Preparing to begin A-level Chemistry

at

Hereford Sixth Form College

Introduction

These pages aim to make sure that your knowledge and understanding from GCSE have put you in a good place to begin the A-level work next year. They focus on topics from GCSE and key skills (especially Maths skills) which will continue to be really important as you move on to a broader and deeper understanding at A-level.

The work is in 2 sections:

Section A focuses on maths skills and includes questions in boxes like this one. **The answers are at the end of the document so you can self-mark**. The questions at the end need to be answered **on paper and ready to hand in when you come along to your first Chemistry lesson**.

Section B covers a number of topics from Chemistry which should (mainly, at least) be familiar from your GCSE work. This section also includes lots of questions and the **answers are at the end of the document so you can self-mark**. **You won't be handing any of Section B in to us.**

It's REALLY important to do the self-marking questions and to mark them so that you are up to speed once term begins. We won't be spending much time practising these key skills and there is a risk of being left behind if you don't at least attempt all of them.

Finally…

Don't worry too much if you can't do the odd question or if there's the occasional bit of content you've actually never heard of! We will recap some things that may not have appeared in your GCSE during the first year of the course.

Contents

Section A: *Maths Skills for Chemistry*

Chemistry at A-level makes use of many maths skills. The examples we will see below are not Chemistry related but they do show the kind of manipulations you will have to be able to carry out. When you've learned the relevant chemistry, you'll be applying the skills to new chemical scenarios.

If you don't yet have these skills, please be aware that it's going to require a lot of extra work to develop them. We will be assuming you can do these things.

At the end of this section, there are a number of questions for you to have a go at. For this section only, you should **answer on paper and bring your answers in to hand to your teacher when you have your first lesson**.

It is very important to **show your working fully** in answering these questions.

A1 Rearranging Equations

This is a very important thing to be able to do, as it comes up a lot in chemical problems. The important thing is to remember that **you can do whatever you like to one side of an equation, as long as you do the same to the other side**. This allows both sides to remain equal and it's still an *equa*tion.

A1.1 Adding and subtracting

If we're **adding** two quantities, we can **subtract** one of them. It vanishes from one side and appears on the other side. Take the following:

$$A = B + C$$

Let's subtract C from both sides:

$$A - C = B + C - C$$

Which gives us the following, because the "$+ C - C$" part vanishes:

$$A - C = B$$

We have made B the subject of our equation and could now calculate B.

Instead, if there's a **subtraction** in the original equation, we do the opposite and **add** to both sides.

A1.2 Multiplying and dividing

If there's a **division** we can **multiply** both sides. Let's consider the very familiar equation for *velocity*:

$$velocity = \frac{distance}{time}$$

If we want to make *distance* the subject, we need to spot that *distance* is currently being **divided** by *time*. So, let's **multiply** both sides by time. This is what happens:

$$velocity\ x\ time = distance\ x\ \frac{time}{time}$$

But *time/time* is just equal to 1. So we end up with:

$$velocity\ x\ time = distance$$

Which is what we wanted to do.

What if we wanted to make *time* the subject?

Look at the last equation above. We are **multiplying** *time* by *velocity*. Let's **divide** both sides by *velocity* to get rid of it from the left-hand side. We end up with:

$$\frac{velocity}{velocity} x\ time = \frac{distance}{velocity}$$

But again, divide a quantity by itself gives you 1. So we end up with:

$$time = \frac{distance}{velocity}$$

Which is what we wanted to do.

A1.3 A simpler way…

Notice that quantities can "move diagonally" across the equals sign. So, something that is at the bottom (we're dividing by it) can move to the other side and end up at the top (so we're multiplying by it. Let's move *time* diagonally:

$$velocity = \frac{distance}{time}$$

This gives us:

$$velocity\ x\ time = distance$$

Alternatively, we can move two quantities at the same time:

$$velocity = \frac{distance}{time}$$

So *velocity* can move downwards and *time* can move upwards. This leads to:

$$time = \frac{distance}{velocity}$$

A1.4 Your turn.

Make C the subject of the following equations. Keep applying addition, subtraction, multiplication or division to each side until you get it. Or, move quantities diagonally if they're multiplying or dividing.

$i.\ A = B + C$ ....

$ii.\ A = B + C + D$ ....

$iii.\ A = BxC$ ....

$iv.\ A = \frac{B}{C}$ ....

$v.\ A = \frac{BD}{C}$ ....

$vi.\ A = BD + C$ ....

$vii.\ A = BC^2 + D$ ....

A2 Significant Figures

If my ruler only has 1 mm increments but I report a measurement as 2.34 mm, the ".34" part is meaningless because I can't actually see those numbers on the ruler. I've no idea if my measurement is 1.34 or in fact any other number between 1 and 2 mm. The ".34" part isn't **significant.**

Furthermore, if I take a measurement of 2 mm from my ruler and do a calculation with it, although the calculator may return a number to 7 decimal places, most of those are meaningless and therefore not significant. Remember I only put 1 figure into the length part of the calculation.

So, when we measure something in Chemistry – probably more likely to be something like mass or temperature – we can't ever measure it exactly because we'll be limited by the **precision** of the device we're using (the balance or the thermometer). So when we report these measurements and use them in calculations, we use the idea of **significant figures** to indicate how precisely we know something.

A2.1 An example: measuring percentage mass

For example, here are the results of an experiment to measure what % of the mass of a rock sample is chromium:

- Mass of rock (measured on a balance measuring grams to 2 decimal places) = 5.30g
- Mass of chromium in the rock (measured by a chemical experiment) = 1.07g

Notice, first, that we wrote 5.30g and not just 5.3g. This is to show that we know the mass to the nearest 0.01g. We think that the zero is probably pretty accurate. So although it might not be **exactly** 5.30 it's closer to 5.30 than to 5.29 or 5.31. **We know something about that last number** and therefore we have given the mass to 3 significant figures.

So we can calculate the % by mass of chromium

= (mass of chromium / total mass) x 100 = (1.07 / 5.30) x 100

= 20.1886792 % (says your calculator)

We can't really know the number this precisely because we are limited by precision of the numbers we fed in. So, **since the mass values we used were known to 3 significant figures**, we should round our answer to the same precision, to get:

% by mass of chromium = 20.2% (to 3 s.f.)

Notice that the unrounded number had a 1 just before the point where we rounded off, but it got rounded up to 2 because the number following was greater than or equal to 5.

Remember too that for numbers less than one, the first significant figure is the first non-zero digit.

e.g. if you need to round 0.007438993 to 2 significant figures, you get 0.0074 (because that 7 was the first significant figure).

A2.2 Your turn.

Write the following numbers to the specified number of significant figures.

i. 1.345 *to 3 significant figures* ……………………………………………………………………………………….

ii. 3.14159265 *to 5 significant figures* ……………………………………………………………………………………….

iii. 3.026×10^6 to *3 significant figures* ……………………………………………………………………………………….

A2.3 Some rules about significant figures:

- Zeroes before the first non-zero number don't count. 0.0023 is 2 significant figures.
- Trailing zeroes count, otherwise we why write them? 1.2300 is **4** significant figures. We know something about those zeroes and it tells us that our instrument can measure to 4 decimal places. Otherwise, we'd write 1.23.
- 100 is **difficult**: do we know that all the zeroes are in fact zero? If so, this is **3** significant figures. If we are rounding the second zero, it's **2** significant figures because we know the first zero is zero. If we're rounding to the nearest hundred, it's **1** significant figure: we only know that the "1" is accurate. So to avoid confusion, we use standard form: $1.0 \text{x} 10^2$ tells us that we are working to two significant figures. $1 \text{x} 10^2$ says we are working to one.

A2.4 Your turn. How many significant figures are in the following?

$i.$ 2.3 ………………………………………………………………………………………

$ii.$ 15.890 ………………………………………………………………………………………

$iii.$ 0.001 ………………………………………………………………………………………

$iv.$ 0.20 ……………………………………………………………………………………

$v.$ $1.50x10^3$ ………………………………………………………………………………………

$vi.$ 2000 ………………………………………………………………………………………

A3 Standard form

A3.1 What is standard form?

It is very common in science to give numbers – especially very small or very big ones – in standard form:

e.g. 52171 can be written as 5.2171×10^4 and 0.0000337 is 3.37×10^{-5}

They all have an **index**: this is the number written as a superscript. This tells us how many tens are in the multiplication we do to form that number. For example, 10^2 tells us that we have carried out 10x10. 10^3 says 10x10x10 and so on.

The number in front tells us how many "lots" of this we have. So, 1.2×10^2 tells us that we have carried out 10x10 to arrive at 100, then we have 1.2 "lots" of this, so (1.2 x 100) or 120.

When the index is negative, it tells us how many times ten appears in the **division** we do: so, 10^{-2} tells us that we have divided by 10 twice e.g. 1/(10x10) or 1/100. Again, the number in front tells us how many lots of this we have. So $12.x10x^2$ is the same as 1/100 or 0.01 then we have 1.2 **lots** of this, so (1.2x0.01) or 0.012.

A3.2 Your turn.

Write the following standard form numbers out in full.

$i.\ 1x10^4$ ……………………………………………………………………………

$ii.\ 1.3x10^6$ ……………………………………………………………………………

$iii.\ 2\,x10^{-2}$ ……………………………………………………………………………

$iv.\ 5.6x10^{-4}$ ……………………………………………………………………………

A3.3 Making it even simpler

For a big number, count how many places to the left the decimal point has to move to the left, to leave just one digit in front of it. If it has to move x places, the number will be … $x\ 10^x$

e.g. for 52171, the point moves 4 places left to leave just 5 in front, so we get 5.2171×10^4

For a small number, count how many places to the right the decimal point must move to the right, to leave just one non-zero digit in front of it. If it has to move x places, the number will be … $x\ 10^{-x}$.

e.g. for 0.0000337, the point moves 5 places right to leave just 3 in front,
so we get 3.37×10^{-5}

A3.4 Your turn

Write the following in standard form. **If it's not clear, assume 2 significant figures.**

$i.$ 100,000 ……………………………………………………………………

$ii.$ 130 ……………………………………………………………………

$iii.$ 150,000 ……………………………………………………………………

$iv.$ 0.001 ……………………………………………………………………

$v.$ 0.000067 ……………………………………………………………………

A4 Converting units

A4.1 Multiply or divide?

To save writing out lots of zeroes we sometimes **change the unit**. Rather than measure the distance from town to town using a metre stick and thereby having to write lots of numbers, it's better to effectively use a longer stick with length 1 km. 150 km is more convenient than writing 150,000 m. It gets even more convenient when we're measuring the distance to the nearest galaxy or the width of a molecule of water!

Let's say we need to convert a number in one unit, say A to a number in another unit, say B. We need to consider two things:

- How many of A are in 1 lot of B? e.g. how many g in a kg?
- In doing this conversion are we making the number **larger** or **smaller**? Do we therefore need to multiply or divide? E.g. in going from 3 km to the same number in metres, are we expecting the number to get larger by a factor of 1000 or smaller by a factor of 1000? Is 3 km a lot of metres or not many? Do we **multiply** 3 km by 1000 or **divide** 3 km by 1000?

A4.2 Your turn. Convert the following numbers. Think about whether you are making the number larger or smaller.

$i.$ 5 km to m ……………………………………………………………………

$ii.$ 0.3 m to km ……………………………………………………………………

$iii.$ 15,020 g to kg (4 s.f.) ……………………………………………………………………

$iv.$ $1.6x10^5$ kg to g ……………………………………………………………………

$v.$ 1320 mg to g (4 s.f.) ……………………………………………………………………

(*1 mg is 1* ***milli****gram*)

A4.3 Converting volume units.

How many cm^2 are in 1 m^2? It's not 100!

Remember that we calculate area by multiplying two lengths together. So an area of 1m2 is calculated by multiplying: 1 m x 1 m. This is actually 100 cm x 100 cm so in cm^2, the area is 100 cm x 100 cm = 10,000 cm^2. The conversion factor is 10,000.

So, convert the lengths before working out the area.

Now, volumes are calculated by multiplying three lengths. 1 m^3 is calculated by multiplying: 1 m x 1m x 1 m. This is 100 cm x 100 cm x 100 cm = 1,000,000 cm^3. 1 m^3 is a **lot** of cm^3! The conversion factor is 1,000,000.

So, convert the lengths before working out the volume.

A4.4 Your turn. Convert the following volumes. Bear in mind what you learned earlier: when you convert, are you making the number bigger or smaller? Are you multiplying or dividing?

$i.\ 3\ m^3\ to\ cm^3$ ....

$ii.\ 30{,}000\ cm^3\ to\ m^3$ ....

$iii.\ 1.6x10^5\ cm^3\ to\ m^3$ ....

$iv.\ 1.6x10^5\ kg\ to\ g$ ....

A5 Questions for handing in

Now here are the questions for handing in: they aren't specifically related to Chemistry but use the key skills that you'll need to deal with A-level chemical calculations. **Be careful with units!**

Please answer on paper and have them ready to hand to your teacher in your first lesson:

1.

$$velocity = \frac{distance}{time}$$

Use this relationship to find how far a cormorant can travel in 1200 seconds, if it is flying at 5 ms^{-1}.

....
....
....
....

2. Use the same relationship to find out how long it takes a car travelling at 50 ms^{-1} to travel 50 km.

....
....
....
....

3. If :

$$\frac{AB^2}{C} = k$$

write an expression for B.

....
....
....

4. If a Labrador grows at a steady rate, we can write an expression for its mass:

$$mass\ (m) = birth\ mass\ (m_0) + growth\ rate\ (R) x\ time\ (t)$$

where *mass* is in kg, *growth rate* is in kg per day (kg day^{-1}) and *time* is in days.

If a labrador has a mass of 17.6 kg after 180 days and its growth rate is 0.0950 kg day^{-1}, what was its birth mass, m_0?

....
....
....
....

5. What is 3034560 to 4 significant figures?

....
....
....

6. Put the following numbers into standard form:

i. 356,000

....

ii. 0.00714

....

7. Express the following as conventional numbers (write out in full):

i. 1.04×10^5

....

ii. 2.22×10^{-4}

....

iii. What is 247 g in kg?

………
………
………

iv. What is 4.45 g per cm^3, in mg per cm^3?

………
………
………

v. Calculate:

$28 \div (-4)$ ……………………………………………………………………………

$(-50) \times (-6)$ ……………………………………………………………………………

$(-10) - ((-20) / 10)$ ……………………………………………………………………………

vi. $A = B - CD$. Find the value of C when $A = -10$, $B = -80$, $D = +2$

………
………
………

vii. A percentage question! If I have eaten 30% of my Stinking Bishop cheese, and I now have 140g left, what mass of Stinking Bishop did I have to start with? Think about how to turn 30 % into 100 %: what are we multiplying by?

………
………
………

A6 Answers to self-marked Maths skills questions

A1.4 Your turn.

Make C the subject of the following equations. Keep applying addition, subtraction, multiplication or division to each side until you get it. Or, move quantities diagonally if they're multiplying or dividing.

$i.\ A = B + C$	$C = A - B$	$ii.\ A = B + C + D$	$C = A - B - D$
$iii.\ A = BxC$	$C = \frac{A}{B}$	$iv.\ A = \frac{B}{C}$	$C = \frac{B}{A}$
$v.\ A = \frac{BD}{C}$	$C = \frac{BD}{A}$	$vi.\ A = BD + C$	$C = A - BD$
$vii.\ A = BC^2 + D$	$C = \sqrt{\frac{A-D}{B}}$		

A2.2 Your turn.

Write the following numbers to the specified amount of significant figures.

1.345 *to 3 significant figures*	*1.35*
3.14159265 *to 5 significant figures*	*3.1416*
3.026 x 10^6 to *3 significant figures*	*3.03x10^6*

A2.4 Your turn. How many significant figures are in the following?

$i.\ 2.3$	2	$ii.\ 15.890$	5
$iii.\ 0.001$	1	$iv.\ 0.20$	2
$v.\ 1.50x10^3$	3	$vi.\ 2000$	4,3,2,or 1

A3.2 Your turn.

Write the following standard form numbers out in full.

$i.\ 1x10^4$	10,000	$ii.\ 1.3x10^6$	1,300,000
$iii.\ 2\,x10^{-2}$	0.02	$iv.\ 5.6x10^{-4}$	0.00056

A3.4 Your turn

Write the following in standard form. **If it's not clear, assume 2 significant figures.**

$i.\ 100{,}000$	$1x10^5$	$ii.\ 130$	$1.3x10^2$
$iii.\ 150{,}000$	$1.5x10^5$	$iv.\ 0.001$	$1x10^{-3}$
$v.\ 0.000067$	$6.7x10^{-5}$		

A4.2 Your turn. Convert the following numbers. Think about whether you are making the number larger or smaller.

$i.\ 5\ km\ to\ m$	5000 m	$ii.\ 0.3\ m\ to\ km$	0.0003 or $3x10^{-4}$ km
$iii.\ 15{,}020\ g\ to\ kg\ (4\ s.f)$	15.02 kg	$iv.\ 1.6x10^5\ kg\ to\ g$	$1.6x10^8$ g
$v.\ 1320\ mg\ to\ g$ (4 s.f.)	1.320 g		

(*1 mg is 1 **milli**gram*)

A4.4 Your turn. Convert the following volumes. Bear in mind what you learned earlier: when you convert, are you making the number bigger or smaller? Are you multiplying or dividing?

$i.\ 3\ m^3\ to\ cm^3$	3,000,000 cm^3
$ii.\ 30{,}000\ cm^3\ to\ m^3$	0.03 m^3
$iii.\ 1.6x10^5\ cm^3\ to\ m^3$	0.16 m^3
$iv.\ 1.6x10^5\ kg\ to\ g$	$1.6x10^8\ g$

Section B: *Important Ideas for Chemistry*

- This section covers a number of important topics that you may already have learned something about at GCSE. Don't worry if you haven't covered any of these things: we will be going through them all in the early stages of your A-level course.

- Each topic has a few questions at the end for you to have a think about. These are NOT for handing in to your teacher. Instead, you can get worked answers to them from us, if you send an email to jared.lewis@hereford.ac.uk. These answers will let you check what you've done but also help you see what to do if you get stuck.

B1 The Periodic Table

B1.1 How is it organised?

The Periodic Table is a miracle of thought and a reflection of our progress in understanding the atom and trends in atomic structure. The good news is that you don't have to learn it by heart! You will have a data sheet in lessons, and in your exams, which includes the Table and important numbers for each element. The more you use the table, the quicker you'll become at finding particular elements, and seeing the relationships between them. The Table, as it appears on the A-level data sheet, is included on the next page.

Key features of the Periodic Table include:

- It is organised with atoms placed **in order of increasing atomic number**.

- Atoms of elements in the same **period** (a horizontal row) have at least one electron in the shell corresponding to their period number. For example, you'll find calcium (atomic number 20) in the fourth row down (period 4), so you know that a calcium atom has at least one electron in the fourth shell.

- Elements in the same **group** (vertical column) all have the same number of electrons in their outermost shell. This number is the same as the group number shown in brackets across the top of the table (eg nitrogen is in group 5 so it has 5 outer shell electrons). The consequence is that elements in the same group have similar physical and chemical properties. Their reactivity changes as you descend the group.

- At A level you will learn about the way electrons behave in atoms and this will give you an understanding of why the Periodic Table has the shape it does.

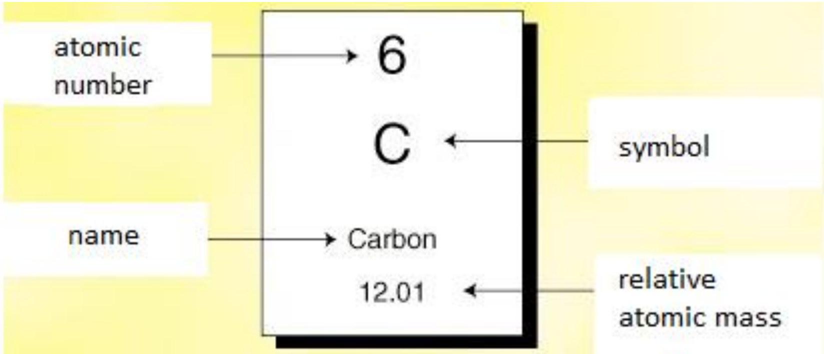
The Periodic Table of the Elements

Key
atomic number
Symbol
name
relative atomic mass

(1)	(2)											(3)	(4)	(5)	(6)	(7)	(0)
1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**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

Each element has its own square on the table which includes three key pieces of information:



The **atomic number** is the number of protons in the nucleus of an atom of that element. It defines the element. This number does not change in **chemical** reactions. Changing the nucleus is in the domain of physicists.

The **relative atomic mass** is the average mass of an atom of an element. More about this later.

B1.2 Your turn

1. You've probably never met any rubidium (atomic number 37) but see if you can predict what sort of properties it might have, from its position in the table.

i. How many electrons in the outer shell? ………………………………………………….

ii. How many shells does it have? …………………………………………………………

iii. Will it react by giving electrons or receiving them? (think about elements in the group)………………………….…………………………………………………….….…...

iv. Will it tend to form a **cation** or an **anion**? ……………………………………………….

v. Will the products of its reaction with water be acidic or alkaline? ………………………..

vi. Will it be **reactive** or relatively **inert**? …………………………………………………..

2. Which elements are these?

Group 14 (old group 4), Period 3 …………………………………………………………………..

2 shells, 7 electrons in the outer shell…………………………………………………………….

B2 Atoms and Ions

B2.1 What's inside an atom or ion?

Atoms are of **protons** and **neutrons** (which are packed together in the nucleus) and **electrons**, which occupy most of the space in the atom, outside the nucleus. You need to know the basic properties of these particles:

PARTICLE	RELATIVE MASS	CHARGE
Proton	1	+1
Neutron	1	0
Electron	approx. 1/2000	-1

A few key points:

- The number of electrons in a neutral atom is equal to the number of protons. Atoms have no charge because there are equal numbers of positive and negative charges, and protons and electrons have equal and opposite charges.

- The sum of the number of protons and the number of neutrons in an atom is called the **mass number**. This is written top left of the symbol when we write down a specific isotope's symbol.

  e.g. an atom with 16 protons is a sulfur atom (atomic number 16). If this atom has 17 neutrons, its mass number is 33, so it can be written as ^{33}S.

- A simple subtraction: *sum of mass number – atomic number* will determine the number of neutrons.

- The chemical properties of atoms are dominated by the electrostatic forces between the positively charged protons and negatively charged electrons. So neutrons, which have no charge, have only a small effect on chemical properties. For this reason you can get atoms of the same element, with virtually the same chemical properties but different numbers of neutrons. These are called **isotopes**.

Ions are particles formed when atoms lose or gain electrons (note that the number of protons and neutrons never changes in a *chemical* process). Since electrons are negatively charged, we can see that:

- Positive ions (also called **cations**) are formed when electrons are lost from an atom.
  e.g. Mg^{2+} is a magnesium ion, formed when a Mg atom loses two electrons.
- Negative ions (also called **anions**) are formed when extra electrons are gained by an atom.
  e.g. Cl^{-} is a chloride ion, formed when a chlorine atom gains an electron. (Notice – and try to get in the habit of following this pattern – the change in the spelling when an atom becomes a negative ion).

B2.2 Your turn.

1. Complete the table.

	Protons	Neutrons	Electrons	Symbol
i	1	1	1	
ii	26	30	26	
iii	92	143		$^{235}_{92}U$
iv				^{15}N

2. Write the name and the symbol for what you get when:

i. An oxygen atom gains two electrons. ……………………………………………………..

ii. An aluminium atom loses three electrons ……………………………………………

B3 Chemical Calculations

B3.1 Relative atomic mass

The **mass number** refers to a specific isotope. However – with rare exceptions – we are usually working in Chemistry not with individual atoms but with enormous numbers of atoms (e.g. a mugful of neon gas contains around 10,000,000,000,000,000,000,000 atoms). These will generally include a **mixture of isotopes** (our mugful of neon is a mixture of ^{20}Ne, ^{21}Ne and ^{22}Ne atoms, for example). These all have the same atomic number (same number of protons) and the same number of electrons but **different numbers of neutrons**.

So in most practical situations in Chemistry, what we care about is the **average** mass of an atom of an element in the mixture. Fortunately, the relative amounts of each isotope are fairly constant everywhere on Earth (although the small differences that exist are interesting and sometimes useful to measure). This means that once we've measured the average atomic mass, it will be valid for pretty much any calculation.

The problem is that the mass number doesn't strictly tell us about the mass. It just counts particles in the nucleus. It's not a mass but a number. ^{12}C doesn't have a mass of 12 g or 12 kg. Also, atoms with different numbers of particles in the nucleus don't always have a mass that you'd expect if you just added the particles' masses up. The study of these unexpected changes is part of atomic physics.

So we need a way of measuring mass. Measuring it in kg or g would lead to a ridiculous number of zeroes and would get tedious to handle pretty quickly. So, instead of using the kg as our unit, we use a much smaller unit. We haven't chosen the mass of a hydrogen atom as "1" in our new scale (which might have been perhaps the most obvious move) but we've chosen ^{12}C. 1/12 of the mass of a ^{12}C atom is "1" in our **relative atomic mass scale**.

So we define **relative atomic mass**:

- **Relative atomic mass (A_r) is the weighted mean mass of an atom of an element relative to 1/12 of the mass of an atom of ^{12}C.**

Take carbon. It has two naturally-occurring isotopes in meaningful amounts: ^{12}C and ^{13}C. The **relative** masses are 12.000 and 13.003. In any given sample, the average mass will be between the two.

The idea of a "weighted mean" is that the more abundant it is on Earth the more the isotope contributes towards the average. For example, to calculate the A_r of carbon, which is a mixture of ^{12}C and the heavier ^{13}C, the following calculation:

$$A_r = \frac{12.000 + 13.003}{2} = 12.5$$

can't be correct because there is far more ^{12}C than ^{13}C. The average should be closer to 12.

Here's a procedure that will work for us. Imagine we have 100 Cl (chlorine) atoms. If 75% of the sample are ^{35}Cl isotope then we have 75 atoms of ^{35}Cl. The remainder, 25 atoms, are ^{37}Cl. We'd expect the average mass to be closer to 35 than 37.

- For each isotope take the **mass number** (which will be close enough to the relative mass for our purposes, so for example, ^{13}C can be counted as 13, ^{35}Cl as 35 and ^{37}Cl as 37) and multiply the mass number x the number of each isotope.
- Add these terms together
- Divide by the 100 (the number of all atoms).

For example: in a sample of chlorine, 75% of atoms are ^{35}Cl and 25% are ^{37}Cl.

The average mass is therefore:

Mass number   no of atoms in a sample of 100

$$A_{\mathrm{r}} = \frac{(35 \times 75) + (37 \times 75)}{100} = 35.5$$

B3.2 Your turn

1. Sulfur occurs as a mixture of 3 isotopes: ^{32}S: 94.9% ^{33}S: 0.80% ^{34}S : 4.3%.

Calculate the relative atomic mass of sulfur.

………
………
………

B3.3 The mole

The Mole is a quantity that's used in Chemistry to define an amount of a substance. The definition of a mole looks a bit strange:

- **One mole is the amount of substance containing the Avogadro number of atoms, molecules or groups of ions.**
- **Avogadro's number is = 6.02×10^{23} mol^{-1}** (ie 602 000 000 000 000 000 000 000)

What we're doing is a bit like counting boxes of bananas instead of individual bananas. Rather than state that we have 15,500 bananas it would be better to say that we have 310 boxes of bananas, each with 50 bananas. As long as everyone knows how many are in a box, then it's a more convenient way of counting them. We all know what 153 boxes of bananas means. It's just that the Chemistry "box" is defined as quite a lot of "bananas" (atoms)! Otherwise, we'd be dealing with huge numbers all the time.

The oddly-defined "box" is actually a convenient size because we're saying that 12 g of ^{12}C is 1 box. Now because ^{24}Mg atoms are twice as heavy, 1 box would have a mass of 24 g. ^{48}Ti is four times a s heavy so its box would have a mass of 48 g. All the boxes have the same number of atoms. IT's N_A in each box.

We call the box a **mole**. One mole of any substance contains the same number (N_A) of particles. So we can make this incredibly important general statement:

One mole of atoms of any element has a mass (in grams) that is numerically equal to the relative atomic mass.

For example, Iron has a relative atomic mass of 55.8 (find it in the Periodic Table), so 1 mole of iron has a mass of 55.8g.

B3.4 Your turn

1 . What is the mass of 3 moles of fluorine atoms?

…………………………………………………………………………………………………….……
…………..……………………………………………………………………………………………

2. How many atoms are there?

……
……………………………………………………………………………………………………

3. A sample of Ne (neon) atoms has a mass of 4 g. How many moles is this?

……
………………………………………………………………………………………………………

4. How many atoms are in the sample of Neon?

……
……………………………………………………………………………………………………...

B3.5 Relative formula mass, M_r

Relative formula mass lets us work with substances with more than one atom in their formula by simply **adding up the individual relative atomic masses** to get the relative formula mass (M_r).

Oxygen exists as molecules made up of pairs of oxygen atoms, so its formula is O_2. The relative formula mass is therefore:

$$M_r = 2 \times A_r = 2 \times 16.0 = 32.0 \text{ for } O_2.$$

Sodium oxide has the formula Na_2O. So its M_r is given by:

$$2 \times A_r \text{ of sodium} + A_r \text{ of oxygen} \quad = (2 \times 23.0) + 16.0 \quad = \quad 62.0$$

So 1 mole of oxygen molecules has a mass of 32.0 g. This isn't the same as 1 mole of oxygen **atoms** – this would have a mass of 16.0, because **we're not counting the same thing**. 1 mole of N **atoms** has a mass of about 14.0. 1 mole of N_2 **molecules** has a mass of about 28 g though.

- **One mole of any substance has a mass (in grams) equal to the relative formula mass (M_r) of the substance.**

B3.6 Your turn

1. What is the relative formula mass of CH_4?

....
....

2. What would be the mass of 2 moles of CH_4 molecules?

....
....

3. A molecule containing only carbon and hydrogen has a relative formula mass of 44.0. It has 3 carbon atoms. How many H atoms are there?

....
....

B4 Structure and Bonding

There are basically 3 ways in which atoms can form strong bonds to other atoms:

- **Ionic bonding is** electrostatic attraction between oppositely charged ions. It most commonly occurs when a metal forms a compound with a non-metal. For example, NaF and $FeCl_3$ are ionic.

- **Covalent bonding** happens when two atoms share a pair of electrons. The bond is the attractive force between the shared pair and the nuclei of the two atoms. This type of bonding is common when non-metals react with one another. For example, H_2O has covalent bonding.

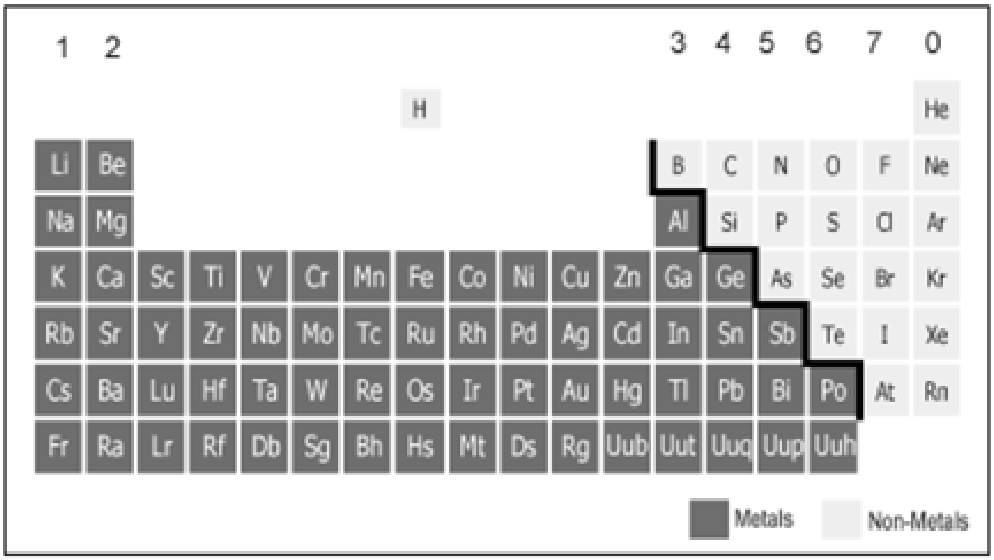
- **Metallic bonding** occurs between the atoms of metallic elements – it is the attractive force between a lattice of positive ions and a sea of delocalised electrons surrounding them.

B4.1 Ionic bonding

This most often occurs when a **metal forms a compound with a non-metal** (this is not **always** the case, but it's a good guideline for when bonding will be ionic. So the correct definition is:

- **Ionic bonding is** electrostatic attraction between oppositely charged ions.

You can predict whether a compound will be ionic if you know which elements are metals and which are not. There is a rough boundary in the Periodic Table that makes this possible and you'll need to know where this boundary is:



When metals in the main groups of the Periodic Table (it's more complex for the transition elements) react with non-metals, the metal atoms usually **lose** all of their outer shell electrons to form positive ions. As a result the size of the positive charge is the same as the group number. e.g. calcium is in group 2, so a Ca atom has 2 outer shell electrons and when it reacts it loses both, to form a Ca^{2+} ion.

Non-metals will tend to **gain** electrons to complete their outer shell. For example, Cl is in Group 17 ("Group 7"), has 7 electrons in its outer shell and gains one to become Cl^-. Other non-metals may gain more than one electron and therefore have a more negative charge.

B4.2 Your turn

1. Which ions are formed by the following elements?

i.	Na	ii.	O ....
iii	Mg	iv.	I ....
v.	Al	vi.	S ....

An ionic bond is the attraction between oppositely charged ions formed in this way. Ionic compounds are neutral overall, so **the total charge of the positive and negative ions must cancel each other out**. This enables you to work out the formula of an ionic compound

e.g. sodium forms Na^+ ions while oxygen forms O^{2-} ions. So to balance out the charges you need twice as many Na^+ as O^{2-} ions. The formula of sodium oxide is therefore Na_2O.

B4.3 Your turn

1. What will be the formula of the ionic compound formed by ions of the following elements?

ii.	Na and Cl	ii.	Mg and F ....
iii	Al and Br	iv.	Li and N ....
vi.	Al and O	vi.	Al and N................................

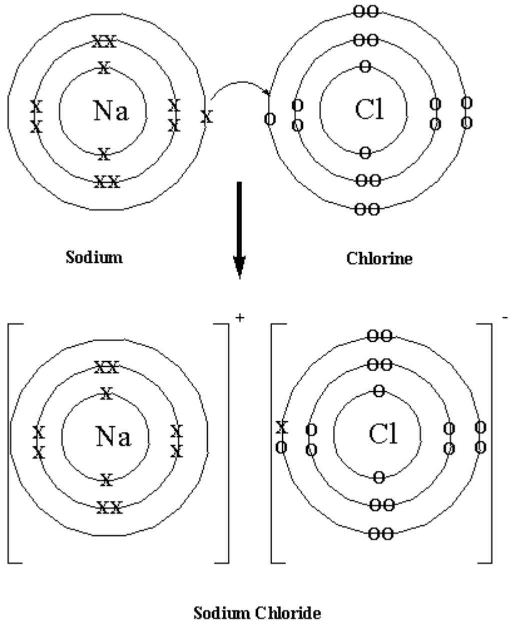
These websites may help you, if you're not yet comfortable predicting the formulae of ionic compounds:

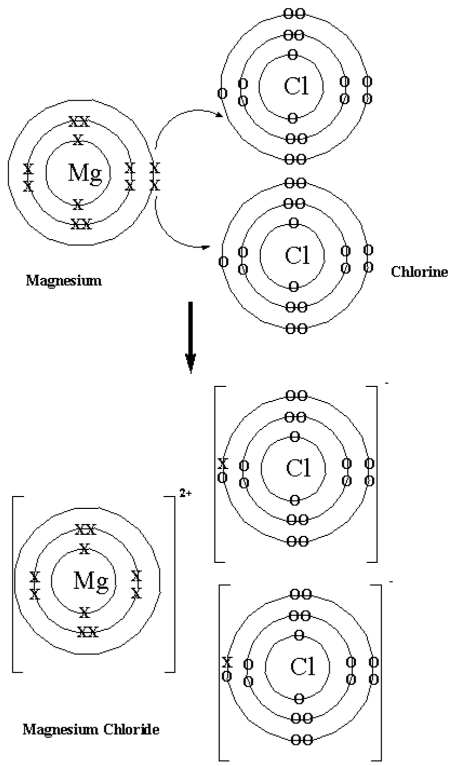
http://freesciencelessons.co.uk/charges-on-ions/

and

http://freesciencelessons.co.uk/formula-of-ionic-compounds/

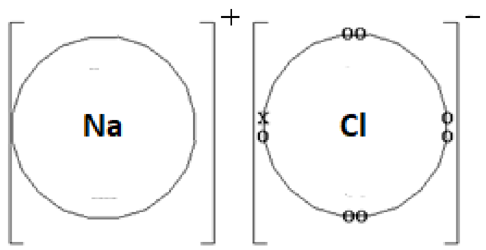
You should remember how to draw **dot and cross diagrams** from GCSE for ionic compounds. Here are some examples:





Notice the **square brackets** and the **charges** which show that these are no longer atoms: they're **ions**. At A-level we will still use these diagrams, but keep things simpler: we just show the electron arrangement once the bonds are formed (not what the atoms were like before this) and we just show **just the outer shell electrons** in the compound.

e.g. for sodium chloride:



Some ionic compounds contain **polyatomic ions**: these are ions that are made up of a group of atoms joined together by covalent bonds, and the whole thing carries a charge. There is a short list of common ones that you'll need to learn, so get ahead of the game by <u>learning them now</u>:

- **Polyatomic Ions to memorise:**

Hydroxide OH^- **Sulfate SO_4^{2-}** **Nitrate NO_3^-**

Carbonate CO_3^{2-} **Ammonium NH_4^+**

Note: you have to be careful when writing the formula of a compound if it contains **more than one polyatomic ion** *e.g*. calcium hydroxide is made up of Ca^{2+} and OH^-, and it obviously takes two OH^- for each Ca^{2+} to balance the charges. However, CaO_2H_2 would be an unclear way of representing it. Is the anion made of two O atoms and two H atoms all joined together? Are there two anions: one containing two O atoms and one containing two H atoms? Clearly not. $CaOH_2$ is wrong because it suggests only one O atoms and two H atoms.

To show that there are two lots of the whole OH^- ion, you put a bracket around it in the formula:

$$Ca(OH)_2$$

B4.4 Your turn

Write down the formulae of the following compounds:

i. Magnesium sulfate	....
ii. Ammonium fluoride	....
iii. Strontium nitrate	....
iv. Ammonium carbonate	....
v. Aluminium sulfate	....

B5 Balancing Equations

At A-level you will need to be able to write balanced symbol equations for most of the reactions you come across, so practising balancing – and being able to do it quickly – is a skill that's worth developing at this point. It WILL become hugely important very quickly after you start the course.

B5.1 How to balance an equation

The first rule is that you must always make sure the individual substance formulae are correct before you start trying to do your balancing act. So:

- First decide if it's an ionic or covalent substance. Metal plus non-metal, or just non-metals?.
- If it's ionic, think about the ionic charge and then use the charge-balancing rule (see above) to get the formula.
- If it's covalent, there are various ways you can get the correct formula – but we won't get into that at the moment.

Now you can start putting your equation together:

1. Write down the correct chemical formula for each reactant and product
2. Check the equation to see if there are **the same number of atoms of each element in the reactants and products**. If there are not, it's not strictly an equation. You need to adjust it.
3. To set about balancing, start with an element that is only in one reactant and one product. Put a number IN FRONT of the particular reactant or product WITHOUT changing the formula, to make that element balance. Remember, these numbers in front are balancing

numbers – they are not part of the formula of the individual substance. Any number in front multiplies every atom in that compound.

4. Now do the same for a different element. Remember you can't change any balancing number you have already fixed. Continue to do this for each element, until both sides contain the same number of each type of atom.

e.g. the reaction of iron (III) oxide with magnesium:

$$\ldots\ldots Fe_2O_3 + \ldots\ldots Mg \rightarrow \ldots\ldots MgO + \ldots\ldots Fe$$

Let's start with Fe: at the moment, we have 2 on the left but 1 on the right. We can fix this by putting in a balancing number:

$$\underline{\mathbf{1}}\ Fe_2O_3 + \ldots\ldots Mg \rightarrow \ldots\ldots MgO + \underline{\mathbf{2}}\ Fe$$

Note we didn't write "Fe_2": this would be changing the formula. The underlined numbers are now fixed – we mustn't change them. So we can see that since we're stuck with **1** Fe_2O_3 , we'll have to have 3 oxygens on the right too. To make that happen, we need a balancing number for MgO:

$$\underline{\mathbf{1}}\ Fe_2O_3 + \ldots\ldots Mg \rightarrow \underline{\mathbf{3}}\ MgO + \underline{2}\ Fe$$

Now we're almost there. We just need to make Mg balance, and that's easily done:

$$\underline{\mathbf{1}}\ Fe_2O_3 + \underline{\mathbf{3}}\ Mg \rightarrow \underline{\mathbf{3}}\ MgO + \underline{\mathbf{2}}\ Fe$$

We don't generally write "1" in an equation – it's just implied if there is no other balancing number. So we would write our final equation as:

$$Fe_2O_3 + \underline{\mathbf{3}}\ Mg \rightarrow \underline{\mathbf{3}}\ MgO + \underline{\mathbf{2}}\ Fe$$

If you would like more help in learning to balance equations, try these websites:

http://freesciencelessons.co.uk/balancing-chemical-equations/ and

http://www.wikihow.com/Balance-Chemical-Equations

B5.2 Your turn

1. Balance the following by adding balancing numbers.

i.	C	+	H_2	→	CH_4		
ii.	Si	+	Cl_2	→	$SiCl_4$		
iii.	H_2	+	F_2	→	HF		
iv.	CH_4	+	O_2	→	CO_2	+	H_2O
v.	Fe	+	Cl_2	→	$FeCl_3$		
vi.	Al	+	O_2	→	Al_2O		

For the following reactions, you first have to write formulae for the reactant and products before balancing. Use what you learned earlier: remember to match the charges of the ions in an ionic compound's formula and remember the formulae of the polyatomic ions.

i. burning magnesium in oxygen to give magnesium oxide

………………………………………………………………………………………………

ii. sodium + chlorine → sodium chloride

(chlorine is a diatomic molecule)

………………………………………………………………………………………………

iii. sodium + water → sodium hydroxide + hydrogen

………………………………………………………………………………………………

iv. sodium carbonate + hydrochloric acid → sodium chloride + carbon dioxide + water

………………………………………………………………………………………………

v. sodium hydroxide + sulphuric acid → sodium sulphate + water

………………………………………………………………………………………………

vi. burning C_5H_{12} in oxygen to give carbon dioxide and water

………………………………………………………………………………………………

vii. reacting sodium carbonate solution with magnesium chloride solution to give a precipitate of magnesium carbonate and another compound left in solution.

………………………………………………………………………………………………

B6 Answers to self-marked Chemistry questions

B1.2 Your turn

1. You've probably never met any rubidium (atomic number 37) but see if you can predict what sort of properties it might have, from its position in the table.

   i. How many electrons in the outer shell? ....1 (group 1)...........................
   ii. How many shells does it have? ....5 (Period 5)..............
   iii. Will it react by giving electrons or receiving them? (think about elements in the group)

   ....Losing ....
   iv. Will it tend to form a **cation** or an **anion**? ....Cation................................
   v. Will the products of its reaction with water be acidic or alkaline? ......alkaline: think NaOH and LiOH
   vi. Will it be **reactive** or relatively **inert**? ....Low in group 1 so (very) reactive......

2. Which elements are these?

   i. Group 14 (old group 4), Period 3 ....Si................................
   ii. 2 shells, 7 electrons in the outer shell.................Fluorine (F)...(Period 2, group 17)...

B2.2 Your turn.

1. Complete the table.

	Protons	Neutrons	Electrons	Symbol
i	1	1	1	$^{2}_{1}H$
ii	26	30	26	$^{56}_{26}Fe$
iii	92	143	92	$^{235}_{92}U$
iv	7	8	7	^{15}N

2. Write the name and the symbol for what you get when:

   (i) An oxygen atom gains two electrons. ....O^{2-}................

   (ii) An aluminium atom loses three electrons ....Al^{3+}...............

B3.2 Your turn

1. Sulfur occurs as a mixture of 3 isotopes: ^{32}S: 94.9% ^{33}S: 0.80% ^{34}S : 4.3%.
   Calculate the relative atomic mass of sulfur.

$$\frac{(32x94.9) + (33x0.8) + (34x4.3)}{100} = 32.09$$

B3.4 Your turn

1. What is the mass of 3 moles of fluorine atoms?
   1 mole has a mass of 19.0 g so 3 moles has a mass of 3x19.0 g = 57.0 g
2. How many atoms are there?
   Each mole has N_A atoms, so 3 moles has 3xN_A or 3x6.022x10^{23} = 18.066x10^{23} (or 1.8066x10^{24})
3. A sample of Ne (neon) atoms has a mass of 4.0 g. How many moles is this?
   One mole has a mass of 20.2 g. So 4.0 g is 4/20.2 = 0.20 mole
4. How many atoms are in the sample of Neon?
   1 mole has N_A atoms so 0.20 mole has 0.20 x N_A = 1.2x10^{23}

B3.6 Your turn

1. What is the relative formula mass of CH_4?
   12.0 + 4(1.0) = 16.0
2. What would be the mass of 2 moles of CH_4 molecules?
   One mole would have a mass of 16.0 g, so 2 moles has a mass of 2x16.0 = 32.0 g
3. A molecule containing only carbon and hydrogen has a relative formula mass of 44.0. It has 3 carbon atoms. How many H atoms are there?
   3 C atoms have a combined mass of 3x12.0 = 36.0. This leaves a mass of 44.0-36.0 = 8.0 which is 8.0/1.0 = 8 hydrogens. 3 carbons and 8 hydrogens together have a formula C_3H_8.

B4.2 Your turn

1. Which ions are formed by the following elements?

i.	Na ....Na^{+}.....................		ii.	O ....O^{2-}.......................	
ii.	Mg ....Mg^{2+}..................		iii.	I ....I^{-}.........................	
iii.	Al ....Al^{3+}..............		vi.	S ....S^{2-}......................	

B4.3 Your turn

1. What will be the formula of the ionic compound formed by ions of the following elements?

i.	Na and Cl ....$NaCl$...........	ii.	Mg and F ....MgF_2...........
ii.	Al and Br ....$AlBr_3$..............	iv.	Li and N ....Li_3N............
iv.	Al and O ....Al_2O_3................	vi.	Al and N.............AlN.............

B4.4 Your turn

1. Write down the formulae of the following compounds:

i.	Magnesium sulfate	.....................$MgSO_4$.............................
ii.	Ammonium fluoride	.....................AlF_3................................
iii.	Strontium nitrate	.....................$Sr(NO_3)_2$..........................
iv.	Ammonium carbonate	.....................$(NH_4)_2CO_3$........................
v.	Aluminium sulfate	......................$Al_2(SO_4)_3$.......................

B5.2 Your turn

1. Balance the following by adding balancing numbers.

i.	C	+	$2H_2$	→	CH_4		
ii.	Si	+	$2Cl_2$	→	$SiCl_4$		
iii.	H_2	+	F_2	→	$2HF$		
iv.	CH_4	+	$2O_2$	→	CO_2	+	$2H_2O$
v.	$2Fe$	+	$3Cl_2$	→	$2FeCl_3$		
vi.	$4Al$	+	$3O_2$	→	$2Al_2O_3$		

2. For the following reactions, you first have to write formulae for the reactant and products before balancing. Use what you learned earlier: remember to match the charges of the ions in an ionic compound's formula and remember the formulae of the polyatomic ions.

i. burning magnesium in oxygen to give magnesium oxide (oxygen is a diatomic molecule)

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

ii. sodium + chlorine → sodium chloride (chlorine is a diatomic molecule)

..............$2Na + Cl_2 \rightarrow 2NaCl$..

iii. sodium + water → sodium hydroxide + hydrogen

...............$2Na + 2H_2O \rightarrow 2NaOH + H_2$...

iv. sodium carbonate + hydrochloric acid → sodium chloride + carbon dioxide + water

..............$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$..

v. sodium hydroxide + sulphuric acid → sodium sulphate + water

..............$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$..

vi. burning C_5H_{12} in oxygen to give carbon dioxide and water

..............$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$..

vii. reacting sodium carbonate solution with magnesium chloride solution to give a precipitate of magnesium carbonate and another compound left in solution.

..............$Na_2CO_3 + MgCl_2 \rightarrow MgCO_3 + 2NaCl$..