1A - The changing atom

History of the atom

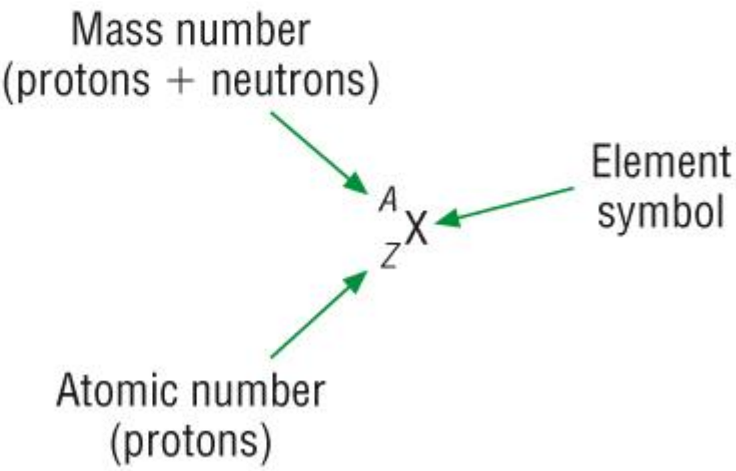
- The model of the atom has changed as our observations of its behavior and properties have increased.
- A model is used to explain observations. The model changes to explain any new observations.
- A map gives you an overview of a town/city. As technology has improved our maps have become more accurate and detailed. This is the same for scientific models.

Atomic structure

Protons, electrons and neutrons

Sub-Atomic Particle	Atomic Mass	Atomic Charge	working it out
Proton	1	+1	**bottom**
Electron	1/2000	-1	**bottom**
Neutron	1	0	**top - bottom**

- From GCSE this table shows that only protons and neutrons have a mass.
- Since different elements have different atoms.
- These atoms have different numbers of protons and neutrons (and



- So an atom of one element **must** have a different mass from another element, we call this the **Mass Number**.
- The number of protons determines which element an atom is and the bottom number tell us this, we call this the **Atomic number**.

Examples:-

1) Lithium

$^{7}_{3}\text{Li}$

Protons = 3
Electrons = 3
Neutrons = 4 (7-3)

2) Nitrogen

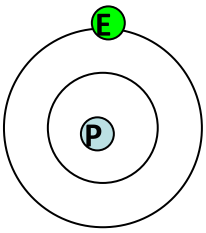
$^{14}_{7}\mathrm{N}$

Protons = 7
Electrons = 7
Neutrons = 7 (14-7)

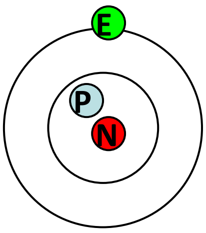
Isotopes

An atom of the same element that has the same number of protons and electrons but a different number of neutrons.

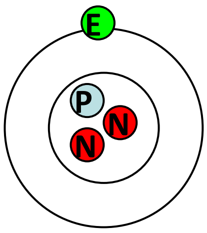
- As the atom has the same number of protons and electrons it will have the same chemical properties.
- They are all hydrogen atoms because they all have the **same number of protons**
- Hydrogen can be used as an example:-
-



Hydrogen – 0 Neutrons



Hydrogen – 1 Neutron (deuterium)



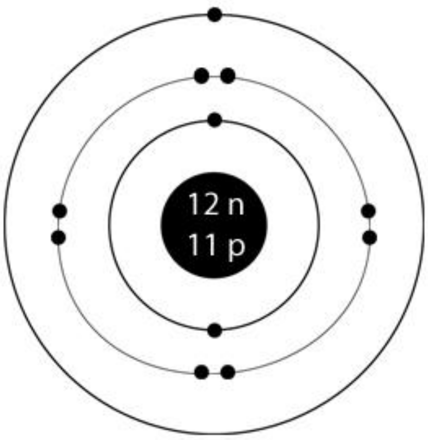
Hydrogen – 2 Neutrons (tritium)

Atomic structure of ions

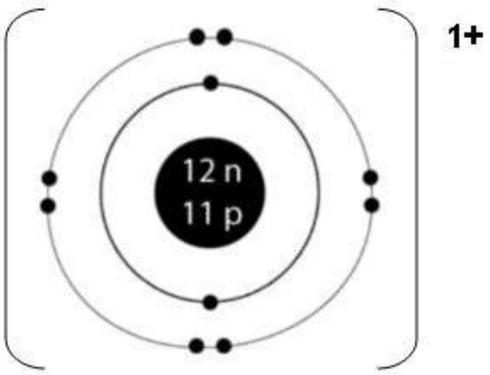
- When an atom becomes an ion it must have gained or lost electrons.
- This will have an effect on its atomic structure:

Sodium ion

- The sodium ion has one less electron than its number of protons therefore the charge is 1+ (as electrons are negative)



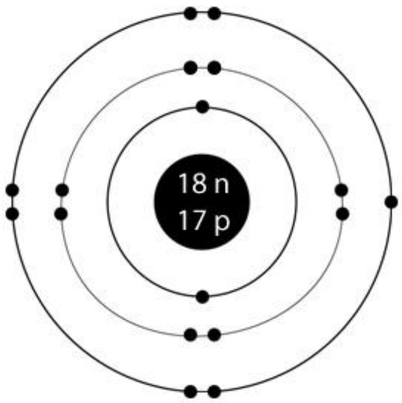
Mass number, A = 23
Atomic number, Z, P = 11 (+)
Electrons, e = 11 (-)
Neutrons, N = 12
charge is neutral, 11P (+) = 11e (-)



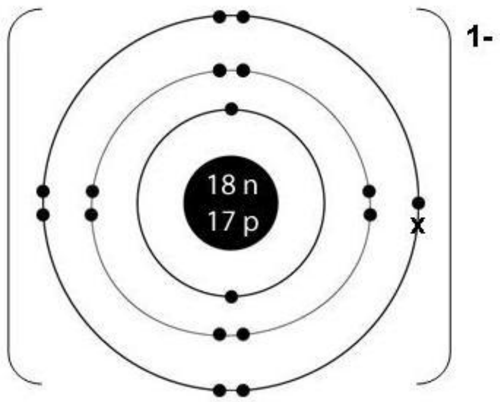
Mass number, A = 23
Atomic number, Z, P = 11 (+)
Electrons, e = 10 (-)
Neutrons, N = 12
charge is 1+, 11P (+) = 10e (-)

Chloride ion:

- The chloride ion has one more electron than its number of protons therefore the charge is 1- (as electrons are negative)



Mass number, A = 35
Atomic number, Z, P = 17 (+)
Electrons, e = 17 (-)
Neutrons, N = 18
charge is neutral, 17P (+) = 17e (-)



Mass number, A = 35
Atomic number, Z, P = 17 (+)
Electrons, e = 18 (-)
Neutrons, N = 18
charge is 1-, 17P (+) = 18e (-)

- **Questions 1, 2 p7 / 1 p35 / 1p36**

Chemical equations:

Writing formulae and balancing equations

- The first part of which is chemical symbols and writing formulae.
- Some names are straightforward whereas others are unexpected.

Writing formula

- A formulae is a shorthand way of describing a chemical substance.
- The composition and elements present are represented in the chemical formulae.
- Following a set of rules the chemical formula for a substance can deduced:-

Periodic Table Group	Charge on ion	Other ions
1	1+	H^+ NH_4^+ Ag^+
2	2+	Co^{2+} Cu^{2+} Fe^{2+}
3	3+	Fe^{3+}
6	2-	SO_4^{2-} CO_3^{2-}
7	1-	OH^- NO_3^-

- The ions in a chemical formula must **add up to zero**.
- Use subscripts after an ion in a formula to double/triple that ion so the sum=0. eg. $CuCl_{\mathbf{2}}$
- If you are double/tripling ions that consist of more than one element brackets must be used. eg. $Ca\mathbf{(}OH\mathbf{)}_2$
- If Roman numeral numbers follow a metal ion in brackets, that tells you the positive charge of that metal ion. They are usually Transition Metals as they can have more than one oxidation state. eg. Copper (II) Chloride. Copper can also exist as copper (I).
- Water of crystallization can be added to compounds after a **dot**. eg. $CuSO_4\mathbf{.}7H_2O$

How to work out formulae
Sodium Chloride –

Na^+	:	Cl^-	Write the ions with charges
1	:	1	Write the ratio of the charges
			Scale up if necessary to =0
$NaCl$			Bring together omitting the charges

Copper (II) Chloride

Cu^{2+}	:	Cl^-	Write the ions with charges
2	:	1	Write the ratio of the charges
Cu^{2+}	:	$Cl^-{}_2$	Scale up if necessary to =0
$CuCl_2$			Bring together omitting the charges

Calcium Hydroxide

Ca^{2+}	:	OH^-	Write the ions with charges
2	:	1	Write the ratio of the charges
Ca^{2+}	:	$(OH^-)_2$	Scale up if necessary to =0
$Ca(OH)_2$			Bring together omitting the charges

Balancing equations

- One of the most important concepts in chemistry is that mass is always conserved.
- You always have exactly the same at the end as what you started.
- If you follow 4 steps – you cant go wrong:-

Step 1 Write out the word equation.

Iron + Hydrochloric acid → Iron (II) chloride + Hydrogen

Step 2 Write the correct formula underneath.

Iron + Hydrochloric acid → Iron (II) chloride + Hydrogen
$Fe + HCl \rightarrow FeCl_2 + H_2$

Step 3 Balance using large numbers to scale up the number of molecules. A rule of thumb to help you balance is to balance the elements in this order - **MACHO:**

Metal
Any other element
Carbon
Hydrogen
Oxygen
A table helps.

Iron + Hydrochloric acid → Iron (II) chloride + Hydrogen
$Fe + \mathbf{2}HCl \rightarrow FeCl_2 + H_2$

Atoms			**Balanced**
Fe	1	1	Yes
Cl	1**x2=2**	2	Yes
H	2	2	Yes

Put the **2** before the molecule with the element you are scaling up. You now have the same number of atoms on each side.

Step 4 All that remains is to add the state symbols:

Iron + Hydrochloric acid → Iron (II) chloride + Hydrogen

$Fe_{(s)} + 2HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$

- **You now have a balanced chemical equation.** This can be used to record an experiment or calculate amounts to mix.

Questions 1-2 p19 / 11 p35

Acids and Bases

- pH less than 7
- Turn litmus red.
- Is neutralised by alkali and bases.

Strong acids - Hydrochloric acid:

Strong acids dissociate fully in water releases H^+ ions into the solution:

- When hydrogen chloride is bubbles through water it dissolves. With water, it completely ionises:

$$HCl_{(g)} + aq \rightarrow \mathbf{H^+}_{(aq)} + Cl^-_{(aq)}$$

Acids can be defined as proton donors

Weak acids - Ethanoic acid

Weak acids partially dissociate in water

- Weak acids such as ethanoic acid only dissociate a small amount giving a small amount of hydrogen ions:

$$CH_3COOH_{(l)} + aq \rightleftarrows \mathbf{H^+}_{(aq)} + CH_3COO^-_{(aq)}$$

Bases

- These are chemical opposites to acids which include: metal oxides, hydroxides, ammonia and amines.
- They neutralise acids

Bases can be defined as proton acceptors

Alkalis

- These compounds dissolved in water gives a solution with a pH of greater than 7
- Some common alkalis: sodium hydroxide, potassium hydroxide and ammonia.
- An alkali forms hydroxide ions in aqueous solution:

$$NaOH_{(s)} + aq \rightarrow Na^+_{(aq)} + \mathbf{OH^-}_{(aq)}$$

Alkalis dissociate to give hydroxide ions in solution

- Weak bases such as ammonia only dissociate a small amount giving a small amount of hydroxide ions:

$$NH_{3(g)} + H_2O_{(l)} \rightleftarrows NH_4^+{}_{(aq)} + \mathbf{OH^-}{}_{(aq)}$$

Biological acids and bases:

Acids: fatty acids, amino acids, nucleic acids
Amphoteric: This means they have acidic and basic properties - amino acids - NH_2 and CO_2H

Questions 1-2 p23 / 12 p35

<u>Salts</u>

All salts contain the following:

- A positive cation - usually a metal or ammonium (NH_4^+)
- A negative anion derived from an acid:

Acid	Cation...Salt
Sulphuric	...sulphate
H_2SO_4	...SO_4^{2-}
Nitric	...nitrate
HNO_3	...NO_3^-
Hydrochloric	...chloride
HCl	...Cl^-

- These salts are formed with one of the following reactions:

1	**Metal**	**+**	**Acid**	**→**	**Salt**	**+**	**Hydrogen**			
2	**Metal Oxide**	**+**	**Acid**	**→**	**Salt**	**+**	**Water**			
3	**Metal Hydroxide**	**+**	**Acid**	**→**	**Salt**	**+**	**Water**			
4	**Metal Carbonate**	**+**	**Acid**	**→**	**Salt**	**+**	**Water**	**+**	**Carbon dioxide**	

Constructing balanced chemical equations:

Example: Sodium hydroxide reacts with sulphuric acid, write a balanced chemical equation for this reaction:

Step 1 Use the 4 reactions above to identify your reaction - Sodium **hydroxide** tells us it is reaction **(3)**

Sodium hydroxide + Sulphuric acid → Salt + Water

Step 2 Work out the name of your salt:

Sodium hydroxide + Sulphuric acid → **Sodium sulphate** + Water

Step 3 Write the correct formula underneath.

Sodium hydroxide + Sulphuric acid → Sodium sulphate + Water

$$\mathbf{NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O}$$

Step 4 Balance the equation, in this order - works 90% of the time

Metal
Any other element
Carbon
Hydrogen
Oxygen

Sodium hydroxide + Sulphuric acid → Sodium sulphate + Water

$NaOH$ + H_2SO_4 → Na_2SO_4 + H_2O

Element			Balanced
Na	**1**	**2**	**NO**
S			
H			
O			

Sodium hydroxide + Sulphuric acid → Sodium sulphate + Water

2$NaOH$ + H_2SO_4 → Na_2SO_4 + H_2O

Element			Balanced
Na	1 x **2** = 2	2	Yes
S	1	1	Yes
H	**4**	**2**	**NO**
O			

Sodium hydroxide + Sulphuric acid → Sodium sulphate + Water

$2NaOH$ + H_2SO_4 → Na_2SO_4 + **2**H_2O

Element			Balanced
Na	1 x **2** = 2	2	Yes
S	1	1	Yes
H	4	2 x **2** = 4	**YES**
O	6	6	YES

Step 4 Add the state symbols:

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$$

You now have a balanced chemical equation.

Acid salts:

- Sulphuric acid has 2 H^+ ions, we call these **diprotic** acids as they can both be replaced:

  H_2SO_4 → **$NaHSO_4$** Sodium hydrogen sulphate

- $NaHSO_4$ is known as an acid salt as it can **donate another proton:**

  $NaHSO_4$ → Na_2SO_4 Sodium sulphate

Ammonium salts and fertilisers

- All fertilisers contain Nitrogen in the form of ammonium, NH_4^+ and / or Nitrate, NO_3^-.
- Ammonium salts are formed when acids react with ammonia, NH_3:

$$\mathbf{NH_{3(aq)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}}$$

Calculating the % of N in a fertiliser:

- As it is the N element that is important, we need to be able to calculate the % of that element in the compound.
- Use this formula to calculate the %:

$$\textbf{\% Element} = \frac{\textbf{No of atoms of that element in the formula x Ar of the element x 100}}{\textbf{Mr of the compound}}$$

Worked example:

- Calculate the % of N in ammonium nitrate, $\mathbf{NH_4NO_3}$:

1) Write out the formula:

$$\text{\% Element} = \frac{\text{No of atoms of that element in the formula x Ar of the element x 100}}{\text{Mr of the compound}}$$

2) Calculate the Mr of the compound:

Element	**Ar**	**No that element**	**Sub total**
N	14	2	28
H	1	4	4
O	16	3	48
		Mr =	**80**

3) Fill in the rest of the formula and calculate the answer

$$\text{\% Element} = \frac{\text{2 (N) x 14 (Ar) x 100}}{80}$$

% Element = 35%

Questions 1-2 p25 / 13 p35

<u>Oxidation numbers</u>

- It is used to describe the number of electrons used to bond with another atom.
- It is also used for combining powers of atoms.
- It is a type of 'book keeping' for electrons.
- It is a number describing the movement of electrons and is found by the application of certain rules:-

<u>Rules for assigning Oxidation Numbers</u>

1) Ox. No. of an element = 0

2) Ox. No. of each atom in a compound counts separately. Sum = 0

3) Ox. No. of an ionic element = charge on ion.

4) In a polyatomic ion (SO_4^{2-}), The sum of the Ox. No.'s of the atoms = charge on ion.

5) H = +1 except with metals (metal hydrides = -1).

6) Gp 7 (Halogens) = -1 (except with oxygen)

7) O = -2 except in peroxides (H_2O_2, O = -1)

Examples:

Compounds -

CO_2 — **O = - 2 each (2 x -2 = - 4)**
The molecule is neutral so C + - 4 = 0
C = + 4

Formula -

$MgCl_2$ — **Mg = + 2**
Cl = - 1 each (2 x -1 = - 2)

Molecular ions -

NO_3^- — **O = - 2 each (3 x -2 = - 6)**
The molecule = - 1 so
N + - 6 = -1
N = + 5

Oxidation numbers in chemical names:

- Some elements form compounds where they could have a different charge / oxidation number.
- You will already be familiar with the Transition metals for this from GCSE:

Transition metals:

Compound	**Name**	**Element with different Ox. No.**	**Ox. No. of that element**
$FeCl_2$	Iron (II) chloride	Fe	+2
$FeCl_3$	Iron (III) chloride	Fe	+3

- Roman numerals indicate the oxidation number of the element before it.
- This also occurs with oxyanions:

Oxyanions:

Oxyanion	Name	Element with different Ox. No.	Ox. No. of that element
NO_2^-	Nitrate (III)	N	+3
NO_3^-	Nitrate (V)	N	+5

- These are negative molecules that contain an oxygen atom.
- Oxyanions usually end in '**ate**' to indicate the presence of oxygen.

Questions 1-4 p31 / 14 p35

<u>**Redox reactions**</u>

- Our understanding of oxidation and reduction was limited to reactions involving oxygen and hydrogen:

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

- **Magnesium** has been **oxidised** as it has **gained oxygen**.
- But it has also **lost 2 electrons**:

$$Mg_{(s)} - 2e^- \rightarrow Mg^{2+}_{(s)}$$

- **Oxygen** has been **reduced** as it has **lost oxygen**.
- But each oxygen has also **gained 2 electrons**:

$$O_{2(g)} + 4e^- \rightarrow 2O^{2-}_{(s)}$$

- A new definition could be used involving electrons and this could be applied to all reactions:

Oxidation - is addition of oxygen / loss of hydrogen /
Is
Loss of electrons

Reduction - loss of oxygen / addition of hydrogen /
Is
Gain of electrons

- Oxidation and reduction must occur simultaneously as all reactions involve a movement of electrons.
- These reactions are given the shorthand term of **REDOX** reactions. As they involve
  REDuction and **OX**idation
- Redox reactions can now be applied to reaction that do not involve oxygen or hydrogen:

$$\mathbf{2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}}$$

- Looking at each species often makes it easier to decide which has been oxidised and reduced.

- Adding the oxidation number makes it even easier:

$$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$
$$\underset{0}{Fe_{(s)}} \rightarrow \underset{+3}{Fe^{3+}_{(s)}} + 3e^-$$

- To go from 0 to +3 means electrons have been lost = **Oxidation (increases Ox No's)**

$$\underset{0}{1/2Cl_{2(g)}} + e^- \rightarrow \underset{-1}{Cl^-_{(s)}}$$

- To go from 0 to -1 means electrons have been gained = **Reduction (reduces Ox No's)**

Oxidising and reducing agents:

$$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$

$$\underset{0}{Fe_{(s)}} \rightarrow \underset{+3}{Fe^{3+}_{(s)}} + 3e^- \quad \textbf{Oxidation}$$

$$\underset{0}{1/2Cl_{2(g)}} + e^- \rightarrow \underset{-1}{Cl^-_{(s)}} \quad \textbf{Reduction}$$

- As chlorine accepted the electrons from iron for it to be oxidised we say that **chlorine** is the **oxidising agent**
- As iron gave the electrons to chlorine for it to be reduced we say that **iron** is the **reducing agent:**

Oxidation **- Reducing agent**
Is
Loss of electrons

Reduction = **Oxidising agent**
Is
Gain of electrons

Examples of oxidising and reducing agents

- Iron ions can exist as Fe^{2+} and Fe^{3+}:
- Fe^{2+} solutions are green and Fe^{3+} solutions are brown.

$$\underset{\text{Green}}{Fe^{2+}} \rightarrow \underset{\text{Brown}}{Fe^{3+}} + e^- \quad \textbf{Oxidation (requires an oxidising agent)}$$

- And:

$$\underset{\text{Brown}}{Fe^{3+}} + e- \rightarrow \underset{\text{Green}}{Fe^{2+}} \quad \textbf{Reduction (requires a reducing agent)}$$

Using oxidation numbers with equations:

- When assigning oxidation numbers to elements it is a good idea to set it out as below:

$Mg_{(s)} \quad + \quad 2HCl_{(aq)} \quad \rightarrow \quad MgCl_{2(aq)} \quad + \quad H_{2(g)}$

	$Mg_{(s)}$	+	$2HCl_{(aq)}$		$\rightarrow MgCl_{2(aq)}$		+	$H_{2(g)}$
Ox No's	0		+1	-1	+2	-1		0

$0 \rightarrow +2$ **$+1 \rightarrow 0$**   **(x2)**

Up 2   **Down 2**

Things to note:

- The **total increase** in oxidation numbers by one element is always **equal** to the **total decrease** in oxidation numbers by another element.
- Cl^- has played no part inn the reaction. It has not lost / gained electrons. It has not changed its oxidation number. These types of ions are called **spectator ions**.
- Reactions can be rewritten without these **spectator ions** to make an **ionic equation:**

$Mg_{(s)} \quad + \quad 2H^+_{(aq)} \quad \rightarrow \quad Mg^{2+}_{(aq)} \quad + \quad H_{2(g)}$

Qu 1-2 p33 / Qu 1, 11, 12, 13, 14, 18 P35 / Qu 1, 2a, 5a, 6a, 6b P36, 37