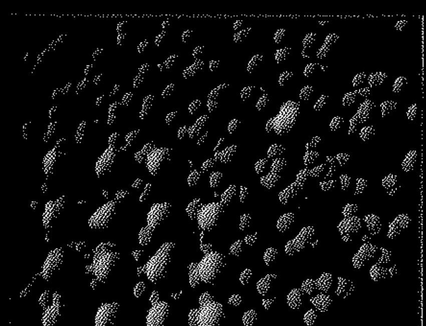
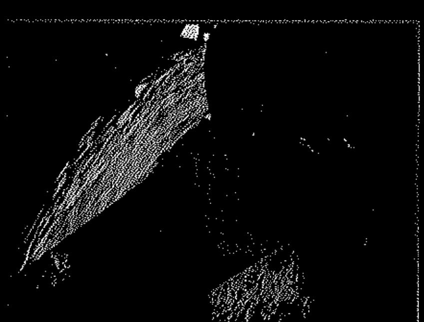
CHEMICAL BONDING series

Inner Forces:

Bonding in Metals
Ionic Bonding
Bonding in Molecules
Bonding Between Molecules





Includes:
4 Videos and Resource Guide

VEA

THE CHEMICAL BONDING SERIES
INNER FORCES

INTRODUCTION

This four program series provides a comprehensive coverage of chemical bonding. The programs use computer generated graphic models and footage of examples from everyday life to illustrate and explain the principles of chemical bonding.

While these programs are most directly relevant to the chemical bonding section of any senior secondary course, they could also be used to review bonding prior to the commencement of other sections of a course which require a basic grasp of bonding.

PROGRAM NOTES AND WORKSHEETS

The accompanying program notes provide a summary, in point form, of the material presented in each program, along with suggestions for further discussion and activities. As the programs are very compact, it may be appropriate to pause the program at various points to provide further discussion and explanation of the material presented. Some suggested points for such pauses are given in the program notes.

Worksheets which accompany the program notes are designed to provide students with a focus for recording major concepts covered in each program. In addition, they are designed to direct students to clarify and extend their understanding of each concept, by providing examples not directly presented in the program. (Where examples are to be recorded directly from the program, these sections are shaded on the worksheets). Depending upon where in the teaching sequence this material is used, the program and worksheets may serve to initiate a topic, or to revise work previously completed in classroom and laboratory activities.

A glossary of terms is included. This could be used directly in its current form, or could serve to provide a list of terms for students to generate their own glossary. Suggestions for the use of concept maps to review each area of chemical bonding are included in the worksheets. Some terms are provided, but students could be encouraged with the aid of the glossary to add terms to their concept maps.

GLOSSARY

alloy	a mixture of metals, or a metal mixed with a nonmetal such as carbon.
anion	a negatively charged ion.
atomic number (symbol Z)	the number of protons in the nucleus of an atom.
bond	an attraction which holds atoms or ions together.
brittleness	tendency to shatter when placed under a force. A property of a substance where the structure cannot be modified without significant disruption to the bonding.
cation	a positively charged ion.
compound	a pure substance composed of two or more elements which have chemically reacted together.
covalent bond	a chemical bond formed between two atoms due to their sharing of one or more pairs of electrons. Covalent bonds may be polar or nonpolar.
covalent network lattice	a solid formed when atoms covalently bond in a continuous lattice.
delocalised electrons	electrons in a substance which are free to move because they are not tightly held by atoms or ions.
dipole-dipole bond	an electrostatic attraction between polar molecules which weakly binds the molecules together.
dispersion forces	bonding between molecules due to the electrostatic attraction between momentary dipoles in the molecules. A very weak bonding, which increases as the size of the molecules increases.
double bond	a covalent bond involving the sharing of two pairs of electrons between two atoms.
electron	a negatively charged subatomic particle. Electrons are located around the nucleus.
electron dot formula	a representation of the sharing of electrons by atoms in a molecule.

electron shell	electrons around the nucleus of an atom are arranged in energy levels or shells. Electrons within one shell have similar energy.
electron subshells	electrons within shells have similar energies. Sets of electrons within a shell may have the same energy. Such a set of electrons makes up a subshell. e.g. shell 2 has 2 subshells (2s and 2p).
electron transfer	the process where one or more electrons are moved from one atom to another to form a cation and an anion.
electronegativity	a measure of the ability of an atom to attract electrons.
element	a pure substance which cannot be broken into simpler substances by chemical means.
empirical formula	a formula which shows the types of atoms in a compound and their simplest numerical ratio.
energy level	an energy value which an electron in an atom is permitted to have.
group	a vertical column of atoms in the periodic table. Elements in a group have the same number of outershell electrons.
hydrogen bonding	a type of intermolecular bonding, a special case of dipole-dipole bonding between molecules where H is bonded to a very electronegative element such as N, O, F.
ionic bond	a chemical bond which occurs between cations and anions where electrostatic forces hold the ions together.
intermolecular bonding	a weak bonding between molecules. Three types are dipole-dipole, dispersion forces and hydrogen bonding.
intramolecular bonding	strong bonding within molecules. The covalent bonds between the atoms which make up a molecule.
lattice	a solid where the particles (atoms, ions or molecules) are arranged in an ordered pattern.
mass number (symbol A)	refers to the total number of protons and neutrons in the nucleus of an atom.

metallic bond	a bond formed when metal atoms release electrons to form cations and a sea of delocalised electrons. Electrostatic attraction between the cations and the electrons binds the structure together.
molecule	a group of atoms covalently bonded together.
neutron	an electrically neutral subatomic particle. Neutrons are found in the nucleus of an atom.
nucleus	positively charged centre of the atom, made up of protons and neutrons. Contains most of the mass of the atom.
orbital	a region of space where electrons may be found. The maximum number of electrons in an orbital is two. Electrons in an orbital have the same energy.
period	a horizontal row of elements in the periodic table. Elements are listed in order of increasing atomic number.
polar covalent bond	a covalent bond where the electrons are not equally shared between the atoms. One atom has a partial positive charge, the other a partial negative charge.
polar molecule	a molecule which has one side with a slight positive charge compared to the other. This is due to the uneven distribution of electrons within the molecule.
proton	a positively charged subatomic particle. Protons are found in the nucleus of an atom.
single bond	a covalent bond which involves the sharing of one pair of electrons between two atoms.
structural formula	a representation of the arrangement of atoms within a molecule.
triple bond	a covalent bond which involves the sharing of three pairs of electrons between two atoms.

PROGRAM ONE: BONDING IN METALS

OVERVIEW

Program 1, Bonding in Metals (22 mins) provides the foundation for the series. It begins with an examination of atomic structure and the idea that bonding is related to the filling of electron shells to create more stable particles. Examples of metals and their usefulness are presented, and the relationship between properties and structure is examined.

PRE-PROGRAM ACTIVITIES

The program presents atomic structure in terms of the three basic sub-atomic particles. Electron arrangement is described in terms of shells and sub-shells. Given that many students (especially at Year 11 level) with be unfamiliar with subshells, worksheet 1.1 could be completed to review atomic structure and to establish an awareness of the subshell model presented in the program.

PROGRAM STRUCTURE

- The program begins with examples of metals, their usefulness and properties. Bonding is introduced by electron micrographs showing crystalline structure.

- Atomic structure is presented in terms of protons, neutrons and electrons. Electrons are seen to be arranged in shells and subshells containing electrons of similar energy. The stability of filled shells compared with partially filled shells is highlighted.

  NOTE TO TEACHERS:
  The program could be paused here to review or complete the material on worksheet 1.1.

- Formation of metallic bonds by the delocalisation of outer shell electrons is presented, using aluminium as an example. These delocalised electrons explain the conductivity of metals.

- Change of state from solid to liquid is presented via models. The differences in melting points of metals are explained in terms of ion size and charge. e.g. Na (m.p. 98 °C), K (m.p. 63 °C) - same charge, different size.

- Cooling of molten metal to form crystals is shown. Bonding between crystals forms the 'weak links' which account for metal cracking.

- The effect of rate of cooling is shown, and its use is illustrated e.g. cast iron, forged tools.

NOTE TO TEACHERS:

The effects of heating and cooling could be further examined by a practical activity. For example, Chemistry One Teachers Resource Book (2nd edition), Heinemann ,1995, page 7, or Key Chemistry Book One, MUP, 1991, page 70, or Chemical Connections (2nd edition), Jacaranda, 1996, pages 63-67. Worksheet 1.2 could be used to summarise the effects of heat treatment.

- Properties such as electrical and thermal conductivity, malleability and ductility are illustrated and explained.

- Patterns within the lattice are shown to depend on ion size and charge. '8 coordination' e.g. Fe, and '12 coordination' e.g. Cu and Al, are shown.

- Alloying is presented and illustrated by examples.

NOTE TO TEACHERS

Worksheet 1.3 could be used here to build on the examples used in the program.

- A summary of the levels of organisation of substances (atomic, lattice and crystalline) is presented.

NOTE TO TEACHERS

To review 'Bonding in Metals' students could generate a concept map. One suggestion, (particularly useful for ESL students), is to provide "key words" (presented on coloured card) and "connecting phrases" (presented on different coloured card). Students, or groups of students them make connections and create further cards as needed. Some typical key words and connecting phrases are presented on worksheet 1.4. Further discussion on concept maps can be found in Chemistry One (2nd edition), Heinemann, 1994, pages 20-21 and Chemical Connections (2nd edition), Jacaranda, 1996, pages 39-43.

PROGRAM TWO: IONIC BONDING

OVERVIEW

Program 2, Ionic Bonding (16 mins) uses the idea of filling of electron shells to explain the formation of cations and anions. Electrostatic forces are then used to explain lattice formation, and the relationship between properties and structure is examined.

PROGRAM STRUCTURE

- Ionic bonding is introduced by comparing the malleability of silver with the brittleness of rock salt. The regular crystalline shape of salt suggests a regular lattice structure.

- Formation of sodium and chloride ions from their respective atoms is shown in terms of the attainment of more stable outer electron shells. The sizes of atoms and ions are compared.

  NOTE TO TEACHERS
  The sodium nucleus is presented as being composed of 11 protons and the same number of neutrons. The most common isotope of sodium (the only isotope listed in the Aylward and Findlay SI Data book) has 12 neutrons, mass number 23. While the mass number is not central to the discussion, this might cause some confusion for students familiar with ^{23}Na.
  The stability of the chloride ion is presented in terms of the filling of the 3p subshell.
  To extend the idea of ion formation and ion size, worksheet 2.1 could be completed at this point. While the program does not present electron transfer in equation form, this does provide useful practice in writing electron configurations and reinforces the idea of stable outer shell configuration achieved by electron transfer.

- Lattice formation is explained in terms of electrostatic repulsion and attraction. The one to one ratio of NaCl is explained. Formulae for several other ionic compounds are shown.

  NOTE TO TEACHERS
  During the program several ionic compounds are shown, along with their chemical formulae. Students could complete worksheet 2.2 to practice writing formulae.

- Models of different lattice structures are illustrated and explained in terms of the relative sizes of ions. e.g. 6-6 coordinated NaCl, 4-4 coordinated AgCl, 8-8 coordinated CsCl.

NOTE TO TEACHERS
All examples presented in the program to this point are 1:1 ratio. Ion ratio could also be mentioned as another factor in lattice formation.

- Molecular ions are illustrated using the fertiliser components NH_4^+, NO_3^- and PO_4^{3-} as examples.

NOTE TO TEACHERS
The use of fertiliser components as examples introduces one important use of ionic compounds. Worksheet 2.3 could be completed either before, during or after the program to provide further examples of everyday uses of ionic compounds.

- Properties such as brittleness, melting point and electrical conductivity of solids and liquids are illustrated and explained.

- The program concludes with a summary comparing the structure of ionic and metallic compounds. Reference to another type of bonding within molecular ions introduces program 3.

NOTE TO TEACHERS
To review 'Ionic Bonding' students could generate a concept map. Some typical key words and connecting phrases are provided on worksheet 2.4.

PROGRAM 3: BONDING IN MOLECULES

OVERVIEW

Program 3, Bonding in Molecules (16 mins) considers the bonds formed when atoms share electrons (covalent bonding). The idea of filling electron shells to reach a more stable state is used to explain the formation of simple molecules such as water and methane. Shapes of molecules are explained by considering the repulsion between molecular orbitals.

PROGRAM STRUCTURE

- The formation of water molecules is explained in terms of the filling of the electron shells in oxygen and hydrogen. The exothermic nature of the reaction is shown by the use of the reaction in propelling rocket engines.

  NOTE TO TEACHERS
  To reinforce the idea of the filling of electron shells by electron sharing students could complete worksheet 3.1 at this point. Molecules may be drawn using both the electron dot and structural formulae.

- The shape of the water is shown using models. Molecular orbitals and their repulsion explain water's V-shape. Bonding and non-bonding electron pairs are considered.

- Further shapes are illustrated using CH_4, NH_3 and NH_4^+. Double bonds (O_2) and triple bonds (N_2) are introduced. Quadruple bonds do not form.

  NOTE TO TEACHERS
  Shapes of molecules could be investigated further by completion of worksheet 3.2. All examples used in the program involve only period 2 elements, and refer to the filling of the outer electron shell. For molecules involving period 3 elements, some explanation of the stability of eight electrons in the outer shell (in contrast to a 'full' outer shell) could be given. Non-octet molecules such as SO_2 are not considered in the program.
  Students could use the molecular model kits to build these simple molecules.

- More complex molecules, and methods used to simplify diagrams of molecules are introduced using glucose as an example.

  NOTE TO TEACHERS
  Models of glucose could be built to further visualise the complexity of the three dimensional shape of molecules. Students might then try joining their models to make cellulose.

- Diamond is introduced as a "very large molecule", a covalent network lattice. Such lattices usually only form under special conditions of very high temperature and pressure.

  NOTE TO TEACHERS
  To emphasise the fundamentally different nature of diamond (and other covalent network lattices) worksheet 3.3 could be completed here.
  To further extend the idea of "very large molecules" students could be asked to research one or more of the covalently bonded structures such as graphite, bucky balls, silicon, silicon dioxide e.t.c.

- The program concludes by distinguishing between intermolecular and intramolecular bonding. This introduces program 4.

  NOTE TO TEACHERS
  To ensure students are clear on the two types of bonding, worksheet 3.4 could be completed, and students asked to consider the type of bonding which might occur between molecules.

PROGRAM 4: BONDING BETWEEN MOLECULES

OVERVIEW

Program 4, Bonding between Molecules (17 mins) examines the bonding between molecules. Bonding is explained in terms of dipole-dipole bonding for/polar molecules, and dispersion forces for both polar and non-polar molecules. Changes of state for water are examined and the decreased density of ice when compared with liquid water is explained in terms of intermolecular bonding.

PROGRAM STRUCTURE

- The program commences with a review of the formation and shape of water molecules. Formation of a dipole in the water molecule is explained in terms of the unequal pull on the shared electrons, giving a polar covalent bond.

- Change of state of water is explained in terms of the influence of the kinetic energy of molecules and the dipole-dipole bonding between them.

- Expansion of water on forming ice is explained in terms of the formation of an open, regular pattern or lattice in ice.

  NOTE TO TEACHERS
  Using worksheet 4.1, further discussion of the "special" bonding in water could be conducted here. Comparison of the boiling points of the group 6 hydrides could be used to introduce "hydrogen bonding" - a special case of dipole-dipole bonding. While the program does not discuss the term hydrogen bonding it could usefully be considered here.

- Other polar molecules are introduced. Polarity is explained in terms of electronegativity. Elements like O, N, F, Cl and Br have high electronegativity.

  NOTE TO TEACHERS
  Electronegativity could be further considered here using worksheet 4.2.

- Formation of dipole-dipole bonds is used to explain the solubility of glucose, sucrose and other polar molecules such as HCl and NH_3.

- Bonding in non-polar molecules is explained in terms of dispersion forces, resulting from the interaction of momentary dipoles.

- Non-polar molecules which contain polar bonds are illustrated by the examples of CO_2 and CH_4.

NOTE TO TEACHERS
Worksheet 4.3 could be completed here to help reinforce the idea of the overall polarity of molecules.

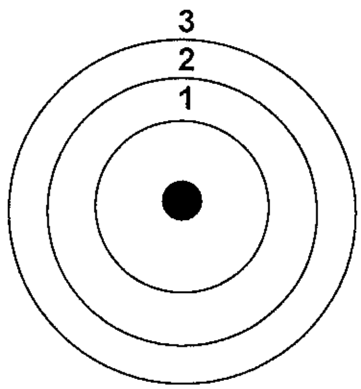
- Increasing strength of dispersion forces with increasing molecular size is illustrated by the trend in boiling points of alkanes and the solidity of plastics at room temperature.

- Intermolecular forces (dipole-dipole and dispersion forces) explain the properties of molecular substances.

NOTE TO TEACHERS
To review bonding within and between molecules, students could generate a concept map. Some typical key words and connecting phrases are provided on worksheet 4.4.

WORKSHEET 1.1 REVIEW OF ATOMIC STRUCTURE

Atoms are too small to be seen, even with the most powerful microscopes. However, the results of many experiments allow us to develop models of atoms. In the model shown below, electrons are grouped in regions around the nucleus. Each region in which a group of electrons move is called an **electron shell.** Electrons in the same shell are **approximately the same distance** from the nucleus, and have **approximately the same energy**.



In which shell do electrons have the lowest energy?

Which shell do electrons fill first?

In this model, different shells hold different numbers of electrons. **Shell n holds a maximum of $2n^2$ electrons**. The **outermost shell never holds more than eight electrons**, regardless of the maximum number possible.

Use this model to write electronic configurations for the elements listed below.

Na (Z=11) ________________________________

Ar (Z=18) ________________________________

Ca (Z=20) ________________________________

This model has some limitations. A more complex model shows that shells are composed of **subshells**. i.e. not all electrons in the third shell have exactly the same energy, so we may distinguish subshells within this shell. Subshells are given the names s, p, d, f. Just as shells hold maximum numbers of electrons, so to do subshells. A set of rules governs the filling of subshells, just as we have rules for shells. The table below summarises the subshells for the first four shells.

shell	1	2	3	4
subshells	1s	2s 2p	3s 3p 3d	4s 4p 4d 4f
max. number of electrons in the subshell	2	2 6	2 6 10	2 6 10 14
max. number of electrons in the shell	2	8	18	32

Using these rules electron configurations may be written using the notation $1s^2$ meaning shell 1, subshell s has 2 electrons.

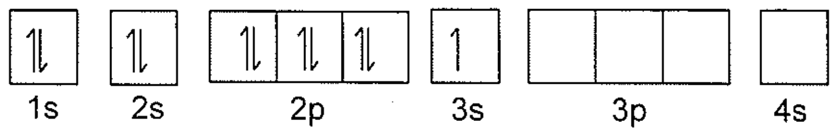
Na (Z=11) $1s^22s^22p^63s^1$

Ar (Z=18) $1s^22s^22p^63s^23p^6$

Ca (Z=20) $1s^22s^22p^63s^23p^64s^2$ but what happened to 3d?

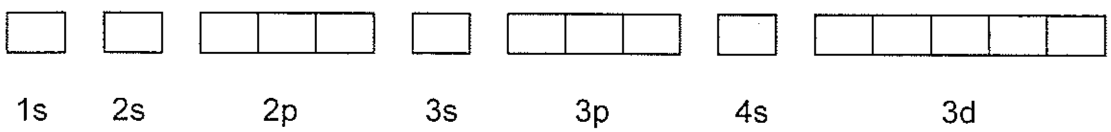
It is found that the shells and subshells begin to 'overlap' i.e. the 4s subshell actually has less energy than the 3d! Other rules exist to show the order of filling of all subshells, but we are only concerned with filling up to the 3d subshell at this stage. One way to represent this filling of subshells is to use the notation shown below.

For Na (Z = 11)

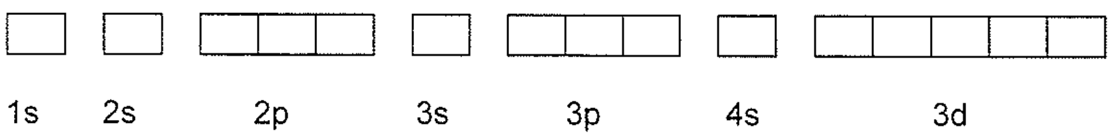


Complete the electronic configurations for the following elements using this notation.

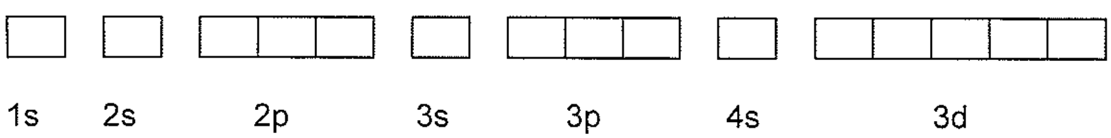
Ar (Z = 18)



Ca (Z = 20)



Fe (Z = 26)



For many of the properties of atoms we wish to explain, the simple shell model is adequate. For other properties, the more complex subshell model is needed.

WORKSHEET 1.2 HEAT TREATMENT OF METALS

Treatment type	Explanation	Effect of treatment on the metal	Examples of use of the treatment
Forging			
Casting			
Annealing			
Tempering			
Quenching			

WORKSHEET 1.3 ALLOYING OF METALS

Type of alloy	Metals alloyed	Properties of the alloy
Aluminium used in aircraft	Al,	
'Gold' Jewellery	Au,	
Stainless steel	Fe,	
Solder		
Bronze		
10 cent 'silver' coin		

WORKSHEET 1.4 CONCEPT MAP FOR METALS

KEY WORDS

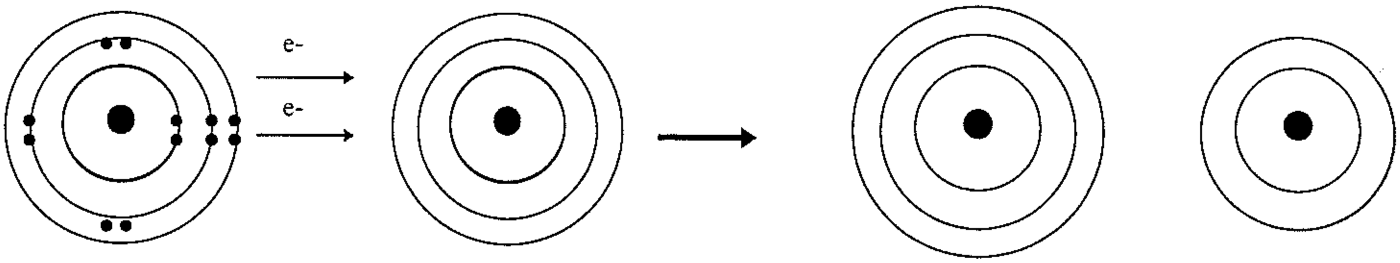
DELOCALISED ELECTRONS	CATION	METAL ATOM	HIGH MELTING POINT
MALLEABLE	HARD	LATTICE OF IONS	CONDUCTOR OF ELECTRICITY

CONNECTING PHRASES

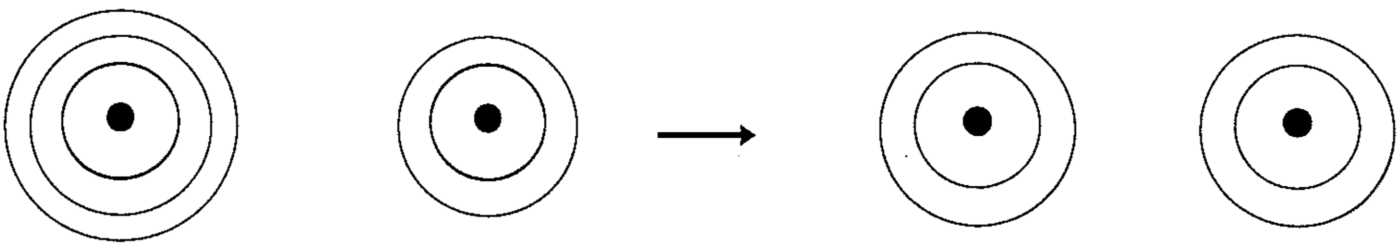
IS MADE UP OF	PRODUCE	CONSISTS OF	ATTRACT
CAUSE	BECAUSE	THEREFORE	HOLDS PARTICLES TOGETHER
DETERMINES			

WORKSHEET 2.1 ELECTRON TRANSFER EQUATIONS

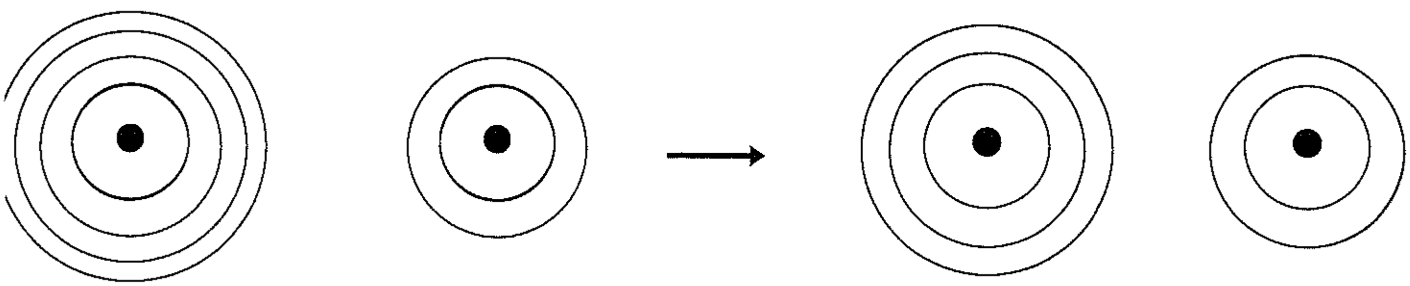
Complete the following electron transfer equations by drawing electrons, showing electron transfer and completing the electronic configurations.



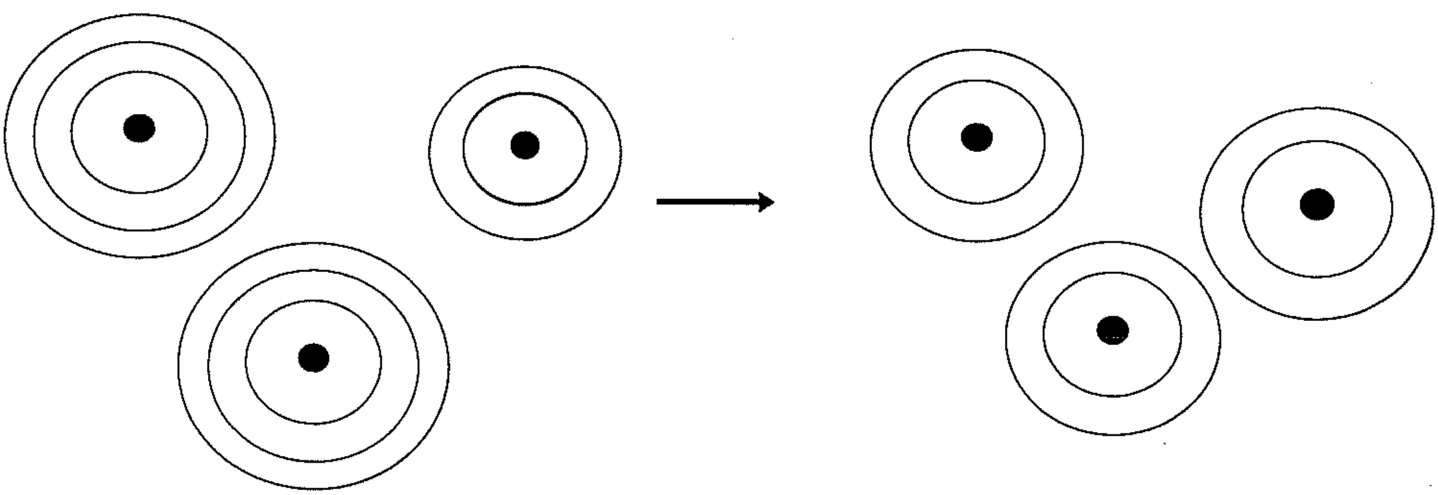
Mg(2,8,2) S(2,8,6) S^{2-}(2,8,8) Mg^{2+} (2,8)



Na( ) F( ) F^{-}( ) Na^{+}( )



Ca( ) O( ) Ca^{2+}( ) O^{2-}( )



2Na( ) O( ) 2Na^{+}( ) O^{2-}( )

WORKSHEET 2.2 WRITING EMPIRICAL FORMULAE

Write an empirical formula for each of the ionic compounds listed (you may wish to refer to the list of cations and anions given in your textbook).

sodium sulfate	________	lithium fluoride	________
potassium oxide	________	ammonium sulfide	________
aluminium dichromate	________	zinc iodide	________
iron(II) nitrate	________	silver nitride	________
calcium chloride	________	tin(II) bromide	________
nickel(II) phosphate	________	copper(I) hydroxide	________

WORKSHEET 2.3 USES OF IONIC COMPOUNDS

IONIC COMPOUND	FORMULA	USES
ammonium nitrate		
calcium carbonate		
sodium hydrogen carbonate		
calcium fluoride		
sodium chloride		
calcium phosphate		
sodium carbonate		

WORKSHEET 2.4 IONIC BONDING CONCEPT MAP

KEY WORDS

IONS	CATION	ANION	METAL ATOM
NONMETAL ATOM	HARD	BRITTLE	HIGH MELTING POINT
MOLTEN	MOBILE IONS	LATTICE OF IONS	CONDUCTOR OF ELECTRICITY
EMPIRICAL FORMULAE	ELECTROSTATIC FORCES		

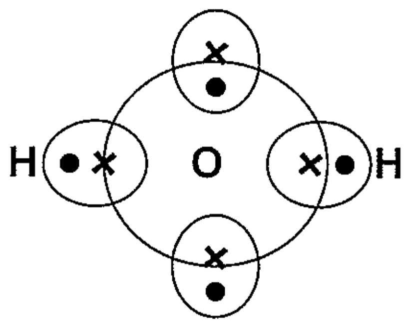
CONNECTING PHRASES

TRANSFER ELECTRONS	LOSE ELECTRONS	GAIN ELECTRONS	ATTRACT
CAUSE	PRODUCE	THEREFORE	BECAUSE
DETERMINES	IS MADE UP OF	HOLDS PARTICLES TOGETHER	

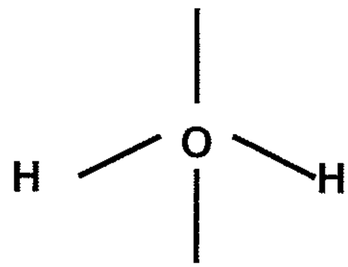
ADDITIONAL CARDS

WORKSHEET 3.1 FORMING MOLECULES

The formation of molecules when atoms share electrons can be illustrated using diagrams. Two types of diagrams are shown for the water molecule.



Electron dot formula



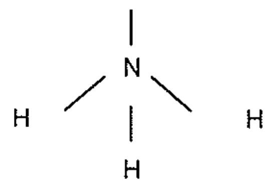
Structural formula

Draw similar diagrams for the molecules listed in the table below.

Molecular formula	Electron dot formula	Structural formula
CF_4		
NH_3		
NH_4^+		
H_2S		
N_2		
CO_2		

WORKSHEET 3.2 SHAPES OF MOLECULES

Complete the following table showing the shapes of molecules.

Molecular Formula	Structural formula	Orbital arrangement	Shape of molecule
NH_3		Tetrahedral	Pyramidal
HCN			
H_2O			
CH_4			
OF_2			
C_2H_2			

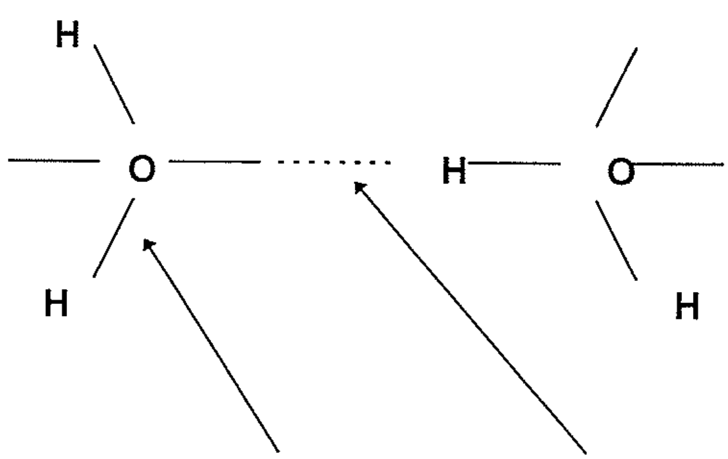
WORKSHEET 3.3 COVALENT LATTICES

	DIAMOND	CARBON DIOXIDE
STRUCTURE	a continuous lattice of carbon atoms	a discrete molecule
MELTING POINT		
HARDNESS		a gas at room temperature

WORKSHEET 3.4 TWO TYPES OF BONDING

Label the two types of bonding shown for water, using these terms;

within, between, strong, weak, intermolecular, intramolecular



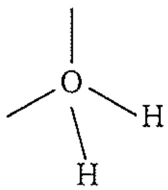
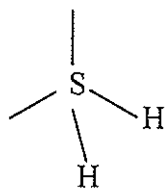
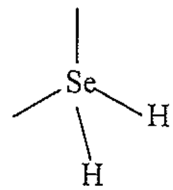
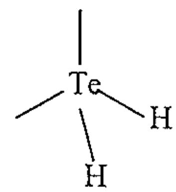
Location of bond ______________ ______________

Strength of bond ______________ ______________

Type of bond ______________ ______________

WORKSHEET 4.1 INTERMOLECULAR BONDING IN WATER

Water forms a polar molecule, and between molecules dipole-dipole bonding occurs. To examine this intermolecular bonding in water more closely, consider the data given below for the hydrides of four group six elements. H_2O, H_2S, H_2Se and H_2Te all form polar V-shaped molecules.

	H_2O	H_2S	H_2Se	H_2Te
B.P	100°C	– 41°C	– 62°C	– 2°C

Explain why the H-O, H-S, H-Se and H-Te bonds are all polar.

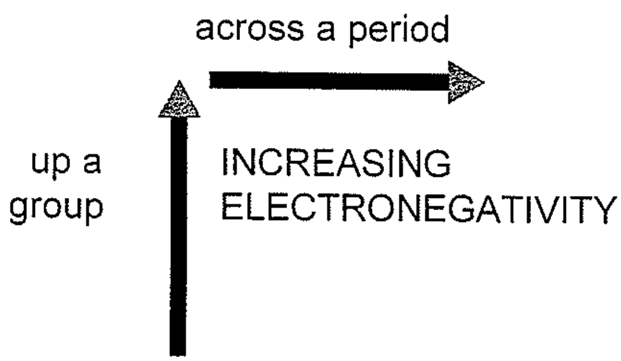
Which types of intermolecular bonding would you expect to occur between water molecules? Would the same types of intermolecular bonding be expected for the other hydrides listed above?

Considering the types of intermolecular bonding which occur, how can the range of boiling points for these hydrides be explained?

It is clear that the intermolecular bond in water is a "special" case. This "special" case is given the name hydrogen bonding, and it occurs for molecules where hydrogen is bonded to one of the very electronegative elements, N, O or F. Describe another property of water (other than boiling point) which can be explained in terms of hydrogen bonding.

WORKSHEET 4.2 ELECTRONEGATIVITY

Electronegativity is a measure of the power of an atom to attract electrons. Like many other properties, electronegativity shows a pattern within the periodic table. While numeric values for electronegativity can be assigned to elements, the patterns illustrated below enable the polarity of many bonds to be determined.



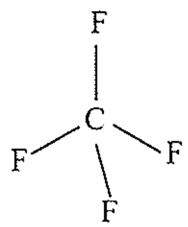
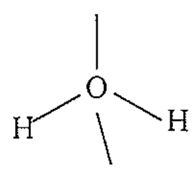
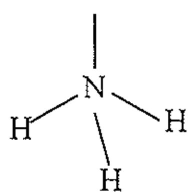
Using the patterns shown above, predict the polarity of each of the bonds shown below, and label each to show partial charges.

Br — Cl   O — N   C — F

Br — H   Cl — S   N — P

WORKSHEET 4.3 POLAR MOLECULES

Molecules may contain polar bonds, but may themselves not be polar. The molecule will be non-polar if the net overall electron distribution is even. Consider CH_4 - each bond is slightly polar, but the overall effect of the four polar bonds is an even distribution of charge, giving a non-polar molecule. Complete the examples given in the table below. It may be useful to make models of these molecules to help determine whether the partial charges are evenly distributed across the molecule.

MOLECULE	BOND POLARITY	IS THE MOLECULE POLAR?
	δ- F —— H δ+	The net effect of the four polar bonds is an even distribution of charge, hence non-polar
		
		
H —— Cl		
O ═ C ═ O		

WORKSHEET 4.4 CONCEPT MAP FOR MOLECULES

KEY WORDS

NONMETAL ATOMS	LATTICE OF MOLECULES	COVALENT BOND	LOW MELTING POINT
OUTER SHELL ELECTRONS	SOFT	BOND POLARITY	NON CONDUCTOR OF ELECTRICITY
REPULSION BY MOLECULAR ORBITALS	DISPERSION FORCES	POLAR MOLECULE	NONPOLAR MOLECULE
DIPOLE-DIPOLE BOND	MOLECULAR SHAPE		

CONNECTING PHRASES

SHARE ELECTRONS	PRODUCE	CONSISTS OF	ATTRACT
CAUSE	BECAUSE	THEREFORE	HELD TOGETHER BY
IS MADE UP OF	DETERMINES	HOLDS PARTICLES TOGETHER	