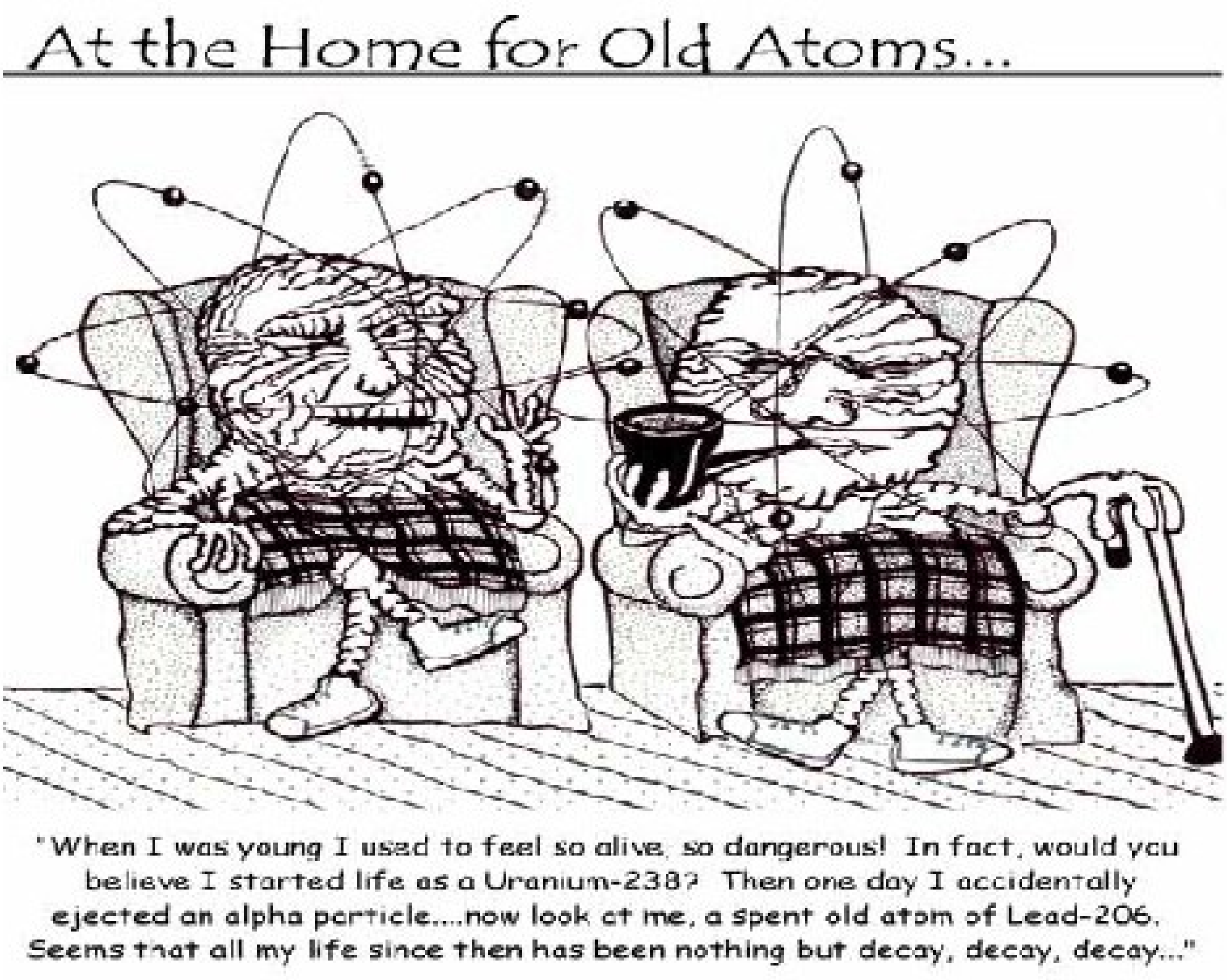
3AB Chemistry

Atomic structure
And
bonding



http://www.siraze.net/chemistry/sezennur/subjects/comics/comics21.htm

Tyson

Name: ___________________________

Atomic structure and Periodic Table

- Explain the structure of the atom in terms of protons, neutrons and electrons
- Write the electron configuration using the shell model for the first twenty elements e.g. Na. 2, 8, 1
- Explain trends in ionisation energy, atomic radius and electronegativity across periods and down groups (for main group elements) in the Periodic Table
- Describe and explain the relationship between the number of valence electrons and an element's.
  - Bonding capacity
  - Position on Periodic Table
  - Physical and chemical properties

Bonding

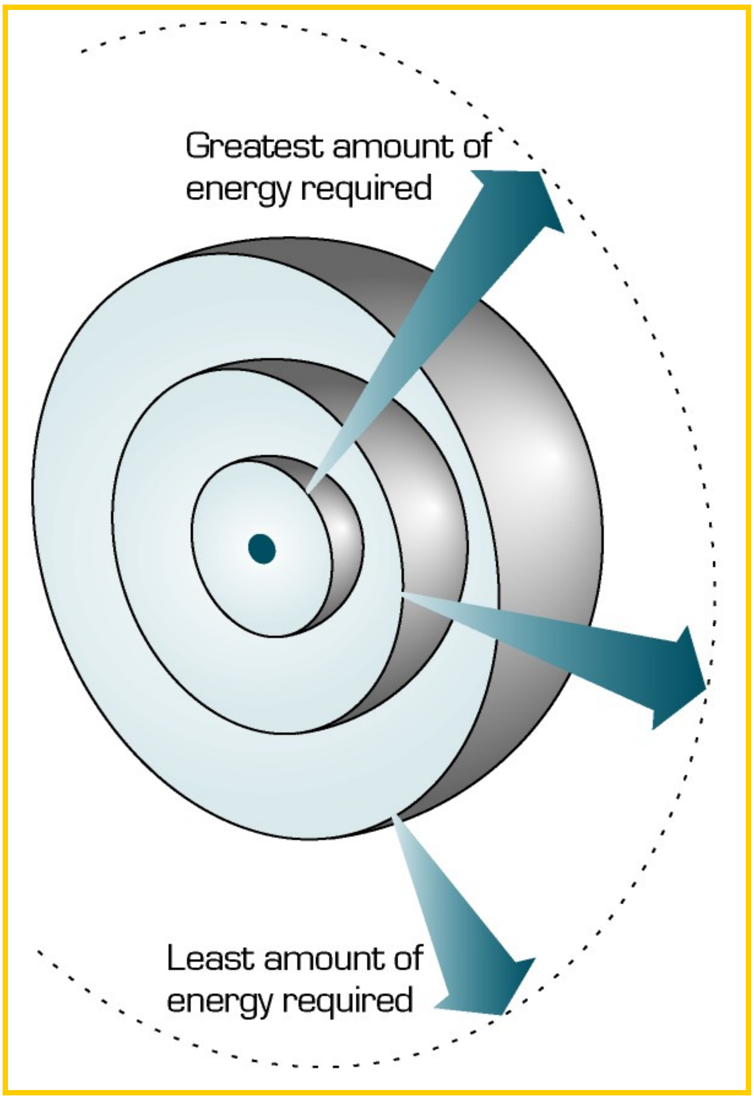
- Describe and apply the relationships between the physical properties and the structure of ionic, metallic, covalent network and covalent molecular substances
- Use the Valence Shell Electron Pair Repulsion (VSEPR) theory and Lewis structure diagrams to explain and predict and draw the shape of molecules and polyatomic ions (octet only)
- Explain polar and non-polar covalent bonds in terms of the electronegativity of the atoms involved in the bond formation
- Use the relationship between molecule shape and bond polarity to predict and explain the polarity of a molecule
- Explain the differences between intermolecular and intramolecular forces
- Describe and explain the origin and relative strength of the following intermolecular interactions for molecules of a similar size:
  - Dispersion forces
  - Dipole-dipole attractions
  - Hydrogen bonds
  - Ion-dipole interactions such as solvation of ions in aqueous solution
- Explain the relationships between physical properties such as melting and boiling point, and the types of intermolecular forces present in substances of similar size
- Apply an understanding of intermolecular interactions to explain the trends in melting and boiling points of hydrides of groups 15, 16 and 17 accounting for the anomalous behaviour of NH_3, H_20 and HF
- Describe and explain the nature of the interaction between solute and solvent particles in a solution
- Use the nature of the interactions, including the formation of ion-dipole and hydrogen bonds to explain water's ability to dissolve ionic, polar and non-polar solutes

Atomic radius

Ionisation Energy

Ionisation energy is defined as the energy required to remove the most loosely bound electron form an atom. (Note that the most loosely bound electron is one which is located in the outer shell i.e. a valence electron)

Why do we need energy to remove an electron from an atom?
Because the nucleus of an atom is positively charged (because it contains protons) and electrons are negatively charged. Opposite charges attract – this is called an electrostatic force of attraction. Hence we need energy to pull electrons away from the positively charged nucleus i.e. we need to overcome the electrostatic force of attraction.



As we move from left to right across a period in the periodic table the ionization energy increases. Why?
The increase in ionization energy in moving across a period is due to the fact that each successive element has one more proton in its nucleus. There is also an extra valence electron and hence a greater number of electrostatic forces of attraction between the nucleus of the atom and the valence electrons. Note that this increased electrostatic force of attraction between the nucleus and the valence shell causes the atomic radius to decrease across a period.

As we move down a group in the periodic table the ionization energy decreases. Why?
As we move down a group in the periodic table the number of shells that each atom has increases. This means that the distance between the nucleus and the valence electrons increases. It also means that the positively charged nucleus has a greater number of layers of negative electrons shielding the valence electrons from its positive charge.

Ref: Anderton,J.D., Garnett, P.J., Liddelow, W.R., Lowe, R.K., and Manno, I.J. (1996). Foundations of Chemistry 2ndEd. Longman, Australia. p84

Explain the shape of the graph drawn above. Why does it go up then down?

The 3 factors that influence the ionisation energy of an atom are:

1.

2.

3.

Explain how each of these factors influence ionisation energy.

Successive ionization energies

It is possible to remove more than one electron from an atom. The energy required to remove the first electron form an atom is called the **first ionization energy**. The energy required to remove another electron (i.e. a second one) is called the **second ionization energy** etc.

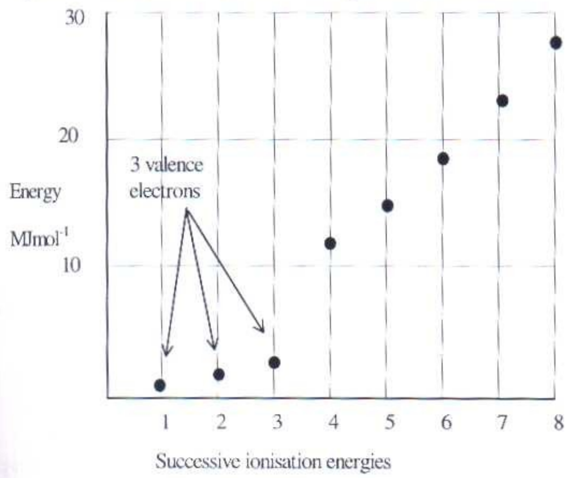
Will the second ionization energy be greater than, less than or equal to the first ionization energy?

When we remove an electron form an atom is becomes positively charged. If we then try to take a second electron away we are trying to remove a negatively charged particle from something that is positively charged and opposite charges attract! Hence the second ionization energy for any atom will always be greater than the first ionization energy.

Why is their such a big jump between the third ionization energy and the fourth ionization energy for an aluminium atom?

If you look at the graph of the successive ionization energies for aluminium you will notice a big jump between the 3rd and the 4th ionization energies. If you consider that the electronic configuration for aluminium is 2,8,3 can you figure out why this is so?

Figure 3 **Successive ionisation energies for aluminium**



Electronic structure of aluminium:

Al **$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^1$**

The first three ionisation energies for Al are relatively low with only small increases in energy needed to remove successive electrons. These energies correspond to the removal of aluminium's three valence electrons (**$3s^2$ $3p^1$**). The fourth ionisation energy shows a marked increase and corresponds to the removal of an electron from the second quantum level. This pattern of ionisation energy shows that Al has three valence electrons.

Note that the first 3 electrons are removed from the third shell whereas the fourth electron is removed from the second shell which is located closer to the positively charged nucleus. Not only is this electron located closer but it has less shielding by negatively charged electrons.

Sample TEE questions

(2000 TEE)

1. Which one of the following electron configurations represents an element that forms a monoatomic ion with a charge of –2?

   (a) $1s^22s^2$

   (b) $1s^22s^22p^2$

   (c) $1s^22s^22p^4$

   (d) $1s^22s^22p^6$

(1987 TEE)

4. Two elements have the following electron configurations

   X: $1s^2\ 2s^22p^6\ 3s^23p^1$

   Y: $1s^2\ 2s^2\ 2p^4$

   The formula of a compound formed between X and Y is most likely to be

   (a) X_3Y_2.

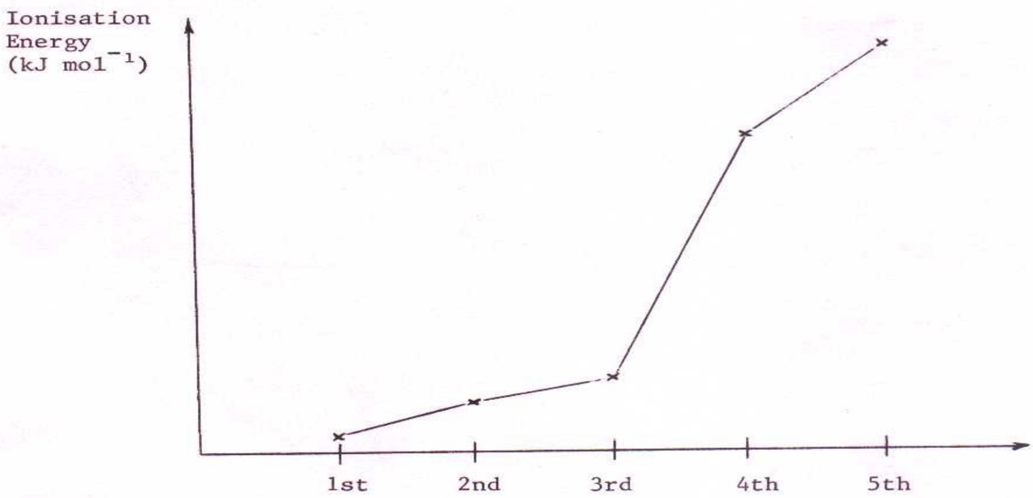
   (b) X_2Y_3.

   (c) X_2Y.

   (d) XY_2.

   (e) X_4Y.

5. The graph below shows the first, second, third, fourth and fifth ionisation energies of a particular element.



   The element is most likely to be

   (a) sodium.

   (b) magnesium.

   (c) aluminium.

   (d) silicon.

   (e) phosphorus.

1. The first four successive ionisation energies for element X are

   0.637 MJ mol^{-1} 1.24 MJ mol^{-1} 2.40 MJ mol^{-1} 7.10 MJ mol^{-1}

   Which of the following formulae is most likely for the chloride of element X?

   (a) $XC\ell$

   (b) $XC\ell_2$

   (c) $XC\ell_3$

   (d) $X_2C\ell_3$

   (e) $X_3C\ell_2$

Other sample questions

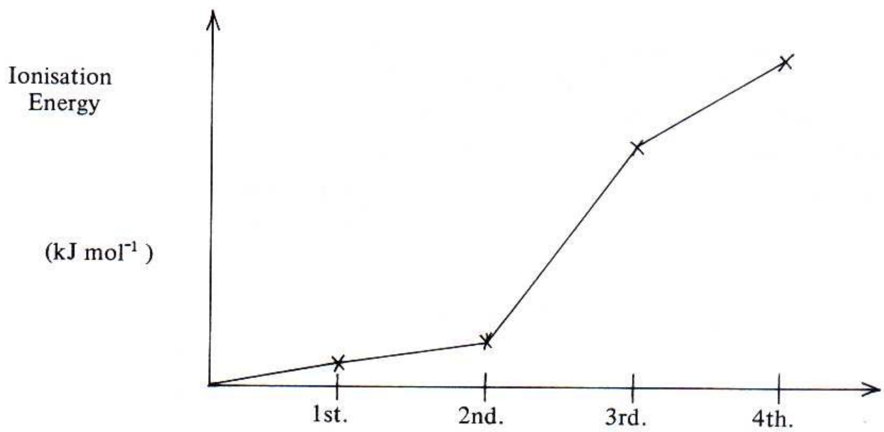
1. The atoms of an element have the electron configuration

   $1s^2 2s^2 2p^6 3s^1$

   The number of valence electrons is:

   a) 1
   b) 3
   c) 7
   d) 9
   e) 11

9. The graph below shows the first, second, third and fourth ionisation energies of a particular element.



The element is most likely to be

(1) Carbon.
(2) Oxygen.
(3) Magnesium.
(4) Chlorine.
(5) Potassium.

Review exercise 1.1 page 6 textbook

1. Nuclear power can come from the fission of uranium, plutonium or thorium, or the fusion of hydrogen into helium. In most nuclear power plants today, uranium-235 is the main fuel used. Uranium-235 undergoes fission, when it absorbs an additional neutron, according to the following equations:

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{236}_{92}\mathrm{U} \rightarrow {}^{92}_{36}\mathrm{Kr} + {}^{141}_{56}\mathrm{Ba} + 3\,{}^{1}_{0}\mathrm{n} + \text{energy}$$

The neutrons produced in this process can then be absorbed by other uranium-235 atoms to cause more fission reactions. Other nuclear reactions also occur when the neutrons bombard the other kinds of atoms present in the nuclear fuel (which is usually only about 3% uranium-235). The heat produced from these fission reactions is used to form steam that the drives the turbines and finally the electrical generators. Control rods, containing a material that is able to absorb neutrons, limit the rate of these nuclear reactions.

The fission of an atom of uranium-235 produces 10 million times the energy produced by the combustion of an atom of carbon from coal. So can nuclear energy be regarded as a more sustainable alternative to coal and other fossil fuels in Australian power plants? The answer to this question depends on two safety issues: the safe operation of nuclear power plants, and the safe disposal of the highly radioactive by-products produced by the plants.

a) $^{235}_{92}\mathrm{U}$ and $^{236}_{92}\mathrm{U}$ are isotopes of uranium. What is meant by this statement?
b) Determine the number of protons and neutrons in the four different atoms given in the nuclear equations at the beginning of the question.
c) During the fusion reaction of hydrogen, four hydrogen-1 atoms combine to form a helium atom with two neutrons, together with positrons (positively charged electrons).
This fusion reaction is a possible source of energy for nuclear power stations and it is the reaction that produces most of the sun's energy. Write an equation, like the one given for the fission of uranium, to represent the fusion reaction for hydrogen.

2. Copy and complete the following table.

Formula	Atomic number	Mass number	Number of protons	Number of electrons	Number of neutrons
Cs		133			
Sn^{2+}					70
	35	81		36	
			24	21	28

Review exercise 1.2   page 11 textbook

1. Give the electron configurations of the following atoms and ions.

   a) beryllium
   b) silicon
   c) potassium ion, K^+
   d) nitride ion, N^{3-}
   e) calcium
   f) lithium ion, Li^+

2. Give the formula of an example of an atom or ion for each of the following descriptions:

   a) A neutral atom that has 6 electrons in its second shell
   b) Another negative ion that has the same electron configuration as F^-
   c) A positive ion that has the same electron configuration as F^-
   d) A neutral atom that has a ‘full’ second shell
   e) An atom with electrons in only one shell
   f) An ion composed of one proton and no electrons

3. For boron, the five successive ionisation energies are as follows.

   807   2433   3666   25 033   32 834 kJ mol^{-1}

   a) Why is the second ionisation energy larger than the first ionisation energy?
   b) Explain why the difference between the third and fourth ionisation energies is much larger than the difference between the second and third ionisation energies.

4. The first six successive ionisation energies, in MJ mol-1, of several elements are given below. Predict how many electrons there would be in the outermost shell of the atoms of each element.

   a) 0.793, 1.583, 3.283, 4.362, 16.098, 19.791
   b) 0.502, 4.569, 6.919, 9.550, 13.356, 16.616

5. The graph in Figure 1.7 shows the successive ionisation energies for neon.

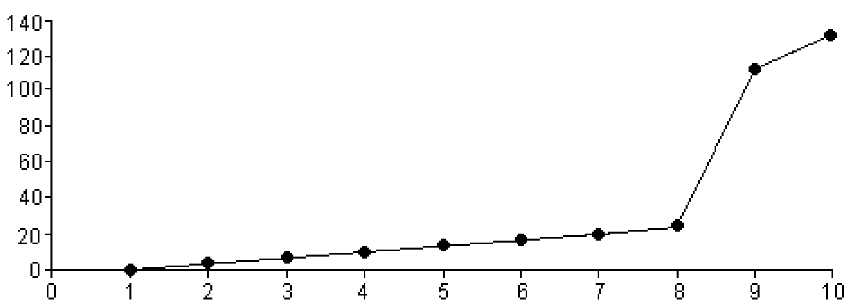


Figure 1.7 Successive ionisation energies for neon

   a) What is the electrons configuration of neon?
   b) Explain the shape of the graph shown in Figure 1.7.
   c) i Write the electron configuration of argon.
      ii Make a sketch of the graph of successive ionisation energies, like that drawn in Figure 1.7 you would expect for argon. Include figure s on the horizontal axis.

Review exercise 1.3 page 15 textbook

1. Give the name of:

   a) a gaseous element, at room temperature, in period 3 of the periodic table
   b) a group 14 element that is a non-metal
   c) the group 2, period 2 element
   d) a transition metal that forms blue-coloured compounds
   e) the period 4 element with 7 electrons in its outermost shell
   f) a group 2 element whose atoms have more protons in their nuclei than calcium atoms

2. The first six successive ionisation energies of element X are listed below (in kJ mol^{-1}).

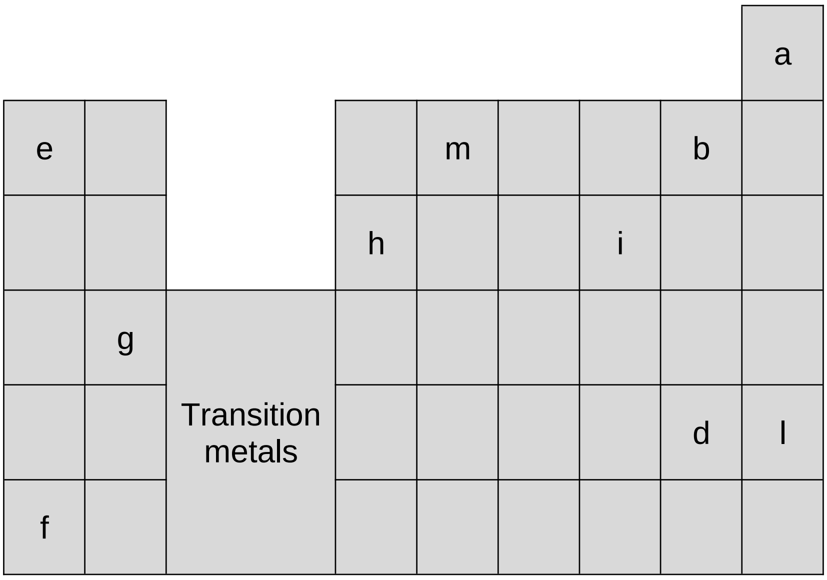
   580 1820 2750 11 600 14 800 18 400

   a) To which group of the periodic table does this element belong?
   b) What evidence given in the ionisation energy data establishes that X cannot be a period 2 element?

3. Compare the properties of the group 1 and group 17 elements listed in Table 1.5.

Review exercise 1.4   page 23 textbook

1. The following diagram is a representation of part of the periodic table.



Choose the letter that best represents the position of each of the elements described below.

a) The metal with the largest atomic radius
b) The group 17 element with the higher electronegativity
c) A non-metal with 6 electrons in its outermost shell
d) An element in group 13
e) The element with a core charge of +4
f) The period 2 element with the highest first ionisation energy
g) An element in group 2
h) A metal in period 3
i) The element with the highest first ionisation energy
j) The element with the lowest electronegativity
k) The element with the highest electronegativity

2. Explain why, going down a group, the atomic radius increases even though the number of protons in the nucleus that can attract the electrons increases.

3. a) Would a sodium ion, Na+, have a larger or smaller atomic radius than a sodium atom? Explain your answer.
b) The ionic radii of the phosphide ion (P^{3-}), sulfide ion (S^{2-}) and chloride (Cl^{-}) are respectively 212, 190 and 181 pm (1 pm = 10-12m). Explain this trend.

4. How does the concept of core charge help to explain:

a) the change in atomic radii of the elements across period 2 of the periodic table?
b) the change in first ionisation energy of the elements down group 17?

5. a) What is the first ionisation energy a measure of?
   b) Define the word ‘electronegativity’.
   c) Explain why both the first ionisation energies and the electronegativities increase across a period.
   d) Explain why both the first ionisation energies and the electronegativities decrease down a group.

Questions page 24 textbook

1. Identify each of the following as true or false:

   In all neutral atoms made up of protons, electros and neutrons:

   a) the number of neutrons minus the number of electrons is zero
   b) the number of protons equals the number of electrons
   c) the number of protons plus the number of neutrons equals the number of electrons
   d) the number of neutrons equals the number of protons
   e) the mass number minus the atomic number equals the number of neutrons
   f) the atomic number equals the number of electrons
   g) the number of protons plus the number of electrons equals the mass number

2. For the questions below, choose your answers from the following species (the symbols are ficticious):

   $^{70}_{31}Z^{3+}$ $^{74}_{32}Y$ $^{73}_{32}X$ $^{80}_{34}T$ $^{65}_{30}R^{2+}$ $^{76}_{33}Q$ $^{79}_{35}M$

   a) Which of the species are isotopes of the same element?
   b) Which two have the same number of electrons?
   c) Which have the same number of neutrons?
   d) Which is a group 17 element?
   e) Which would have similar properties to sulfur?

3. Which of the following entities have an electron configuration of 2, 8, 8?

   Ne Cl^- Sr^{2+} Ca^{2+} Ar Al^{3+} S^{2-} K^+

4. Write the electron configuration for:

   a) the sodium atom
   b) the element in group 16, period 3
   c) the Al^{3+} ion
   d) the -3 ion of the element in group 15, period 3
   e) a halogen with electrons in only two shells

5. Water in which both hydrogen atoms in each molecules are the isotope deuterium is called heavy water. Its formula is often written as D_2O, where D stands for the isotope $^{2}_{1}H$. Heavy water is commonly used in nuclear power stations to slow down the neutrons that will cause the fission reaction in uranium. Normal water and heavy water have very similar chemical properties; however, some of their physical properties differ slightly. For example, at 20°C, the density of H_2O is 1.00 g mL^{-1} but the density of D_2O is 1.11 g mL^{-1}. Propose an explanation for the difference in the density of these two forms of water.

6. Describe and explain the trend in first ionisation energy down a group in the periodic table.

7. The chloride ion, Cl^-, is isoelectronic (has the same number of electrons) with the argon atom. How does the radius of a chloride ion compare with that of an argon atom? Explain your answer.

8. Explain why the successive ionisation energies of a given element increase.

9. The successive ionisation energies, in MJ mol^{-1}, for carbon are 1.093, 2.359, 4.627, 6.229, 37.838 and 47.285. What does the dramatic increase between the fourth and fifth ionisation energies tell you about the reality of the electron shell model of the atom?

10. Consider the following first eight successive ionisation energies, in MJ mol^{-1}, for elements X, Y and Z.

X	Y	Z
0.509	1.15	0.79
0.969	2.11	1.58
3.43	3.5	3.23
4.7	4.57	4.36
6.0	5.77	16.1
7.7	8.56	19.8
9.0	9.94	23.8
10.7	18.6	29.2

a) Using these data, determine the group in the periodic table to which each element belongs.
b) Use the following information about these elements to identify X, Y and Z.
One element is the second most abundant in the Earth's crust, another is a dark red liquid at room temperature, and the name of the remaining one comes from the Greek 'barys', meaning heavy.

11. The electronegativities of three elements are:
carbon 2.5, nitrogen 3.0, phosphorus 2.2.

a) Explain why the electronegativity of carbon is less than that of nitrogen.
b) Explain why the electronegativity of nitrogen is larger than that of phosphorus

12. Account for the following observations.

a) The second ionisation energy of lithium is higher than the first ionisation energy of helium even though Li^{+} and He have the same number of electrons.
b) The second ionisation energy of sodium is much higher than the second ionisation energy of magnesium.

13. What is the relationship between first ionisation energy and electronegativity?

14. The first three elements of period 3 are sodium, magnesium and aluminium. The first ionisation energies (in kJ mol^{-1}) of these three elements, in no particular order, are 736, 494, 576, while the ionic radii (in pm), again in no particular order, are 68, 100 and 50. Match these values to each of the three elements and explain how you reached this conclusion.

15. a) Determine the electron configuration and core charge (effective nuclear charge) of the following period 3 ions:
Na^{+} Mg^{2+} Al^{3+} P^{3-} S^{2-} Cl^{-}
b) Sketch a graph showing the trend you would expect in the radii of these common ions formed by period 3 elements (except for Si) going across the period.
c) Explain your reasoning in estimating the relative ionic sizes.

16. Consider the graphs shown in Figure 1.19.

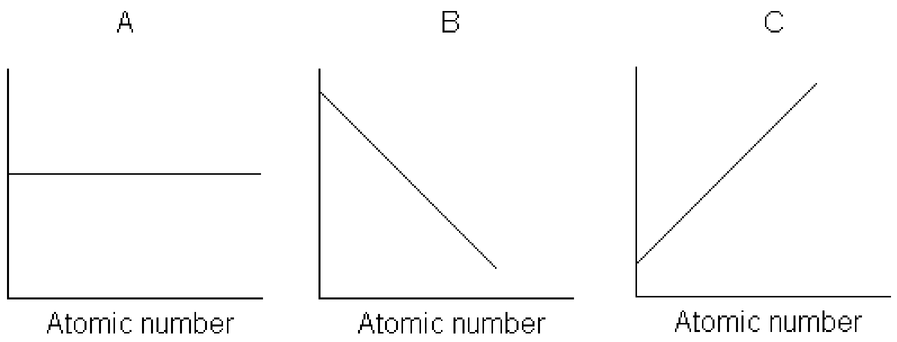


Figure 1.19

Which of these graphs could approximately represent each of the following descriptions?

a) The plot of first ionisation energy for As, Se and Br in that order
b) The plot of atomic radius for Al, Si and P in that order
c) The plot of number of outer-shell electrons for Cl^{-}, Ar and K^{+}
d) The plot of ionic radius for N^{3-}, O^{2-}, and F^{-} in that order
e) The plot of core charge of O, S and Se
f) The plot of electronegativity of Cl, Br and I

17. Lithium salts are toxic, but in controlled doses they have proved to be remarkably effective in the treatment of certain depressive illnesses. In the human body, the most commonly found group 1 and 2 elements (in the form of ions) include sodium, potassium, magnesium and calcium. The following table provides some data relating to the atomic and ionic radii of these elements.

Element	Atomic radius (pm)	Ionic radius (pm)
Lithium	152	60
Sodium	186	95
Potassium	227	133
Magnesium	160	65
Calcium	197	99

a) Account for the difference in atomic and ionic radii for lithium, sodium and potassium.
b) Why are the ionic radii of these elements smaller than the atomic radii?
c) Why is the ionic radius of a calcium ion so much smaller than the ionic radius of a potassium ion?
d) From the information provided suggest which one of these elements lithium might be replacing in its biological action. Justify your answer.

18. Of the elements whose atomic radii have been measured, the atom with the largest atomic radius is caesium, Cs. Its atomic radius is determined to be between 260 and 273 pm.

a) Why does caesium have such a large atomic radius?
b) Your answer to part a) suggests that there should be another element with a larger atomic radius. What is this element?

Metals (this is year 11 revision)

The characteristic **physical properties** of metals include

- _________________ conductors of electricity in both the solid and liquid states
- _________________ conductors of heat
- _________________ when freshly cut or cleaned (another word for shiny is lustrous)
- __________________ which means that they can be beaten and bent into different shapes without breaking
- __________________ which means they can be drawn out into thin wires
- __________________ in colour except for copper and gold
- Melting points vary over a wide range but most are ________________
- Hard, tough and relatively ________________ solids

What is the difference between a physical property and a chemical property of a substance?

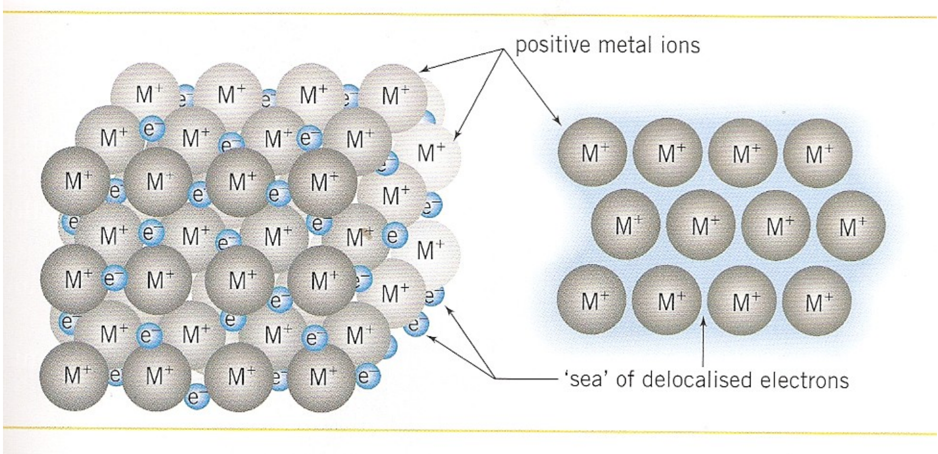
____

____

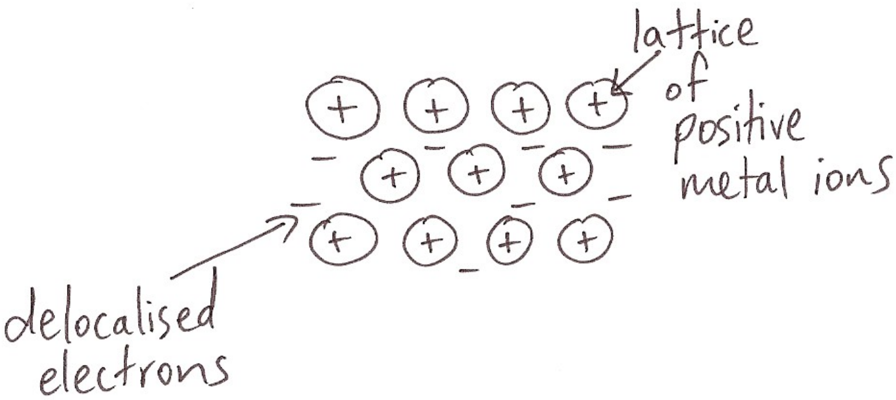
____

The structure of metals

A metal consists of a lattice of _____________________ ions surrounded by a ‘sea’ of mobile delocalised valence electrons.



What would a quick diagram of metal structure look like if you were drawing it in an exam?



The bonding in metals

Metallic bonding is the ________________________ attraction between the delocalised electrons and the positive ions in the metallic lattice. This is a strong force of attraction

Uses of metals

Aluminium is used to make drink cans, cooking foil and window frames.
Copper is used to make hot water pipes and electrical wiring.

Explaining the physical properties of metals – a summary

Property	**Explanation**
Relatively high density	Positive ions tightly packed in the lattice
Malleability and ductility	Layers of positive ions can slip over one another without disrupting the bonding because metallic bonding is non-directional
Conductivity of heat and electricity	Mobile delocalised electrons transfer charge and heat energy
High melting and boiling points	Strong metallic bonding exists throughout the entire lattice

Review exercise 2.1 page 33 textbook

1. In terms of the valence electrons, what distinguishes metals from non-metals?

2. a) When sodium reacts with a non-metal, it only forms ions with a +1 charge, but iron forms two different ions with +2 or +3 charges. Use the ionisation energy data given in Table 2.1 to propose a reason for this difference in behaviour of the two metals.
   b) A chunk of sodium can be easily cut with a knife and yet a bar of iron is much harder. Suggest a reason for the difference in hardness of these two metals.

3. a) The melting and boiling points of three metals are listed in Table 2.4. Suggest an explanation for the trends in these values.

   Table 2.4

Element	Melting point (°C)	Boiling point (°C)
Calcium	839	1484
Strontium	764	1384
Barium	725	1140

   b) Table 2.5 lists the boiling points of the group 1, 2 and 13 metals of period 4. Explain the trend in these values.

   Table 2.5

Element	Boiling point (°C)
Potassium	774
Calcium	1484
Gallium	2403

4. What properties would you look for in a metal used for the following items?

   a) A water tank for collecting rainwater from the roof for use in the garden
   b) A fuel tank for a car
   c) Electrical transmission lines from a power plant to the city
   d) A saucepan for cooking food
   e) The body of a space shuttle

5. The following passage is a description of thallium.

Thallium is a blue-white metal. It is soft enough to be cut with a knife and is also very malleable. Being quite reactive, thallium readily reacts with oxygen in the air to produce a dark grey oxide. Thallium is highly toxic and has been used in murders. It has the nickname of ‘the poisoner’s poison’. However, now that its effects are well known and an antidote for it has been found, thallium has lost its popularity as the ‘poison-of-choice’ for murderers.

a) How many valence electrons would you expect thallium to have in its neutral atoms?
b) Describe the bonding and structure of thallium. In your answer, include a simplified sketch of the particles present and their arrangement.
c) Explain why thallium is malleable.
d) Many metals are hard, but thallium is a soft metal. Propose an explanation for why thallium is a soft metal.
e) Would you expect thallium to conduct electricity? Explain your answer in terms of the bonding and structure of thallium.
f) Suggest why thallium reacts so readily with oxygen.

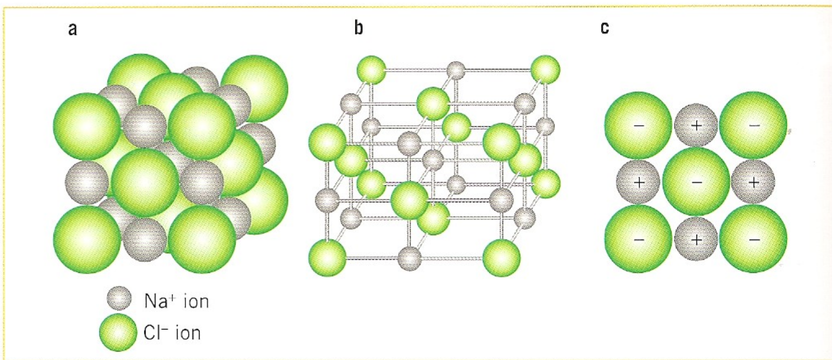
Ionic compounds (this is revision from year 11)

Physical properties of ionic compounds include

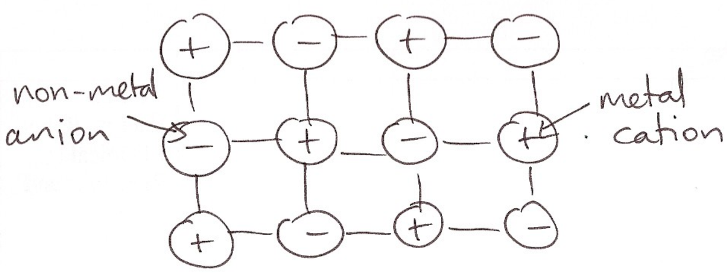
- ________________ melting and boiling points
- They are made up of _______________ crystals
- They are _______________ which means that they shatter easily when hit
- They are ___________ conductors of electricity in the solid state
- They ___________ conduct electricity in the aqueous and liquid states

The structure of ionic compounds

Ionic substances consist of positive ions and negative ions arranged in a regular lattice.



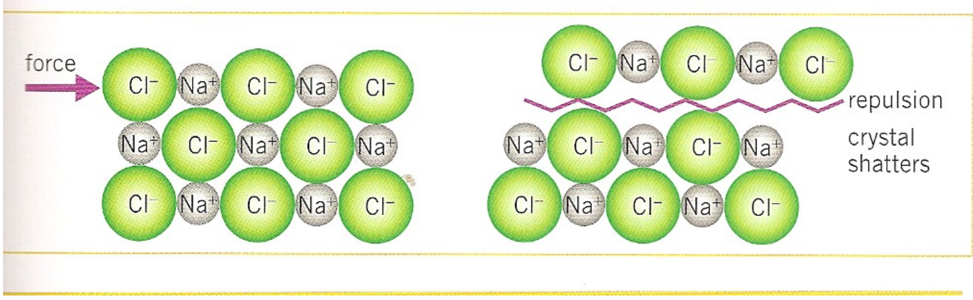
What would a quick sketch in an exam look like?



The bonding in ionic compounds

Ionic bonding is the ___________________ attraction between oppositely charged particles. This is a strong force of attraction.

Explaining hardness and brittleness of ionic solids



In your own words

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Explaining high melting and boiling points

In your own words

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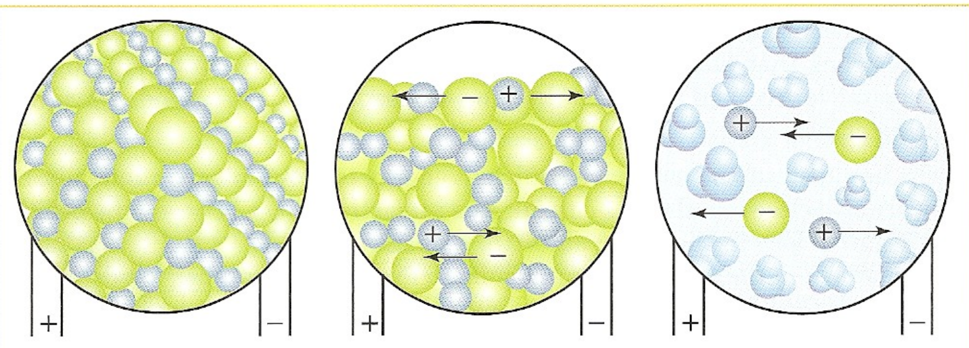
What happens when ionic solids dissolve?

Explaining the electrical conductivity of ionic substances



a In solid state, ions are trapped in fixed positions, so no current will flow.

b In molten state, mobile ions are free to move towards oppositely charged electrodes and carry a current.

c In aqueous solution, the dissolved ions are mobile and so can carry the current.

In your own words

Explaining the physical properties of ionic compounds – a summary

property	explanation
High melting and boiling points, hardness	Strong attractive forces between the positive and negative ions
Brittleness	When a layer of ions moves, repulsion between ions of similar charge occurs and the crystal shatters
Non-conductor of electricity in solid state	Positive and negative ions are held in fixed positions by strong ionic bonds in a lattice, there are no moving charged particles to conduct the current
Conductor of electricity in molten and aqueous states	Positive and negative ions are no longer arranged in a rigid lattice, they are free to move and conduct electric current

Electron dot diagrams for ionic compounds

Review exercise 2.2 page 38 textbook

1. Explain, in terms of the transfer of electrons, why when potassium reacts with phosphorus the formula of the compound formed is K_3P. Use an electron dot diagram as part of your answer.

2. Propose an explanation for why the melting point of magnesium oxide (2850°C) is much higher than the melting point of sodium fluoride (993°C), despite both substances being ionic compounds.

3. Both calcium and calcium chloride are composed of continuous arrays of charged particles.

   a) Compare the types of charged particles present in these two substances.
   b) The continuous arrays of charged particles are held together by bonds. Compare the bonding in the two substances.
   c) Explain the differences and the similarities you would expect in the electrical conductivity of these two substances.

4. Why does a crystal of copper sulfate break into many small pieces if hit by a hammer, yet if a piece of copper is hit, only a dent appears in its surface.

5. Explain why ionic compounds are solids at room temperature.

6. If a normal AA battery or an alkaline battery is cut open, a paste of ionic compounds will be found among the contents. What general property of ionic compounds is being utilised by having the paste inside the battery?

7. a) Name the following ionic compounds.
      i $(NH_4)_2SO_4$
      ii $Ca(HCO_3)_2$
      iii $Fe(H_2PO_4)_3$

   b) Write the formulas for the following ionic compounds
      i magnesium chlorite
      ii sodium dichromate
      iii copper(II) phosphide

Covalent molecular substances revision from year 11

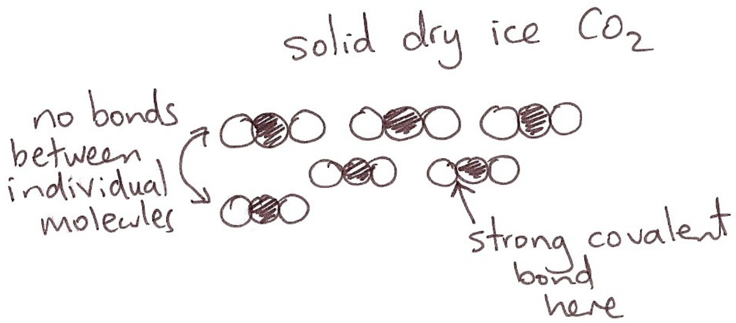
Physical properties of covalent molecular substances include

- They have ________________ melting and boiling points
- They are ______________ conductors of electricity in the solid and liquid states
- They are usually _____________ conductors of electricity in the aqueous phase although some are weak conductors of electricity.
- They form solids that are generally quite _________________, are easily scratched and have a waxy appearance

The structure of covalent molecular substances

Covalent molecular substances consist of discrete atoms or molecules. This means that the molecules are not bonded to each other.

What would a quick sketch look like in exam?



The bonding in covalent molecular substances

Within molecules atoms are held together by the electrostatic force of attraction between shared electrons and the nuclei of atoms. This form of bonding is called covalent bonding. **Covalent bonds are strong bonds**

Explaining physical properties of covalent molecular substances – a summary

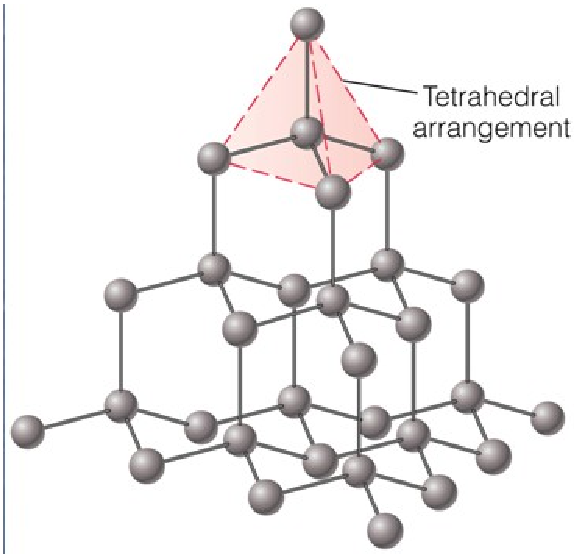
property	explanation
Low melting and boiling points	Forces between molecules are weak so not much energy is needed to separate the molecules DO NOT SAY THAT COVALENT BONDS ARE WEAK
Non- conductors of electricity when solid or liquid	The molecules are uncharged and electrons are localised in covalent bonds or in the atoms, because there are no moving charged particles there is nothing to conduct the current
Solids are generally soft	Forces between molecules are weak

Physical properties include

- Very ________________ melting and boiling points
- ______________ of electricity in the solid state (except graphite)
- Extremely _______________

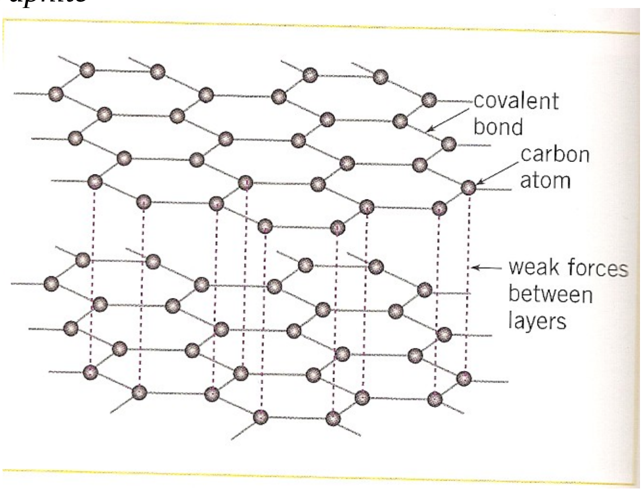
The structure of covalent network substances
In covalent network substances every atom is covalently bonded to other atoms forming a giant network lattice. No separate molecules can be distinguished. (There are no discrete molecules)

The structure of diamond

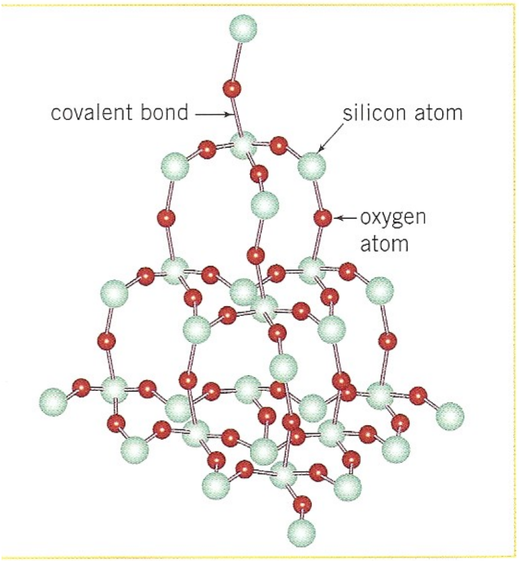


http://www.teacherweb.com/NY/JohnJay/varian/Unit10L03CovalentProperties.ppt#303,6,Crystal Structure of Diamond

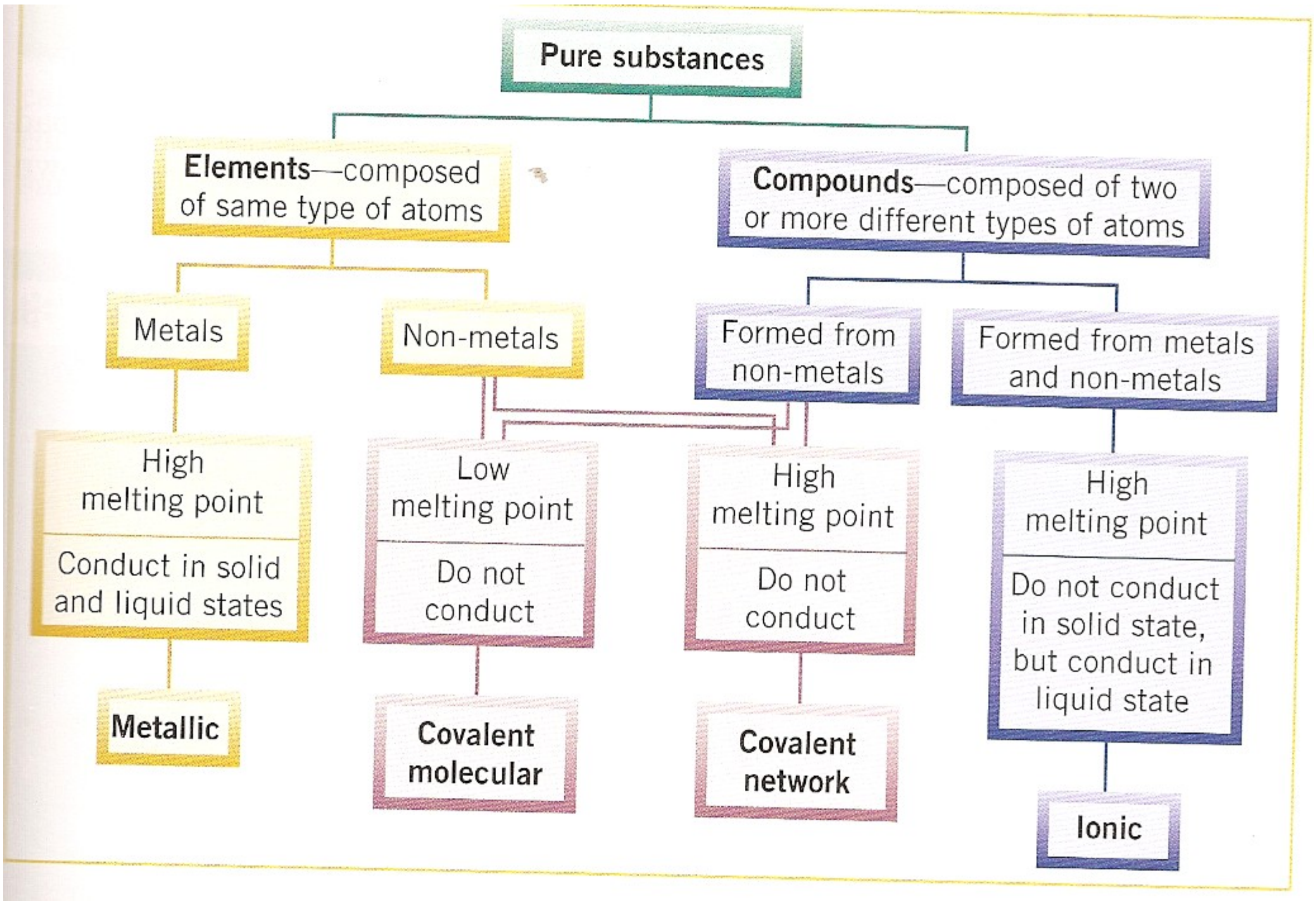
The structure of graphite



The structure of silicon dioxide



The bonding in covalent network substances


Pure substances
Elements—composed of same type of atoms
Compounds—composed of two or more different types of atoms
Metals
Non-metals
Formed from non-metals
Formed from metals and non-metals
High melting point
Conduct in solid and liquid states
Low melting point
Do not conduct
High melting point
Do not conduct
High melting point
Do not conduct in solid state, but conduct in liquid state
Metallic
Covalent molecular
Covalent network
Ionic

Electron dot diagrams for covalent molecular substances

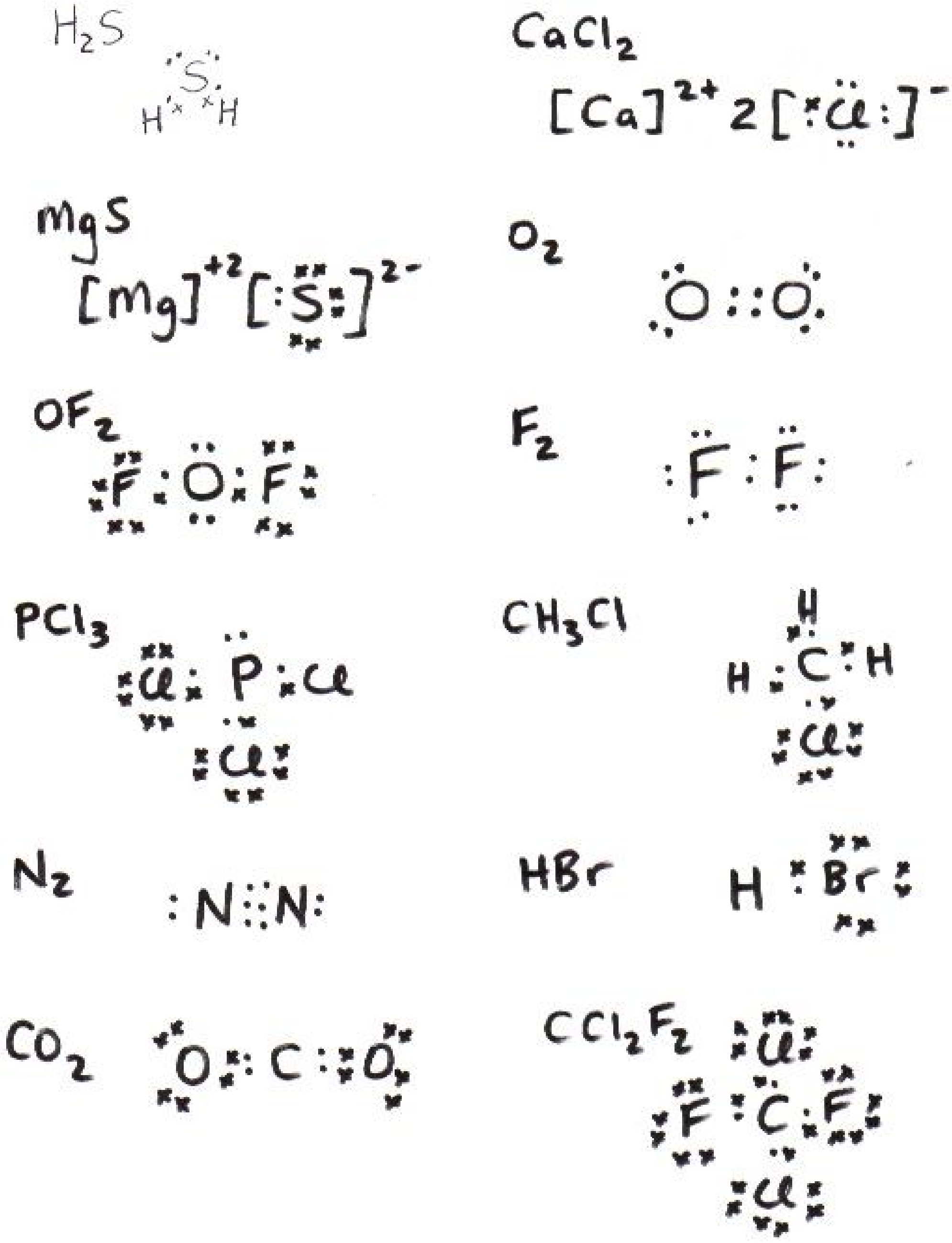
Year 11 Mastery Test
Electron Dot Diagrams

Name: ____________________

Draw electron dot diagrams for the following:

H_2S	$CaCl_2$
MgS	O_2
OF_2	F_2
PCl_3	CH_3Cl
N_2	HBr
CO_2	CCl_2F_2

Electron Dot Diagrams- Answers



FIDELITER

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Year 11 Chemistry

Mastery Test : Bonding

Name: ____________________ Date: _______________________

1. Classify each of the following substances as either ionic, covalent molecular, metallic or covalent network

a) calcium chloride ___________________________

b) aluminium ___________________________

c) carbon dioxide ___________________________

d) ammonium chloride ___________________________

e) ammonia gas ___________________________

f) silicon dioxide ___________________________

g) diamond ___________________________

h) graphite ___________________________

2. Explain why a metal such as aluminium is a good choice to make a frypan

____

____

3. Aluminium is a good conductor of electricity but solid aluminium chloride is not. Explain why.

____

____

____

4. Why are mothballs (contain C, H and O) soft and waxy?

____

____

Mastery Test : Bonding – Answers

1.
a) ionic
b) metallic
c) covalent molecular
d) ionic
e) covalent molecular
f) covalent network
g) covalent network
h) covalent network

2. Properties that you would look for in a frypan would be a high melting point (this is a physical property) so that it would not melt when you put it on a hotplate. You would not want it to react with water or oxygen (these are chemical properties) – you need to be able to wash it up and leave it exposed to air!

3. Aluminium is a metallic substance that has a delocalised sea of electrons to conduct an electric current. Solid aluminium chloride is an ionic substance that consists of anions and cations in fixed positions in a lattice. Because it does not contain any moving charged particles it is not capable of conducting an electric current.

4. Substances that contain only non-metals are covalent molecular substances. In the solid state they consist of molecules packed closely together with no bonds between the molecules. It is therefore easy to separate the molecules or have them slide over each other and so they can be considered ‘soft’.

Review exercise 2.3 page 44 textbook

1. Draw electron dot diagrams for each of the following molecules.

   a) NI_3
   b) $CClF_3$
   c) N_2O
   d) HBr
   e) CH_3COOH
   f) HCN

2. Draw line structures for the following molecules and polyatomic ions.

   a) SiH_4
   b) NO_2^-
   c) PO_4^{3-}
   d) PH_3
   e) N_2H_4

3. Explain the differences in the properties of the following two compounds in terms of their bond and structure.

	Melting point (°C)	Hardness of solid
CO_2	-57 (under pressure)	Easily broken into smaller pieces
SiO_2	1710	Very hard

4. a) How many bonding electrons are in a molecule of H_2O?
   b) How many pairs of bonding electrons are in a molecule of $SiCl_4$?
   c) How many lone pairs of electrons are in a molecule of NF_3?

5. a) When water boils, which bonds are broken?
   b) When electricity is passed through water (containing a little dissolved salt), hydrogen oxygen gases are formed. Which bonds are broken in this electrolysis process?
   c) When potassium nitrate, KNO_3, is melted, which bonds are broken?
6. Explain why carbon tetrachloride, CCl_4, has a melting point of -23°C yet calcium chloride has a melting point of 772°C.

7. a) Why would you not expect liquid hydrogen chloride, HCl, to conduct electricity?
   b) However, when hydrogen chloride, HCl, is dissolved in water, the resulting solution is a conductor of electricity. Propose an explanation for this observation.

8. The hydrocarbons ethane, ethene and ethyne are all gases at room temperature. Despite being composed of the same types of atoms, their C-C bond lengths and bond strengths differ, as shown in Table 2.8.

Table 2.8

Hydrocarbon	Molecular formula	C-C bond length (pm)	C-C bond strength ($kJ\ mol^{-1}$)
Ethane	C_2H_6	153	247
Ethane	C_2H_4	134	614
Ethane	C_2H_2	120	839

a) Draw electron dot diagrams for ethane, ethene and ethyne.
b) Why are these three compounds gases at room temperature?
c) Explain the differences in the C-C bond lengths of these three compounds?
d) Propose an explanation for why the C-C bond strengths should be different for these three hydrocarbons.

9. Read the following information about silicon carbide, then answer the questions.

Silicon carbide
Silicon combines with carbon to form a crystalline compound, silicon carbide, with a rigid diamond-like structure. The arrangement of atoms is similar to that in diamond except alternate carbon atoms are replaced by silicon atoms.

Silicon carbide (SiC), known as carborundum, is commonly used as an abrasive in sharpening and grinding tools and on ‘sandpaper’. Because it can withstand extreme temperature, it is used for manufacturing high-performance brake discs in sports cars. Silicon carbide heating elements are used in industrial processes such as the melting of non-ferrous metals and in the production of glass, ceramics and electronic components. A gem-quality synthetic form of silicon carbide is also sold under the tradename of moissanite, as a low-cost alternative to diamond. It is almost as hard as diamond but has a larger index of refraction that gives it more fire and brilliance.

a) What type of substance is silicon carbide?
b) Describe the type of bonding that exists between the atoms in silicon carbide.
c) Draw a small portion of the crystal lattice of silicon carbide.
d) Discuss the uses of silicon carbide in terms of its properties, bonding and structure.

Review exercise 2.4 page 46 textbook

1. Explain each of the following observations in terms of the bonds or forces present and the type of particles in the substances.

   a) Silver is a malleable element.
   b) Diamond is a very hard substance.
   c) Acetone, C_3H_6O, is a volatile liquid.
   d) Calcium carbonate is a brittle solid.
   e) Graphite is a good conductor of electricity.

2. Consider the following processes.

   i Sublimation of dry ice (solid CO_2)
   ii $Cl_2(g) \rightarrow 2Cl(g)$
   iii Bending an iron nail until it breaks in half
   iv Shaterring a crystal of potassium nitrate
   v Breaking the ‘lead’ (graphite) of a pencil

   For each of these processes, identify the types of:

   a) bonds being broken
   b) particles (molecules, atoms or ions) being separated

3. Consider the following substances:

lead(II) bromide	sodium chloride	sulfur	ethanol (C_2H_5OH)
hydrogen bromide	potassium iodide	zinc	silicon carbide

   From these substances name those you would expect to:

   a) conduct electricity in the solid state
   b) conduct electricity both in the liquid state and as an aqueous solution
   c) not conduct electricity in their pure form at room temperature but produce a conducting liquid when dissolved in water
   d) not readily conduct electricity under any conditions
   e) have a melting point below 120°C

4. Copy and complete the following table to summarise the properties of different types of substances.

Type of substance	Particles present (atoms, ions, electrons, molecules)	Melting point (high or low)	Electrical conductivity	
			of solid	of liquid
Ionic				
Metallic				
Covalent molecular				
Covalent network				

Review exercise 2.5 page 50 textbook

No need to do these questions

1. RESEARCH
   a) What are the properties of soda glass?
   b) Discuss the ‘pluses and minuses’ of using this type of glass, relating these to the properties of the glass.

2. RESEARCH
   a) What are optical fibres?
   b) Relate the properties of optical fibres to their uses?
   c) How are optical fibres made?
   d) Discuss the contribution of fibre optic cables to modern society.

3. RESEARCH
   a) Give some examples of composite materials.
   b) For each example:
      i name the materials present in the composite
      ii compare the properties of these separated materials with the properties of the composite

4. RESEARCH
   How are liquid crystals used in calculator displays to provide the black and clear regions where keyboard buttons are pressed? What features of the liquid crystal molecule are important in this process?

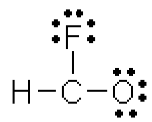
QUESTIONS page 51 textbook

1. The following compounds may be found around the home. Draw the electron dot diagrams for each of these.

Common name	Use	Formula
a) Bicarbonate of soda	In cooking	$NaHCO_3$
b) Sulfate of ammonia	Fertiliser	$(NH_4)_2SO_4$
c) Caustic soda	Drain cleaner	$NaOH$
d) Sulfuric acid	Car battery acid	H_2SO_4
e) Ethanol	Wine	CH_3CH_2OH
f) Hydrogen peroxide	Teeth whitener	H_2O_2

2. As can be seen in Figure 2.4, which shows the many pipes and towers that are part of an oil refinery, metals are used extensively for these constructions. Suggest reasons why metals are used for this purpose and not ionic, covalent network or covalent molecular substances.

3. Consider the structure of fluoromethanal drawn below.



a) Explain what is wrong with this structure.
b) Redraw the structure correctly.

4. Why is the boiling point of a metal a better measure of the strength of the metallic bond than the melting point?

5. Explain the uses of each of the following materials in terms of their bonding and properties.

a) An alloy of magnesium is used in the wheels of some cars.
b) Aluminium is used for window frames.
c) Silicon carbide is used in sandpaper.
d) Sodium chloride is present in sports drinks, such as Gatorade.
e) Copper is used in electrical wiring.
f) Limestone is used in the construction of buildings, fences and retaining walls.
g) Hydrocarbons are used as propellants in spray cans.
h) Graphite is used in 'lead' pencils.
i) Ammonium nitrate is a better 'fast acting' fertiliser than urea, $CO(NH_2)_2$.

TABLE 2.9 IONISATION ENERGIES AND SOME PHYSICAL PROPERTIES OF A VARIETY OF METALS FROM THE SAME PERIOD OF THE PERIODIC TABLE

Metal	Group	Melting point (°C)	Density (g cm^{-3})	Hardness (Brinell scale, MPa)	Successive ionisation energies (MJ mol^{-1})			
					1st	2nd	3rd	4th
Potassium	1	254	0.86	0.363	0.425	3.058	4.418	5.883
Calcium	2	839	1.55	167	0.596	1.152	4.918	6.480
Scandium	Transition metal	1538	3.0	750	0.637	1.241	2.395	7.095
Titanium	Transition metal	1660	4.50	716	.0664	1.316	2.659	4.181
Iron	Transition metal	1535	7.86	490	0.766	1.567	2.964	5.29

6. Use the data given in Table 2.9 to suggest explanations for the following observations.

   a) Iron has a higher melting point than calcium.
   b) Potassium is a softer metal than scandium.
   c) Potassium only forms +1 ions, but titanium forms +2 and +3 ions.
   d) The density of titanium is greater than the density of calcium.

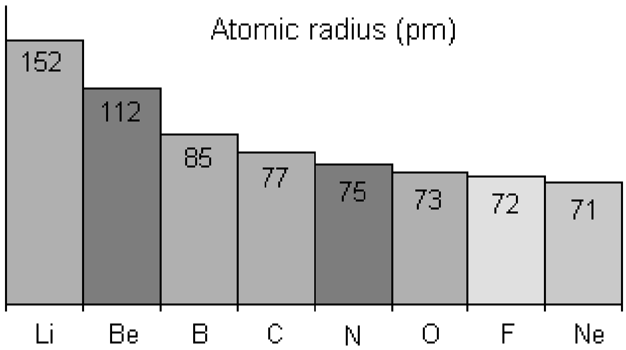
7. Explain why metals are always reductants in oxidation-reduction reactions.

8. Propose explanations for the differences in the covalent bond energies (the energy required for $X_2(g) \rightarrow 2X(g)$) given below:

   F_2 160 kJ mol^{-1}
   O_2 498 kJ mol^{-1}
   N_2 945 kJ mol^{-1}

9. Lead can be flattened into a thin sheet by an applied force, but lead nitrate shatters. Account for these two observations.

10. The atomic radii of the atoms of the period 2 elements are shown in Figure 2.26. Explain the trend shown in this data.



11. Superconductors are materials through which electricity can flow with little or no resistance. As a consequence, when electricity flows through superconductor, there should be little or no heating effect. Superconductors offer the promise of benefits such as power transmission without power loss and superfast electronic circuits. However, the few metals that can be superconducting, such as aluminium and mercury, only exhibit this property at extremely low temperatures, just above absolute zero. Research is being carried out to develop superconducting materials that can work at room temperatures.

    Suggest why metals only behave as super conductors at extremely low temperatures.

12. At times in bonding topics, reference is made to the 'octet rule'. Essentially this 'rule' proposes that since there are eight electrons in the valence shell of the unreactive noble gases (except helium) chemical bonding in molecular substances will result in all atoms (except hydrogen) having share of this stable number of eight electrons.

    While this 'rule' is true for many covalent molecular compounds, in particular those involving period elements as they are only able to accommodate eight electrons in their outermost shell, it does not apply to all possible molecular substances.

    Each of the compounds listed below forms a non octet structure. Draw an electron dot diagram for each.

    a) SCl_6
    b) TeF_4
    c) BCl_3
    d) IF_7
    e) PCl_5

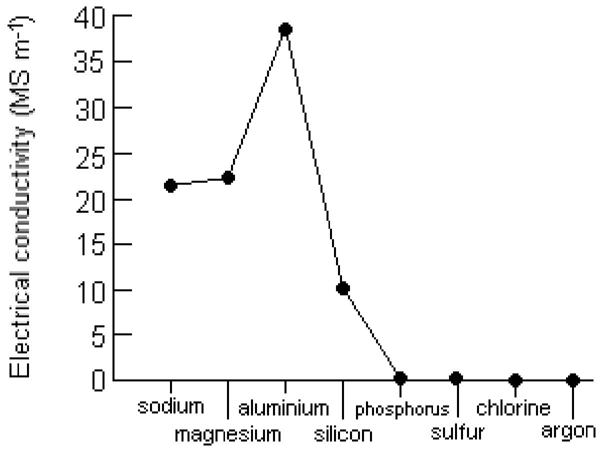


Figure 2.27 The trend in electrical conductivities of the period 3 elements

13. The trend in electrical conductivity of the period 3 elements, in their solid state, is shown in the graph in Figure 2.27. Explain this trend in terms of the bonding and structure of the elements.

14. If energy is required to ionise metal atoms and to convert non-metal molecules to anions (negative ions), account for the fact that ionic compounds form so readily.

15. The ionisation energies for carbon and sodium are given in the table below.

Ionisation energy (kJ mol^{-1})						
	1st	2nd	3rd	4th	5th	6th
Carbon	1100	2372	4660	6271	38 093	47 580
Sodium	502	4569	6919	9550	13 356	16 616

a) Give the electron configuration of carbon.
b) Write equations to represent the first and the second ionisation of carbon.
c) Explain why the second ionisation energy is larger than the first ionisation energy for carbon.
d) Explain why there is a very large difference between the fourth and the fifth ionisation energies for carbon.
e) Suggest a reason that might explain why the first ionisation energy of carbon is larger than the first ionisation energy of sodium.
f) Using the ionisation energy data, discuss why metallic bonding holds the particles together in a piece of sodium, yet covalent bonding holds the atoms together in a diamond.

16. One of the reasons gold, silver and copper are used extensively in the jewellery industry is that they are reasonably soft metals, and are ductile and malleable. Gold, in particular, is so ductile and malleable that 1.0 g of it can be drawn into a wire 165 m long and 20μ m thick or hammered into a 1.0m2 sheet only 0.050 μ m thick.

a) What are the meanings of the words ‘ductile’ and ‘malleable’?
b) Use a model of metallic bonding and structure to explain why metals are malleable and ductile.
c) If 1.0g of gold is shaped into a cube, what would the dimensions of this cube be? The density of gold is 19.3 g cm^{-3}.
d) How many atoms of gold would there be in this 1.0g cube of gold?
e) How many atoms thick is the sheet described above, that can be made from 1.0g of gold?
f) Suggest several other reasons, apart from softness, malleability and ductility, why gold, silver and copper are used to make jewellery.

17. One type of dry-chemical fire extinguisher suitable for use with electrical and flammable liquid fires contains powdered sodium hydrogencarbonate is blown onto the fire, it decomposes at 60°C to form carbon dioxide and sodium hydrogencarbonate and then the carbon dioxide and the newly formed solid sodium carbonate all smother the fire.

    a) Why is it not safe to use a water-based fire extinguisher on an electrical fire or a flammable liquid fire?
    b) Give the formulas of sodium hydrogencarbonate and sodium carbonate.
    c) Draw electron dot diagrams for sodium hydrogencarbonate, sodium carbonate and carbon dioxide.
    d) Write a balanced equations for the decomposition reaction sodium hydrogencarbonate undergoes when heated.
    e) Suggest an explanation for the way in which each of the three substances puts out a fire.
    f) Explain why, at the temperature of the fire, carbon dioxide is a gas, but sodium carbonate is a solid.

18. Magnesium is a very versatile metal. It is a lightweight soft metal and even though it is quite reactive, it forms a tough impervious oxide layer that prevents further reaction with oxygen. It can be easily rolled, pounded, welded and riveted in virtually any shape, but more importantly, it forms low-density alloys that are strong. When alloyed with other metals such as aluminium and zinc, magnesium is used in objects such as aircraft bodies, car wheels, jet-engine parts, portable power tools, mobile phones, computer laptops and cameras.

    Even though magnesium is not scarce (it is the eighth most abundant element in the Earth's crust), its extraction from the relevant minerals requires a significant amount of energy. This is because of magnesium's reactivity.

    Commercially, two methods are used to produce magnesium – an electrolytic process and a thermal reduction process. In the thermal process, magnesium oxide, obtained from the mineral dolomite, is heated to about 1400°C with an alloy of silicon and iron, to form magnesium vapour:

$$2MgO(s) + Si(s) \rightarrow SiO_2(s) + 2Mg(g)$$

    The vapour is then condensed to form high-purity magnesium.

    In the electrolytic method, the raw material used can either be seawater or a slurry of dolomite. In either case, the first step is the precipitation of magnesium ions from the liquid mixture in a reaction with calcium hydroxide. The magnesium hydroxide formed is then reacted with hydrochloric acid to produce a solution of magnesium chloride. This solution is evaporated and the solid obtained is heated strongly to form anhydrous magnesium chloride.

The magnesium chloride is mixed with calcium chloride and sodium chloride and the mixture is electrolysed in a cell using carbon anodes and steel cathodes. The electrode reactions that occur are:

$Mg^{2+}(l) + 2e^- \rightarrow Mg(l)$
and $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

The molten magnesium floats to the top of the electrolytic mixture, from where it is collected and cast into ingots.

a) The first four successive ionisation energies (in MJ mol^{-1}) of magnesium are 0.744, 1.457, 7.739 and 10.547. Use this data to:

   i explain why magnesium can be easily shaped
   ii explain why magnesium forms an ionic compound when it reacts with non-metals
   iii predict the formula of the ions formed by magnesium in these reactions. Explain your answer.

b) Use electron dot diagrams to show what happens to the valence electrons during the reaction of magnesium with oxygen to form magnesium oxide.
c) Suggest why magnesium is much less dense (1.7g cm^{-3}) than some other commonly used metals such as iron (7.9g cm^{-3}).
d) Why is magnesium not found in its elemental form in the Earth's crust?
e) Explain why magnesium becomes a stronger metal when it is alloyed with other metals.
f) In the high-temperature conditions used in the thermal reduction production of magnesium when magnesium exists as a gas, but silicon dioxide is a solid. Discuss, in terms of bonding and structure concepts, why this is the case.
g) Why can magnesium not be obtained by carrying out the electrolysis of seawater?
h) Write equations for the two reactions that are used to produce magnesium chloride from magnesium ions in seawater.
i) When a solution of magnesium chloride from magnesium ions in seawater.
j) i Draw a labelled sketch of the electrolytic cell that might be used in the industrial production of magnesium from magnesium chloride described in the previois column.
   ii Explain, in terms of their structure and bonding, why steel and carbon can be used as electrodes in this cell.
   iii What product is formed at the anode and what product is formed at the cathode?
   iv Estimate the approximate temperature at which the electrolytic cell operates. Explain your reasoning.
   v Why is it necessary to operate the cell at this temperature?
   vi What is the purpose of mixing sodium and calcium chlorides with the magnesium chloride in the cell?
k) To make the electrolytic production of magnesium more environmentally friendly and economical, suggest what might be done with the chlorine formed in the electrolysis cell.
l) About 1 kg of magnesium can be recovered from 800 L of seawater; so the oceans are a virtually inexhaustible source of this metal.

However, what is the major disadvantage with the production of magnesium by electrolysis?

19. Prior to the 1860s, the major industrial process for the manufacture of sodium carbonate, now used for making glass, water softeners, soaps and detergents, and as a cheap base, was the Leblanc process. The chemistry of this process is summarised in the equations below:

$2NaCl(s) + H_2SO_4(l) \rightarrow Na_2SO_4(s) + 2HCl(g)$
$Na_2SO_4(s) + 4C(s) + CaCO_3(s) \rightarrow Na_2CO_3(s) + CaS(s) + 4CO(g)$

a) The substances involved in these two equations exhibit ionic and covalent bonding. Classify each as either an ionic, covalent molecular or covalent network substance.
b) From an environmental perspective, these two reactions are not regarded as 'friendly'. What are the issues that would ultimately see this process eliminated by the end of the 1860s?

Today, the Solvay process is used to manufacture sodium carbonate. The series of reactions involved in this process is summarised in the following equations:

$NaCl(aq) + CO_2(g) + NH_3(g) + H_2O(l) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$
$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$
$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
$2NH_4Cl(aq) + CaO(s) \rightarrow 2NH_3(g) + CaCl_2(aq) + H_2O(l)$

c) As with the Leblanc process, the compounds involved in these reactions exhibit both ionic and covalent bonding. Classify the reactants and products under the two headings (ionic or covalent molecular compounds) and draw electron dot diagrams for each.
d) Why is it that the reactions summarised by these four equations are more environmentally friendly than the two reactions in the Leblanc process?

Electronegativity

Electronegativity is the electron attracting power of an atom.

Non-metals have a strong attraction for electrons and so have higher electronegativities

What is the trend in electronegativity as you move across the period from lithium to fluorine?

Why is electronegativity of group 18 zero?

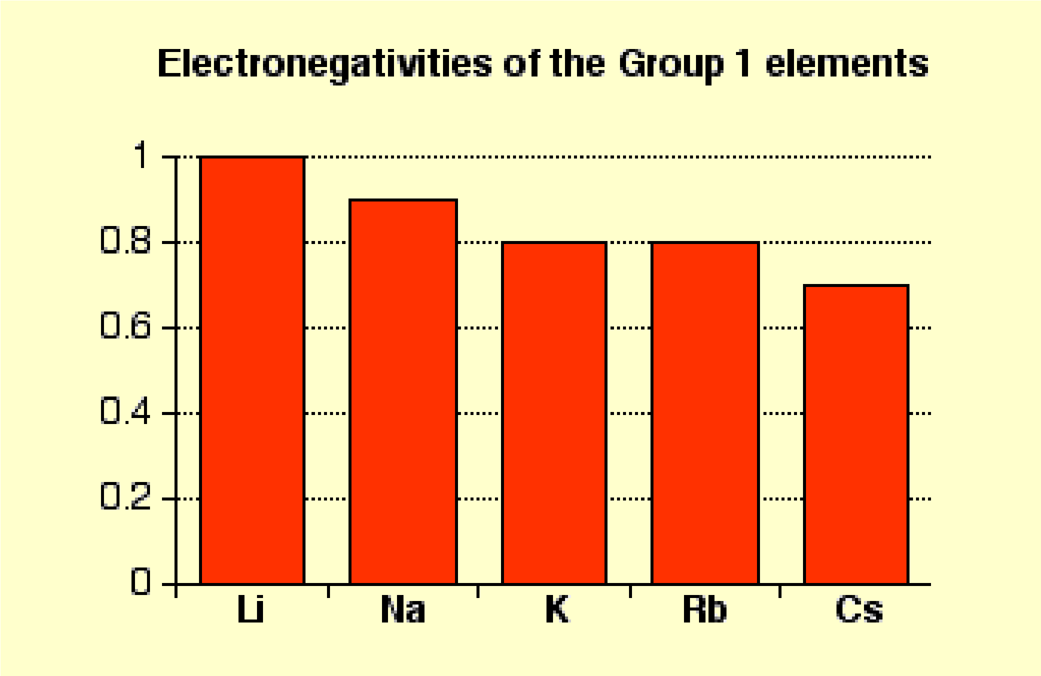
Because they have full outer shells – don't form bonds;don't attract electrons

Table 1: Characteristics of Elements Data Set

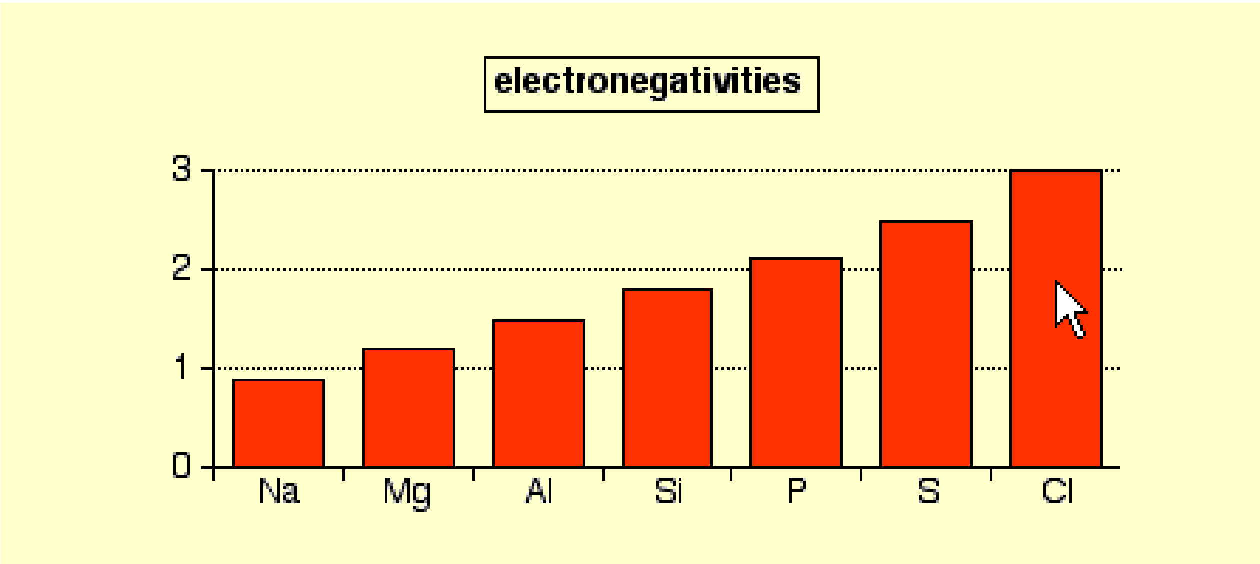
Element	Atomic Number	Period	Group	First Ionization Potential (volts)	Electro-negativity	Boiling Point (°K)
H	1	1	IA	13.598	2.1	20.28
He	2	1	0	24.587	0	1.216
Li	3	2	IA	5.392	0.98	1615
Be	4	2	IIA	9.322	1.57	3243
C	6	2	IVA	11.26	2.55	5100
F	9	2	VIIA	17.422	3.98	85
Ne	10	2	0	21.564	0	27.1
Na	11	3	IA	5.139	0.93	1156
Mg	12	3	IIA	7.646	1.31	1380
Si	14	3	IVA	8.151	1.9	2630
Cl	17	3	VIIA	12.967	3.16	239.18
Ar	18	3	0	15.759	0	87.45

Explain difference in electronegativity between H and Li

Explain difference in electronegativity of Li and F



Trend across a period



Electronegativity increases →

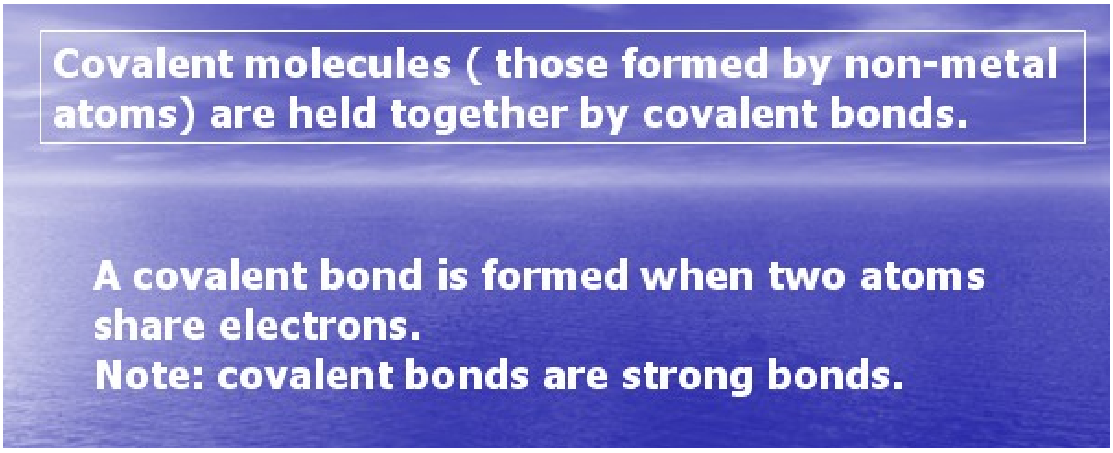
Electronegativity decreases ↓

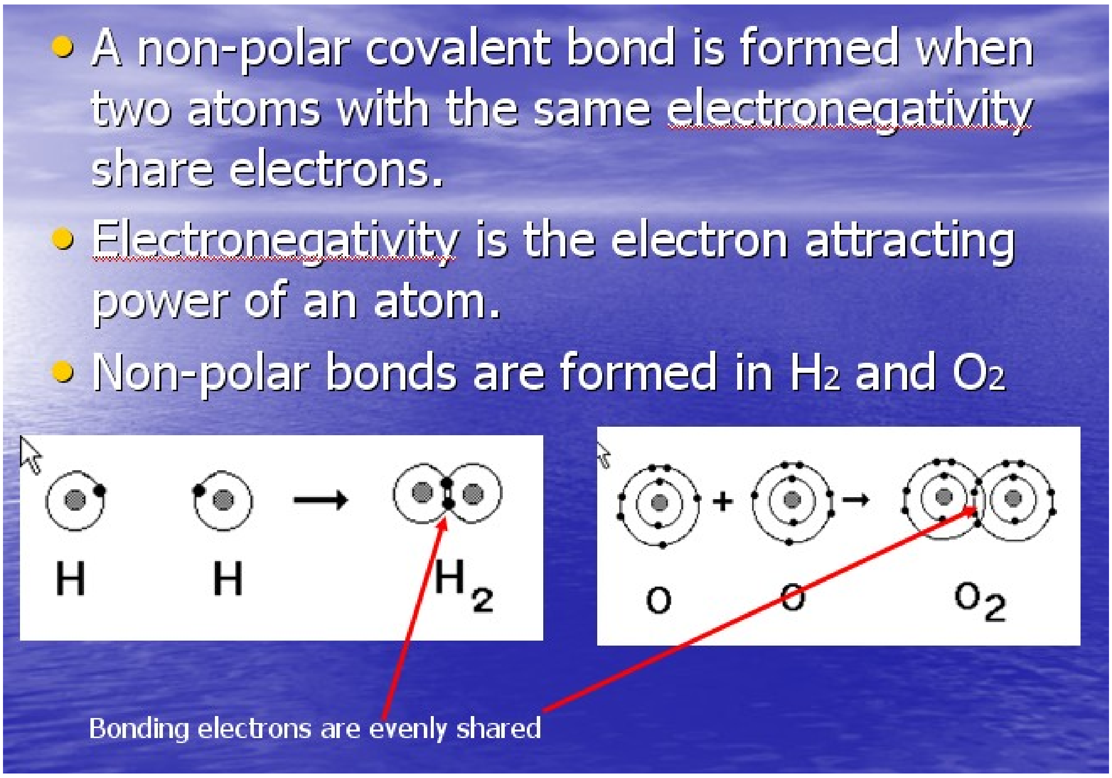
2.20 H	electronegativity / symbol

Period \ Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1																		He
2	0.98 Li	1.57 Be											2.04 B	2.55 C	3.04 N	3.44 O	3.98 F	Ne
3	0.93 Na	1.31 Mg											1.61 Al	1.90 Si	2.19 P	2.58 S	3.16 Cl	Ar
4	0.82 K	1.00 Ca	1.36 Sc	1.54 Ti	1.63 V	1.66 Cr	1.55 Mn	1.83 Fe	1.88 Co	1.91 Ni	1.90 Cu	1.65 Zn	1.81 Ga	2.01 Ge	2.18 As	2.55 Se	2.96 Br	3.00 Kr
5	0.82 Rb	0.95 Sr	1.22 Y	1.33 Zr	1.6 Nb	2.16 Mo	1.9 Tc	2.2 Ru	2.28 Rh	2.20 Pd	1.93 Ag	1.69 Cd	1.78 In	1.96 Sn	2.05 Sb	2.1 Te	2.66 I	2.6 Xe
6	0.79 Cs	0.89 Ba	La*	1.3 Hf	1.5 Ta	2.36 W	1.9 Re	2.2 Os	2.20 Ir	2.28 Pt	2.54 Au	2.00 Hg	1.62 Tl	2.33 Pb	2.02 Bi	2.0 Po	2.2 At	Rn
7	0.7 Fr	0.9 Ra	Ac**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

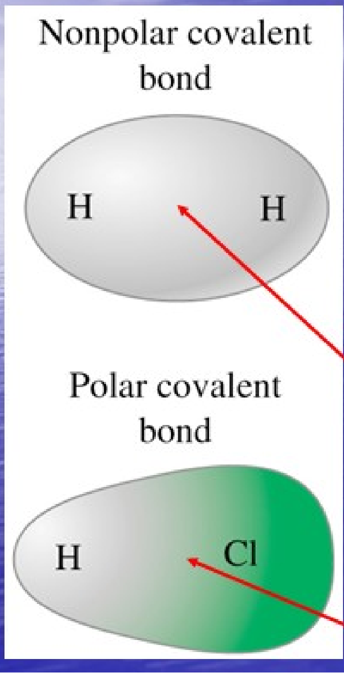
Lanthanides *	1.12 Ce	1.13 Pr	1.14 Nd	1.13 Pm	1.17 Sm	1.2 Eu	1.2 Gd	1.1 Tb	1.22 Dy	1.23 Ho	1.24 Er	1.25 Tm	1.1 Yb	1.27 Lu
Actinides **	1.3 Th	1.5 Pa	1.38 U	1.36 Np	1.28 Pu	1.13 Am	1.28 Cm	1.3 Bk	1.3 Cf	1.3 Es	1.3 Fm	1.3 Md	1.3 No	Lr

Polar and non-polar molecules





Polar bonds

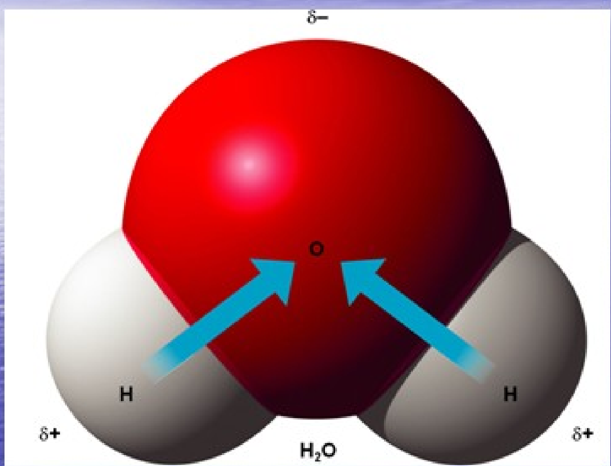


- A polar covalent bond is formed when two atoms with different electronegativities share electrons. The atom with the stronger electronegativity will pull the shared electrons closer to it and become slightly negative.

Electrons evenly shared = non-polar bond

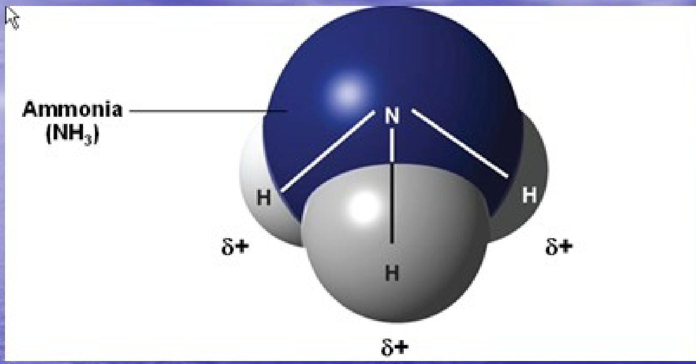
Uneven sharing = polar bond

Polar bonds in water



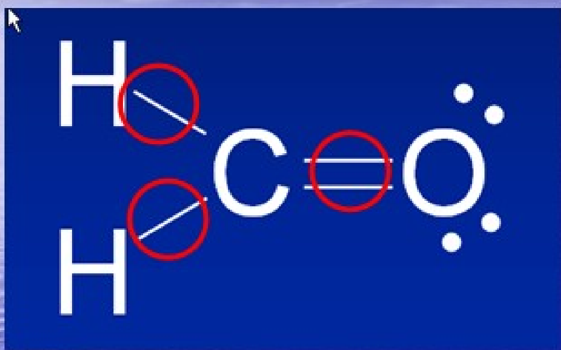
Oxygen atoms are more electronegative than hydrogen atoms so the bonding electrons are pulled closer to the oxygen atom – a polar bond is formed and the oxygen atom becomes slightly negative.

Polar bonds in ammonia



Nitrogen is more electronegative than hydrogen. This means that the bonding electrons lie closer to the nitrogen atom and it becomes slightly negative.
A polar covalent bond is formed.

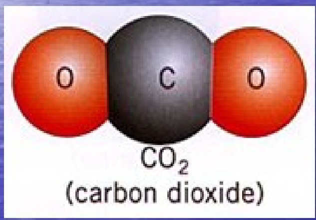
Circle the polar covalent bonds in the diagrams below.



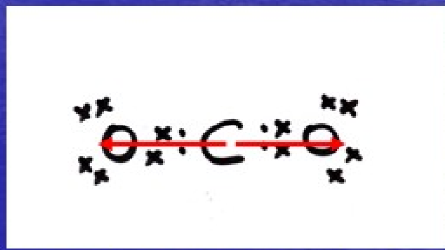
All of these bonds are between two different atoms so they will all be polar bonds.

Non-polar molecules

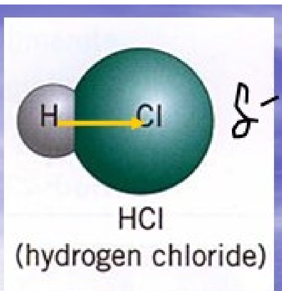
- A non-polar molecule is one where the electrons are symmetrically (evenly) distributed throughout the molecule.



CO_2
(carbon dioxide)



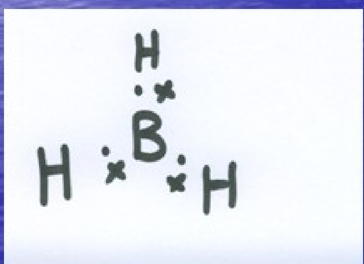
Oxygen is more electronegative than C and so the bonding electrons are pulled towards the O atom. Even though polar bonds are formed the electrons in the molecule are still symmetrically distributed.

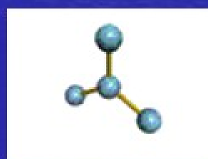


HCl
(hydrogen chloride)

Hydrogen chloride is a polar molecule because the bonding electrons are pulled towards the more electronegative Cl atom. A permanent negative dipole is formed on the Cl atom

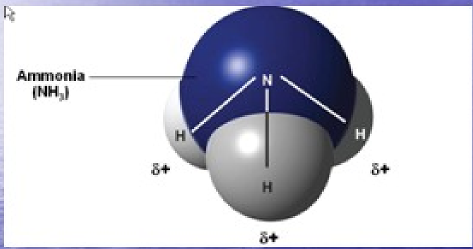
BH_3 is a non-polar molecule. B is more electronegative than H so 3 polar bonds are formed. However the arrangement of electrons in the molecule is symmetrical so the molecule is non-polar.





Test Yourself

- Are the molecules below polar or non-polar?



polar

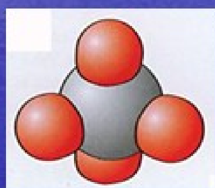
$H-C\equiv N:$

polar

Test yourself

- Are the following molecules polar or non-polar?

  Methane (CH_4)



Non-polar

Dichloromethane (CH_2Cl_2)

Polar

Intermolecular forces ref page 57 textbook

So far we have studied **strong bonds**

- attraction between positive and negative ions in a lattice ___________________
- attraction between positive metal ions and delocalised valence electrons__________________
- attraction between two nuclei and one or more shared electron pairs __________________________

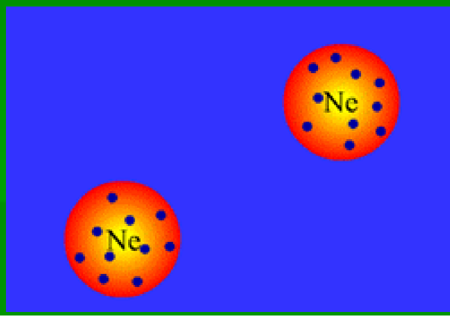
Now we are going to consider
Weak intermolecular forces (van der Waal forces)

- attraction between all types of particles resulting from temporary dipoles ______________________________
- attraction between polar molecules ______________________
- attraction between polar molecules containing a H atom directly bonded to an N, O or F atom _______________________
- attraction between polar solvent molecules and dissolved ions _____________________________

Dispersion forces

- Intermolecular forces between non-polar molecules arise from the fact that at any instant a molecule may have a temporary dipole. The temporary dipole occurs at any moment when the electrons in an atom or molecule are not symmetrically distributed.

The animation labels the force as Van der Waal forces – we call them Dispersion forces

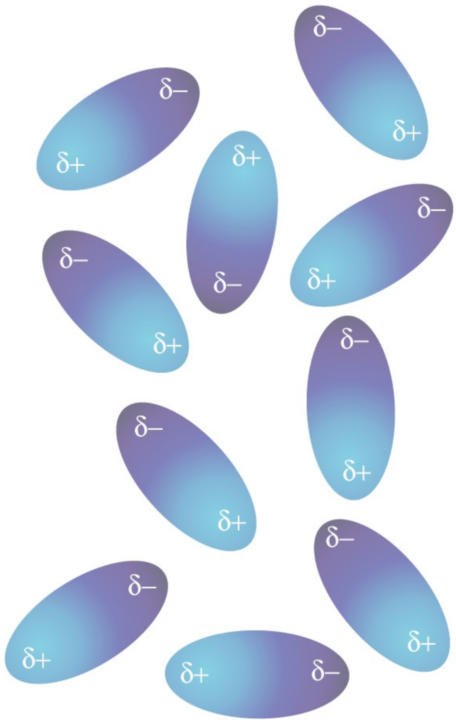


- Dispersion forces between molecules increase in strength with the number of electrons present in the molecule.

Halogen	Number of electrons	Melt point ($_0$C)
F_2	18	-220
Cl_2	34	-101
Br_2	70	-7
I_2	106	114

What are dipole-dipole forces?

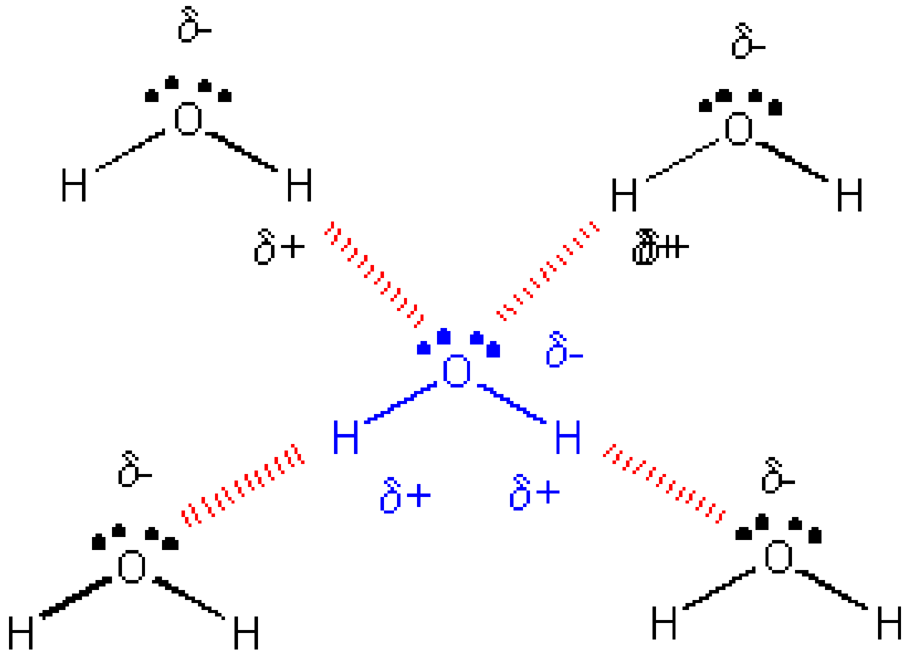
- They are forces of attraction between the oppositely charged ends of neighbouring polar molecules.
- All polar molecules have dipole-dipole forces.
- Non-polar molecules do not have dipole-dipole forces.

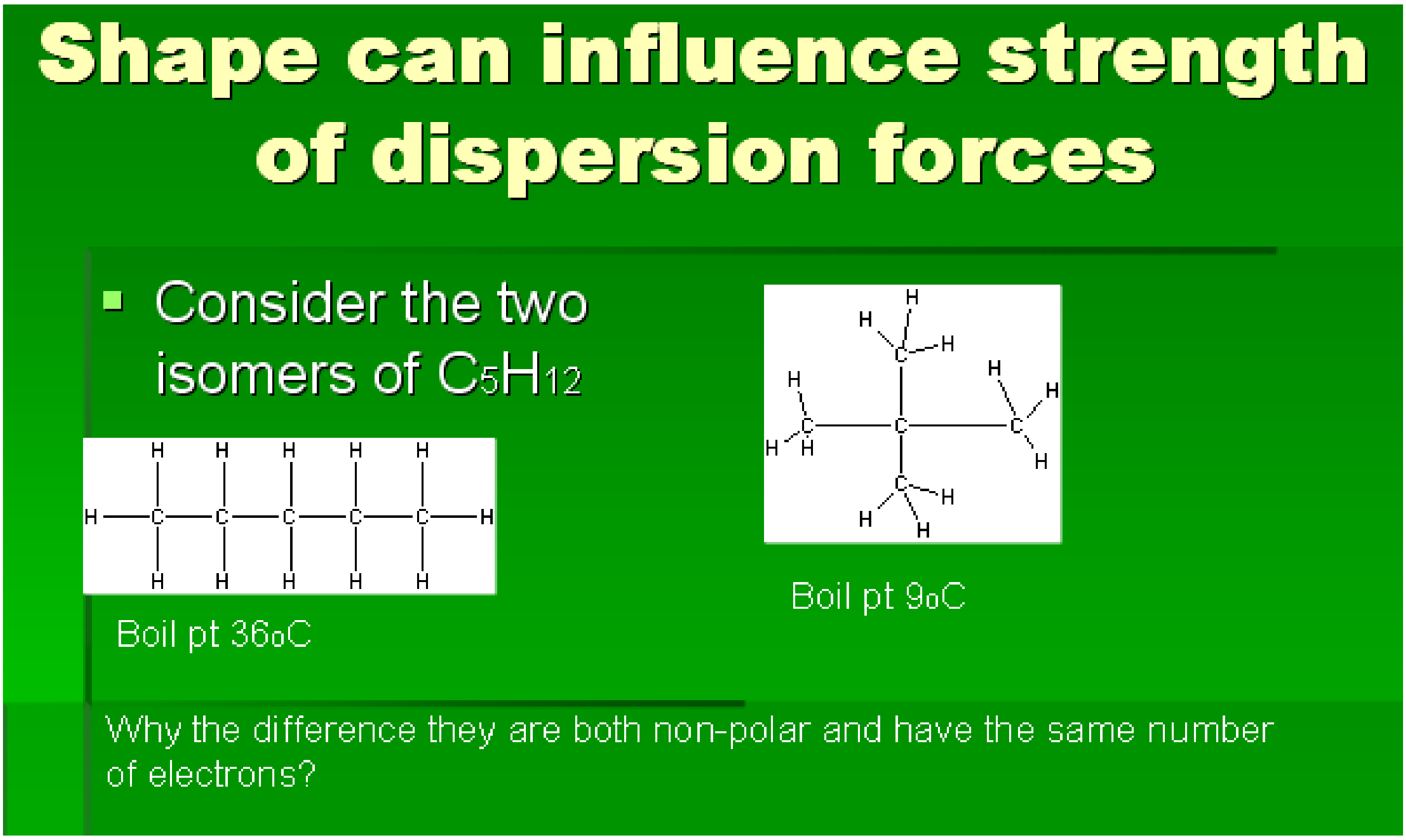


In a dipole substance, the molecules are oriented so the positive region of one molecule faces the negative region of another molecule.

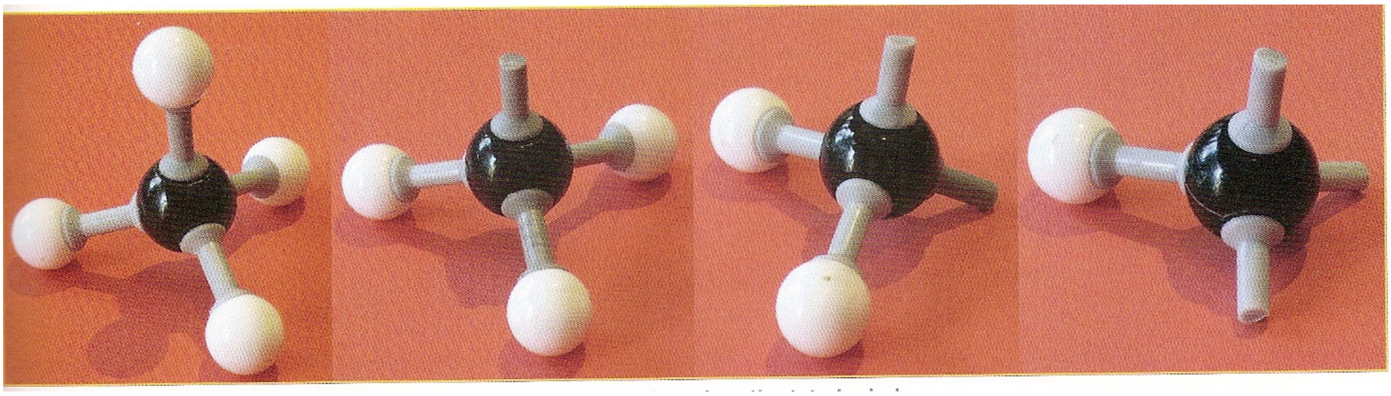
Hydrogen bonding

- This intermolecular force occurs in molecules that have a hydrogen atom directly bonded to an atom of N,O or F. It is sometimes referred to as H-bonding.
- The N,O or F atoms are very electronegative and this leads to a significant positive charge developing on the H atom. This charge is attracted to a lone pair of electrons on a neighbouring molecule. This attraction is referred to as a H-bond.





Shape can influence strength of dispersion forces
Consider the two isomers of C5H12
Boil pt 36oC
Boil pt 9oC
Why the difference they are both non-polar and have the same number of electrons?



12 Chemistry – Shapes of Molecules

1. Make a model of dichloromethane CCl_2H_2

   a) Sketch a diagram of the model

   b) Is the bond angle between the Cl atoms ever 180^o? Explain your answer.

   c) Draw an electron dot diagram for this structure

   d) What shape is this molecule?

2. Make a model of NH_3

   a) Sketch a diagram of the model

   b) Are the hydrogen atoms all in one plane?

   c) Draw an electron dot diagram for this structure

   d) What shape is this molecule?

3. Make a model of ethene CH_2CH_2

a) Sketch a diagram of the model

b) Are all the atoms in one plane?

c) Draw an electron dot diagram for this structure

d) What shape is this molecule?

4. Make a model of H_2O

a) Sketch a diagram of the model

b) Is the bond angle between the two H atoms ever 180^o? Explain your answer.

c) Draw an electron dot diagram

d) What shape is this molecule?

Review exercise 3.1 page 58 textbook

1. The boiling points of three liquids used as non-aqueous solvents are:
   ethanol 78°C hexane 69°C tetrachloroethene 121°C

   a) Which of these substances would have the strongest intermolecular forces?
   b) Which of these liquids will therefore be the most volatile? Explain your reasoning.

2. Of the following changes, which involve the breaking of only intermolecular forces?

   a) Methane is burnt in air
   b) A block of dry ice (solid carbon dioxide) sublimes
   c) A piece of sodium metal is cut in half
   d) Sodium chloride dissolves in water
   e) Sugar dissolves in water
   f) Cooking oil is boiled
   g) A puddle of water evaporates
   h) Electricity is passed through water to form oxygen and hydrogen
   i) A sticky note is removed from a piece of paper

3. Explain why, when crushing ice, the ice breaks up but the metal blades of the ice-crusher are not affected.

Review exercise 3.2 page 63 textbook

1. What is the basic electrostatic principle upon which the VSEPR hypothesis is based?

2. Draw line structures and state the shape of the molecules of the hydrides of the element silicon to chlorine in period 3 of the periodic table.

3. a) Draw an electron dot diagram of ethyne, C_2H_2, and hydrogen peroxide, H_2O_2.
   b) Describe the differences in the intermolecular bonding in these two molecules.
   c) Describe the difference in shape of these two molecules. Explain why there is a difference.

4. Line diagrams of various molecules are depicted in Figure 3.6 (these diagrams do not necessarily indicate the correct bond angles). For each of these molecules describe the correct shape and provide an estimate of the bond angle(s).

a) b) c) d) e)

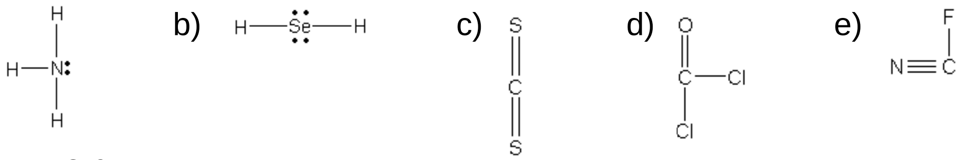


Figure 3.6

5. Draw line structures of the following molecules and ions, then identify their shape and give the approximate bond angles.

   a) CH_3F
   b) HBr
   c) OH^-
   d) C_2F_2
   e) OCl_2
   f) SO_4^{2-}
   g) PO_3^{3-}

Review exercise 3.3 page 66 textbook

1. Draw line structures of the following molecules and identify the covalent bonds as either non polar or polar:

   a) HBr
   b) $SiCl_4$
   c) SO_2
   d) N_2H_4

2. The following molecules have polar covalent bonds. Draw line structures of the molecules and indicate which end of each bond will have a partial positive charge and which end will have partial negative charge.

   a) HI
   b) NBr_3
   c) OCl_2

3. a) Describe how electronegativity changes across a period and down a group in the periodic table.
   b) Explain, using the concept of ‘core charge’, why fluorine is more electronegative than oxygen.
   c) Explain why chlorine is less electronegative than fluorine.

Review exercise 3.4 page 72 textbook

1. What is the difference between the terms 'polar covalent bond' and 'polar molecule'?

2. a) Which of the following molecules and ions would be regarded, in chemistry, as symmetric.

   CBr_4 H_2S HI CO_2 $CHCl_3$ SO_4^{2-} NCl_3 HBr
   NO_2^- SO_3^{2-} CH_3F SCl_2 C_2H_4

   b) Name the shapes of the molecules listed in part a).

3. Classify each of the following as an ionic, a polar covalent molecular or a non-polar covalent molecular substance.

   $SiCl_4$ $MgCl_2$ CH_3Br Cl_2O PH_3 SrO F_2
   $CH_3CH_2CH_2CH_3$

4. A positively charged Perspex rod is held near a thin stream of water from a tap. The stream is attracted towards the rod. Explain this observation and draw a diagram to show what is happening on the molecular scale to produce this phenomenon.

Review exercise 3.5 page 74 textbook

1. Which of the following substances will have dipole-dipole forces between their molecules?

   Br_2 CH_3Br CH_2Br_2 SBr_2 CBr_4 HBr NBr_3 $SiBr_4$

2. Explain the following observations.

   a) Acetone, $(CH_3)_2CO$, is a liquid at room temperature, but butane, C_4H_{10}, is a gas.
   b) Carbon monoxide has a higher melting point (-205°C) than oxygen (-219°C).

3. Predict the substances with the higher boiling point, in each pair. Give a reason for your answer.

   a) SiH_4 or PH_3
   b) PH_3 or NH_3
   c) Na_3P or PH_3

Review exercise 3.6 page 79 textbook

1. Which of the following substances have only dispersion forces between their molecules?

   O_2 HCl H_2O CO_2 SO_2 C_2H_2 CH_3-CH_3 PF_3 F_2

2. Why are covalent bonds stronger than dispersion forces?

3. Use the data Table 3.6 to explain the nature of the bonding present between the molecules and any trends in that bonding.

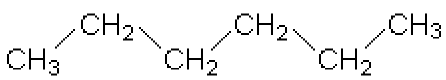
   TABLE 3.6

Compound	Boiling point (°C)
CHF_3	-84
CF_4	-128
$CHCl_3$	62
CCl_4	77

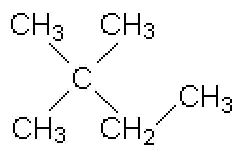
4. Describe the similarities and differences between dispersion forces and dipole-dipole forces.

5. a) Which of bromine, Br_2, or xenon, Xe, would be predicted to have the higher boiling point? Explain your answer.
   b) Use the arguments developed in answering part a) to make a prediction about the relative boiling points of fluorine, F_2, and argon, Ar.

6. Predict which of hexane and 2,2-dimethylbutane, shown below, will have the higher boiling point. Explain your answer.



hexane



2,2-dimethylbutane

Review exercise 3.7 page 84 textbook

1. Identify the type or types of intermolecular forces in the following substances.

   $CHCl_3$ CH_2F_2 CBr_4 CH_3OH $CH_3CH_2NH_2$ HCN S_8 SO_3 H_2S

2. Explain the origin of each of the three types of can der Waals forces and explain why one is stronger than the other two.
3. The melting points of the group 17 hydrides are listed below:

Substance	HF	HCl	HBr	HI
Melting point (°C)	-85	-114	-87	-51

   Explain the following trends shown in these figures, in terms of the intermolecular forces in the substances.

   a) Going down the group from HCl to HBr to HI, the melting points increase.
   b) HF has a higher melting point than HCl.

4. For each of the following pairs of substances, predict the one with the higher melting point. Give a brief reason for your answer.

   a) F_2 or HF
   b) HF or H_2O
   c) NaCl or HCl
   d) NH_3 or PH_3
   e) F_2 or Cl_2
   f) CH_3CH_2OH or CH_3OH
   g) CH_4 or C_2H_6

5. The formula of sulfuric acid can be written as $(OH)_2SO_2$ to give a clearer idea of how the individual atoms are connected in the molecule. Sulfuric acid boils at 320°C. Fluorosulfuric acid has a boiling point of 170°C, while sulfuryl fluoride, SO_2F_2, has a boiling point of -55°C.

   a) Draw line structures of the molecules of these three substances.
   b) Account for the differences in the boiling points in terms of the intermolecular force present.

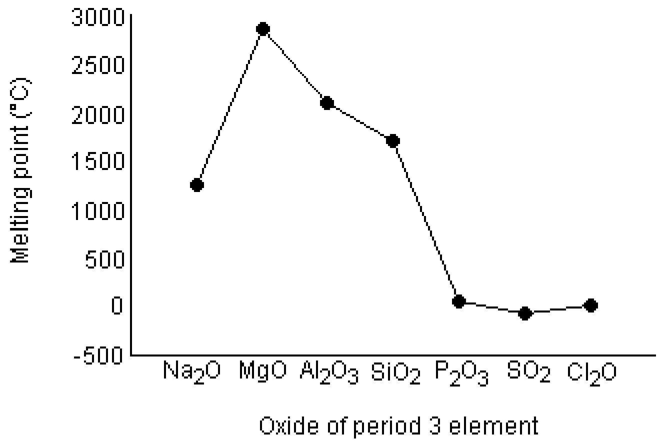
6. Between which of the following pairs of substances will hydrogen bonding occur?

   a) Water and ethanol (CH_3CH_2OH)
   b) Ammonia and water
   c) Phosphine, PH_3, and water
   d) CH_3OCH and HF
   e) CH_3COCH_3 and H_2NNH_2
   f) CF_4 and water

Review exercise 3.9 page 87 textbook

1. The trend in the melting points of oxide of the period 3 elements are shown in the graph in Figure 3.27.

   a) Draw electron dot or line representations of each compound.
   b) Explain the melting point trends shown in this group of compounds.



QUESTIONS page 88 textbook

1. For each of the following pairs of substances predict which one of the pair will have the higher boiling point and explain your answer.

   a) C_3H_8 or C_4H_{10}
   b) SiH_4 or CH_4
   c) HCl or HI
   d) H_2S or H_2O

2. a) Predict the shapes and bond angles of the following molecules or ions.

i	$SiCl_4$	ii	NI_3
iii	SO_3	iv	SO_4^{2-}
v	H_2O	vi	CO_2
vii	SO_2	viii	NO_3^-
ix	AsH_3		

   b) For the neutral molecules in part a), which ones have polar bonds?
   c) For the neutral molecules in part a), which ones are polar?

3. a) Which of the following molecules or ions have a tetrahedral shape?

   SiH_4 PH_4^+ SO_4^{2-} NH_3 CH_2Br_2 Cl_2 H_2S

   b) Which of the following molecules or ions have a linear shape?

   CH_4 HCl PCl_3 SO_2 N_2 Br_2 H_2S CO_2 NO_2^-

   c) Which of the following molecules have a V-shape?

   SO_2 PH_3 H_2S OF_2 CO_2 CS_2 CCl_4

   d) Which of the following molecules or ions have a pyramidal shape?

   NCl_3 NCl_3 $CHCl_3$ PH_4^+ H_3O^+

   e) Which of the following molecules or ions have a trigonal planar shape?

   PCl_3 BBr_3 CCl_4 NH_3 H_3O^+ BCl_3 SO_3

4. The order of boiling points of water, ethanol, CH_3CH_2OH, and ethoxyethane, $CH_3CH_2OCH_2CH_3$, is ethoxyethane < ethanol < water. However, for the sulfur analogues the order is H_2S < CH_3CH_2SH < $CH_3CH_2SCH_2CH_3$. Explain this difference.

5. Draw a line structure of a molecule that fulfils each of the following descriptions.

   a) A tetrahedral polar molecule formed between a period 2 element, iodine and hydrogen.
   b) A V-shaped molecule formed between a period 3 element and chlorine.

c) A molecule, formed between a period 2 elements and hydrogen, that contains four pairs of valence electrons but just one non-bonding pair of electrons.
d) A molecule, formed between a period 3 element and hydrogen, that has four valence pairs of electrons with three non-bonding pairs of electrons.

6. Describe the attractive forces that have to be overcome in each of the following changes.

a) $Br_2(l) \rightarrow 2Br(g)$
b) $KCl(s) \rightarrow K^+(l) + Cl^-(l)$
c) $HF(l) \rightarrow HF(g)$
d) $SO_2(l) \rightarrow SO_2(g)$

7. What is the name of the force or forces that opposes vaporisation of the following substances?

a) hexane
b) fluoroethane, CH_3CH_2F
c) hydrogen peroxide, H_2O_2

8. Consider the following melting points.

Substance	CCl_4	CF_4	NaCl	O_2	NO
Melting point (°C)	-23	-187	801	-219	-164

Discuss why:

a) CCl_4 has a higher melting point than CF_4.
b) NaCl has a higher melting point than CCl4.
c) O2 has a lower melting point than NO.

9. Consider the following melting points.

Substance	H_2	F_2	HF	He	Ne	H_2O	CH_4
Melting point (°C)	-259	-220	-83	-272	-249	0	-183

a) Propose a reason for the difference in melting points of hydrogen and fluorine.
b) Discuss why hydrogen fluoride has a higher melting point than either fluorine or hydrogen.
c) Explain the difference in the melting points of fluorine and neon.
d) Propose an explanation for the observation that CH_4 has a lower melting point than HF and H_2O but a similar melting point to the four elements.

10. Dispersion forces are the only intermolecular forces present in paraffin wax yet it has reasonably high melting and boiling points for a covalent molecular compound. Suggest a possible explanation for this.

11. The boiling points of the hydrides of group 15 elements are shown below.

Hydride	NH_3	PH_3	AsH_3	SbH_3
Boiling point	-33°C	-90°C	-59°C	-21°C

a) Why do the boiling points increase with increasing molecular mass for the hydrides PH_3, AsH_3 and SbH_3?
b) Why does ammonia have a much higher boiling point than expected from the trend shown by the other hydrides?

12. Explain why:

a) bromine is a liquid at room temperature but chlorine is a gas
b) ethanol, C_2H_5OH, is a liquid at room temperature but carbon dioxide is a gas.

13. a) What does the heat of vaporisation, ΔH_{vap}, measure?
b) The heats of vaporisation of three group 15 hydrides are shown below.

Hydride	ΔH_{vap} (kJ mol^{-1})
NH_3	23.4
PH_3	14.6
AsH_3	17.5

Plot a graph of ΔH_{vap} against number of electrons for each of these hydrides.
c) Explain the observed values for the heats of vaporisation of the three hydrides.
d) If ammonia did not behave anomalously what value of ΔH_{vap} would you predict for NH_3 if it continued the trend shown by the other two hydrides?
e) Determine the strength of the hydrogen bond in ammonia from this information.
f) The strength of the hydrogen bonding interaction in water has been determined to be around 22 kJ mol^{-1}. Explain the difference between the value for the strength of the hydrogen bonding interaction in ammonia and that in water.

14. In which of the following solutions would there be hydrogen bonding between the solute and solvent molecules?

a) HCl dissolved in water
b) CH_3CH_2OH dissolved in water
c) Acetic acid, CH_3COOH, dissolved in ethoxyethane, $CH_3CH_2OCH_2CH_3$
d) NH_3Cl dissolved in water
e) NH_3 dissolved in CH_3OH
f) HF dissolved in CH_3COCH_3
g) CCl_4 dissolved in $CH_3CH_2CH_2OH$

15. Explain why it is possible to separate oxygen and nitrogen by fractional distillation of liquid air. Which of oxygen and nitrogen will have the higher boiling point? Justify your response.

16. In which of groups 1, 2, 14, 15 and 18 of the periodic table would you expect to find an element that satisfies the conditions stated below?

    a) An element having the highest possible first ionisation energy
    b) An element having the lowest possible ionisation energy
    c) An element that forms the molecular compounds XO2 and XH4
    d) An element that forms the ionic compounds ZCl2 and ZS
    e) An element having an atomic number of 33
    f) An element with one electron in its valence shell
    g) An element that reacts with one hydrogen to form a molecule that has a pyramidal shape

17. Explain why the melting points and boiling points of group 1 elements decrease down the group, but the melting points and boiling points of group 17 elements increase down the group.

18. Figure 3.28 shows some chlorides of elements from period 3 at room temperature. From left to right these are sodium chloride, magnesium chloride, aluminium chloride, silicon tetrachloride, phosphorus trichloride and phosphorus pentachloride.

    a) Write the formulas of the six chlorides.
    b) Describe the bonding in each of these six chlorides, including an electron dot diagram for each compound. (The last chloride is a non-octane compound.)
    c) Explain why the first three chlorides are solid but the next two chlorides are liquids.
    d) Explain why one of the phosphorus chlorides is a liquid but the other is a solid at room temperature.

19. When you use a pencil to write on a piece of paper, layers of graphite from the pencil are left on the paper during the writing process.

    a) Graphite is regarded as non-polar. Why is this?
    b) Paper is regarded as polar because it essentially composed of cellulose, which is polymer of glucose. (See Chapter 5, page 151, for the structural formula of glucose.) What type of bonding exists between graphite and paper?
    c) Why can a pencil mark on paper be easily removed with an eraser?

    One of the dyes in a particular brand of ink has the structure shown on the right.

    d) What type of intermolecular bonding could this dye exhibit with paper? What feature in the dye molecule leads you to this conclusion?
    e) Comment on whether it will be easier or harder to remove this dye from the paper than to remove the graphite.

Several years ago erasable pens were in fashion. These pens worked by leaving a layer of synthetic rubber (impregnanted with a dye strongly attracted to the rubber) on the paper. The rubber in one of the patented pen varieties is made by joining together many molecules having the structure shown below.

f) Explain how erasable pens work in terms of the intermolecular forces between the layer of black rubber and the paper.
g) What would you predict about the polarity of the black dye used to impregnate the rubber?

20. The industrial process for the production of ammonia involves the reaction between hydrogen and nitrogen in the presence of a suitable catalyst to form ammonia:

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

During the reaction, the products are cooled under pressure and the ammonia liquefies. The unreacted hydrogen and nitrogen are recycled.

a) Draw electron dot diagrams of the three molecules involved in this reaction.
b) Which will have the higher boiling point, hydrogen or nitrogen? Explain your answer.
c) How is it possible that the ammonia can be separated so easily from the other two gases?

21. In the laboratory production of hydrogen chloride, solid sodium chloride is heated with liquid sulfuric acid and immediately fumes of hydrogen chloride gas are evolved. In terms of the intermolecular bonding involved, explain why sulfuric acid is a liquid but hydrogen chloride is a gas under these conditions.

22. The production of methanol is achieved by the reaction of carbon monoxide with hydrogen in the presence of a suitable catalyst. The overall process can be represented simply as:

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

Using the knowledge of bonding, suggest the methanol can be separated from any unreacted starting materials and explain your answer in terms of the bonding interactions involved.

23. Analysis of a recently discovered natural gas source shows the following composition:

Component	Percentage by volume
Methane	79.5
Ethane	7.2
Nitrogen	8.4
Carbon dioxide	0.5
Propane	2.8
Butane and other alkanes	1.6

In order to produce a commercially useful natural gas supply, some of these components have to be separated.

a) Knowing that carbon dioxide can be classified as an acidic oxide, suggest how it might be separated chemically from the other gases.
b) In devising a method for separating the methane from the other components, there is no need to remove the nitrogen. Why is this the case?
c) How would ethane and other alkanes be separated from the methane? Explain your answer in terms of the intermolecular bonding between the molecules of the alkanes.

24. It has been argued that there is no such thing as an ionic bond. What are the reasons that are proposed to support this statement? Is it them possible for something to have a ‘pure’ covalent bond? Explain your answer.

Chemistry 12 2009 **Name**
Bonding Assignment

Answer these questions in the spaces provided.

1. This question asks you to predict the properties of elements from their positions in the Periodic Table. The symbols for ten elements are shown. Answer the questions about these elements.

Li																	Ne
																Cl	
				V							Zn		Ge			Br	
																I	
	Sr																
Cs																	

a) Write the symbol of the element with the highest electronegativity.

b) Write the formula for a covalent molecular compound that could be formed by combining two of the elements

c) Write the symbol for the element with the lowest first ionisation energy.

d) Write the formula for the carbonate of Cs

e) Write the symbol for the element that exists as a covalent molecular solid at room temperature

(5 marks)

2. Place the following substances in the appropriate column based on the most significant type of intermolecular force present. (4 marks)

C_2H_5OH, CH_3Cl, H_2O, CH_2F_2, BH_3, NI_3, CS_2, HF

Hydrogen bonding	**Dipole-dipole interactions**	**Dispersion forces**

4. For each of the species listed in the table below draw the structural formula, including all valence electrons, and sketch the shape.

Species	Structural formula	Shape
Amide ion NH_2^-		
Ammonium ion NH_4^+		
Azide ion N_3^-		

(6 marks)

5. For each of the following pairs of substances, state which has the higher melting point and indicate which force is responsible for the difference.

Pair of substances	Higher melting	Force
Cl_2 and P_4		
CH_3OH and C_2H_6		
CO_2 and SiO_2		
Mg and Ba		

(6 marks)

6. Fill in the boxes in the table below with a species, chosen from the list provided, that matches the description in the box. Only one answer per box is required.
CH_4 CH_2O CH_2Cl_2 NO_2^- HF Na^+ SO_3 Cl_2 H_2O NH_3

A bent, polar species	A non-polar species	A species that can form hydrogen bonds between its molecules
A tetrahedral, polar species	A species that contains only non-polar bonds	A pyramidal species

(6 marks)

7. The following table gives some data about three elements in the fourth row of the Periodic Table.

Element	**Melting point °C**	**First ionisation energy MJ mol^{-1}**	**Electrical conductivity MSm^{-1}**
Potassium	63	0.43	14
Germanium	937	0.77	10^{-6}
Bromine	-7	1.15	10^{-16}

Account for the way in which the values relate to the structure of the elements at the atomic level.

(6 marks)

Chemistry 12 2009 Name
Bonding ANSWERS

1. This question asks you to predict the properties of elements from their positions in the Periodic Table. The symbols for ten elements are shown. Answer the questions about these elements.

Li																	Ne
																Cl	
				V							Zn		Ge			Br	
																I	
	Sr																
Cs																	

a) Write the symbol of the element with the highest electronegativity. Cl

b) Write the formula for a covalent molecular compound that could be formed by combining two of the elements
ClBr

c) Write the symbol for the element with the lowest first ionisation energy.
Cs

d) Write the formula for the carbonate of Cs Cs_2CO_3

e) Write the symbol for the element that exists as a covalent molecular solid at room temperature I

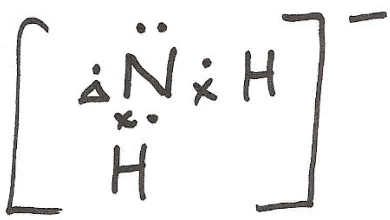
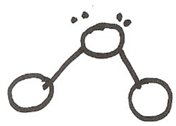
(5 marks)

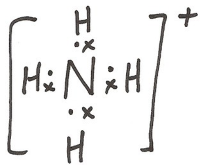
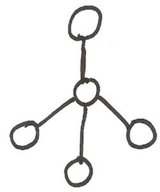
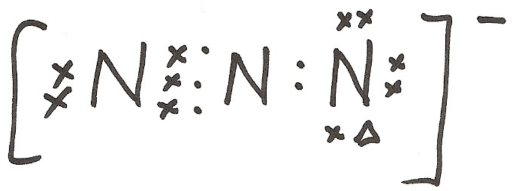
2. Place the following substances in the appropriate column based on the most significant type of intermolecular force present. (4 marks)

C_2H_5OH, CH_3Cl, H_2O, CH_2F_2, BH_3, NI_3, CS_2, HF

Hydrogen bonding	Dipole-dipole interactions	Dispersion forces
C_2H_5OH	CH_3Cl	BH_3
H_2O	CH_2F_2	CS_2
HF	NI_3	

4. For each of the species listed in the table below draw the structural formula, including all valence electrons, and sketch the shape.

Species	Structural formula	Shape
Amide ion NH_2^-	2 singles 2 lone pairs 	Bent 

Ammonium ion NH_4^+	4 singles 	Tetrahedral 
Azide ion N_3^-	1 triple, 1 single 	Linear 

(6 marks)

5. For each of the following pairs of substances, state which has the higher melting point and indicate which force is responsible for the difference.

Pair of substances	Higher melting	Force
Cl_2 and P_4	P_4	Dispersion force
CH_3OH and C_2H_6	CH_3OH	Hydrogen bond
CO_2 and SiO_2	SiO_2	Covalent bond
Mg and Ba	Mg	Metallic bond

(6 marks)

6. Fill in the boxes in the table below with a species, chosen from the list provided, that matches the description in the box. Only one answer per box is required.

CH_4 CH_2O CH_2Cl_2 NO_2^- HF Na^+ SO_3 Cl_2 H_2O NH_3

A bent, polar species H_2O	A non-polar species Cl_2	A species that can form hydrogen bonds between its molecules HF
A tetrahedral, polar species CH_2Cl_2	A species that contains only non-polar bonds Cl_2	A pyramidal species NH_3

(6 marks)

7. The following table gives some data about three elements in the fourth row of the Periodic Table.

Element	**Melting point °C**	**First ionisation energy MJ mol^{-1}**	**Electrical conductivity MSm^{-1}**
Potassium	63	0.43	14
Germanium	937	0.77	10^{-6}
Bromine	-7	1.15	10^{-16}

Account for the way in which the values relate to the structure of the elements at the atomic level.

(6 marks)

Bromine is a covalent molecular liquid at room temperature. Dispersion forces are easily disrupted. The lack of charged species accounts for the lack of conductivity.
Germanium appears to be a covalent network solid as it has a high melting point and poor conductivity.
Potassium is a metal. The attraction of delocalised valence electrons for the small nucleus accounts for the low melting point and the conductivity.

ST MARY'S ANGLICAN GIRLS' SCHOOL

FIDELITER

FOUNDED 1921

YEAR 12 CHEMISTRY TEST 3 (2009)

PUT A CROSS (X) THROUGH THE CORRECT ANSWER.

1.	a	b	c	d
2.	a	b	c	d
3.	a	b	c	d
4.	a	b	c	d
5.	a	b	c	d
6.	a	b	c	d
7.	a	b	c	d
8.	a	b	c	d
9.	a	b	c	d
10.	a	b	c	d
11.	a	b	c	d

PART A / 11	
PART B / 40	
TOTAL / 51	

PART B - Answer all questions in the spaces provided.

Species	Electron Dot Diagram	Shape (sketch or name)
Sulfur dioxide, SO_2		
Sulfate ion, SO_4^{2-}		
Nitrogen trichloride, NCl_3		

1. For each species listed in the table below
   a) Draw the electron dot diagrams showing **all** valence shell electron pairs.
   b) Indicate the shape of the species by either a sketch or a name.

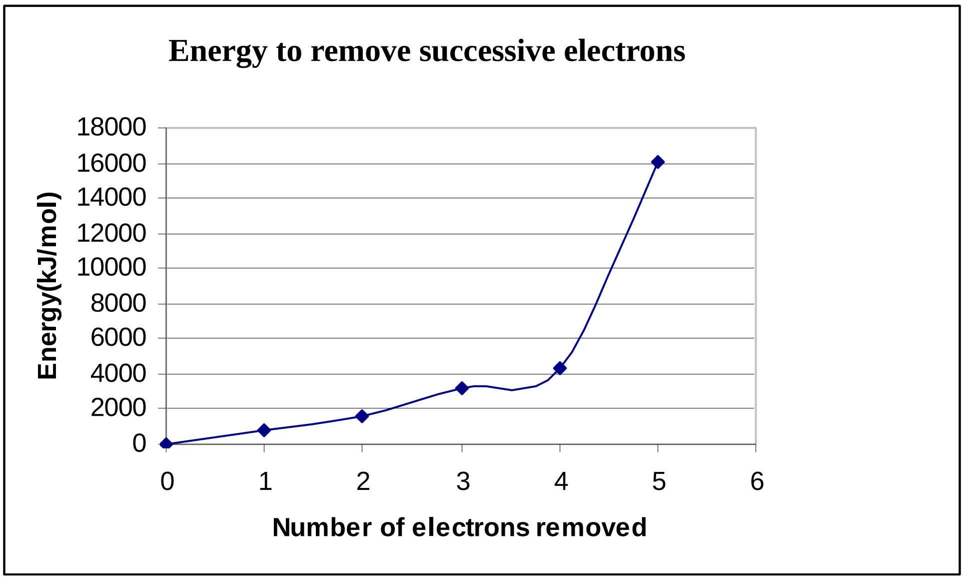
(9 marks)

2. Limestone contains the ionic compound, $CaCO_3$. Limestone decomposes when it is heated strongly, forming an ionic compound, CaO and a covalent compound, CO_2.
   a) State what is meant by ionic bonding.

   b) Draw the electron dot diagrams to show the bonding in CaO and CO_2.

(5 marks)

3. An element from the third row of the periodic table has five successive electrons removed. The energy required is shown on the graph below.



a) Name and describe the type of bonding forces in this element. Justify your choice.

b) Give three properties of this type of material.

(6 marks)

4. a) Which element would you expect to have the higher first ionisation energy — sodium Na or chlorine Cl? Explain your reasoning.

b) Explain why potassium atoms readily form K^+ ions in chemical reactions, while
chlorine atoms readily form Cl^- ions.

(4 marks)

5.

Name a group of elements for which the outermost energy level contains only *s* electrons.	
Name a period of elements for which the outermost energy level contains only *s* and *p* electrons.	
Name a period of elements for which the outermost energy level contains only *s*, *p* and *d* electrons.	
Name a metal that forms complex ions.	
Name an ionic compound made from non-metallic elements.	

(5 marks)

6. Ethanol (CH_3CH_2OH) is fully miscible with water. It is also fully miscible with petrol (a mixture of hydrocarbons). Give chemical reasons why ethanol can mix with both water and petrol.

(3 marks)

7. The table below shows some physical and chemical properties of the chlorides of some Period 3 elements.

	NaCl	$MgCl_2$	PCl_3	SCl_2
Melting point (ºC)	800	710	-90	-80
Boiling point (ºC)	1470	1420	80	60
Electrical conductivity of solid	Poor	Poor	Poor	Poor
Electrical conductivity of liquid	Good	Good	Poor	Poor
pH of water solution	7	7	<7	<7

a) Explain the large difference in the melting points of $MgCl_2$ and SCl_2.

b) Explain the difference in electrical conductivity of solid and liquid $MgCl_2$.

c) Would you expect each of these to be an electrical conductor? Explain in each case.

i) a water solution of NaCl

ii) a water solution of PCl_3

(7 marks)

PART A - Answer all questions on the answer sheet.

1. **As you go down Group II from Be to Ra, which of the options below correctly describes the trends in the characteristics of the elements?**

	Electronegativity	1st Ionisation Energy	Electrical Conductivity
a)	**Decreases**	**Decreases**	**Increases**
b)	**Increases**	**Decreases**	**Decreases**
c)	**Decreases**	**Increases**	**Increases**
d)	**Increases**	**Increases**	**Decreases**

2. An element has the electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^5$.
In which group and period of the Periodic Table is the element located?
a) Group III, period 4
b) Group V, period 4
c) Group IV, period 1
d) Group VII, period 3

3. Which one of the following has only dispersion forces between its molecules in the liquid phase?
a) CO_2
b) NH_3
c) C_2H_5OH
d) H_2O

4. Which of the lists below indicate the boiling points of the compounds in either descending order?
a) HI, HBr, HCl
b) NH_3, PH_3, AsH_3
c) CH_4, C_3H_8, C_2H_6
d) C_3H_8, CH_3CH_2COOH, CH_3COCH_3

5. Which of the following best describes the shape and polarity of a molecule whose formula is CF_4 ?
a) tetrahedral, non polar
b) pyramidal, polar
c) pyramidal, non polar
d) tetrahedral, polar

6. Which of the following solids contains discrete molecules?
a) Lead.
b) Calcium oxide.
c) Diamond.
d) Iodine.

7. When compounds are formed between the following pairs of elements, which ones are most likely to form predominantly covalent bonds?

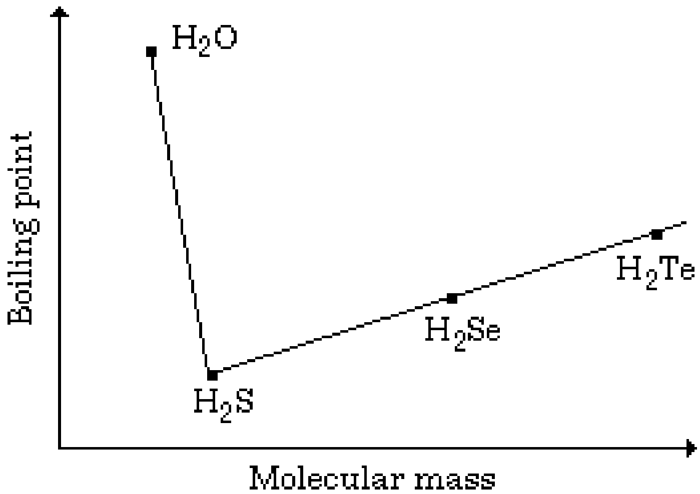
I Potassium - chlorine.
II Oxygen - carbon.
III Hydrogen - carbon.
IV Caesium - fluorine.
V Chlorine - fluorine.

a) **II**, **III** and **V**.
b) **II** only.
c) **I** and **III**.
d) **I** and **IV**.

8. At room temperature, methane is a gas while carbon tetrachloride is a liquid. This is best explained by which one of the following?

a) There is appreciable hydrogen bonding in carbon tetrachloride, but not in methane.
b) The carbon tetrachloride molecule is polar, while the methane molecule is non-polar.
c) Carbon tetrachloride has an appreciably higher molecular mass than methane.
d) The bonds in carbon tetrachloride are polar covalent.

9. The diagram below compares the boiling points of H_2O, H_2S, H_2Se and H_2Te.



The boiling point of water is much higher than expected because:

a) water is an ionic compound.
b) weak Van der Waal's forces exist between water molecules.
c) hydrogen bonding occurs between water molecules.
d) water is a liquid.

10. A water molecule is polar because:

a) hydrogen and oxygen have the same electronegativity.
b) the molecule is linear with a net dipole.
c) oxygen has a higher electronegativity than hydrogen and it is non-linear.
d) the hydrogen atoms acquire a slight negative charge while the oxygen atom acquires a slight positive charge.

11. Which of the following molecules possesses a molecular dipole?

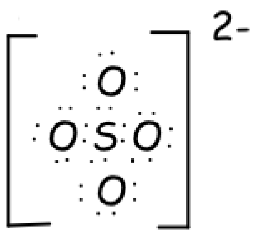
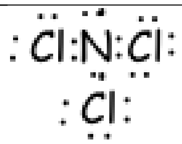
I F_2O
II BeF_2
III BF_3
IV NF_3
V CF_4

a) **II**, **III** and **V**.
b) **I** only.
c) **IV** only.
d) **I** and **IV**.

ANSWERS

1.	a	b	c	d
2.	a	b	c	d
3.	a	b	c	d
4.	a	b	c	d
5.	a	b	c	d
6.	a	b	c	d
7.	a	b	c	d
8.	a	b	C	d
9.	a	b	c	d
10.	a	b	c	d
11.	a	b	c	d

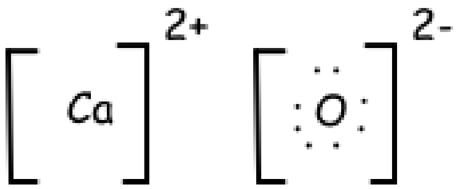
1.

Species	**Electron Dot Diagram**	**Shape (sketch or name)**
Sulfur dioxide, SO_2		Bent
Sulfate ion, SO_4^{2-}		Tetrahedral
Nitrogen trichloride, NCl_3		Pyramidal

Diagrams = 2 and shape =1 mark (9 marks)

2. a) It is the electrostatic attraction between oppositely charged ions. [1]

b)



[2] &

[2]

(5 marks)

3. a) Covalent.
In the form of a covalent network.
This element loses 4 e^- relatively easily and the 5th is removed from the non-valence shell. Element is non-metal Si.
Covalent bonding is the attraction between adjacent nuclei for the shared e^-'s between them. [4]

b) High m.pt and b.pt
Non-conductors of electricity
Hard
Brittle [1 for 1 or 2 props] [2 for 3 props]

(6 marks)

4. a) Cl has more protons so there is a greater attraction force- which causes the atom to be smaller and there is a shorter distance between the nucleus and e-. For these 2 reasons it takes more energy to remove an e-. [2]

K has a low ionisation E and electronegativity so will easily lose an e- but Cl has high electronegativity and ionisation E so is more likely to gain an e-. [2]

(4 marks)

5.

Name a group of elements for which the outermost energy level contains only *s* electrons.	**I or II**
Name a period of elements for which the outermost energy level contains only *s* and *p* electrons.	Period 2
Name a period of elements for which the outermost energy level contains only *s*, *p* and *d* electrons.	Period 3
Name a metal that forms complex ions.	Iron, copper, etc
Name an ionic compound made from non-metallic elements.	NH_4Cl

(5 marks)

6. Ethanol (CH_3CH_2OH) has non-polar hydrocarbon end and a polar H-bonded end. [1] The non-polar end has enough attraction to petrol to break the petrol-

petrol forces (dispersion) and dissolve in the petrol. [1] It also has H-bonding attractions to water and can break the water-water H-bonds and is fully miscible with water. [1]

(3 marks)

7.

a) $MgCl_2$ is ionic and has many strong ionic bonds between Mg^{2+} and Cl^- ions that need a lot of energy to break. [1] Whereas SCl_2 is a CM compound with dipole-dipole forces between molecules that require less energy to break. [1]

b) Solid $MgCl_2$ has ions but these charged particles fixed in place and cannot carry charge. [1] Liquid form now allows the ions Mg^{2+} and Cl^- to move and carry the charge. [1]

c) i) Yes. [1/2]

$NaCl \rightarrow Na^+ + Cl^-$

The solution has mobile charged particles, $Na^+ + Cl^-$, ions to carry the charge. [1]

ii) No. [1/2] As dissolves PCl_3 in water (which it will do to a small degree because it is a polar molecule) it will not form ions and not conduct electricity. [1]

(7 marks)