MID-COURSE REVISION QUESTIONS

The following questions are designed both as revision and to allow you to assess your understanding of some of the concepts dealt with in Topics 1 - 8. They might be used in the tutorial sessions or more likely, can be attempted at other times.

TOPIC 1.

1. Explain the difference between

(a) an atom and a molecule

(b) an element and a compound

(c) a chemical change and a physical change

TOPIC 2.

1. Explain the meaning of the following:

(a) isoelectronic species

(b) isotopes of a given element

(c) atomic emission spectrum

2. Explain the difference between:

(a) mass number and atomic number

(b) ground state and excited state

(c) a cation and an anion

3. What determines which element an atom represents?

4. In what way(s) do different isotopes of an element differ?

TOPIC 3.

1. Explain why the statement “metals want to lose electrons” is incorrect.

2. What fundamental rule must be observed in writing a correct formula for a salt?

3. Outline the rules to be observed in writing the name for a binary ionic compound.

TOPIC 4.

1. Why is the molecule of hydrogen H_2 and not H_3 or any other formula?

2. What is the difference between an ionic bond and a covalent bond?

3. Draw an electron orbit diagram and a structural formula for the water molecule.

4. Explain the meaning of the following terms:

(a) valence

(b) lone-pair of electrons

(c) bonding pair of electrons

(d) unsaturated compound

(e) binary compound

(f) polyatomic ion

5. Explain why the bonds between atoms of two non-metals as exemplified in the compound SO_2 are covalent while the bonds between a metal and non-metal such as in the compound MgO are ionic.

6. Outline the rules relating to naming covalent binary compounds.

TOPIC 5.

1. Explain the difference between empirical, molecular and structural formulas for a compound that exists as molecules.

2. Describe:

(a) metallic bonding. (b) network covalent bonding

3. Explain why covalently bonded molecules have specific 3-dimensional shapes.

4. Which elements usually exist as diatomic molecules at normal pressure and temperature?

5. State the fundamental principle that underlies the writing of a properly balanced equation for a reaction.

TOPIC 6.

1. Explain the meaning of the following terms:

(a) solvent

(b) solute

(c) polar bond

(d) polar molecule

2. Explain the process by which an ionic solid such as sodium chloride dissolves in water.

3. Explain the process by which an ionic compound can be reclaimed from a solution.

4. Complete the following:

(a) an acid + a reactive metal forms ....

(b) an acid + a metal oxide forms ....

(c) an acid + a metal hydroxide forms ....

(d) an acid + a carbonate forms ....

TOPIC 7.

1. Why do atoms of different elements have different masses?

2. Define each of the following:

(a) (relative) atomic weight

(b) gram atomic weight

(c) the mole

(d) molar mass

3. Explain how to express a given mass of a pure compound as moles of that compound.

4. Why is it that reactants in a chemical reaction do not combine in simple numeric ratios by mass?

TOPIC 8.

1. Explain the principle involved in calculating the percentage by mass of the constituent elements in a compound.

2. Explain the principle involved in deducing an empirical formula from percentage by mass data.

3. Outline the steps involved in obtaining the molecular formula for a compound that exists as molecules.

MID-COURSE REVISION QUESTIONS - ANSWERS

TOPIC 1.

Question 1

(a) An atom is the smallest particle of an **element** while a molecule, which may be of an **element or a compound**, always has more than one atom bonded together by chemical bonds. A molecule of a compound always contains atoms of more than one element.

(b) An element is a substance that cannot be broken down into simpler component substances and it consists of just one type of atom. A compound consists of smallest units which contain more than one element and thus more than one type of atom.

(c) A chemical change occurs if that change is accompanied by the breaking and/or the forming of chemical bonds. Physical changes do not involve bonds breaking or forming - e.g. separation of liquids in a mixture by distillation.

TOPIC 2

Question 1

(a) Two species are said to be isoelectronic if they have different numbers of protons in their nuclei but the same arrangement of electrons. e.g. He and Li^+ are isoelectronic because both have two electrons in their first orbit and no other electrons.

(b) Atoms of a given element always have the same number of protons but the number of neutrons can vary. Isotopes of a given element differ in the number of neutrons present in their nuclei. A particular isotope can be specified by its mass number which is the total of protons plus neutrons in its nucleus - e.g. ^{12}C represents the isotope of carbon with 6 neutrons while the ^{13}C isotope has 7 neutrons and ^{14}C has 8 neutrons. All three isotopes of carbon necessarily contain 6 protons in their nuclei.

(c) Electrons in an atom can absorb energy and be elevated to briefly occupy a higher energy orbit. These excited electrons soon return to lower energy orbits and in so doing, emit energy equivalent to the energy difference between the orbits used. The emitted energy is in the form of electromagnetic radiation and often occurs at frequencies corresponding to the visible part of the spectrum. A pattern of spectral lines is observed, called the atomic emission spectrum of that element and it is a characteristic for each different element.

Question 2

(a) The mass number is the total of the number of protons plus neutrons in an atom's nucleus while the atomic number is just the number of protons present. Atoms of a given element always have the same atomic number but their mass number varies, depending on the particular isotope considered.

(b) When all the electrons in an atom occupy the lowest available energy orbits, the atom is in its ground state. If exactly the appropriate amount of energy is supplied to an atom from some external source, electrons can move into higher energy orbits and while there, the atom is said to be in the excited state. The excited state is necessarily only short-lived before the excited electron(s) fall back to the ground state with the accompanying release of energy as electromagnetic radiation.

(c) Cations and anions differ from each other in the sign of the electrical charge which they bear. Cations are species which bear an overall positive electrical charge. Atoms can be converted to cations by the loss of one or more electrons, leaving an overall excess of protons and hence positive charge on the species. Anions are species that bear an overall negative electrical charge which can be acquired by an atom gaining one or more electrons. The formation of both cations and anions from electrically neutral atoms always involves the gain or loss of electrons, a process that involves relatively little energy, and never the transfer of protons from the nucleus.

Question 3

The number of protons in the nucleus of an atom (its atomic number) is the determinant of which element it is.

Question 4

The various isotopes of a given element differ only in properties related to their mass. Such properties include the rate at which the various isotopes diffuse through a small hole. Chemically, all isotopes of a given element behave identically.

TOPIC 3

Question 1

To remove one or more electrons from an atom of a metal, energy (the ionization energy) must be supplied to overcome the electrical attraction between the electrons and the atomic nucleus. Consequently metal atoms do not "want to lose electrons". Indeed, the opposite applies, electrons must be dragged, kicking and screaming to remove them from their atom. The glimmer of truth in the statement lies in the fact that the ionization energy of metals is relatively small compared with non-metals.

Question 2

A salt consists of cations and anions combined in the solid state to form a crystal lattice in which cations and anions are packed in such a way that attractions between opposite charged species is maximised while repulsions between like-charged species is minimised. Overall, the total positive charge must equal the total negative charge and this will be achieved only if the formula for the salt contains an electrically balanced total cation and total anion charge - i.e. the sum of the charges on the cations must equal the sum of the charges on the anions.

Question 3

Binary ionic compounds are named as two words, the first being the cation and the second the anion. Cations are given the same name as the metal from which they were formed. If that metal can form more than one type of cation, the charge is appended as Roman numerals in brackets and without a space. For example, copper(II) for Cu^{2+}. Simple anions use a stem from the name of the element, to which is added the "ide" suffix - e.g. oxide (O^{2-}). Only lower case letters are used.

TOPIC 4

Question 1

The H atom has a single electron in its first electron orbit. That orbit can contain a maximum of two electrons. In forming a molecule, overlap of the orbits of two H atoms allows each to obtain a share of two electrons via a covalent bond and an H_2 molecule results. As the overlapped orbits cannot contain any more electrons, it is not possible for a third or more H atom to join together to form H_3, H_4 etc.

Question 2

Ionic bonds result from the complete transfer of electrons from metal atoms to non-metal atoms, the resulting oppositely charged ions then experiencing a strong electrostatic attraction. In a covalent bond, electrons are shared between atoms which are sufficiently close to allow overlap of their outer electron orbits. In practice, ionic bonds are seldom totally ionic and only covalent bonds between identical atoms such as Cl_2 are totally covalent. Instead covalent bonds are usually at least slightly polar due to one atom being more electronegative than the other.

Question 3

See Page IV-2 of Topic 4.

Question 4

(a) The valence of an atom is the number of bonds that it forms in a compound which may be ionic or covalent. Valence has no sign attached to it. Some elements such as hydrogen and oxygen can only display a single valence state while other elements such as copper and sulfur can have more than one valence state.

(b) Among the valence level electrons (i.e. outer level electrons) of atoms in a molecule, some of them may not be participating in the molecule's covalent bonds. Such electrons are called "non-bonding electrons" and as electrons are usually found to be in pairs, the term "lone pair" of electrons is often used.

(c) Electrons that are involved in the covalent bonds within a molecule are referred to as "bonding electrons". Each covalent bond requires a pair of electrons in order to be stable, so the term "bonding pair" is applied to them. Multiple bonds on a given atom involve more than one bonding pair. Thus a double bond will require four electrons or two bonding pairs and a triple bond requires three bonding pairs.

(d) An unsaturated compound contains atoms that are joined by one or more double or triple bonds.

(e) A binary compound is one that contains only two elements, e.g. water (H_2O), methane (CH_4), ammonia (NH_3).

(f) A polyatomic ion contains more than a single atom. The atoms may be of the same element (e.g. peroxide ion, O_2^{2-}, or different elements such as in the sulfate ion, SO_4^{2-}.

Question 5
Electrons can be removed from metal atoms to form cations which have the noble gas structure by supplying relatively small amounts of energy. Likewise, non-metals are relatively easily converted into anions with the noble gas structure by the gain of one or more electrons. Consequently compounds of metals such as magnesium with non-metals such as oxygen are usually ionic. However, to convert a non-metal such as sulfur into a cation with the noble gas structure would require each S atom to lose 6 electrons and form a S^{6+} ion. Such a cation would require very large amounts of energy to form and would not be stable. Instead, the S atom in combining with oxygen can more easily attain the noble gas structure by sharing electrons with the O atom as covalent bonds in a molecule. Thus non-metals combining with other non-metals form covalently bonded molecules rather than ionic compounds.

Question 6
Naming binary covalent compounds follows similar rules to those used for binary ionic compounds except that there is no cation to name as the first word. Instead, the element which is the least electronegative is named first - at this stage, that can be taken as the element from the lower number family or Periodic Table Group. To cater for the variable valency often encountered among non-metals, there are two methods that can be used where ambiguity could arise. One method is to use prefixes such as mono, di, tri, etc ahead of each of the two parts of the name to indicate how many of each of that atom is in the formula. The second method uses Roman numerals to indicate the valency of the first element in the name. For example, the compound SF_6 can be named as sulfur hexafluoride or as sulfur(VI) fluoride.

TOPIC 5

Question 1

The empirical formula gives the number of atoms of each of the component elements in any compound, expressed as the simplest whole-number ratio. Every compound, be it ionic or covalent, has an empirical formula.

Compounds that exist as molecules also have a molecular formula which gives the actual number of atoms of each of the component elements present in the molecule. This formula is always a simple multiple of the empirical formula. For example, the compound hydrogen peroxide has the empirical formula HO and the molecular formula H_2O_2. Molecular formulas do not indicate the structure of the molecule. Ionic compounds do not exist as molecules and therefore do not have molecular formulas.

The structural formula for a compound shows in some way how the atoms are arranged within the molecule. For example, the compound hydrogen peroxide has the structure H–O–O–H where each dashed line represents a bonding pair of electrons. Another way of showing this structure without the dashed lines is simply to write it as HOOH.

Question 2

(a) Metallic bonding is the characteristic type of bonding that holds atoms of metals together. The outer electrons of metal atoms are sufficiently loosely held by the nuclei for them to be able to move from atom to atom, thus effectively being shared by many atoms. The mobility of outer level electrons in metals is the basis for metallic properties such as malleability and ductility (bonds break but then re-form when the metal's nuclei move relative to each other), conduction of electricity (electrons jump from atom to atom through the metal when a voltage is applied), and metallic lustre when freshly cut (electrons absorb and then re-radiate electromagnetic energy in the form of light from surroundings). Metals have the property of conducting electricity in both the solid and molten states, without any degradation of the element.

(b) Network covalent bonding occurs between atoms of non-metals and may be present in elements such as carbon (e.g. diamond) as well as in compounds such as silicon dioxide (beach sand). This type of bonding consists of covalent bonds joining the various atoms with no defined limit on the number of atoms involved. Covalent network bonding results in the substance having a very high melting point as sufficient heat energy must be applied in order to break the very strong covalent bonds between the atoms. Likewise, these solids are not soluble in any type of solvent. Normally they are not good conductors of electricity but graphite, an allotrope of carbon, does conduct electricity due to the weakness of the bonds between the planar layers of C atoms.

Question 3

Covalent bonding arises from the overlap of electron orbits containing the outer (valence level) electrons of the bonded atoms. Consequently the bonded atoms are not free to pack into the spatial arrangement which minimises electrostatic repulsions and maximises electrostatic attractions as is the case with ionic compounds, but instead they must remain attached by their bonds to each other in the molecule. The particular 3-dimensional shape of any given molecule results from repulsions between the valence level electrons, both bonding and non-bonding, on each atom.

Question 4
The elements hydrogen, nitrogen, oxygen, plus the halogens, fluorine, chlorine, bromine and iodine, exist as diatomic molecules at normal temperatures and pressures.

Question 5
The Law of Conservation of Matter states:
"Matter can be neither created nor destroyed, but merely converted from one form to another."
Thus in a balanced equation all the atoms on the left hand side must appear on the right hand side in the same numbers.

TOPIC 6
Question 1
(a) A solvent is a substance (usually but not necessarily a liquid) in which another substance which may be gaseous, liquid or solid, has been dissolved to form a homogeneous solution.

(b) The substance dissolving in a solvent to form a solution is called the solute.

(c) The electrons in covalent bonds between any two atoms are only shared equally when the bonded atoms are of the same element. Otherwise, one atom usually has greater electron attracting power (said to be more electronegative) and consequently the bonding electrons are displaced closer to it. The covalent bond then has a slight ionic character and is said to be a polar covalent bond.

(d) If when the individual polarities of all the bonds in a molecule are combined, an overall symmetry of the charge is not achieved, the molecule will have an overall molecular dipole and it is said to be a polar molecule. Water is an angular molecule and is an example of a polar molecule while carbon dioxide, due to its linear shape, is a non-polar molecule.

Question 2
Water is a highly polar molecule and the oxygen atom end carries a slight negative charge which is attracted to the positively charged Na^+ ions in the sodium chloride crystal. Similarly, the hydrogen atoms of water molecules carry a slight positive charge and are attracted to the negative charge of the Cl^- ions. This interaction of the Na^+ and Cl^- ions with water molecules releases much energy (called the crystal lattice energy) and allows the ions to break away from the crystal, entering into solution as aquated ions, each being surrounded by a sheath of up to six water molecules.

Question 3
If a solution containing an ionic compound as the solute is heated sufficiently to boil off the water which is a volatile solvent, the non-volatile ionic compound will remain. When sufficient water has been removed, the ions will reform the ionic solid which can be reclaimed from any remaining solution by filtration.

Question 4
(a) An acid + a reactive metal forms a salt and hydrogen gas.

(b) An acid + a metal oxide forms a salt and water.

(c) An acid + a metal hydroxide forms a salt and water.

(d) An acid + a carbonate forms a salt and water and carbon dioxide gas.

TOPIC 7.
Question 1
Atoms of each element have their own unique number of protons plus neutrons. The mass of an atom is mostly due to its protons and neutrons, so atoms of different elements will have different masses, increasing as their atomic number increases.

Question 2
(a) Relative atomic weight: the mass of an atom of an element expressed relative to a particular datum. Initially the datum used was to take the H atom as exactly 1 but currently the datum used is to take 1/12 the mass of the C^{12} atom as 1. [Note that "weight" and "mass" are often used interchangeably in this context.]

(b) The gram atomic weight of an element is its relative atomic weight expressed as grams and is the mass of 1 mole of atoms of that element. For example, the relative atomic weight of carbon is 12.01 (no units) and its gram atomic weight is therefore 12.01 g mol^{-1}.

(c) The chemical mole can be defined in two ways:
(i) as an Avogadro number of any species. This definition can apply not only to the constituent particles of an element or compound that have a chemical formula, but also to objects that have no chemical formula such as electrons.
(ii) as the amount of material in a gram formula weight or molar mass of any element or compound. For example, a mole of the element carbon would be the amount of carbon in 12.01 g of that element.

(d) The molar mass of an element or compound is its formula mass expressed as grams. For water, the molecular weight is 18.02 so 1 mole of water would have a mass of 18.02 g, its molar mass.

Question 3.
One mole of any compound is the amount of material contained in a mass equal to its formula weight expressed in grams, i.e. its molar mass. Thus to convert a given mass of a compound to moles, calculate its molar mass by adding the relative atomic masses of all its constituent atoms and divide the given mass by the molar mass.

Question 4
Atoms can only combine in integer ratios as they cannot be subdivided - e.g. H_2O has the ratio of H:O of 2:1, HCl has the ratio of H:Cl of 1:1.
However, the masses of atoms are determined mainly by the masses of their constituent protons and neutrons which are not equal. Also the existence of isotopes for most elements results in weighted averages for atomic masses which are even further removed from being integers.
In hydrogen chloride for example, one Cl atom combines with one H atom in the ratio of 1:1 forming an HCl molecule. However, the average masses of H and Cl atoms are not in a simple numerical ratio (1.01:35.45) so they cannot combine in simple numerical ratios by mass.

TOPIC 8.
Question 1
Any given pure compound always has the same formula - i.e. the number of each constituent atom of each element in that compound is invariant. Likewise, the relative atomic weight of each element is fixed. Thus knowing the number of atoms of any given element in a compound's formula, the % by mass of that constituent element can be derived from its total contribution to the molar mass.
e.g. for the compound of formula H_2O, molar mass = $1.01 \times 2 + 16.00 = 18.02$ g mol^{-1} and the contribution to that total from H atoms = $1.01 \times 2 = 2.02$ g mol^{-1}.
Thus the % by mass of H in H_2O = $(2.02 / 18.02) \times 100 = 11.2$ %.

Question 2
The percentage by mass of each constituent element in a pure compound can be obtained by experiment. The empirical formula for the compound can be deduced by obtaining the relative number of moles of atoms of each of those constituent elements because the relative number of constituent atoms is in the same ratio as that relative number of moles.

Question 3
Firstly the empirical formula for the compound must be obtained from the experimental % composition data. Then the molecular weight for that compound must also be obtained by experiment. Comparing the molecular weight with the empirical formula weight, the former will be found to be a simple multiple of the latter. That multiple is then applied to the empirical formula, multiplying the number of each constituent atom by the multiple gives the molecular formula.