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 _____

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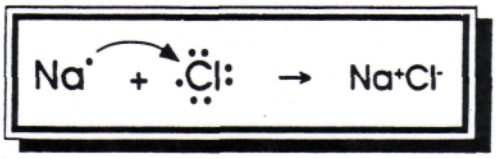
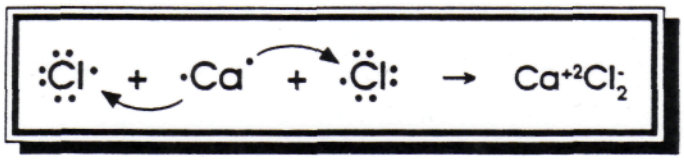
IONIC BONDING

Name _____

Ionic bonding occurs when a metal transfers one or more electrons to a nonmetal in effort to attain a stable octet of electrons. For example, the transfer of an electron . . . 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

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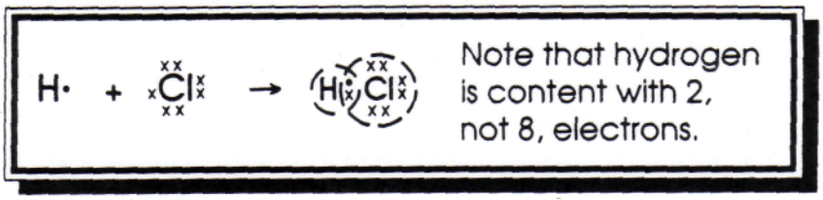
COVALENT BONDING

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:



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)

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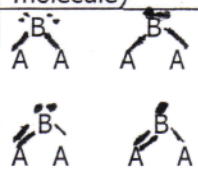
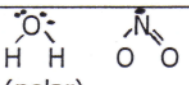
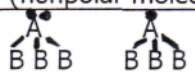
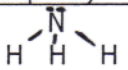
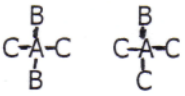
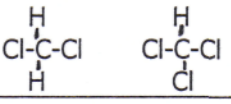
IONIC AND COVALENT BONDING

Directions: Complete the table below:

COMPOUND	BOND TYPE	LEWIS ELECTRON DOT
NaCl		
$MgCl_2$		
Br_2		
N_2		
HCl		
H_2O		
CO_2		
NH_3		
CH_4		
NH_4^+		
H_3O^+		

SHAPES AND POLARITY OF MOLECULES

A molecule is the smallest discrete particle of an element or compound formed by covalently bonded atoms. Molecules may be polar or nonpolar. A nonpolar molecule is a molecule that is symmetrical (identical parts on each side of an axis), while a polar molecule is asymmetrical (lack of identical parts on each side of an axis). Use the following chart to associate various shapes with their molecular polarity.

Number of atoms in molecules	**Structural formula and molecular polarity**	**Predicted shape (bond angle)**	**Examples**
2 atoms	A--A (non polar molecule) A--B (polar molecule)	Linear molecule (180 degrees)	H--H and O=O (nonpolar) H--Cl and Br--Cl (polar)
3 atoms	A=B=A AΞB-A (nonpolar molecules) AΞB-C (polar molecule)	Linear molecule (180 degrees)	O=C=O (nonpolar) NΞC-H (polar)
3 atoms	 (polar molecules)	Bent or angular molecule (105.4 degrees)	 (polar)
4 atoms	 (nonpolar molecule)	Trigonal planar molecule (120 degrees)	 (nonpolar)
4 atoms	 (polar molecule)	Tigonal pyramidal molecule (90 degrees)	 (polar)
5 atoms	 (nonpolar molecule)  (polar molecules)	Tetrahedral molecule (109.5 degrees)	 (nonpolar) 

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

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POLARITY OF MOLECULES

Name _____

Determine whether the following molecules are polar or nonpolar.

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

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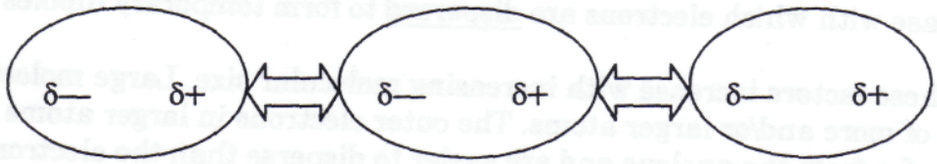
INTERMOLECULAR FORCES
(Van der Waal forces)

Molecules are much closer to one another in liquids and solids. In the gaseous state, ten molecular diameters or more typically separates particles. In liquids and solids the particles are in touch with one another. Thus, in general, liquids and solids have greater densities than gases. Similarly, liquids and solids are much less compressible than gases.

Intermolecular forces (forces of attraction between molecules) play a very large role in describing certain properties of liquids and gases; more than they do in gases. For example, surface tension shown by several liquids is an important phenomenon illustrating the effect of intermolecular forces. These intermolecular forces have been broken into several categories based on the nature of the attractions between molecules.

DIPOLE FORCES or DIPOLE-DIPOLE ATTRACTION:

Dipole forces are the results of the interaction between polar molecules. Adjacent molecules line up or orient themselves so that the negative pole of one molecule is close to the positive pole of its neighbor. This is an electrical attractive force between adjacent molecules:



THE HYDROGEN BOND:

The Hydrogen Bond is an unusually strong type of dipole force. The hydrogen bond is a force exerted between an H atom covalently bonded to an F, O or N atom in one molecule and an unshared pair of electrons on the F, O, or N atom of a neighboring molecule:

:X—H -------:X—H -------:X—H where X is F, O, or N

There are two reasons that hydrogen bonds are stronger than ordinary dipole forces:

1) The difference in electronegativity between Hydrogen (2.2) and Fluorine (4.0), Oxygen (3.5), or Nitrogen (3.0) is quite large causing the electrons to be primarily associated with the more electronegative atom (F,O, or N). Since the hydrogen atom has no shielding electrons the hydrogen atom behaves almost like a bare proton.

2) The small size of the H atom allows the unshared pair of electrons of an F, O, or N atom of a neighboring molecule to approach the H atom very closely. It is significant that hydrogen bonding occurs only with these three nonmetals, all of which have small atomic radii.

LONDON (DISPERSION) FORCES:

The most common type of intermolecular force, found in all molecular substances, is often referred to as a dispersion force. It is basically electrical in nature, involving an attraction between temporary or induced dipoles in adjacent molecules.

On the average, electrons in a nonpolar molecule, such as H_2 orCl_2 are as close to one nucleus as to the other. However, at a given instant, the electron cloud may be concentrated at one end of the molecule. This momentary concentration of the electron cloud on one side of the molecule creates a temporary dipole of the molecule. One side of the molecule acquires a partial negative charge; the other side has a partial positive charge of equal magnitude.

This temporary dipole induces a similar dipole in an adjacent molecule. The positive end of one molecule is attracted to the negative side of the adjacent molecule. These temporary dipoles, both in the same direction, lead to an attractive force between the molecules. This is the "dispersion" force.

All molecules have dispersion forces. The strength of these forces depends upon two factors:

1) The number of electrons in the atoms that make up the molecule
2) The ease with which electrons are dispersed to form temporary dipoles

Both of these factors increase with increasing molecular size. Large molecules are made up of more and/or larger atoms. The outer electrons in larger atoms are relatively far from the nucleus and are easier to disperse than the electrons in smaller atoms. In general, molecular size and Molar Mass parallel one another. Thus, as molar mass increases, dispersion forces become stronger, and the boiling point of non-polar molecular substances increases.

These Intermolecular Forces are relatively weak when compared with ordinary covalent bonds. When these intermolecular forces are compared with each other, however, in general, the weakest of these three is the London (dispersion) force and the strongest of these is the Hydrogen bond.

Adapted from
Chemistry: Principles and Reactions 3rd edition
By Masterton and Hurley
Publishers: Harcourt Brace College Publishers
pp. 240-244

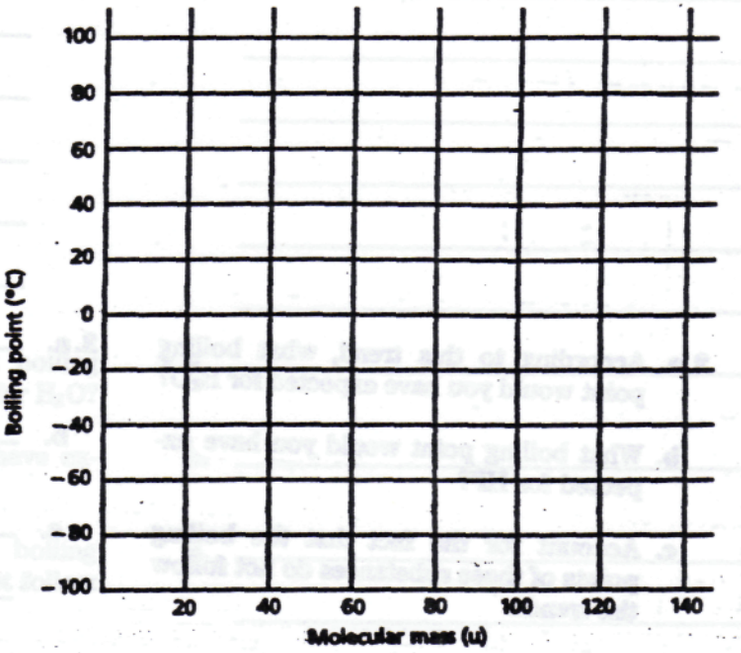
Name _____ Date _____ Class _____

Hydrogen Bonding and Boiling Point

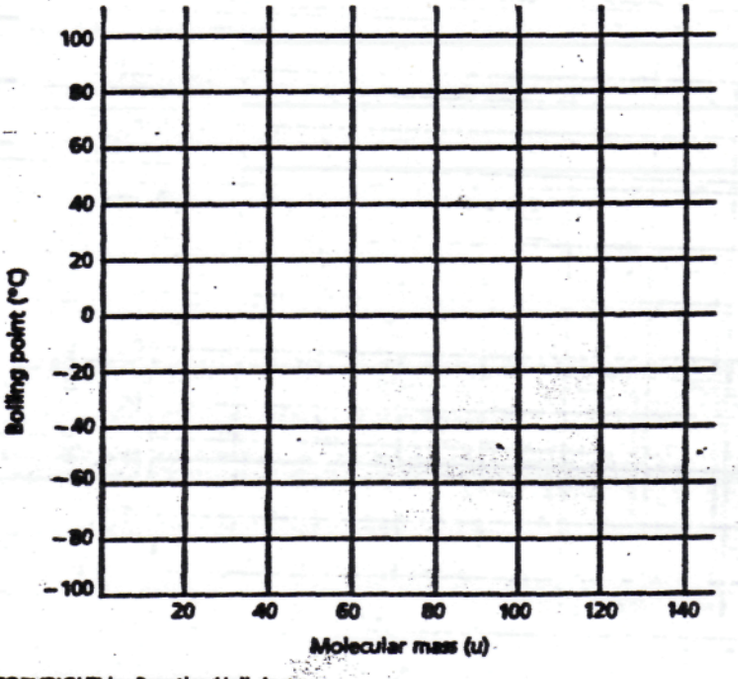
The boiling point of a substance is related to the strength of the attractive forces between its molecules: the stronger the attraction, the higher the boiling point.

On the two grids below, graph the boiling points of the hydrogen compounds listed versus the molecular masses of the compounds. (Note that in the first group, the elements combined with hydrogen have six valence electrons. In the second group, they have seven valence electrons.) Then answer the questions that follow.

Substance	Molecular Mass (u)	Boiling point
H_2O	18.0	100°C
H_2S	34.1	−62°C
H_2Se	81.0	−42°C
H_2Te	129.6	−2°C



Substance	Molecular Mass (u)	Boiling point
HF	20.0	19°C
HCl	36.5	−84°C
HBr	80.9	−67°C
HI	127.9	−35°C



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REVIEW ACTIVITY Chapter 15

Hydrogen Bonding and Boiling Point (continued)

1. a. What trend do you observe in the boiling points of H_2S, HSe, and H_2Te, and in the boiling points of HCl, HBr, and HI?

1. a. _____

b. Account for this trend.

b. _____

2. a. According to this trend, what boiling point would you have expected for H_2O?

2. a. _____

b. What boiling point would you have expected for HF?

b. _____

c. Account for the fact that the boiling points of these substances do not follow the trends.

c. _____

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REVIEW ACTIVITY Chapter 15

Hydrogen Bonding and Boiling Point (continued)

1. a. What trend do you observe in the boiling points of H_2S, HSe, and H_2Te, and in the boiling points of HCl, HBr, and HI?

1. a. _____

b. Account for this trend.

b. _____

2. a. According to this trend, what boiling point would you have expected for H_2O?

2. a. _____

b. What boiling point would you have expected for HF?

b. _____

c. Account for the fact that the boiling points of these substances do not follow the trends.

c. _____

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Name:

SHAPES AND MOLECULAR ATTRACTION

Complete the following table:

Molecule	Bond type	E dot/Structural Formula	Shape	Polarity of Molecule	Molecular Attraction
H_2Se					
PH_3					
SiO_2					
$CHCl_3$					
HI					
O_2					
NH_3					

Crystal Type	Particles in Crystal	Principal Attractive Forces Between Particles	Melting Point	Electrical Conductivity of Liquid	Characteristics of Crystal	Conditions for Formation	Examples
Ionic crystals	Positive and negative ions	Electrostatic attractions between ions. Very strong: 600–4000 kJ/mol	High	High. They also conduct electricity in solution.	Hard, brittle. Most dissolve in polar solvents.	Formed between atoms with a large difference (≥ 1.7) in electronegativity.	All salts All metal hydrides KCl, CaF_2, CsBr, MgO, $BaCl_2$, NaCl
Covalent network crystals	Atoms	Covalent bonds. Very strong: 300–800 kJ/mol	Very high	Poor	Very hard. Insoluble in most ordinary liquids. Covalent bonds extend from one atom to another in a continuous pattern.	Most formed by two elements of Group 14 or by elements whose average periodic group number is 14.	Diamond, AlN, SiC, $CuCl_2$, BeO, Mg_2Sn, Mg_2Si
Metallic crystals	Positive metal atom kernels and mobile valence electrons	Metallic bonds. Strong: 50–800 kJ/mol	Most are high.	Very high	Most are hard, malleable, ductile. High thermal conductivity. Generally insoluble in liquids. Usually soluble in molten metals.	Formed by elements with low first ionization energies and vacant valence levels.	Cu, Bi, Zn, Fe, Li, Pt, Ca, Pb, La, V, Cr, $CuZn_3$
Molecular crystals (a) Polar	Polar molecules	Dipole-dipole forces. Intermediate strength. Can be strengthened by hydrogen bonds.	Intermediate	Very low	More fragile than ionic crystals. Most are soluble in polar solvents. Usually liquids or solids at room temperature.	Formed from molecules with asymmetrical charge distributions. Polar covalent bonds are formed between atoms having a moderate difference (0.4–1.6) in electronegativity.	All acids Many organic compounds PF_3, H_2O, NH_3, SO_3, $CHCl_3$
(b) Nonpolar	Atoms or nonpolar molecules	Dispersion forces. Weak: 0–5 kJ/mol	Low	Extremely low	Very soft. Most are soluble in nonpolar or slightly polar solvents. Usually gases at room temperature.	Formed from atoms or from molecules containing symmetrical charge distributions. Nonpolar covalent bonds are formed between like atoms or atoms having a small difference (0–0.3) in electronegativity.	All diatomic molecules Ar, Cl_2, H_2, S_8, CH_4, I_2, Ne, CCl_4, CO_2, N_2, SF_6, BF_3

Name _____ Date _____ Class _____

CHAPTER 15 REVIEW ACTIVITY

Text Reference: Section 15-13

Types of Bonds

Choose words from the list to fill in the blanks in the paragraphs.

Word List

bond energy
electronegativity
hydrogen bond
ionic crystal
crystal lattice
network solid
polar molecule
polarity
potential energy
van der Waals force

__(1)__ is a measure of the attraction of an atom for electrons in a chemical bond. This attraction accounts for the degree of ionic or covalent character, or __(2)__, of a bond between atoms. An electrically neutral, chemically bonded combination of atoms that has excess positive charge at one end and excess negative charge at the other is called a(n) __(3)__. In a(n) __(4)__, covalent bonds extend from one atom to another in a continuous, highly extended pattern. In a(n) __(5)__, charged particles occur in a regular pattern called a(n) __(6)__.

A bond such as that between a hydrogen atom of one water molecule and the oxygen atom of another is called a(n) __(7)__. A weak force of attraction between molecules that arises because of shifting positions of electrons is called a(n) __(8)__.

__(9)__ is a measure of the strength of a chemical bond. Generally, a chemical change will tend to occur if it leads to a lower state of __(10)__.

1. _____
2. _____
3. _____
4. _____
5. _____
6. _____
7. _____
8. _____
9. _____
10. _____

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Activity 3-6
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 nonpolar 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

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

16. Describe bonding in metallic solids. _____

_____

17. What are the significant physical properties of metallic solids? _____

_____

Hydrogen bonding

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 characteristics of a molecule determine the magnitude of its van der Waals forces?

_____

_____

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

_____

_____