AS Chemistry

Atoms, Periodicity & Structure

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Atomic Theory

1. Summarise how the experiments below contributed to our understanding of atomic structure.

Electrolysis 1891

Cathode Rays

J.J. Thomson e/m 1897

Millikan’s ‘oil drops’

1. Before we had the nuclear model of the atom, the accepted theory was that of the ‘plum pudding.’ How did Rutherford’s experiments show that the plum pudding model was not correct? *Be sure to describe the two opposing models, Rutherford’s experiment and what this implied about atomic structure.*

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1. Complete the table below;

|  |  |  |  |
| --- | --- | --- | --- |
|  | Proton | Neutron | Electron |
| Relative mass |  |  |  |
| Relative charge |  |  |  |
| Actual mass/kg |  |  |  |
| Actual charge/C |  |  |  |

Periodicity

1. What was the main advantage that Mendeleev’s periodic table had over previous attempts to order the elements?

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1. What are the outer shell configurations of:
2. Group 2 elements.
3. Group 7 elements.
4. Group 1 elements.
5. Describe and explain the trend in atomic radius:
6. Down a group.

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1. Across a period.

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Ionisation Energy

Ionisation energies tell us the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous ions. Successive ionisation energies tell us the energy required to remove subsequent electrons from an ion. For example the third ionisation energy tells us the energy required to remove an electron from one mole of an ion with a 2+ charge.

1. Write an equation for:
2. The first ionisation energy of lithium.
3. The third ionisation energy of neon.
4. The second ionisation energy of oxygen.
5. How will the following affect ionisation energies and why?
6. The effect of a large atomic radius on an outer electron.
7. The atomic number of an element.
8. Electrons in shells closer to the nucleus.
9. Why do ionisation energies increase as more electrons are removed?
10. An element in period three (Na-Ar) has the following successive ionisation energies in kj mol-1: 789, 1577, 3232, 4356, 16091, 19785, 23787, 29252.

Identify this element giving your reasons.

1. Sketch a graph for the successive ionisation energies of aluminium.

Shells and Orbitals

Ionisation energies provide evidence for the presence of shells and how many electrons each shell holds. We can describe shells using the *principle quantum number* which we give the label *n.*

 Each shell can hold up to 2n2 electrons, so for the first shell:

n=1

If we use the formula above to find out the number of electrons in the shell;

No of electrons = 2 x 12

= 2

1. Using the formula above calculate the number of electrons in the shells one to five.

We now know that there are rules governing where electrons can be in shells. As such we say that each shell is made up of *orbitals*. These come in four varieties: s, p, d & f.

1. Which orbitals are present at:
2. n = 1
3. n = 2
4. n = 3
5. What is the total number of electrons that can be found in the following orbitals?
6. p
7. s
8. d
9. f

Subshells

An electron shell is made up of atomic orbitals with the same principle quantum number, *n.* Within each shell orbitals of the same type are grouped together as a *sub-shell.*

1. Are the following statements about the way in which electrons fill shells and sub-shells true or false?
2. The highest possible energy level is filled first.
3. Subshells are filled singly before electrons begin pairing.
4. Each orbital can hold two electrons with the same spin.
5. For any statement above that you think false, rewrite it so it is correct.

Electronic configurations tell us the number and placement of electrons in an atom. We write them with the form nxy where n is the shell number, x is the type of orbital and y is the number of electrons in the orbital making up the subshell.

1. Write the electronic configuration for the elements H to Ne.
2. How can the electronic configurations be shortened? (*Think about group 8 elements*)
3. What are the electronic configurations of the following ions?
4. Na+
5. O2-
6. Mg2+
7. Al3+
8. F-

Chemical Bonding

1. (a) State the three main types of bonding.

(b) Describe, in terms of electron sharing or transfer, covalent and ionic bonding.

1. Predict the type of bonding in the following compounds:
2. Sodium chloride
3. Zinc oxide
4. Hydrogen chloride
5. Silver bromide
6. Nitrogen bromide
7. Sulfur dioxide

Ionic Bonding

We can use the group in which an atom is found on the periodic table to predict the ions it is likely to form.

1. Predict the formula of these ionic compounds:
2. Calcium iodide
3. Lithium nitride
4. Draw dot and cross diagrams for the following ionic compounds:
5. MgO
6. CaBr2
7. Na3P
8. Al2O3

Covalent Bonding

1. What is a lone pair?
2. How many covalent bonds can the following elements form?
3. O
4. C
5. H
6. N
7. Draw a dot and cross diagram for the following compounds:
8. F2
9. HF
10. SCl2
11. CS2
12. HCN (the C=N bond is a *triple bond)*
13. Explain what is meant by the term *dative bond.*

Molecular Shapes

1. The 3D shapes of molecules are based on the repulsion between electron pairs. Order the following from highest repulsion to least repulsion.

***Bonded pair/bonded pair lone pair/lone pair bonded pair/lone pair***

1. Draw the 3D representation of the following, labelling the bond angle:
2. Methane, CH4
3. Ammonia, NH3
4. Water, H2O
5. Carbon dioxide, CO2
6. Predict the shape and bond angle of the following; it may help to draw a dot and cross diagram of each:
7. H2S
8. AlCl3
9. SiF4
10. PH3
11. Predict the shape and bond angle of the following; it may help to draw a dot and cross diagram of each:
12. H2S
13. AlCl3
14. SiF4
15. PH3

Electronegativity

Electronegativity is described as the ability of an atom to attract electron pairs to itself. If two bonded atoms have the same electronegativity they will form a non-polar bond,; the electrons in the bond will be evenly distributed. However in molecules where one atom is more electronegative than the one it is bonded too the more electronegative atom will attract the bonded electron pair to itself and form a polar bond, in this case a permanent dipole will be produced.

1. Why is CO2 a non-polar molecule whilst H2O is polar?
2. For the following compounds state which element is the most electronegative;
3. HCl
4. CO2
5. NH3
6. FH
7. HNO3
8. Draw diagrams of the above compounds, labelling any dipoles.

Intermolecular Forces

1. Explain how van der Waal’s forces occur between molecules; use the words *induced, instantaneous & dipole* in your answer.
2. Below are the boiling points for the Group 7 elements. Each exists as a diatomic molecule. Explain the trend in boiling points in terms of intermolecular forces.

**F2, -188oC; Cl2, -35oC; Br2, 59oC; I2, 184oC.**

1. Order the following from weakest to strongest in terms of relative bond strength.

**dipole-dipole forces van der Waal’s forces hydrogen bonding**

1. What makes it possible for hydrogen bonding to occur?
2. Explain how the following properties of water are dependent on the ability of water molecules to form hydrogen bonds to one another:
3. Ice floats on liquid water.
4. Water’s boiling point is higher than similar molecules e.g. H2S
5. Water has a high surface tension.
6. Draw a diagram of two water molecules. On the diagram label any dipoles and draw a hydrogen bond.

Exam Questions

1. Electrons are arranged in energy levels.
2. An orbital is a region in which an electron may be found. Draw diagrams to show the shape of an s-orbital and a p-orbital. [2]
3. Complete the table below to show how many electrons **completely** fill each of the following: [3]

|  |  |
| --- | --- |
|  | Number of electrons |
| A d-**orbital** |  |
| A p-**sub-shell** |  |
| The third **shell** (n=3) |  |

1. The energy diagram below is for the eight electrons in an oxygen atom. The diagram is incomplete as it only shows the two electrons in the 1s level.

1s

Complete the diagram for the oxygen atom by:

1. Adding labels for the other sub-shells. [1]
2. Adding arrows to show how the other electrons are arranged. [1]
3. Successive ionisation energies provide evidence for the arrangement of electrons in atoms. The table below shows the eight successive ionisation energies of oxygen.

|  |  |
| --- | --- |
| Ionisation number | Ionisation energy / kj mol-1 |
| 1st | 1 314 |
| 2nd | 3 388 |
| 3rd | 5 301 |
| 4th | 7 469 |
| 5th | 10 989 |
| 6th | 13 327 |
| 7th | 71 337 |
| 8th  | 84 080 |

1. Define the term ***first*** *ionisation energy.* [3]
2. Write an equation, with state symbols, to represent the **third** ionisation energy of oxygen. [2]
3. Explain how the data in the table above provides evidence for two electron shells in oxygen. [2]
4. Although compounds are usually classified as having ionic or covalent bonding, often the bonding is somewhere in between these two extremes.
5. State what is meant by the terms:
6. Ionic bond. [1]
7. Covalent bond. [2]
8. Compounds with covalent bonding often have polar bonds. Polarity can be explained in terms of electronegativity.
9. Explain the term electronegativity [2]
10. Use a suitable example to show how the presence of a polar bond can be explained in terms of electronegativity.

You may find it useful to draw a diagram in your answer. [2]

1. Some polar molecules are capable of forming hydrogen bonds. Draw a diagram to show an example of hydrogen bonding. [2]

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