Supplemental Activities

Module: States of Matter

Section: Intermolecular Forces - Key

Electrostatic Forces

Αςτινιτγ 1

The purpose of this activity is to practice recognizing the nature of the forces important on the molecular level.

1. The electrostatic force depends on two things. What are they?

Charge and distance.

2. Describe the relationship between electrostatic force and the two variables you listed above.

Electrostatic force depends directly on charge. The greater the charge, the larger the force. Electrostatic force depends inversely on distance. The greater the distance, the smaller the force.



ΑCTIVITY2

The purpose of this activity is to remind you of the scale of electrostatic forces.

1. True or false? Electrostatic forces cannot be observed on the macroscopic scale because they are so small. Give real world evidence to back up your claim.

False. Electrostatic forces are important everywhere in our daily lives. The Scotch tape experiment (see guided practice) shows just how easily we can influence and observe electrostatic forces. Also, consider static cling! Electrostatic forces in action.

Ion-Ion forces

ΑCTIVITY 1

The purpose of this activity is to recognize and predict strengths of ion – ion forces.

1. Give an example of two ions that have the same charge but different charge densities.

Charge density is the magnitude of charge contained in a certain volume. So a small -2 ion such as O^{2-} will have a greater charge density than a large -2 ion such as CO_3^{2-} .

- 2. <u>Lattice</u> energy is the amount of energy required to completely ionize <u>one</u> mole of a solid <u>crystal</u> of an ionic substance.
- 3. What microscopic qualities would you predict for an ionic substance that has a relatively low melting point (compared to other ionic compounds)?

Ionic substances are held together by electrostatic forces in a lattice structure. Electrostatic forces depend on both charge and distance. So a cation and an anion with low charges would have weaker attraction. Also, bigger ions would have weaker attraction. We can summarize these two ideas by saying ions with low charge densities would likely combine to create an ionic substance with a low melting point.

ΑCTIVITY2

The purpose of this activity is to practice your mastery of predicting strengths of ion-ion forces.

- 1. What is the charge on the following ions in the following ionic compounds?
 - a. KI +1 on K and -1 on I
 - b. $(NH_4)_2S +1$ on NH_4 and -2 on S
 - c. BaS +2 on Ba and –2 on S
 - d. $K_2O +1$ on K and -2 on O
 - e. $NaMnO_4 +1$ on Na and -1 on MnO_4
 - f. $KNO_2 +1$ on K and -1 on NO_2
 - g. $TiCl_4 +4$ on Ti and -1 on Cl
- 2. Rank the following in order of increasing lattice energy:
 - a. LiCl, LiI, LiF, LiBr

LiI < LiBr < LiCl < LiF

Here I- is the largest anion and F- is the smallest anion. A smaller distance between ions leads to a greater lattice energy.

b. AlF_3 , Al_2O_3

 $AlF_3 < Al_2O_3$

Here O²⁻ has twice the charge of F⁻. The two ions are relatively close in size so the charge difference has a greater affect. A greater charge between two ions leads to a greater lattice energy.

c. KF, CsF, LiF, NaF

CsF < KF < NaF < LiF

Here Cs⁺ is the largest cation and Li⁺ is the smallest cation. A smaller distance between ions leads to a greater lattice energy.

Dipole-Dipole Forces

ΑCTIVITY 1

The purpose of this activity is to practice recognizing when condensed substances that have dipoledipole interactions

1. What kind of molecules experience dipole-dipole forces in their condensed phases?

Polar molecules experience dipole-dipole forces.

2. Explain the dipole-dipole force on a microscopic scale.

Dipole-dipole forces exist when the permanent dipole of one polar molecule aligns with the permanent dipole of a neighboring polar molecule. The slightly negative end of one polar molecule aligns to the slightly positive end of the neighboring polar molecule.

ACTIVITY 2

The purpose of this activity is to test your ability to predict the presence of dipole-dipole forces.

- 1. Which of the following molecules would have dipole-dipole forces in the condensed phase?
- NH₃ BH₃ H₂O CO₂ PF₃ CF₄ SeF₄

In order to properly assess the polarity of these molecules, we need to draw their Lewis structures.

Compounds that do not exhibit dipole-dipole forces:



 BH_3 has no net dipole moment. It is non-polar and **does not** have dipole-dipole forces in the condensed phase.

 CO_2 has no net dipole moment. It is non-polar and **does not** have dipole-dipole forces in the condensed phase.



CF₄ has no net dipole moment. It is non-polar and **does not** have dipole-dipole forces in the condensed phase.

Compounds that do exhibit dipole-dipole forces:

 $\rm NH_3$ has a net dipole moment. It is polar and **does** have dipole-dipole forces in the condensed phase.



 H_2O has a net dipole moment. It is polar and **does** have dipole-dipole forces in the condensed phase.



 PH_3 has a net dipole moment. It is polar and **does** have dipole-dipole forces in the condensed phase.



 SeF_4 has a net dipole moment. It is polar and ${\bf does}$ have dipole-dipole forces in the condensed phase.

ΑCTIVITY **3**

The purpose of this activity is to practice recognizing a very important and special type of dipoledipole force called hydrogen bonding.

- <u>Hydrogen</u> bonding occurs between molecules where hydrogen is covalently to <u>fluorine</u>, <u>oxygen</u> or <u>nitrogen</u>. These bonds have high <u>electronegativity</u> and pull most of the electron density from the hydrogen atom. The hydrogen atom can then get <u>close</u> to the negative areas of neighboring molecules.
- 2. Which kind of intermolecular force is generally stronger: dipole-dipole forces or hydrogen bonds? Why?

Hydrogen bonds are generally stronger because the highly electronegative F, O, or N atoms have pulled so much of the electron density from the hydrogen atom that it is very small. Its small size allows it to get extremely close to the F, O, or N atoms of neighboring hydrogen bonding molecules. A short distance between oppositely charged particles corresponds to high levels of attraction.

3. Is a hydrogen bond between two molecules a type of covalent bond? How do we often represent hydrogen bonds in a drawing?

No hydrogen bonds are not covalent bonds, they are electrostatic interactions between two molecules. Covalent bonds occur when electrons are shared between two atoms. However, the two molecules experiencing hydrogen bonding do not share electrons across that interaction. We usually represent hydrogen bonds as small dashed lines between the relevant hydrogen atom of one molecule and the relevant F, O or N atom of the other molecule.

4. Circle the compounds below that have H-bonding in the condensed phase. For those that do have H-bonding, draw a molecular view of the condensed phase and indicate the H-bonding with dashed line.

 $\begin{array}{ccc} NF_3 & NH_3 & C_2H_4O & C_2H_4NH_2COOH \ (keep \ C \ in \ a \ chain) & CH_2O \\ C_3H_5OOH & & & \\ \end{array}$

In order to assess the presence of hydrogen bonding we must first draw the Lewis structures. Then, we must look for H–F, H–O and/or H–N bonds.

<u>Compounds that **do not** exhibit hydrogen bonding</u>:



Compounds that **do** exhibit hydrogen bonding drawn in their condensed phases:

(The dark pink dashed lines represent the hydrogen bonding)





Dispersion Forces

ΑCTIVITY 1

The purpose of this activity is to practice recognizing that IMFs exist even when there is no fixed charge or permanent dipole moment.

- 1. van der Waals forces are also called <u>dispersion</u> forces, <u>London dispersion</u> forces and <u>induced dipole-dipole</u> forces.
- 2. True or false? Dispersion forces exist only in non-polar condensed phase matter. Explain your answer.

False. Dispersion forces exist in all condensed phase matter. Dispersion forces result from the temporary induced dipoles from the movement of electrons in a molecule. These induced dipoles occur in every molecule regardless of their other intermolecular forces.

3. How do size and shape influence dispersion forces? Is it possible to have a non-polar molecule with a higher boiling point than a molecule with hydrogen bonding capabilities?

Dispersion forces are very distance dependent. A molecule with high surface areas and long-chain structure can pack in very tightly together which enhances the dispersion force affect. So yes, there are cases where a molecule with hydrogen bonding capability still has a lower boiling point than a molecule with only dispersion forces.

ΑCTIVITY **2**

The purpose of this activity is to practice more problems about IMFs

1. Circle the compounds in which the only IMF is the dispersion force:

PCl ₅	SCl_4	H_2O	02	CH_2I_2	C_6H_6
(benzene)					

Compounds in which the only IMF is the dispersion force must be non-polar. In order to properly assess the polarity of these molecules, we need to draw their Lewis structures.

Compounds that **do** exhibit only dispersion forces:



$\dot{o} = \dot{o}$



<u>Compounds that **do not** exhibit only dispersion forces:</u>



SCl₄ exhibits dispersion forces and dipole-dipole forces.

H₂O exhibits dispersion forces, dipole-dipole forces and hydrogen bonding.



CH₂I₂ exhibits dispersion forces and dipole-dipole forces.

2. Consider the following boiling point data.

Substance	Atomic/Molecular Weight (g/mol)	Boiling Point (°C)
Не	4	–269°C
Ne	20	-246°C
Ar	40	–186°C
Kr	83.8	–152°C
Хе	131.3	–107°C

Rn	222	-62°C
F ₂	38	–188°C
HF	20	+19.5°C
Cl ₂	70.9	−34.1°C
HCl	36.46	-85.05°C
Br ₂	159.8	+59.4°C
HBr	80.9	−66.8°C
I ₂	253.8	+185°C
HI	129	–35.36°C

3. The boiling points of the noble gases increases as you go down the group. Please explain.

As you go down the group from helium to radon, the size of the atoms increases as we can observe through the increasing atomic weights. Therefore, the polarizibility of the atoms also increases. All of these atoms are non-polar because they are just monatomic samples. So dispersion forces are the only intermolecular force at play. As the size and therefore polarizibility of the atoms increases, the strength of the dispersion forces increases and thus the boiling points increase.

4. Ne and HF both have a molecular weight of 20 g/mol but their boiling points are different. Please explain.

HF has hydrogen bonding capabilities while Ne is a non-polar monatomic sample. HF is polar and has hydrogen bonding which explains its much higher boiling point.

5. The boiling point of F_2 is lower than the boiling point of HF. Is this the typical pattern for the other halogens? Please explain.

This is not the typical pattern for the other halogens. For the other halogens the X_2 form has a higher boiling point than the HX form. Yet, the X_2 forms are non-polar and the HX forms are polar! The bulky X_2 forms have greater polarizibility of compared to the smaller, more compact HX forms. Greater polarizibility causes stronger dispersion forces which causes higher boiling points. HF, however, has hydrogen bonding capabilities and dipole-dipole forces. These IMFs are stronger than the dispersion forces in F_2 . Therefore HF has a higher boiling point than F_2 .

6. The boiling point of I_2 is greater than the boiling of HF. Please explain.

Even though HF has hydrogen bonding capabilities and dipole-dipole interactions, it is still much smaller and more compact than I_2 . So, even though I_2 only experiences dispersion forces, its polarizibility is so great that its dispersion forces cause a higher boiling point that the polar HF molecule.