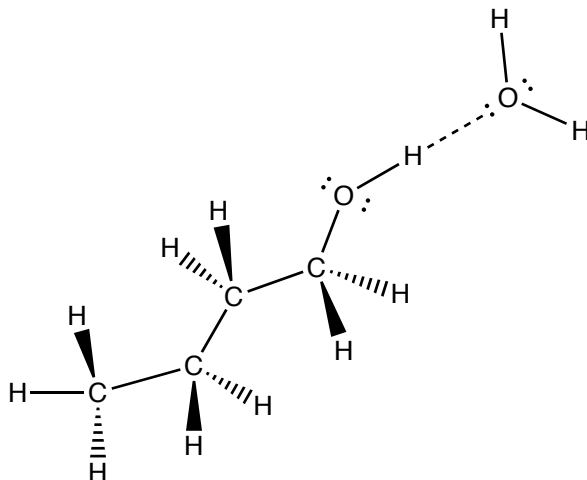
Identify the type of intermolecular forces in

(i) Butanone

(ii) *n*-butane

- B. Butanone is much more soluble in H_2O than is *n*-butane. Account for this difference.

- C. The substance *n*-butanol ($\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$) is shown below forming a hydrogen bond with a water molecule, as represented with a dashed line. There are 2 more locations where a water molecule would hydrogen bond with *n*-butanol. Draw both water molecules in their correct orientation where they would hydrogen bond with the *n*-butanol molecule.



- D. Predict whether the enthalpy of vaporization, $\Delta H_{\text{vap}}^{\circ}$, for *n*-butanol will be greater than 32.2 kJ mol^{-1} ; less than 21.0 kJ mol^{-1} ; or between 21.0 kJ mol^{-1} and 32.2 kJ mol^{-1} . Explain

Substance	$\Delta H_{\text{vap}}^{\circ}$ (kJ mol^{-1})
<i>n</i> -butane ($\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$)	21.0
Butanone ($\text{CH}_3\text{COCH}_2\text{CH}_3$)	32.2
<i>n</i> -butanol ($\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$)	?

- E. Consider the three chloride compounds and the information in the table below.

Substance	Melting Point
PCl ₃	-93.9°C
KCl	776°C
NaCl	801°C

Account for the difference in melting points.

The table below provides information about three of the noble gases.

Substance	Boiling Point (K)	Solubility in Water (cm ³ kg ⁻¹)
Ne	27.3	10.5
Ar	?	?
Xe	166.6	108.1

- F. Neon has a much lower boiling point than xenon. Explain.
- G. Xenon is much more soluble in water than neon. Explain.
- H. Would argon's solubility in water be greater than or less than that of xenon?
- I. Give samples of liquid argon and liquid xenon – in separate, identical, closed containers at the same temperature – which would have the greatest vapor pressure?

AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS

Modified for a Short Response

Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.

- (a) At 25°C and 1 atm, F₂ is a gas, whereas I₂ is a solid.
- (b) Ammonia, NH₃, is very soluble in water, whereas phosphine, PH₃, is only moderately soluble in water.

AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS

Account for each of the following observations about pairs of substances. In your answers, use appropriate principles of chemical bonding and/or intermolecular forces. In each part, your answer must include references to both substances.

- (a) Even though NH₃ and CH₄ have similar molecular masses, NH₃ has a much higher normal boiling point (−33°C) than CH₄ (−164°C).
- (b) At 25°C and 1.0 atm, ethane (C₂H₆) is a gas and hexane (C₆H₁₄) is a liquid.
- (c) Si melts at a much higher temperature (1,410°C) than Cl₂ (−101°C).
- (d) MgO melts at a much higher temperature (2,852°C) than NaF (993°C).

2016 AP Chemistry - Intermolecular Forces
AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS

Modified for a Short Response

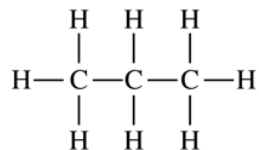
Use the information in the following table to answer the questions below.

Name	Lewis Electron-Dot Diagram	Boiling Point (°C)	Vapor Pressure at 20°C (mm Hg)
Dichloromethane	$\begin{array}{c} \text{H} \\ \cdot\cdot \\ \cdot\cdot \\ \text{:}\ddot{\text{C}}\text{:}\ddot{\text{C}}\text{:}\text{H} \\ \cdot\cdot \\ \cdot\cdot \\ \text{:}\ddot{\text{C}}\text{:} \\ \cdot\cdot \\ \cdot\cdot \end{array}$	39.6	353
Carbon tetrachloride	$\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{:}\ddot{\text{C}}\text{:} \\ \cdot\cdot \\ \cdot\cdot \\ \text{:}\ddot{\text{C}}\text{:}\ddot{\text{C}}\text{:} \\ \cdot\cdot \\ \cdot\cdot \\ \text{:}\ddot{\text{C}}\text{:} \\ \cdot\cdot \\ \cdot\cdot \end{array}$	76.7	89

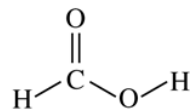
- (a) Dichloromethane has a greater solubility in water than carbon tetrachloride has. Account for this observation in terms of the intermolecular forces between each of the solutes and water.
- (b) In terms of intermolecular forces, explain why dichloromethane has a higher vapor pressure than carbon tetrachloride.

AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS
Modified for a Short Response

Use principles of atomic structure, bonding, and intermolecular forces to answer the following questions. Your responses must include specific information about all substances referred to in each part.



Propane



Methanoic Acid

- (a) The complete structural formulas of propane, C_3H_8 , and methanoic acid, HCOOH , are shown above. In the table below, write the type(s) of intermolecular attractive force(s) that occur in each substance.

Substance	Boiling Point	Intermolecular Attractive Force(s)
Propane	229 K	
Methanoic acid	374 K	

- (b) Use principles of intermolecular attractive forces to explain why methanoic acid has a higher boiling point than propane.