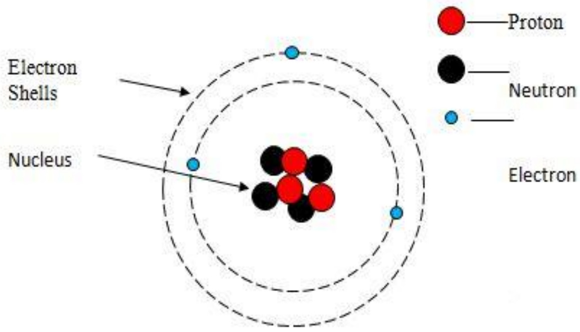
De Lisle College
Chemistry A-level
Transition work
2023-24

Please complete this work over the summer.

Hand it in during your first week studying A-level chemistry at De Lisle.

Name: …………………………………………………………………..

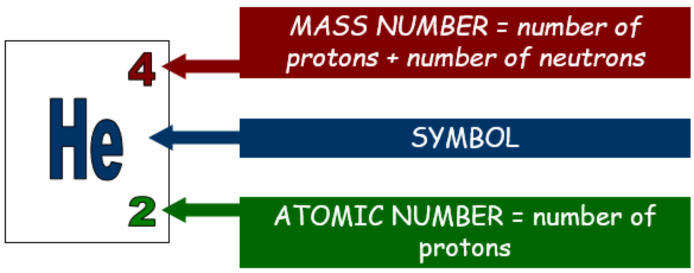
Atoms, ions and compounds



Atoms are made of smaller particles called 'subatomic particles.' These are the proton, neutron & electron. Protons and neutrons are found in the nucleus of an atom which has an overall positive charge. Electrons orbit the nucleus at set distances in energy levels (electron shells)

Subatomic particles have charges and masses. The actual charges and masses are very small so we use relative charges and masses instead. These are given below:

Subatomic particle	**Relative charge**	**Relative mass**
Proton	1+	1
Neutron	0	1
Electron	1-	Very small



Atoms of different elements contain different amounts of subatomic particles. The numbers for each atom can be deduced by looking at the atomic symbol of the element on the periodic table.
The atomic number tells us the number of protons in an atom of an element. It also tells us the number of electrons as number of protons = number of electrons in a neutral atom. The atom is neutral overall as the positive charges from the protons is cancelled out by the negative charges of the electron.
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The number of neutrons can be calculated using the formula:

number of neutrons = mass number – atomic number.

Isotopes

The number of protons in the nucleus of an atom is what determines the element that atom is. An isotope is an atom of an element that has the same number of protons but a different number of neutrons. As the nuclei of isotopes have different numbers of neutrons they will also have different mass numbers. The mass of an isotope is known as the isotopic mass. The average of these masses that takes account of the abundance of the isotopes is known as the relative atomic mass of an element.

Representing isotopes

Isotopes can be represented in a way similar to what we have seen with atoms previously. We can calculate the different numbers of neutrons in isotopes using the formula:

number of neutrons = mass number – atomic number

Besides using element symbols we can also represent isotopes like this: ${}^{12}_{6}C$, ${}^{13}_{6}C$.

C^{12}_{6}	C^{13}_{6}
P: 6 **N:6** **E :6**	**P:** **N:** **E:**

Abundantly clear

The abundance of an isotope is the percentage of the element that is that isotope. For example carbon-12 has an abundance of 98.9% whereas carbon-13 has an abundance of 1.1%. If we know the abundances of all the isotopes of an element then we can calculate the relative atomic mass of the element.

$$\text{Relative atomic mass} = \frac{(isotopic\ mass\ 1\ x\ \%\ abundance)+(isotopic\ mass\ 2\ x\ \%\ abundance)\ldots}{100}$$

For example the relative atomic mass of carbon is worked out like this:

$$\text{Relative atomic mass} = \frac{(12\ x\ 98.9)+(13\ x\ 1.1)}{100}$$

$$= 12.011$$

Electrons

The electrons in an atom are arranged in energy levels. These energy levels are set orbits around the nucleus of an atom. Electrons fill the lowest energy level first and when that is full they begin to fill the second energy level and so on. The energy level closest to the nucleus is energy level 1.

Drawing ions

Ions form when atoms of elements gain or lose electrons. An atom will gain or lose electrons so that its outermost shell is full. This is energetically favourable as having a full outer shell makes an atom more stable. If you can draw the electronic structure of an atom you can draw the electronic structure of an ion and deduce its charge.
To form an ion only the number of electrons changes. Gaining electrons will make an ion negatively charged whereas losing electrons will make an ion positively charged.

Checkpoint 1

Use your periodic table to:

a. Find the relative atomic mass of atoms of each of the following elements.
b. Find the number of protons in atoms of each of the following elements.
c. Now deduce the number of electrons in an atom of each of the elements given.
d. Use ***number of neutrons = mass number – atomic number*** to calculate the neutrons.

Element	Relative atomic mass (Mass number)	No. protons (atomic number)	No. electrons	No. neutrons
H				
C				
Li				
Mn				

For each of the elements below calculate the relative atomic mass from the abundances of the isotopes:

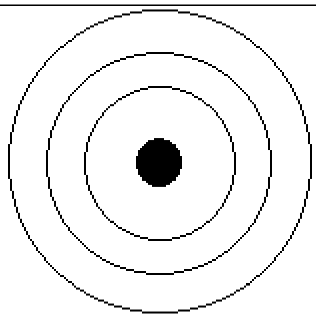
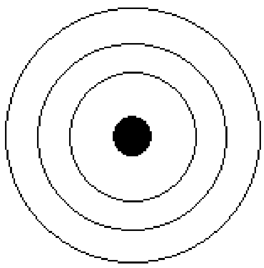
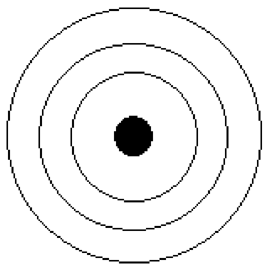
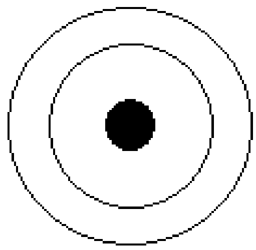
Oxygen-16, 99.76% Oxygen-17, 0.0373% Oxygen-18, 0.20%
Hydrogen-1, 99.972% Hydrogen-2, 0.038%
$^{63}_{29}Cu - 69\%$ $^{65}_{29}Cu - 31\%$
$^{79}_{35}Br - 50.7\%$ $^{81}_{35}Br - 49.3\%$

Ions

Draw the electron configurations for atoms of:

Hydrogen:	Helium:
Beryllium	Lithium
Oxygen	Fluorine

Complete the table to represent ions of the elements provided. Don't forget brackets and charges!

				
Name of element	*Sodium*	Magnesium	Chlorine	oxygen
Number of electrons in an atom	*11*	*12*	*17*	*8*
Number of electrons in an ion				

Quantitative chemistry

Conservation of mass

In a chemical reaction that goes to completion the total mass of products must be equal to the total mass of the reactants.

This is explained by the law of conservation of mass which states:

''***No atoms are lost or gained during a chemical reaction.***''

Sometimes it can appear that the mass involved in a reaction has changed. This is because a gas is involved either as a reactant or a product. If one of the reactants is a gas, then it can seem that the products have gained mass. If a gas is a product it can seem that mass has been lost.

As conservation of mass tells us that no atoms are gained or lost during a reaction we can represent reactions using balanced symbol equations. In a balanced symbol equation, the number of atoms of each element present are the same in both the reactants and the products.

Look at the example below:

$$2Mg + O_2 \rightarrow 2MgO$$

In this balanced symbol equation, a two has been written in front of the magnesium, this means that there are two atoms of magnesium present, the small two after the O tells us that a molecule of oxygen contains two oxygen atoms. A two has also been written in front of the magnesium oxide, MgO, in this case the two doubles everything that follows it. This would mean that there are two magnesium atoms and two oxygen atoms in the products.

What is relative formula mass (M_r)?

Relative formula mass has the symbol M_r. The relative formula mass of a substance is the sum of the relative atomic masses of the atoms that make it up. Look at the example below to see how to work it out:

C_2H_6 contains two carbon atoms and six hydrogen atoms. To calculate its relative formula mass you need to know the relative atomic mass of carbon, which is 12, and hydrogen, which is 1.

M_r C_2H_6 = (2 x 12) + (6 x 1)

= 24 + 6

= 30

What is the mole?

In chemistry we measure amounts in moles, a mole of any substance contains an Avogadro's number of particles; be they atoms, molecules or ions. An Avogadro's number is 6.02×10^{23} per mole. The mass of a mole is a substance's relative formula mass expressed in grams.

- **A mole contains 6.02×10^{23} atoms (or molecules or ions)**

- **A mole is the formula mass in grams**

What does this mean practically?

If we consider methane gas, CH_4, we could calculate its relative formula mass from the sum of the relative atomic masses of its atoms as 12 + (4 x 1) = 16. As the mass of a mole of a substance is the substance's relative formula mass in grams a mole of methane will have a mass of 16g. We also know that a mole of methane contains 6.02×10^{23} molecules of methane. So a mole of methane contains an Avogadro's number of methane molecules and has a mass of 16g.

We can calculate the amount of moles of a substance using the following equation:

$$n = m/M_r$$

Where n = moles in mol, m = mass in grams & M_r = relative formula mass

If we know the amount of a substance in moles and the substance's relative formula mass we can also calculate the mass of the substance. To do this we need to rearrange the equation above to make m the subject.

$$m = n \times M_r$$

Where n = moles in mol, m = mass in grams & M_r = relative formula mass

Checkpoint 2

Add numbers **in front of** reactants/products to balance the equations

1 __H_2O_2 → __H_2O + __ O_2

2 __Na + __Cl_2 → __ NaCl

3 __Mg + __HCl → __$MgCl_2$ + __H_2

4 __$CaCl_2$ + __Na_2SO_4 → __$CaSO_4$ + __NaCl

5 __K + __H_2O → __KOH + __H_2

6 __$MgCl_2$ + __ NaOH → __$Mg(OH)_2$ + __NaCl

7 __$CaCl_2$ + __ KOH → __$Ca(OH)_2$ + __KCl

Calculate the relative formula masses of:

HBr		H_2O		$Ca(OH)_2$	
NaCl		CO_2		H_2SO_4	
KI		PF_3		$(NH_4)_2SO_4$	
LiF		SO_2		HNO_3	
CO		NH_3		$CaCO_3$	
NaBr		SF_6		Ba_2F_6	
CsI		NaClO		MgI_2	

Calculate the amount, in moles, of each substance for the stated mass:

6g of H_2 Mr =

Moles =
72g of C Mr = Moles =
36g of H_2O Mr = Moles =
9g of H_2O Mr = Moles =
450g of $CaCO_3$ Mr = Moles =

Calculate the mass, in g, of the following amounts of substance:

4 moles of carbon monoxide, CO
10 moles of ethane, C_2H_6
2 moles of hydrogen chloride, HC*l*

A sample of water contains 1.505×10^{24} molecules.
Calculate how many moles of water this is:

Calculate the mass of the sample of water:

Assuming water has a density of $1 g/cm^3$ what volume of water is the sample composed of?

How many molecules of sulfur dioxide, SO_2, are in 96g of sulfur dioxide gas? Take Avogadro's constant to be 6.02×10^{23}.

Key
atomic number **Symbol** name relative atomic mass

(1) 1	(2) 2	3	4	5	6	7	8	9	10	11	12	(3) 13	(4) 14	(5) 15	(6) 16	(7) 17	(0) 18
1 **H** hydrogen 1.0																	2 **He** helium 4.0
3 **Li** lithium 6.9	4 **Be** beryllium 9.0											5 **B** boron 10.8	6 **C** carbon 12.0	7 **N** nitrogen 14.0	8 **O** oxygen 16.0	9 **F** fluorine 19.0	10 **Ne** neon 20.2
11 **Na** sodium 23.0	12 **Mg** magnesium 24.3											13 **Al** aluminium 27.0	14 **Si** silicon 28.1	15 **P** phosphorus 31.0	16 **S** sulfur 32.1	17 **Cl** chlorine 35.5	18 **Ar** argon 39.9
19 **K** potassium 39.1	20 **Ca** calcium 40.1	21 **Sc** scandium 45.0	22 **Ti** titanium 47.9	23 **V** vanadium 50.9	24 **Cr** chromium 52.0	25 **Mn** manganese 54.9	26 **Fe** iron 55.8	27 **Co** cobalt 58.9	28 **Ni** nickel 58.7	29 **Cu** copper 63.5	30 **Zn** zinc 65.4	31 **Ga** gallium 69.7	32 **Ge** germanium 72.6	33 **As** arsenic 74.9	34 **Se** selenium 79.0	35 **Br** bromine 79.9	36 **Kr** krypton 83.8
37 **Rb** rubidium 85.5	38 **Sr** strontium 87.6	39 **Y** yttrium 88.9	40 **Zr** zirconium 91.2	41 **Nb** niobium 92.9	42 **Mo** molybdenum 95.9	43 **Tc** technetium	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** palladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
55 **Cs** caesium 132.9	56 **Ba** barium 137.3	57–71 lanthanoids	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.8	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **Tl** thallium 204.4	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
87 **Fr** francium	88 **Ra** radium	89–103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium		114 **Fl** flerovium		116 **Lv** livermorium		

57 **La** lanthanum 138.9	58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium 144.9	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.2	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0
89 **Ac** actinium	90 **Th** thorium 232.0	91 **Pa** protactinium	92 **U** uranium 238.1	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium