ELECTRON CONFIGURATION

In 1913 Niels Bohr proposed that electrons had orbits around the nucleus at a given distance. The electrons had a circular orbit at a fixed energy level. It was possible to add energy to an atom and an electron could absorb this energy and move to a higher energy level. This idea is used to explain fireworks and flame tests.

ELEMENT	*NUMBER OF e^-*	*e^- CONFIGURATION*
H		
Be		
N		
Ne		
Al		
S		
K		

IONISATION ENERGY

When atoms lose electrons to form ____________ ions, the atoms are said to have been ionised. This process requires energy. Why?

The ionisation energy is the amount of energy needed to remove 1 mole of electrons from 1 mole of atoms or ions in their gaseous state.

Which of the following equations corresponds to the correct definition of ionisation energy?

a) $Na\,(g) \rightarrow Na^+\,(g) + e^-$ _______

b) $Na\,(s) \rightarrow Na^+\,(s) + e^-$ _______

c) $Na\,(s) \rightarrow Na^+\,(g) + e^-$ _______

d) $Cl_2\,(g) + 2\,e^- \rightarrow 2\,Cl^-\,(g)$ _______

e) $Cl\,(g) + e^- \rightarrow Cl^-\,(g)$ _______

f) $Cl\,(g) \rightarrow Cl^+\,(g) + e^-$ _______

Give equations, including state symbols, which represent the first ionisation energy process for:

a) K

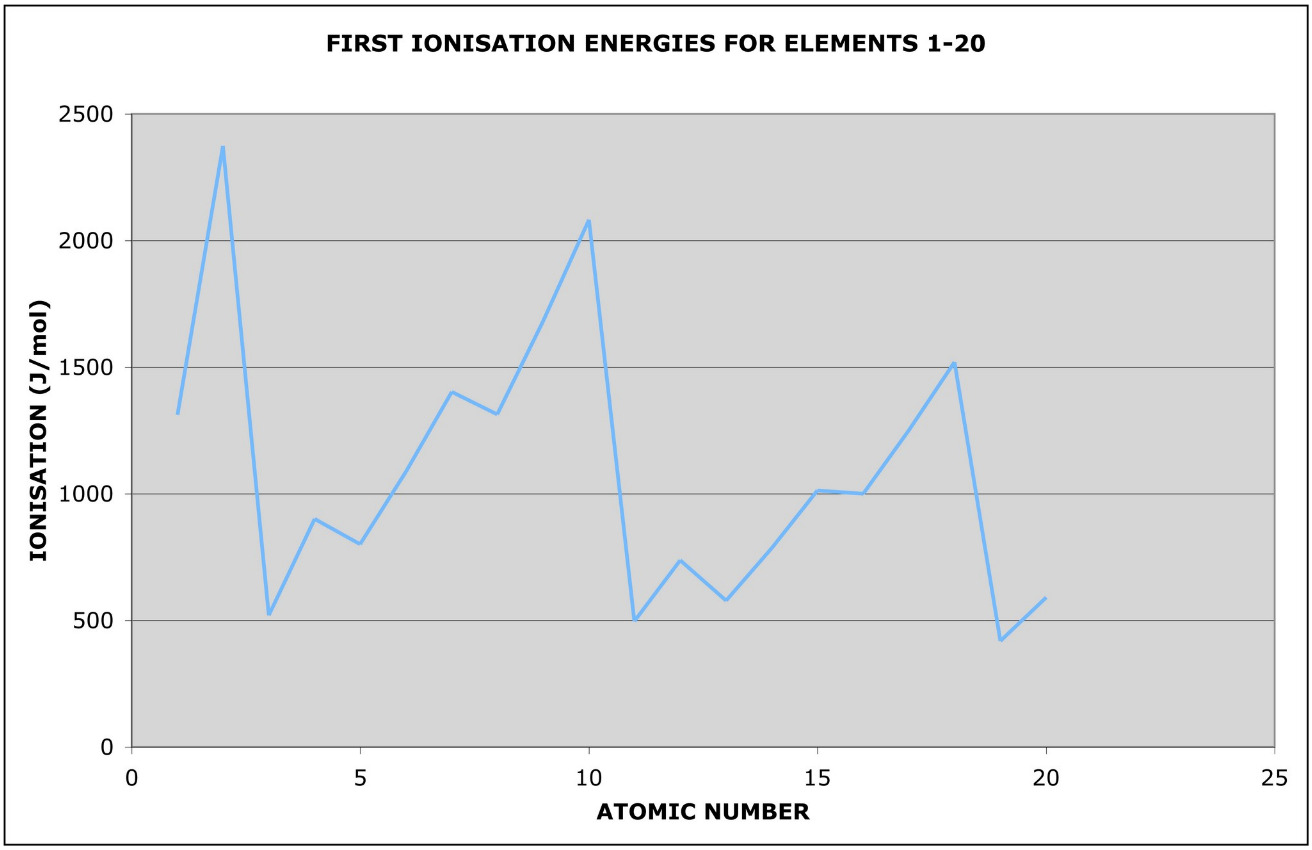
b) Ca

c) Br

Draw a sodium and chlorine atom. Which atom is bigger? Why?________

What features of an atom will affect the size of the ionisation energy?

Look at the below graph of "First ionisation energies".



Why is Group XVIII always higher than Group XVII? (Hint: write the electron configurations of neon and fluorine)

Why does the ionisation energy for each group decrease as the period increases? Explain using group I.

Is the ionisation energy for boron higher or lower than fluorine?

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

1. ____
2. ____
3. ____

Explain how each of these factors influences ionisation energy.

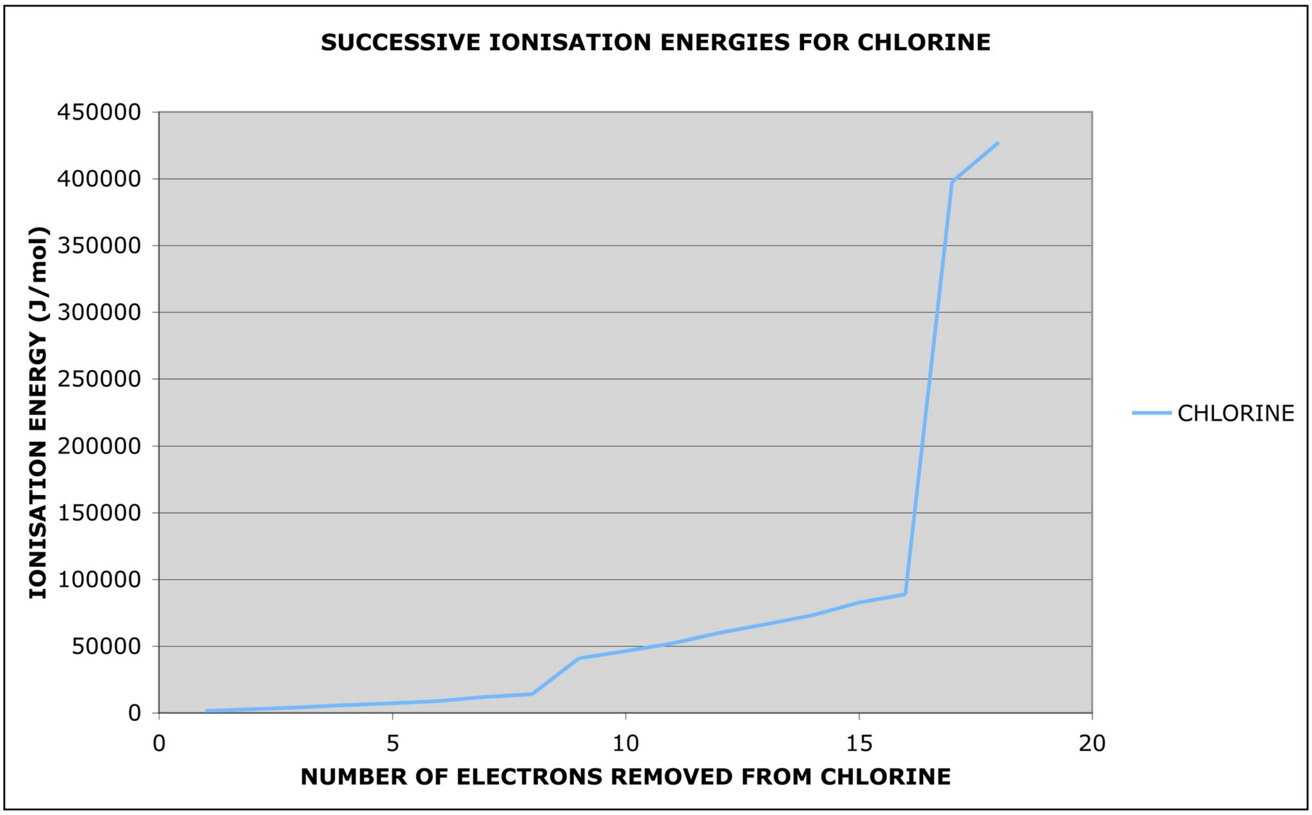
1.

2.

3.

As electrons are moved further from the nucleus more energy is required. For example the second shell electrons have more energy than the first. The second shell electrons are said to have a *<u>higher energy level</u>*. *Think about this as gravitational potential energy, the further you are from the ground the higher your energy.*

Look at the "Successive ionisation energies for chlorine". Explain the changes, trying to use the phrase energy level.



How many valence electrons in each of the following elements?

	1^{ST} E_I	2^{ND} E_I	3^{RD} E_I	4^{TH} E_I	5^{TH} E_I	6^{TH} E_I	7^{TH} E_I	8^{TH} E_I	VALENCE
A	1313.9	3388.3	5300.5	7469.2	10989	13326	71330	84078	
B	418.8	3052	4420	5877	7975	9590	11343	14944	
C	495.8	4562	6910.3	9543	13354	16613	20117	25496	

IONIC BONDING (Metal + Non-metal)

Draw a crystal (ionic lattice) of NaCl.

* *The lattice is made of positive and negative* ***ions*** *held together by electrostatic attraction of the ______________ and ______________ charges.*

* *Each positive ion is surrounded by negative ions and vice versa.*

* *For each Na^+ there is one Cl^-, therefore the formula is NaCl.*

* *In the reaction between Na and Cl_2 there is a loss and a gain of electrons.*

What is the electron configuration of sodium atoms? _____________ and

sodium ions?________________________

What is the electron configuration of chlorine atoms? _____________ and

chloride ions?________________

PROPERTIES OF IONIC COMPOUNDS

PROPERTY	EXPLANATION
1. High melting and boiling point	
2. Hard	
3. Brittle	
4. Non-conductors of electricity as Solids	
5. Conductors of electricity as liquids and in solution	

Bonding between elements in [Groups I or II] and [Group VII or oxygen or sulfur] is the purest form of ionic bonding.

ELECTRON DOT DIAGRAMS

These diagrams show only valence electrons and how they are involved in bonding.

eg: 1) Na Cl $\rightarrow$ Na^+ Cl^-

eg: 2) Mg O

eg: 3) *K* *S*

eg: 4) *Na* *N*

Draw the electron dot diagrams for:

i) *calcium ion*

ii) *bromide ion*

iii) *phosphide ion*

iv) *aluminium ion*

v) *caesium ion*

vi) *carbide ion*

STRUCTURE OF METALS

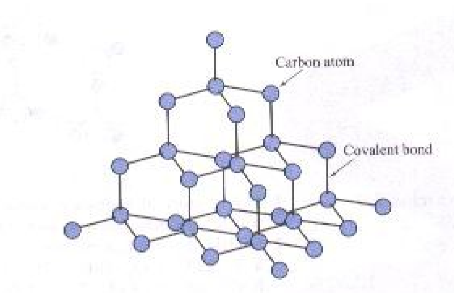
Metals are metal ions that have freed electrons. These electrons are not fixed in one place and are called delocalised electrons. Or metals can be described as metal ions in a "sea of electrons". The metal ions are attracted to the delocalised electrons and this very strong force is called a metallic bond.

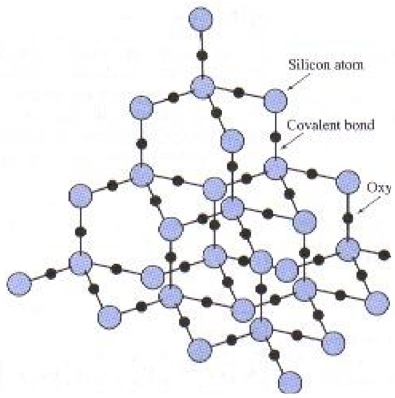
PROPERTY	EXPLANATION
1. High melting and boiling point	
2. Hard	

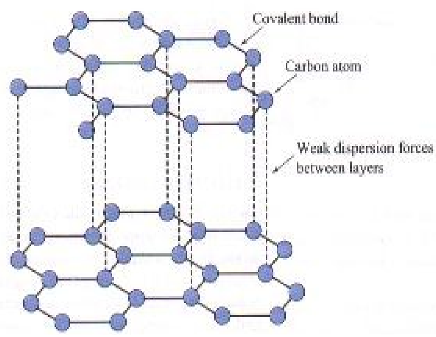
3. *Malleable and ductile*	
4. *Conductors of electricity, as solids and liquids*	

STRUCTURE OF COVALENT NETWORKS (Non-metal + Non-metal)

These substances giant molecules made of atoms that are joined by continuous covalent bonding to other atoms. The main covalent networks are diamond, graphite, silicon and silica (silicon dioxide). Drawn below is each of the covalent networks.







PROPERTY	EXPLANATION
1. *High melting and boiling point*	
2. *Hard*	
3. *Insoluble in water and other solvents*	
4. *Non-conductors of electricity*	
5. *Graphite is a conductor of electricity*	

STRUCTURE OF COVALENT MOLECULES (Non-metal + Non-metal)

*These substances compete equally for electrons. So instead of losing and gaining, they **share** electrons. When electrons are shared by atoms this is called a covalent bond.*

A molecule is a group of atoms which are bonded to each other. They have no bonds to other atoms, just weak forces of attraction. Therefore covalent molecular substances exist as separate (discrete) molecules with <u>strong bonds in the molecule but with weak forces between molecules</u>.

<u>PROPERTIES OF COVALENT MOLECULES</u>

PROPERTY	*EXPLANATION*
1. Low melting and boiling point	
2. Soft	
3. Non-conductors of electricity	
4. Conductors of electricity when the molecules react to make ions in the solution.	

<u>ELECTRON DOT DIAGRAMS</u>

eg: 1) H_2

This bond forms because the nucleus of each atom is attracted to the pair of electrons. Another diagram type shows the covalent bond as a line as two electrons.

How do you know the number of bonds formed by a non-metal? Ask yourself the following questions:

i) What Group of the Periodic Table is the element?

ii) How many e^- *has the element?*

iii) How many more e^- *does the element need to have a stable* e^- *configuration?*

This is the number of bonds the element will USUALLY form.

eg: 2) HCl *How many bonds will H make?* ________

How many bonds will Cl make? ________

eg: 3) Cl_2

eg: 4) NH_3

eg: 5) H_2O

eg: 6) CH_4

eg: 7) O_2

eg: 8) C_2H_6

eg: 9) N_2

eg: 10) CO_2

eg: 11) C_2H_4

eg: 12) CCl_4

In the above molecules, what are the number of electrons shared in a:

a) single bond _______ *b) double bond* _______ *c) triple bond* _______

VALENCE ELECTRON PAIR REPULSION THEORY

This theory assumes that the electron pairs in the valence shells of the atoms will repel one another and try to get as far apart as possible. The different number of pairs of electrons gives molecules different angles between the atoms and different shapes. Make models of the molecules above and assign a shape to each molecule. The choices are; linear, bent, trigonal planar, pyramidal and tetrahedral. Sketch these shapes.

NUMBER OF ATOMS	POSSIBLE SHAPES	EXAMPLES
2		
3		

4		
5		

In polyatomic ions there are covalent bonds holding the atoms together in the ion. For each of the following draw the electron dot diagram and predict the shape.

1) *nitrite ion (NO_2^-)*

2) *ammonium ion*

3) *amide ion (NH_2^-)*

4) *sulfite ion (SO_3^{2-})*

ELECTRONEGATIVITY

Electronegativity is the measure of an atom's tendency to attract an electron.

Not all atoms attract electrons with the same power. The electron attracting ability of an atom is called electronegativity. As different atoms share electrons the atom with the greatest electronegativity will pull the shared electrons closer. The result is a _______ covalent bond with the electrons closer to one end.

A polar bond is one that has 2 poles or one end that is more negative than the other. This other is more positive.
Using the given electronegativities, for the following bonds

i) *Draw the electron dot diagrams.*
ii) *Compare the electronegativities given to decide if the bond is polar.*
iii) *If polar show the positive and negative ends of the bond.*

Electronegativity:
H = 2.2
B = 2 C = 2.5 N = 3.1 O = 3.5 F = 4.1
Si = 1.7 P = 2.1 S = 2.4 Cl = 2.8
Br = 2.7
I = 2.2

Label the polarity on the following atoms (δ+ and δ-):

C-H *O-H* *H-F* *Cl-Cl* *N-O*

What is the pattern in the electronegativities as you move horizontally? Why?

What is the pattern in the electronegativities as you move vertically? Why?

Metals have electronegativity. Why are they not considered important in this section?

INTERMOLECULAR FORCES

Water and iodine are liquid and solid respectively at room temperature. If they are both covalent molecular substances and not gases they must have reasonable forces of attraction between the molecules. If covalent molecular substances have no bonds between molecules then what causes these forces between molecules? These forces are called van der Waals' forces. There are of 3 types.

1) DIPOLE-DIPOLE FORCES

In NO above, the molecule has a polar bond. If the ends of the molecule have a partial charge what do you think happens when lots of these molecules are together? Draw a diagram to show how this happens.

The prefix di means _________. So dipole-dipole forces are forces between

Molecules will have dipole-dipole forces between them when they

- i) *Contain polar bonds.*
- ii) *Are asymmetrical.*

Which of the 12 molecules from pages 12 will have dipole-dipole attractions between their molecules?

The molecules with dipole-dipole attractions draw the molecule showing the separation of charge.

2) HYDROGEN BONDING

A hydrogen bond is a much stronger type of dipole-dipole force. It only occurs when hydrogen is directly bonded to N, O or F. These 3 elements are the most electronegative and will strongly attract the electrons in the covalent bond. This causes an extreme separation of charge and a much stronger force of attraction to neighbouring molecules. Show how hydrogen bonding occurs between water molecules.

For each of the following substances indicate the molecular shape, note whether the molecule is polar or non-polar and whether the molecule will contain hydrogen bonding.

name	melting point (°C)	molecular shape	polarity	hydrogen bonding
H_2O	0			
HCl	-115			
NH_3	-77			
H_2S	-85			
PH_3	-133			
HF	-83			

Compare the melting points of the Period 2 hydrides with its corresponding Period 3 hydride (eg: HF with HCl). What general pattern do you recognise? Why do you think this occurs?

3) DISPERSION FORCES

Every substance has dispersion forces, the weakest of attractions between atoms. In covalent networks, ionic and metallic substances the dispersion forces are so small compared to the covalent, ionic and metallic bond that the dispersion force is ignored but it still exists.

As electrons move around within the atom there are times when the electrons may be on one side of the atom, causing a temporary dipole. This then attracts and repels electrons in neighbouring atoms, inducing dipoles in these atoms. The electrons continue moving and the dipoles change. Dispersion forces are sometimes called temporary dipole forces.

A chemical like HCl has dipole-dipole forces and ___________ forces. Again the dominant force is the ___________ . There are substances that have only dispersion forces. Which 8 from page 12 have only dispersion forces?

The chemicals with only dispersion forces are:

1) noble gases eg ____
2) elemental gases eg ____
3) alkanes eg ____
4) symmetrical molecules eg ____

In general the more electrons a particle has the stronger the dispersion force.
OR
As the formula mass increases so does the strength of the dispersion force.

What are the names and molecular formulae for first 8 alkanes?

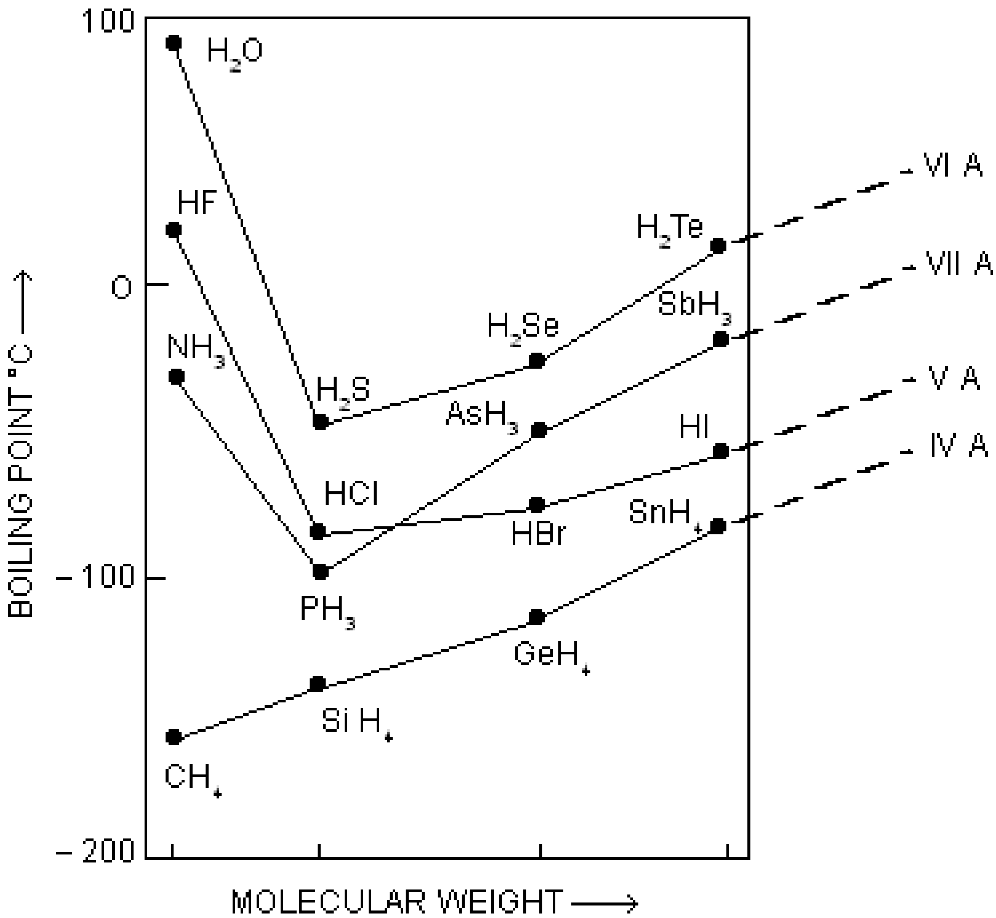
number of carbons	*name*	*molecular formulae*	*melting point (°C)*
1			*-182*
2			*-183*
3			*-190*
4			*-138*
5			*-130*
6			*-95*
7			*-91*
8			*-57*

What happens to the melting point as the number of carbon atoms increases?

Explain why.

There are 3 types intermolecular force. What is the ascending order of relative strengths?

<u>MELTING AND BOILING POINTS OF THE NON-METAL HYDRIDES</u>



(b)

(b) Boiling Point curves of hydrides of IV, V VI and VII group elements.

http://knowledgebin.org/kb/entry/Chemical_-_Bonding-2_212.html

In general what happens to the melting point of all the hydrides as you down the Group? Explain why.

Which 3 compounds are anomalous with regards to the statement above? Why?

What are the bonding capacities of Group XVI __________, Group XVII _________ and Group XVIII _____ _______ ?

SOLVENTS AND SOLUBILITY

When a solvent and solute are mixed to try and form a solution, several things must occur. A force of attraction exists between solute particles. A force of attraction exists between solvent molecules. So there must be a force of attraction between the solute and solvent molecules that is strong enough to

i) Separate the solvent molecules and
ii) Separate the solute molecules.

To achieve this there must be similar intermolecular forces and/or bonding between the solvent and solute particles. The phrase "like dissolves like" is used as memory tool but cannot be used as an explanation of why something dissolves in a solvent.

Explain the following: (as an answer to the following you must state the forces between the solute particles, the forces between the solvent molecules and then explain {or draw} how the forces are/are not strong enough to break each other)

1) *Ammonia is very soluble in water*

2) *Octane is not soluble in water*

3) *Oil is soluble in petrol*

4) *Magnesium oxide is not soluble in water*

5) Iodine is slightly soluble in ethanol (CH_3CH_2OH) and not soluble in water

REVISION

CHEMISTRY FOR WA – STAGE 3

Review Exercise	Pages	Questions
1.1	6	2
1.2	11	1-5
1.3	15	1-3
1.4	23	1-5
2.1	33	1-5
2.2	38	1-7
2.3	44	1-9
2.4	46	1-4
3.1	58	1-3
3.2	63	1-5
3.3	66	1-3
3.4	72	1-4
3.5	74	1-3
3.6	79	1-6
3.7	84	1-6
3.8	85	1-2
3.9	87	1
4.1	100	1-4
4.2	103	1-3

End of Chapter Questions	Pages	Questions
Chapter One	25	1-18
Chapter Two	51	1-19
Chapter Three	88	1-24
Chapter Four	118	

EXPLORING CHEMISTRY – STAGE 3

Sets 9, 10 and 11.

KEYWORDS

1. ionisation energy
2. lattice
3. electron configuration
4. ionic bonding
5. valence electrons
6. delocalised electrons
7. malleable
8. ductile
9. silica
10. covalent bond
11. VSEPR Theory
12. linear
13. bent
14. trigonal planar

15. pyramidal
16. tetrahedral
17. polyatomic ion
18. electronegativity
19. van der Waals' forces
20. dipole-dipole forces
21. asymmetrical
22. polar bond
23. polar molecule
24. hydrogen bonding
25. dispersion force
26. temporary dipole
27. solvent
28. Solute