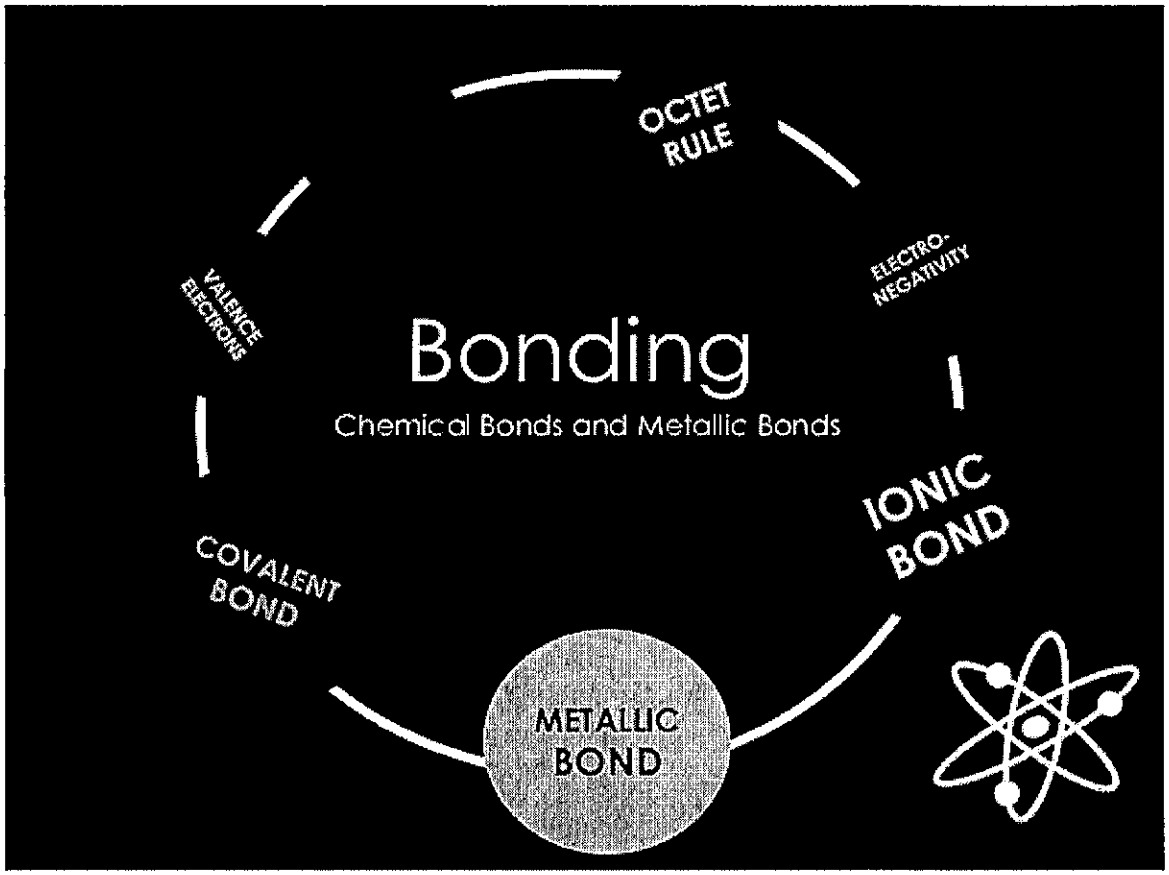
SUPA DOOPA Bonding Packet

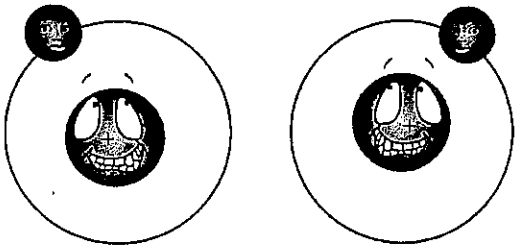


Chemistry: Form WS4.1.1A Name _____

BONDING 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? _____

_____

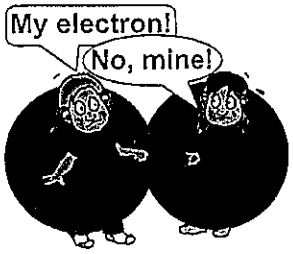
7. What is the total charge on a compound of sodium and chlorine? _____

Chemistry: Form WS4.1.4A Name _____

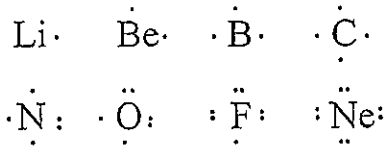
BONDING Date _____ Period _____

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.

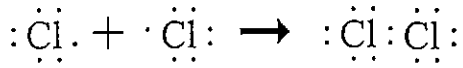


Electron Dot Diagrams Showing Unpaired Valence Electrons (*NOTE*: **When bonding occurs, molecular orbitals form. As a result, the two electrons that are normally paired in the lowest energy orbital move into separate orbitals)**

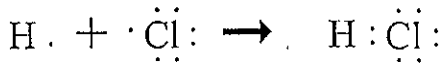


Pairing Electrons:

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



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



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? _____

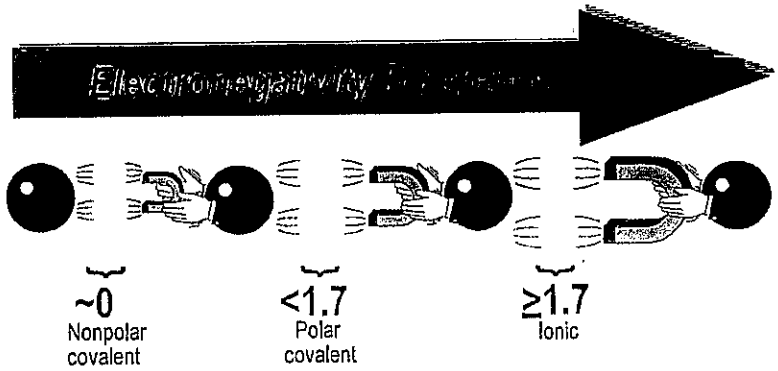
_____

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.

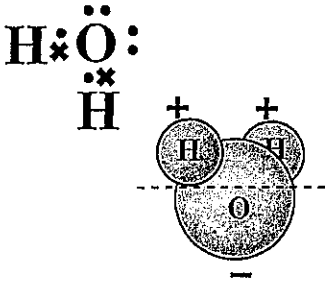
Compound	Electronegativity		Electronegativity Difference	Bond Type Ionic, Polar covalent, Nonpolar covalent
	Metal (low)	Nonmetal (high)		
Example: NaBr	0.9	3.0	2.1	ionic
HCl				
H_2Te				
KI				
SO_2				
H_2O				
CS_2				
N_2O_5				
MgO				

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 _____

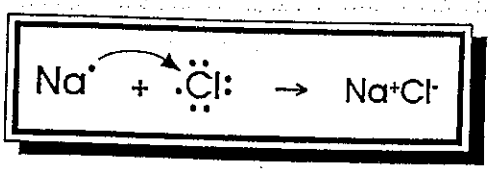
Classify the following compounds as ionic (metal + nonmetal), covalent (nonmetal + nonmetal) or both (compound containing a polyatomic ion).

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 _____

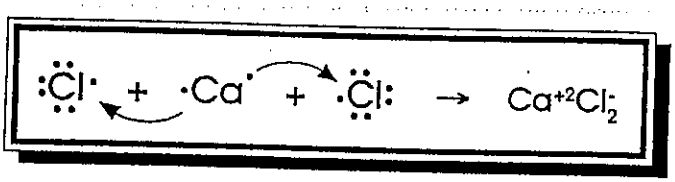
IONIC BONDING

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:

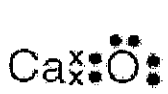
What is the driving force behind bonding?

Questions:

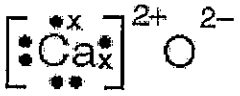
1) What happens to the positive and negative ions created in the process of ionic bonding? Why?

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



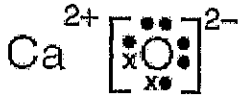
(1)



(3)

Explain each of the other three choices is wrong.

(2)

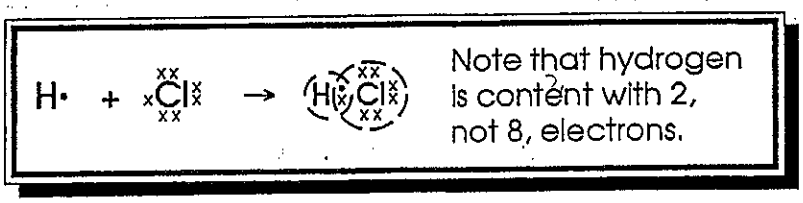


(4)

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)

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

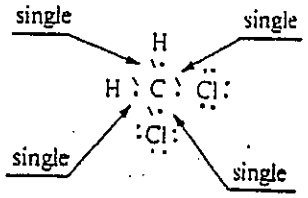
When atoms are held together by covalent bonds, a _____ or a polyatomic _____ is formed. The electron-dot symbols for individual atoms can

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.

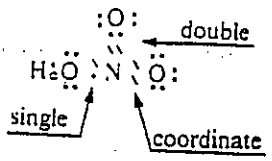
Constructing dot diagrams becomes a trial and _____ process until a reasonable structure is drawn. The following suggestions will help reduce the number of trials and errors.

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

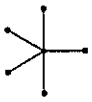
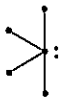
25. HCN

26. HClO

27. C_2H_4

28. C_2H_2

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°	

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

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 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 (weak) --- Dispersion Forces Dipole Hydrogen Bonding (Intermediate)	• Low melting pont and boiling point • Nonconducting, insoluble in H_2O • Soluble in nonpolar solvents --- • Higher melting point and boiling (higher than non-polar covalent solids) • Nonconducting • Likely to be soluble in H_2O	H_2 CCl_4 --- 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	