Chemistry EOY Revision 2012

Atomic Structure

- An atom has the same chemical properties of the element.
- Atoms are made of:
  - Protons – Charge: +1, Relative Mass: 1
  - Neutrons – Charge: 0, Relative Mass 1 (Cannot write neutral charge)
  - Electrons – Charge -1, Relative Mass 0
- Nucleus of an atom, is the centre portion. The nucleus is made of protons and neutrons, while electrons orbit the nucleus. Hence, protons and neutrons collectively are called **nucleons**. As there are no electrons in the nucleus to balance the charge, the nucleus usually has an overall **positive** charge.
- The overall charge of an atom is usually neutral, because there are an equal number of protons and electrons.
- Notation of Atoms in the periodic table:
  - Top Left side of the symbol - Mass number
    - Mass number = Protons + Neutrons
    - Also known as Nucleon number
  - Bottom Left side – Atomic number
    - States the position of the atom on the periodic table
    - Atomic Number = Protons alone (or electrons alone, if it is a neutral atom)
  - Calculations of protons/neutrons/electrons
    - Protons: Atomic number
    - Electrons: Atomic Number
    - Neutrons: Mass Number – Atomic Number
- Isotopes are atoms of the same element with different number of neutrons.
  - Same number of protons, same number of electrons, different number of neutrons.
  - Will have same atomic number, but different mass number (due to increased neutrons)
  - Eg. Hydrogen normally has 1 neutron, but it has 2 isotopes with 2 and 3 neutrons respectively.
  - Isotopes have the same chemical properties, but different physical properties (eg. Mass, boiling point, melting point)
- Bohr's model (Up till 20th element):
  - Has 4 shells
  - By reading off the number of electrons in the 4 shells, you get the **electronic configuration** (eg. 2.8.8.2 is the maximum).
  - The shell must contain the maximum number of electrons, before an electron goes to occupy a new shell.
  - Therefore, from electronic configuration, you can obtain information about the atom.
  - Eg. Of electronic configuration – Magnesium has an electronic configuration of 2.8.8.2

- The outermost shell is known as valence shell, and the electrons in the valence shell are known as valence electrons.
- Elements with the same number of valence electrons tend to have similar characteristics, because valence electrons are responsible for chemical reactions and chemical bonding.

- Visual depiction of atoms using Bohr's model:
  - Fill opposites first
  - Move on to the next shell only when the current shell has reached its capacity.
  - Electrons in the 2^{nd} and 3^{rd} shells need to be paired when there are more than 4 of them.
  - Use dots and crosses (crosses are faster because you have to colour the dots)

Periodic Table

- Elements are arranged in terms of atomic number, or number of protons (or number of electrons, if it is a neutral atom).
- The vertical rows are called **GROUPS**.
  - General Notes:
    - Usually numbered with roman numerals.
    - Metallic properties decrease as group number increases
    - Elements in the same group have the same number of valence electrons (As the number of valence electrons can differ when the shell is full, the group with full valence shell is called group 0, as in 0 valence electrons in the next shell)
  - Alkali Metals: Group 1 (extreme left)
    - Reactive metals (react easily in air/ tarnish)
    - Have generally low density, melting and boiling points
    - **Density increases, reactivity increases** (how vigorous the reaction will be) and **melting/boiling points decrease** as you go down the group
  - Halogens: Group VII (group 7)
    - Non-metals which are **reactive**
    - Low melting/ boiling points.
    - Diatomic Molecules
    - **Melting/boiling points increase, reactivity decreases** as you go down the group.
    - Colour darkens as you go down the group.
  - Noble Gases: Group 0 (extreme right)
    - Non-metals which are not reactive
    - Have noble gas configuration
    - Low melting/boiling points (gases at r.t.p)
    - Colourless and monoatomic
- The horizontal rows are called **PERIODS**
  - The period number refers to the number of electron shells. (Eg. Period 1 elements have 1 shell, period 2 elements have 2 shells, etc.)
- Atom sizes:

- As the group number increases, size of atom decreases, as the number of shells would stay the same across the period, but there would be more protons and electrons as the group number increases, causing stronger forces of attraction between the nucleus and the electrons of the atom.
- As period number increases, size of atom increases (due to increases electron shells)

- A "staircase" can be drawn from Boron (refer to attached periodic table):
    - Everything on the left of the staircase is a metal.
    - Everything on the right of the staircase is a non-metal.
    - Any element that "touches" the staircase line is a metalloid (has some properties of metals, and some properties of non-metals)

<u>Chemical Bonding</u>

- Atoms try to attain a noble gas configuration to become more chemically stable, by going through chemical bonding.
- Ionic bonding:
    - Ion means: An atom or molecule with a net electric charge (no longer neutral at 0 charge) due to gain/loss of electrons.
    - Formed when atom(s) of one element lose one or more electrons, and atom(s) of another element receive them, causing all atoms to reach desired noble gas configuration.
    - As such, the atoms of the two elements bond due to electrostatic forces of attraction between positively and negatively charged ions.
    - Formed between a metal and a non-metal.
    - When an atom loses an electron, the atom has a positive charge (due to loss of negative charge particles).
    - When an atom gains an electron, the atom will have a negative charge (due to gain of negative charge particles).
- Properties of ionic compounds:
    - Hard and crystalline solids at r.t.p
    - High melting and boiling points (therefore solids at r.t.p). This is because ionic bonds are strong, and are fixed in a giant lattice structure, and a lot of energy is required to overcome these strong electrostatic forces of attraction.
    - Conductors of electricity in molten or aqueous states, but not in solid states. This is because the ionic compounds are held in fixed positions by strong ionic bonds, and thus the ions are unable to move freely to conduct electricity by carrying the electrical charges. However, in molten or aqueous state, the ions can move freely, and can carry the electrical charges (therefore being conductors of electricity).
    - Usually soluble in water, but insoluble in organic compounds. This is because the ions attract water molecules which disrupt the crystal structure, causing the ions to separate and go into the solution. They do not attract the molecules of organic compounds (why??).
- Covalent bonding:
    - Formed by the sharing of electrons
    - Between atoms of non-metal elements.
    - All gas molecules are bonded by covalent bonds

- Properties of covalent bonds:
    - Mostly liquid or gaseous at r.t.p (except iodine and sulphur, which are solids).
    - Low melting and boiling points. This is because the molecules are held together by weak intermolecular forces of attraction, so very little energy is needed to overcome the weak intermolecular forces of attraction. Note that the covalent bonds are not broken in state change, but it is the force between the covalent molecules that is broken.
    - Mostly insoluble in water (except sugar, hydrogen chloride, and alcohol), but soluble in organic compounds
- Drawing of ionic compounds:
    - Atoms only lose electrons when they have 1, 2 or 3 valence electrons to begin with.
    - Atoms only gain electrons when they have 5, 6 or 7 valence electrons to begin with.
    - Write the element symbol in the centre of the atom, and if the element is unknown, use X or whatever the question gives you.
    - Bracket the entire drawing of the atom, and write the overall charge of the atom (after the transfer of electrons has taken place) on the top right hand corner outside the bracket.
        - Write 2+, not +2 (number comes first)
    - For all electrons of the first atom to begin with, use a dot, and for all electrons of the second atom to begin with, use a cross.
        - This shows which electron originated from which atom.
        - Try to use crosses for the atom with more electrons that need to be drawn (to save time).
    - When 2 or more atoms of the same element are involved in the ionic compound, indicate this by putting a large "2" or "3" or whatever on the left side of the bracket.
- Drawing of covalent compounds:
    - Draw only the valence electron shell, unless the question tells you not to, because drawing all the shells will cause the diagram to be very messy.
    - The atoms are drawn like a Venn diagram, with the common area containing the shared electrons.
    - If there are more than 2 atoms, alternate them with dots and crosses – no 2 atoms which are sharing electrons should have the same symbol.
    - Structural formula: An alternate method of representing covalent compounds. Instead of dots and crosses, just write the element names, and draw 1 line to connect them for each pair of shared electrons.

Acids and Bases

- An acid is a substance that produces hydrogen ions (H^+) when dissolved in water.
    - Acidic properties are only seen when dissolved in water because acids disassociate in water to form hydrogen ions. However, they do not disassociate to form hydrogen ions in organic solvents.
- Properties of Acids:
    - Sour
    - Turns blue litmus paper red (Acids are red on the universal indicator)
    - React with metals to form a salt and hydrogen

- React with carbonates/ hydrogen carbonates to produce salt, carbon dioxide, and water
- React with metal oxides/ hydroxides (bases) to form salt and water.

- A base is a substance that reacts with an acid to form salt and water.
- An alkali is a substance that disassociates to produce hydroxide ions in water (any metal oxide or hydroxide).
- Properties of alkalis:
  - Feel slippery and soapy (usually taste bitter)
  - Turns red litmus paper blue (Alkalis are blue on the universal indicator)
  - React with acids to form salt and water
  - React with ammonium compounds to form a salt, ammonia gas and water
- pH scale:
  - >7 is alkaline
  - 7 is neutral
  - <7 is acidic
- Methods of pH measurements
  - Methyl Orange (turns from red to yellow at pH 5)
  - Bromothymol blue (turns from yellow to blue at pH 5)
  - Phenolphthalein (turns from colourless to pink at pH9)
  - Universal Indicator Paper (changes colour at every pH)
    - Disposable, therefore reducing risk of contamination
    - Can be used on very small sample size (unlike pH meter, which needs bulb to be covered)
    - Immediate change can be observed (unlike pH meter which takes some time to calibrate)
  - pH meter (shows pH level to 1dp)
- Good Indicators:
  - Observable and show a distinct colour change.
  - Required only in small amounts
  - Do not affect the pH level of the solution being tested.
- Test for gases:
  - Hydrogen – Gas extinguishes a lighted splint with a "pop" sound.
  - Carbon Dioxide – White precipitate forms in the solution when bubbled through limewater.
  - Ammonia Gas – Turns red litmus paper blue, and has a pungent smell
- Neutralization:
  - Happens when acids + alkalis
  - Produces salt + water (Hydrogen ions combine with hydroxide ions to form water $H^+ + OH^- \rightarrow H_2O$
  - Uses:
    - Control pH of soil
    - Treats indigestion, by neutralizing excess hydrochloric acid in stomach
    - Treatment of insect stings by adding a base if the sting is acidic
    - Toothpaste has alkalis to neutralize acids produced by bacteria
    - Treatment of waste water

- Oxides
  - Acidic Oxides:
    - Oxides of non-metals
    - React with water to produce acids
    - Have acidic properties
    - Examples: Sulfur dioxide, Carbon dioxide, Phosphorous V oxide
  - Basic Oxides :
    - Oxides of metals
    - Have basic properties
    - Examples: Copper (II) oxide, Magnesium oxide, Calcium oxide
  - Amphoteric Oxides:
    - Examples: Zinc Oxide, Aluminium oxide, Lead (II) oxide
  - Neutral Oxides:
    - Do not react with acids or bases
    - Examples: Water, carbon monoxide, nitrogen monoxide
- Salts
  - Usually in the aqueous state
  - When naming salts, the name of the metal is put in front, then followed by the other part of the acid.
    - If hydrochloric acid is involved: ____________chloride
    - If sulphuric acid is involved: _____________ sulphate
    - If nitric acid is involved: _____________ nitrate
    - If sulphurous acid is involved: _____________ sulphite
- Reactions:
  - Acid + Metal → Salt + Hydrogen
  - Acid + Base → Salt + Water
  - Acid + Carbonate → Salt + Water + Carbon dioxide
  - Alkali + Ammonium compounds → Salt + Ammonia Gas + Water
- Table of Negative Ions

Charge	Name	Chemical Formula
-1	Fluoride	Fl^-
	Chloride	Cl^-
	Bromide	Br^-
	Iodide	I^-
	Hydroxide	OH^-
	Nitrate	NO_3^-
-2	**Oxide**	O^{2-}
	Sulfide	S^{2-}
	Sulfate	So_4^{2-}
	Carbonate	CO_3^{2-}
-3	**Phosphate**	PO_4^{3-}

- Table of Positive Ions

Charge	Name	Chemical Formula
+1	Hydrogen	H^+
	Lithium	Li^+
	Sodium	Na^+
	Potassium	K^+
	Silver	Ag^+
	Ammonium (Not Ammonia)	NH_4^+
+2	**Magnesium**	Mg^{2+}
	Calcium	Ca^{2+}
	Barium	Ba^{2+}
	Zinc	Zn^{2+}
	Lead (II)	Pb^{2+}
	Copper (II)	Cu^{2+}
	Iron (II)	Fe^{2+}
+3	Aluminum	Al^{3+}
	Iron (III)	Fe^{3+}

- Diatomic Molecules (everuthing else that exists as a single element is monoatomic):

hydrogen (H2)
nitrogen (N2)
oxygen (O2)
fluorine (F2)
chlorine (Cl2)
bromine (Br2)
iodine (I2)

- Acid Formulae:
  - Sulfuric Acid - H_2SO_4
  - Nitric Acid - HNO_3
  - Hydrochloric Acid - HCl

- State symbols
  - Solid – s (For all metal or metalloid ions, oxides, hydroxides and carbonates)
  - Liquid – l (For water)
  - Aqueous – aq (For all acids, alkalis, and salts), aq is used when a substance is dissolved in water
  - Gas – g (For hydrogen and carbon dioxide)