Chemistry: Form WS4.1.1A

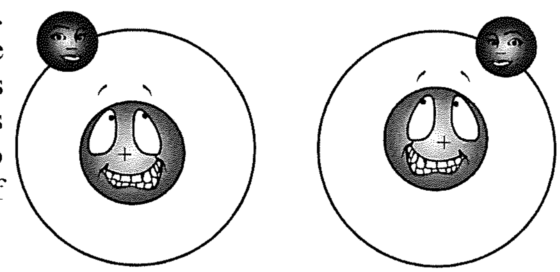
BONDING

Name ______________________

Date ______________ Period ____

How Bonds Form

The electrons of one atom are attracted to the protons of another. When atoms combine, there is a tug of war over the valence electrons. The combining atoms either lose, gain, or share electrons in such a way that they complete their outer shells. Whether atoms gain, lose, or share electrons depends how tightly they hold onto their own electrons and how strongly they pull on the electrons of another atom.



Answer the questions below based on the information above and on your knowledge of chemistry.

1. What is the charge on a proton? ______________________

2. What is the charge on an electron? ______________________

3. Why do an atom's electrons revolve around its protons instead of drifting away? ______________________

4. Why are the electrons of one atom attracted to the protons of another? ______________________

5. What happens when two atoms get near each other that causes them to bond? ______________________

6. How are the elements sodium and chlorine classified? ______________________

7. What would happen during a tug of war between sodium and chlorine over each others outer electrons? Why?

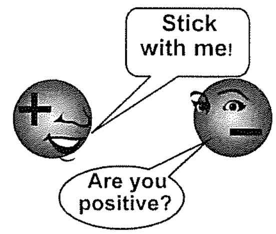
8. How do sodium and chlorine combine? ______________________

Chemistry: Form WS4.1.2A | Name ______________________

BONDING | Date ______________ Period ____

Ionic Bonds

Ionic bonds are caused by the attraction between oppositely charged ions. Ions form as follows: The electrons of one atom are attracted to the protons of another. Metals hold onto electrons loosely while nonmetals hold onto electrons tightly. As a result, metals lose electrons and nonmetals gain electrons in such a way that they complete their outer shells. Atoms that gain or lose electrons become electrically charged. Metals become positively charged ions by losing electrons. Nonmetals become negatively charged ions by gaining electrons. Metal cations and nonmetal anions become ionically bonded because they are oppositely charged.



ION TALK

Answer the questions below based on your understanding of ionic bonds.

1. Draw Bohr-Rutherford diagrams of sodium and chlorine atoms showing the number of protons and neutrons, and the arrangement of electrons.

2. What will happen to sodium and chlorine when they combine (*HINT: Remember how metals and nonmetals combine.)* ______________________

3. Draw Bohr-Rutherford diagrams of sodium and chlorine atoms showing the changes in the arrangement of electrons after they combine.

4. What are the charges on the sodium ion and the chloride ion after they combine? (*HINT: Count the number of protons and electrons of each.)* ______________________

5. What are the oxidation states of sodium and chlorine? ______________________

6. Why do sodium and chlorine become bonded? ______________________

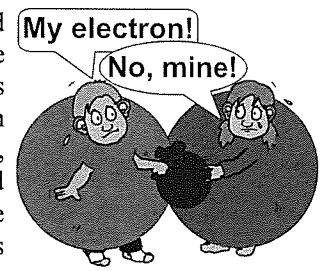
____

7. What is the total charge on a compound of sodium and chlorine? ______________________

Chemistry: Form WS4.1.4A Name ______________________

BONDING Date ______________ Period ____

Covalent Bonds



Covalent bonds are bonds formed by sharing electrons. The electrons of one atom are attracted to the protons of another, but neither atom pulls strongly enough to remove an electron from the other. Covalent bonds form when the electronegativity difference between the elements is less than 1.7 (see the Electronegativity table on the back of the Periodic Table) or when hydrogen behaves like a metal. When a covalent bond forms, no valence electrons are transferred, rather, they are shared. If the electronegativity difference is zero, the electrons are shared equally and the bond is nonpolar. If the electronegativity difference is greater than 0.4 but less than 1.7, the electrons are displaced towards the more electronegative element (nonmetal) and the bond is polar. In a covalent bond, unpaired valence electrons pair up in such a way that the atoms complete their outer shells.

<u>Electron Dot Diagrams Showing Unpaired Valence Electrons</u> (*NOTE*: When bonding occurs, molecular orbitals form. As a result, the two electrons that are normally paired in the lowest energy orbital move into separate orbitals)

Li· Be· ·B· ·C·

·N: ·O: :F: :Ne:

<u>Pairing Electrons</u>:

Nonpolar Covalent Bond: $Cl^0 + Cl^0 \rightarrow Cl_2$:Cl· + ·Cl: → :Cl:Cl:

Polar Covalent Bond: $H^+ + Cl^- \rightarrow HCl$ H· + ·Cl: → H:Cl:

Based on your understanding of covalent bonds, answer the questions below.

1. Draw electron dot diagrams for hydrogen and oxygen.

2. Draw electron dot diagrams showing the pairing of electrons to form water from hydrogen and oxygen. All outer shells should be complete.

3. Are the bonds in water polar or nonpolar. How do you know? ______________________

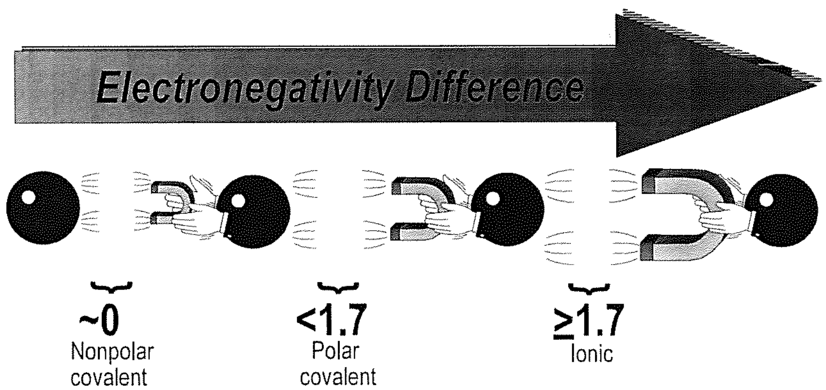
____

Chemistry: Form WS4.1.5A Name ______________________

PERIODIC TABLE AND BONDING Date ______________ Period ____

Bond Type

When atoms combine, there is a tug of war over their valence electrons. The type of bond that forms depends on the outcome of the tug of war. The outcome of the tug of war is determined by the relative strengths of the forces exerted by the atoms. The electronegativity provides a measure of those forces. When the electronegativity difference is greater than or equal to 1.7, the atom with the greater electronegativity gains the electron, and an **ionic bond** is formed. Electronegativity differences below 1.7 result in covalent bonds or sharing. If the electronegativity difference is close to zero (<0.4), the atoms share equally and a **nonpolar bond** forms. Higher electronegativity differences (still below 1.7) result in unequal sharing or **polar bonds**.



Fill in the table below by looking up the electronegativities of the elements in each compound. Determine the electronegativity difference and the bond type.

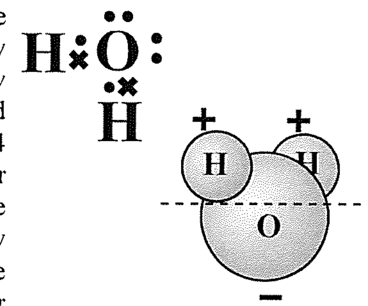
Compound	Electronegativity		Electronegativity Difference	**Bond Type** Ionic, Polar covalent, Nonpolar covalent
	Metal (low)	**Nonmetal (high)**		
Example: NaBr	0.9	3.0	2.1	ionic
HCl				
H_2Te				
KI				
SO_2				
H_2O				
CS_2				
N_2O_5				
MgO				

Chemistry: Form WS4.1.6A Name ______________________

BONDING Date ______________ Period ____

Recognizing Polar Molecules

To determine if a compound is polar, you must consider the electronegativity difference within each bond and the three dimensional shape of the compound. If the electronegativity difference is greater than 1.7 or close to zero, the compound is not polar. Electronegativity differences above 1.7 are found in ionic compounds. Electronegativity differences around zero are found in molecules with nonpolar bonds. Electronegativity differences between 0.4 and 1.7 are found in molecules with polar bonds. These molecules can be polar or nonpolar depending on their shapes. Molecules with polar bonds distributed symmetrically are nonpolar. Asymmetrical molecules with polar bonds are polar. Water is polar. An imaginary line can be drawn through a water molecule separating the positive pole from the negative pole. This is because the charges are distributed asymmetrically. Carbon dioxide is nonpolar because the electronegative oxygens are distributed symmetrically around the carbon. (O=C=O)



Water is polar, because the charges are distributed asymmetrically. The electropositive hydrogens are attached to oxygen's two unpaired electrons..

Determine if each of the compounds listed below, IONIC, POLAR, or NONPOLAR as follows: [1] determine the types of bonds. [2] draw electron dot diagrams to determine the shape.

Compound	Type of Bond: IONIC, POLAR, or NONPOLAR	Electron Dot Diagram	Type of Compound: IONIC, POLAR, or NONPOLAR	Compound	Type of Bond: IONIC, POLAR, or NONPOLAR	Electron Dot Diagram	Type of Compound: IONIC, POLAR, or NONPOLAR
HCl				CCl_4			
CH_4				CH_3Cl			
Cl_2				N_2			
KBr				H_2S			
NH_3				NaBr			

Name ______________________ Class______ Date______________

Activity 4-6
The Chemical Bond III

Polar bonds and polar molecules

1. How does a polar bond differ from a nonpolar bond? ______________________

____

2. How does a polar bond differ from an ionic bond? ______________________

____

____

3. How is electronegativity difference used to help predict bond type? What values separate ionic from polar covalent bonds? ______________________

____

____

4. What is a dipole (polar molecule)? ______________________

____

____

5. How do polar bonds contribute to the polarity of a molecule? ______________________

____

____

6. How can a molecule, such as CO_2 or CH_4, contain polar bonds yet still be a non-polar substance? ______________________

____

____

7. What physical properties are characteristic of dipoles?______________________

____

____

8. Why does water dissolve many ionic compounds? ______________________

____

____

Network solids

9. Describe the bonding in network solids. ______________________

____

10. What are the significant physical properties of network solids? ______________________

____

Properties of ionic solids

11. Ionic solids have relatively ____________ (high/low) melting points.
12. Describe two different conditions under which the ions of ionic solids become free to move.

____

____

13. Describe the electrical conductivity of ionic substances in the solid, liquid, and aqueous solution phases. ____

____

____

14. What two kinds of elements are most likely to react with each other to form binary ionic compounds? ____

The metallic bond

15. Describe bonding in metallic solids. ____

____

16. What are the significant physical properties of metallic solids? ____

____

Hydrogen bonding

17. Draw a diagram to illustrate hydrogen bonding between molecules of HF.

18. Under what circumstances do hydrogen bonds form? ____

____

19. What properties are associated with compounds containing hydrogen bonds? ____________

____

____

Van der Waals forces

20. What is the source of van der Waals forces? ____

____

____

21. What factors determine the magnitude of the van der Waals forces acting between molecules?

____

____

22. What properties of molecules are associated with van der Waals forces? ____________

____

____

TYPES OF CHEMICAL BONDS

Name ____________________

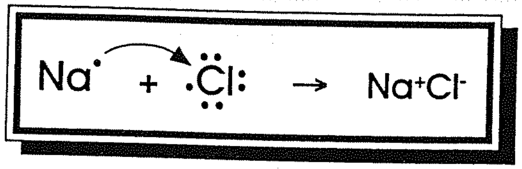
Classify the following compounds as ionic (metal + nonmetal), covalent (nonmetal + nonmetal) or both (compound containing a polyatomic ion).

1. $CaCl_2$ ____________________
2. CO_2 ____________________
3. H_2O ____________________
4. $BaSO_4$ ____________________
5. K_2O ____________________
6. NaF ____________________
7. Na_2CO_3 ____________________
8. CH_4 ____________________
9. SO_3 ____________________
10. LiBr ____________________
11. MgO ____________________
12. NH_4Cl ____________________
13. HCl ____________________
14. KI ____________________
15. NaOH ____________________
16. NO_2 ____________________
17. $AlPO_4$ ____________________
18. $FeCl_3$ ____________________
19. P_2O_5 ____________________
20. N_2O_3 ____________________

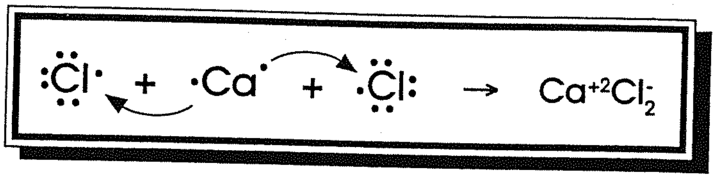
IONIC BONDING

Name ______________________

Ionic bonding occurs when a metal transfers one or more electrons to a nonmetal in an effort to attain a stable octet of electrons. For example, the transfer of an electron from sodium to chlorine can be depicted by a Lewis dot diagram.



Calcium would need two chlorine atoms to get rid of its two valence electrons.



Show the transfer of electrons in the following combinations.

1. K + F

2. Mg + I

3. Be + S

4. Na + O

5. Al + Br

NAME:______________________________ DATE:________________

Hey, Who Took My Electron?

Background:

When atoms unite, attractive forces tend to pull the atoms together. These attractive forces are called chemical bonds, often referred to in shortened form as bonds. When a chemical bond forms, energy is released. When a chemical bond breaks, energy is absorbed. Hence, when two atoms are held together by a chemical bond, the atoms are at a lower energy condition than when they are separated.

In 1916, the American chemist Gilbert Newton Lewis proposed that chemical bonds are formed between atoms because electrons from the atoms interact with each other. Lewis had observed that many elements are most stable when they contain eight electrons in their valence shell. He suggested that atoms with fewer than eight valence electrons bond together to share electrons and complete their valence shells.

While some of Lewis' predictions have since been proven incorrect, his work established the basis of what is known today about chemical bonding. We now know that there are two main types of chemical bonding; ionic bonding and covalent bonding.

In ionic bonding, electrons are completely transferred from one atom to another. In the process of either losing or gaining negatively charged electrons, the reacting atoms form ions. The oppositely charged ions are attracted to each other by electrostatic forces, which are the basis of the ionic bond.

Procedure:

1) Give the electron configuration and Lewis dot diagram for each atom. Use dots for one set of valence electrons and ×s for the other element. You may also use two different colors to differentiate between the electrons form each element.
2) Determine how many electrons will be lost and gained.
3) Determine if any additional atoms are needed to make an even exchange of electrons.
4) Use arrows to show the electrons being transferred.
5) Give the electron configuration, Lewis dot diagram and charge for each ion formed.

Problems:

1) K and F

2) Ba and O

3) Mg and Cl

4) Al and Br

5) K and O

6) Cs and N

7) Al and O

8) Ca and P

9) Sc and I

Reflection:

What is the driving force behind bonding?

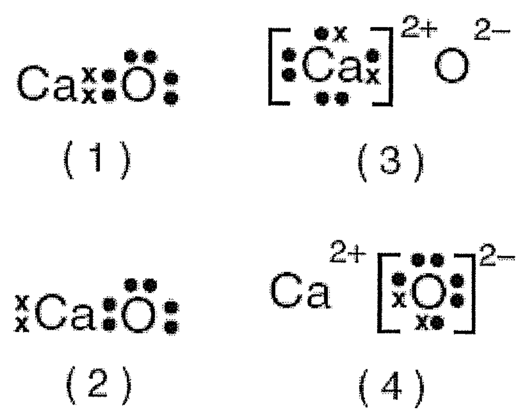
Questions:

1) What happens to the positive and negative ions created in the process of ionic bonding? Why?

2) A Ba^{+2} ion differs from a Ba^{0} atom in that the ion has
(1) More electrons
(2) More protons
(3) Less electrons
(4) Less protons

3) Which Lewis electron-dot diagram to the right represents calcium oxide?
(1) 1
(2) 2
(3) 3
(4) 4

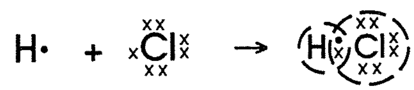
Explain each of the other three choices is wrong.



COVALENT BONDING

Name ______________________

Covalent bonding occurs when two or more nonmetals share electrons, attempting to attain a stable octet of electrons at least part of the time. For example:



Note that hydrogen is content with 2, not 8, electrons.

Show how covalent bonding occurs in each of the following pairs of atoms. Atoms may share one, two or three pairs of electrons.

1. H + H (H_2)

2. F + F (F_2)

3. O + O (O_2)

4. N + N (N_2)

5. C + O (CO_2)

6. H + O (H_2O)

Name ____________________ Class______ Date__________

Activity 4-5

The Chemical Bond II

Covalent bonds

1. What role do valence electrons play in covalent bonding? ____________________

____

____

2. When atoms are bonded together covalently, what two kinds of structures may result?

__________, __________

3. What is a single covalent bond? ____________________

____

____

4. What is a double covalent bond? ____________________

____

____

5. What is a triple covalent bond? ____________________

____

____

6. How is a coordinate covalent bond different from an ordinary covalent bond? __________

____

____

7. What kind of compound frequently shows coordinate covalent bonds? __________

____

Dot diagrams for molecules and polyatomic ions

8. Choose words from the word list to fill in the blanks in the following paragraphs relating to the construction of dot diagrams. The list groups words that have contrasting or related meanings.

Word List

atom(s)/ion(s)/molecule(s)	metal/nonmetal
eight/four	O(oxygen)
error	pairs
kernel/valence	share/transfer

When atoms are held together by covalent bonds, a __________ or a polyatomic __________ is formed. The electron-dot symbols for individual atoms can be used to construct dot diagrams for molecules and polyatomic ________________. The symbol for each element represents the nucleus and ________________ electrons. When atoms form covalent (or coordinate covalent) bonds, each atom must share enough electrons to fill its ________________ shell with at least a share in the total of ________________ electrons, that is, ________________ pairs of electrons.

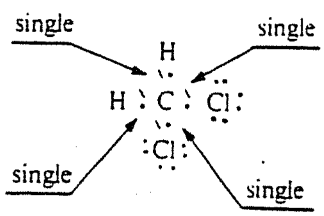
Constructing dot diagrams becomes a trial and ________________ process until a reasonable structure is drawn. The following suggestions will help reduce the number of trials and errors.

a. Choose a central atom, generally a(n) ________________ other than H or O, which is bonded to not more than ________________ other atoms.

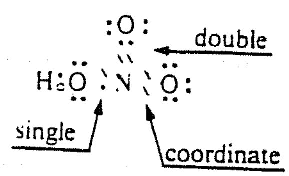
b. In ternary compounds, H atoms are generally bonded to ________________ atoms.

c. Arrange atoms as symmetrically as possible around the central atom; try to represent the ________________ electrons of all atoms as ________________ of shared and unshared electrons.

The diagrams below represent CH_2Cl_2 and HNO_3.



CH_2Cl_2



HNO_3

Molecules

Construct dot diagrams for the following molecules. For molecules 11, 20, 21, 25, and 27, identify bond types as shown above.

9. CH_4

12. Cl_2

10. H_2

11. PH_3

Name ______________________ Class______ Date__________

13. CHI_3

14. CH_3OH

15. H_2Te

16. OF_2

17. H_2S

18. PCl_3

19. SiO_2

20. CO_2

21. $HClO_4$

22. N_2

23. H_2SO_4

24. NH_3

25. HCN

26. $HClO$

27. C_2H_4

28. C_2H_2

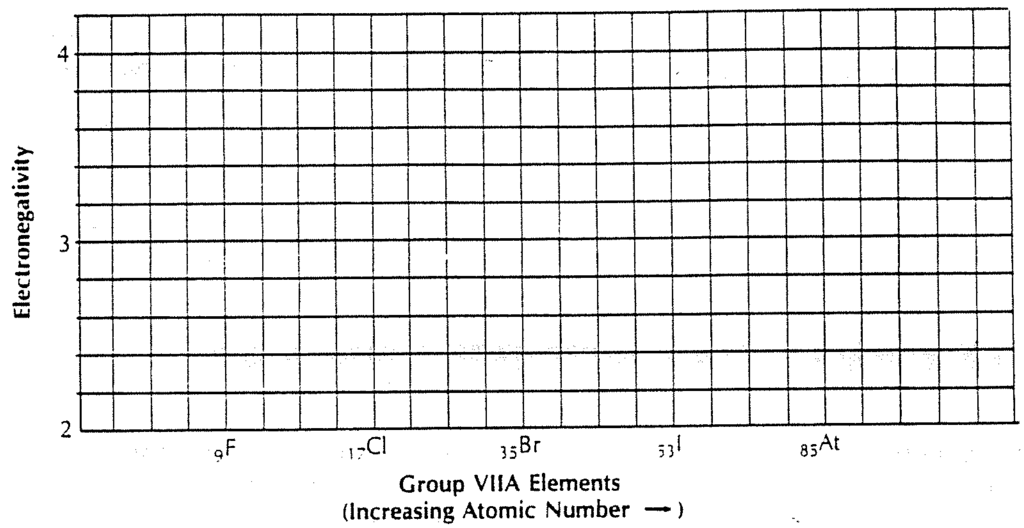
Name ______________________ Class______ Date__________

Activity 4-9
Properties of Atoms II

Electronegativity

1. Define electronegativity. ______________________

2. Plot the electronegativities of the elements of Group VIIA on the axes below. For information you need, use Table I in the Appendix.



3. (a) Describe the trend in electronegativity within Group VIIA. ______________________

   (b) What is the trend in other groups? ______________________

4. Account for the trend stated in question 3. ______________________

5. Determine the electronegativity difference for each pair of atoms below. Use this value to predict the bond type as polar covalent, nonpolar covalent, or ionic.

   a. Na—O ______________
   b. Mg—O ______________
   c. Al—O ______________
   d. Si—O ______________
   e. P—O ______________
   f. S—O ______________
   g. Cl—O ______________

Ionic radius

6. When an atom of a metal forms an ion by ______________ (*gaining/losing*) one or more electrons, its radius ______________ (*increases/decreases*).

7. When an atom of a nonmetal forms an ion by ____________ (*gaining/losing*) one or more electrons, its radius ____________ (*increases/decreases*).

8. The following list of ions is given in order of ____________ (*increasing/decreasing)* ionic radius.

$_{7}N^{3-}$ $_{8}O^{2-}$ $_{9}F^{-}$ $_{11}Na^{+}$ $_{12}Mg^{2+}$ $_{13}Al^{3+}$

9. What is the number of electrons in each ion listed in question 8? ____________

____

10. Account for the trend stated in question 8. ____________

____

____

____

11. What is the trend in ionic radius within a group? Account for your answer. ____________

____

____

____

____

Metallic versus nonmetallic character

Compare metallic and nonmetallic character in terms of these chemical and physical properties.

12. Conductivity. ____________

____

13. Luster. ____________

____

14. Ionization energy. ____________

____

15. Electronegativity. ____________

____

16. Describe the trend in metallic character within a group. ____________

____

____

17. Describe the trend in metallic character within a period. ____________

____

____

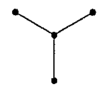
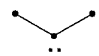
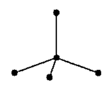
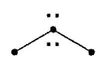
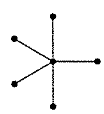
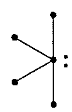
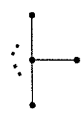
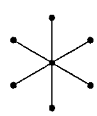
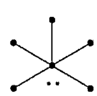
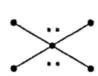
SHAPES OF MOLECULES

Name ____________________

Using VSEPR Theory, name and sketch the shape of the following molecules.

1. N_2	7. HF
2. H_2O	8. CH_3OH
3. CO_2	9. H_2S
4. NH_3	10. I_2
5. CH_4	11. $CHCl_3$
6. SO_3	12. O_2

Molecular Structure

Hybdridization	# of σ Bonds	# of Non-Bonding Pairs	Molecular Shape		Bond Angles	Example
sp	2	0		Linear	180°	
sp^2	3	0		Trigonal planar	120°	
sp^2	2	1		Angular	<120°	
sp^3	4	0		Tetrahedral	109.5°	
sp^3	3	1		Trigonal pyramidal	<109.5°	
sp^3	2	2		Angular	<109.5°	
sp^3d	5	0		Trigonal bipyramidal	120°, 90°	
sp^3d	4	1		Sawhorse (irregular tetrahedron)	<120°, <90°	
sp^3d	3	2		T-shaped	<90°	
sp^3d	2	3		Linear	180°	
sp^3d^2	6	0		Octahedron	90°	
sp^3d^2	5	1		Square pyramidal	<90°	
sp^3d^2	4	2		Square planar	90°	

Student Handout 3 of 3: Intermolecular Forces

Type of Substance	Structural Unit	Force between Units	Properties	Example
Ionic	**ions** **m^+ x^- m^+ x^-** **x^- m^+ x^- m^+** **m^+ x^- m^+ x^-** **x^- m^+ x^- m^+**	Ionic Bonding (strong)	• High melting pont • Conducts electricity only when melted or dissolved • Usually water soluble • Insoluble in non-polar solvents ("like dissolves like")	NaCl MgO
Molecular	**covalent bonds** **a) non-polar molecules**	Dispersion Forces (weak)	• Low melting pont and boiling point • Nonconducting, insoluble in H_2O • Soluble in nonpolar solvents	H_2 CCl_4
	b) polar molecules **x–x x–x x–x** **x–x x–x x–x** **x–x x–x x–x** **x–x x–x x–x**	Dispersion Forces Dipole Hydrogen Bonding (Intermediate)	• Higher melting point and boiling (higher than non-polar covalent solids) • Nonconducting • Likely to be soluble in H_2O	HCl NH_3 H_2O
Covalent Network Solids	**atoms** | | | | –X–X–X–X– | | | | –X–X–X–X– | | | | –X–X–X–X– | | | |	Covalent Bond (strong)	• Hard, solid • VERY high melting point • Non-conductors • Insoluble in common solvents	C (diamond) SiO_2 (glass sand quartz) Si SiC
Metallic	**cations and mobile electrons** **m^+ e^- m^+ e^-** **e^- m^+ e^- m^+** **m^+ e^- m^+ e^-** **e^- m^+ e^- m^+**	Metallic Bond	• Variable melting points (Hg is liquid at room temp. vs. Mg that melts at ~650°C) • Insoluble in common solvents • Malleable, ductile • Good conductors • May react with H_2O	Na Hg Mg Fe

Bonding Hierarchy

Weak ⟶ **Strong**

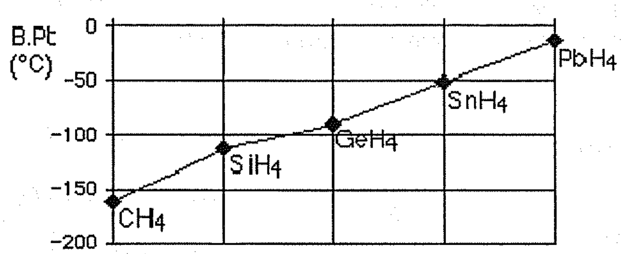
Dispersive	**Dipole**	**H–bonding** ⟶	**Metallic ionic** ⟶	**Covalent bonds**
low MM –> high MM	slightly polar –> very polar	increases with more H atoms	large ions –> small ions low charge –> high charge	

INTERMOLECULAR BONDING - HYDROGEN BONDS

This page explains the origin of hydrogen bonding - a relatively strong form of intermolecular attraction. If you are also interested in the weaker intermolecular forces (van der Waals dispersion forces and dipole-dipole interactions), there is a link at the bottom of the page.

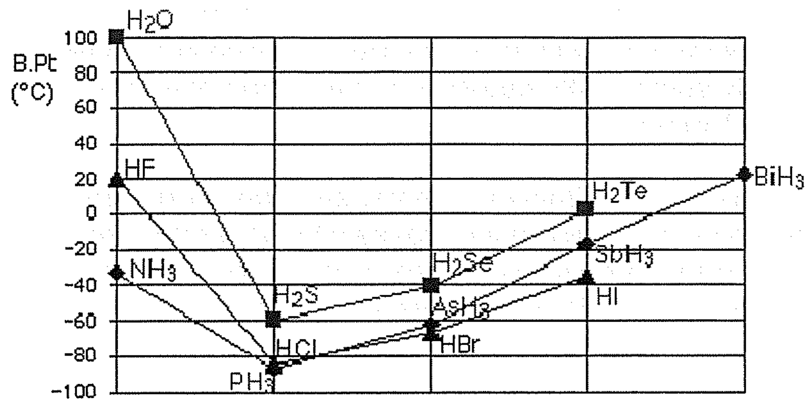
The evidence for hydrogen bonding

Many elements form compounds with hydrogen - referred to as "hydrides". If you plot the boiling points of the hydrides of the Group 4 elements, 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 hydrides of elements in Groups 5, 6 and 7, something odd happens.

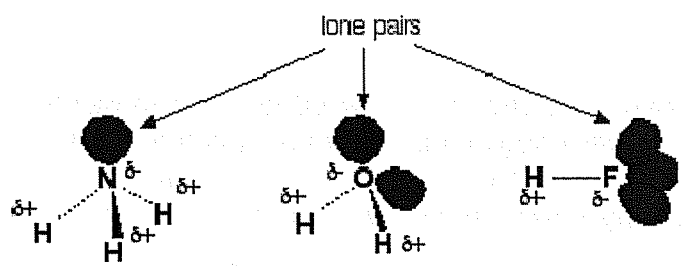


Although for the most part the trend is exactly the same as in group 4 (for exactly the same reasons), the boiling point of the hydride of the first element in each group is abnormally high.

In the cases of NH_3, H_2O 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 origin of hydrogen bonding

The molecules which have this extra bonding are:



Note: 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:

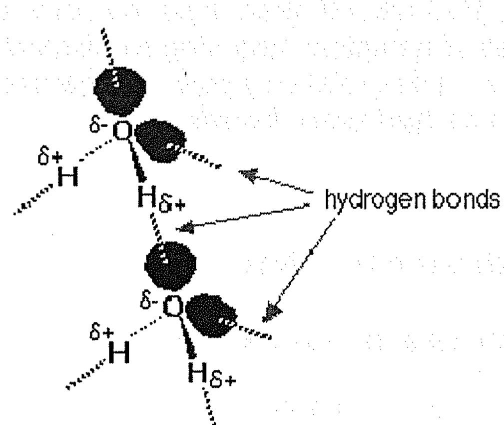
- The hydrogen is attached directly to one of the most electronegative elements, causing the hydrogen to acquire a significant amount of positive charge.

- 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.

> **Note:** If you aren't happy about electronegativity, you should follow this link before you go on.

Consider two water molecules coming close together.



The δ+ 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. On the same scale, van der Waals attractions represent mere passing acquaintances!

Water as a "perfect" example of hydrogen bonding

Notice that each water molecule can potentially form four hydrogen bonds with surrounding water molecules. There are exactly the right numbers of $\delta+$ 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.

Note: You will find more discussion on the effect of hydrogen bonding on the properties of water in the page on molecular structures.

More complex examples of hydrogen bonding

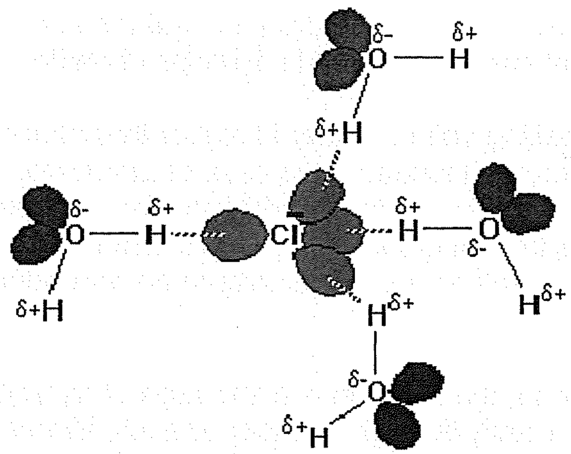
The hydration of negative ions

When an ionic substance dissolves in water, water molecules cluster around the separated ions. This process is called hydration.

Water frequently attaches to positive ions by co-ordinate (dative covalent) bonds. It bonds to negative ions using hydrogen bonds.

Note: If you are interested in the bonding in hydrated positive ions, you could follow this link to co-ordinate (dative covalent) bonding.

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.



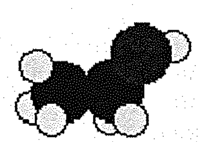
However complicated the negative ion, there will always be lone pairs that the hydrogen atoms from the water molecules can hydrogen bond to.

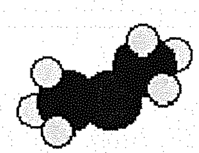
Hydrogen bonding in alcohols

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, CH_3CH_2-O-H, and methoxymethane, CH_3-O-CH_3, both have the same molecular formula, C_2H_6O.

ethanol CH_3CH_2OH 

methoxymethane CH_3OCH_3 

Note: If you haven't done any organic chemistry yet, don't worry about the names.

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 $\delta+$ charge.

In methoxymethane, the lone pairs on the oxygen are still there, but the hydrogens aren't sufficiently $\delta+$ 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 realise 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.

pentane	$CH_3CH_2CH_2CH_2CH_3$	B.Pt:	36.3°C
butan-1-ol	$CH_3CH_2CH_2CH_2OH$	B.Pt:	117°C
2-methylpropan-1-ol	$CH_3\underset{\underset{CH_3}{\vert}}{C}HCH_2OH$	B.Pt:	108°C

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 aren't 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 in organic molecules containing nitrogen

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 CH_3NH_2 (methylamine) to large molecules like proteins and DNA. The two strands of the famous alpha-helix in DNA are held together by hydrogen bonds involving N-H groups.