Chemistry AS90171 Describe chemical reactions

This achievement standard involves the description of chemical reactions, including the carrying out of calculations.

Achievement	Merit	Excellence
Describe chemical reactions.	Interpret information about chemical reactions.	Apply understanding of chemical reactions.

Precipitation reactions These are limited to: formation of chlorides of silver and lead; sulfates of calcium, barium and lead, hydroxides and carbonates of copper(II), iron(II), iron(III), zinc, aluminium, calcium, and magnesium ions.

When two solutions are mixed, two new substances may be formed. If one of these is insoluble, then a precipitate is formed.

nitrates NO_3^-	All **soluble**
chlorides Cl^-	All **soluble** except AgCl, $PbCl_2$
sulfates SO_4^{2-}	All **soluble** except $BaSO_4$, $PbSO_4$, $CaSO_4$
hydroxides OH^-	All **insoluble** except KOH, NaOH
carbonates CO_3^{2-}	All **insoluble** except K_2CO_3, Na_2CO_3

Identify any precipitates (use the solubility rules in Resource Booklet you will receive)
The states of substances (eg (aq) and (s)) will be indicated in the question format, but are not needed in your answers.

precipitates

$$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$$

$$2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

$$MgSO_4(aq) + 2NaCl(aq) \rightarrow MgCl_2(aq) + Na_2SO_4(aq)$$

- Ionic equations. $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$ can be written as an ionic equation. Sodium sulfate is soluble and is left out of the ionic equation.

$$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s).$$

Thermal decomposition reactions These are limited to: hydroxides, carbonates and hydrogen carbonates. Thermal decomposition is a chemical reaction where a single compound breaks up into two or more simpler compounds or elements when heated. Heat is required to break chemical bonds in the compound undergoing decomposition. The decomposition reaction is irreversible.

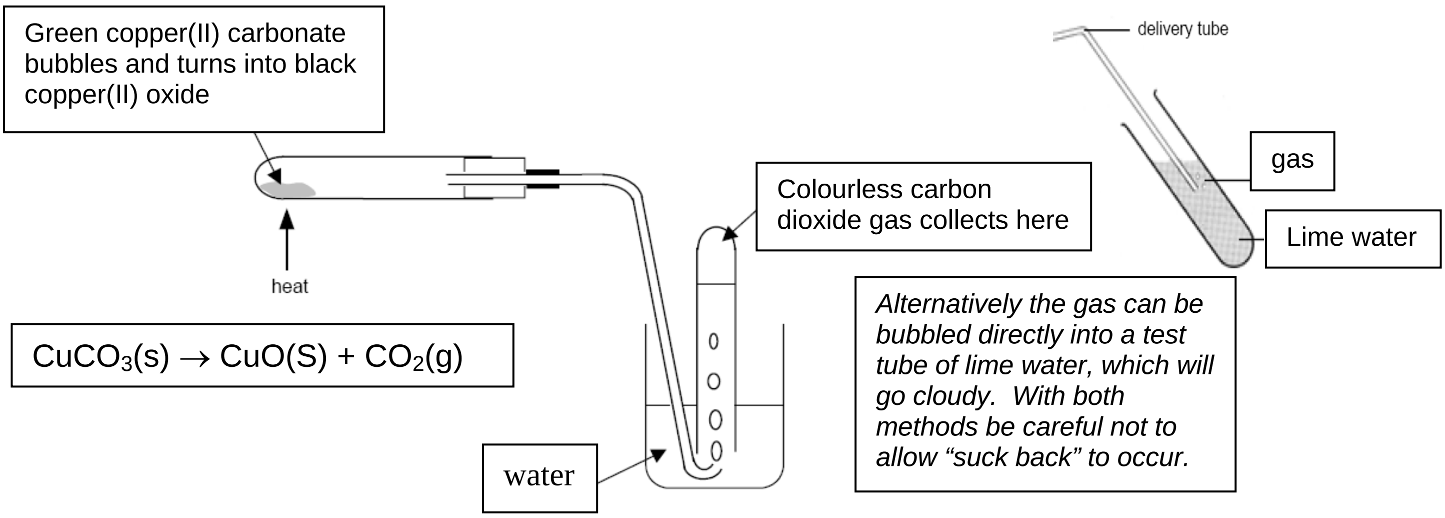
$Cu(OH)_2 \rightarrow CuO + H_2O$
blue — black

$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
white — white

$CaCO_3 \rightarrow CaO + CO_2$
white — white

$ZnCO_3 \rightarrow ZnO + CO_2$
white — white*

* yellow when hot



$CuCO_3(s) \rightarrow CuO(S) + CO_2(g)$

Alternatively the gas can be bubbled directly into a test tube of lime water, which will go cloudy. With both methods be careful not to allow "suck back" to occur.

metal hydroxide → metal oxide and water
metal carbonate → metal oxide + carbon dioxide
metal hydrogen carbonate → metal carbonate + water + carbon dioxide

Oxidation-reduction reactions   Redox = reduction and oxidation

Limited to: simple electron transfer involving elements and monatomic ions (such as Cl_2 /Cl^-, I_2/I^-, Fe^{3+}/Fe^2+, metal/metal ion), simple oxygen transfer (such as between metal oxides and either hydrogen or carbon).

Redox = reduction and oxidation

Simple Oxygen Transfer

Oxidation is the addition of oxygen / removal of hydrogen.

Examples

Copper oxide	+	hydrogen	→	copper	+	water
CuO	+	H_2	→	Cu	+	H_2O

Copper oxide is reduced to copper, and hydrogen is oxidised to water.

Hydrogen is called the reducing agent or reductant. Copper oxide is the oxidising agent or oxidant.

Lead oxide	+	carbon	→	lead	+	carbon dioxide
$2PbO$	+	C	→	$2Pb$	+	CO_2

Lead oxide is reduced to lead, and carbon is oxidised to carbon dioxide.

Carbon is called the reducing agent or reductant. Lead oxide is the oxidising agent or oxidant.

These definitions are not adequate to explain many reactions which chemists recognise as redox reactions.
Eg $2Mg + O_2 \rightarrow 2MgO$. Magnesium has been oxidised and oxygen has been reduced. Simple electron transfer definitions can be used instead.

Simple electron transfer

Loss of electrons = oxidation (Hint: metals always LOSE electrons)
Gain of electrons = reduction

<u>Metal/metal ion</u>

Mg	+	$CuSO_4$	$\rightarrow$	$MgSO_4$	+	Cu
silvery metal		blue solution		colourless solution		pinky orange metal

$Mg \rightarrow Mg^{2+} + 2e^-$ (this is loss of electrons, oxidation)
$Cu^{2+} + 2e^- \rightarrow Cu$ (this is gain of electrons, reduction)

We can combine these two HALF EQUATIONS to make an overall IONIC equation.
$Mg \rightarrow Mg^{2+} + \cancel{2e^-}$
$Cu^{2+} + \cancel{2e^-} \rightarrow Cu$ The $2e^-$ on either side cancel out.

$Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$ (this is balanced for atoms/ions and charge)

SO_4^{2-} ions (sulfate ions) are unchanged in the process – these are called spectator ions. That's why we don't put them in the ionic equation.

<u>**Balancing ionic equations**</u>*: number of atoms and the charge on each side must balance.*

Eg $Mg + Ag^+ \rightarrow Mg^{2+} + Ag$
$\mathbf{Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag}$ [balanced]

Try this easy one: $FeSO_4 + Zn \rightarrow ZnSO_4 + Fe$ [Is this balanced?]

First…… unravel the equation….$FeSO_4$ contains Fe^{2+} ions and SO_4^{2-} ions. You will be given a table of ions….. but make sure you know that the iron ions are 2+ here and not 3+ (Hint look at formula $FeSO_4$ and the charge on the sulfate ion, SO_4^{2-}).
So $Fe^{2+} + 2e^- \rightarrow Fe$

Zn is zinc atoms…. Reacts to become $ZnSO_4$ which contains zinc ions Zn^{2+} and sulfate ions, SO_4^{2-}
$Zn \rightarrow Zn^{2+} + 2e^-$

Combine the 2 HALF EQUATIONS and balance (atoms/ions and charge) if necessary
$\mathbf{Fe^{2+} + Zn \rightarrow Fe + Zn^{2+}}$ [doesn't need balancing]

The spectator ions here are SO_4^{2-}

Now the harder one
$Fe(NO_3)_3 + Mg \rightarrow Mg(NO_3)_2 + Fe$ (this is unbalanced!)

First…… unravel the equation…. $Fe(NO_3)_3$ contains Fe^{3+} ions and NO_3^- ions. You will be given a table of ions….. but make sure you know that the iron ions are 3+ here and not 2+ (Hint look at formula $Fe(NO_3)_3$ and the charge on the nitrate ion, NO_3^-).
So $Fe^{3+} + 3e^- \rightarrow Fe$

Mg is magnesium atoms……. Reacts to become $Mg(NO_3)_2$ which contains magnesium ions Mg^{2+} and nitrate ions NO_3^-
$Mg \rightarrow Mg^{2+} + 2e^-$

$Fe^{3+} + 3e^- \rightarrow Fe$ (x2)	$2Fe^{3+} + 6e^- \rightarrow 2Fe$	$2Fe^{3+} + \cancel{6e^-} \rightarrow 2Fe$
$Mg \rightarrow Mg^{2+} + 2e^-$ (x3)	$3Mg \rightarrow 3Mg^{2+} + 6e^-$	$3Mg \rightarrow 3Mg^{2+} + \cancel{6e^-}$

Combine the 2 HALF EQUATIONS and balance (atoms/ions and charge). The nitrate ions are spectator ions. $\mathbf{2Fe^{3+} + 3Mg \rightarrow 2Fe + 3Mg^{2+}}$

Chlorine/chloride

Chlorine solution: Pale yellow chlorine solution is added to colourless potassium iodide solution. The result is a browny yellow solution due to the formation of iodine, I_2. (KCl is colourless).

Cl_2	+	2KI	$\rightarrow$	2KCl	+	I_2
(pale yellow soln.)		(colorless soln.)		(colourless soln.)		(brown soln.)

$2I^- \rightarrow I_2 + 2e^-$ (this is oxidation) $Cl_2 + 2e^- \rightarrow 2Cl^-$ (this is reduction)

Overall equation is $I^- + Cl_2 \rightarrow I_2 + Cl^-$
Balanced this is $\mathbf{2I^- + Cl_2 \rightarrow I_2 + 2Cl^-}$ The K^+ ions are spectator ions.

And back to MgO – magnesium is oxidised as it burns and oxygen is reduced.

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

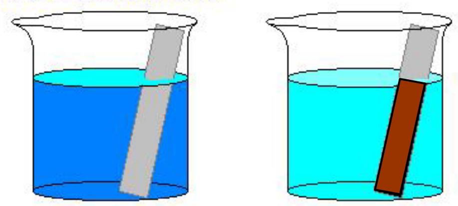
$Mg \rightarrow Mg^{2+} + 2e^-$ (this is loss of electrons, oxidation)
$O_2 + 4e^- \rightarrow 2O^{2-}$ (this is gain of electrons, reduction)

Balanced $2Mg + O_2 \rightarrow 2Mg^{2+} + 2O^{2-}$ (and $2Mg^{2+} + 2O^{2-} = 2MgO$)

Explaining some colour changes

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$
$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Some of the silvery grey zinc metal "dissolves" to form colourless zinc sulfate solution. The blue colour of the copper sulfate solution fades as the Cu^{2+} ion is removed by being reduced to copper metal which is seen as a dark reddish solid.



$$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$$
$$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$$

Copper metal dissolves and, as it forms Cu^{2+} ions, the solution takes on a blue colour (copper nitrate solution). A grey "furry deposit" or "feathery" silvery grey crystals of silver form on the copper metal.

Calculations: Relative atomic mass (A_r) & relative molecular (M_r) sometimes called formula mass.

5 **B** 10.8	6 **C** 12.0	7 **N** 14.0
13 **Al** 27.0	14 **Si** 28.1	15 **P** 31.0

Chemists use a relative atomic mass scale Eg A_r H = 1, A_r C = 12, A_r Si = 28.1, A_r B = 10.8, A_r O = 16 etc. Values will be provided on the supplied periodic Table.

$A_r(H) = 1$ $M_r(H_2O) = 1 + 1 + 16 = 18$ (no units – it's a relative scale).

We can calculate **molar masses**, **M** (the mass of one mole of the substance)

$M(C) = 12 \text{ g mol}^{-1}$
$M(Si) = 28.1 \text{ g mol}^{-1}$
$M(H_2O) = 1 + 1 + 16 = 18 \text{ g mol}^{-1}$
$M(CO_2) = 12 + (16 \times 2) = 44 \text{ g mol}^{-1}$
$M(Ca(NO_3)_2) = Ca + (2 \times N) + (6 \times O) = 164 \text{ g mol}^{-1}$
$M(CuSO_4.5H_2O = Cu + S + (4 \times O) + (5 \times H_2O)$ ……

The mass of one mole of substance is it's A_r or M_r in grams. 1 mole of C has a mass of 12g. 1 mole of CO_2 has a mass of 44g etc.

n = m/M where n = amount (in moles), m = mass (in g) and M = molar mass (in g mol^{-1})

Calculations from equations

- You will be given the balanced equation – write the number of moles of each substance underneath.
- Change moles to masses.
- Scale masses to those in the question.

29 **Cu** 63.6

Copper(II) oxide can be reduced by hydrogen: $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$
What mass of copper can be obtained from 15.9 g of copper(II) oxide?

CuO(s)	+	H_2(g)	→	**Cu**(s)	+	H_2O(g)
1 mole				I mol		
79.6 g				63.6 g		
15.9 g				(15.9/79.6) x 63.6 = 12.7g		

8 **O** 16.0

How much MgO can be made from 12g of Mg?

2**Mg**	+	O_2	→	2**MgO**
2 mol		I mol		2 mol
48.6 g				80.6
12 g				(12/48.6) x 80.6 = 19.9 g

12 **Mg** 24.3

Determine the formula of the compound made when 8.65 g of iron combines with 3.72 g of oxygen. The molar mass of the iron oxide is 159.8.

8	26
O	**Fe**
16.0	55.9

		Fe	O
mass		8.65 g	3.72 g
moles (n)	n = m/M	8.65/55.9 = 0.155 mol	3.72/16.0 = 0.2323 mol
divide by smallest		0.155/0.155 = 1	0.2323/0.155 = 1.5
simplest whole number ratio		2	3

So empirical formula = Fe_2O_3 $M(Fe_2O_3)$ = (55.9 x 2) + (16.0 x 3) = 159.8. Since this is the same as the molar mass of the iron oxide, then the formula of the compound is also $\mathbf{Fe_2O_3}$**.**

Calculate the empirical formula of a compound that is 43.7% phosphorus and 56.3% oxygen. Determine the molecular formula of the compound if its relative molecular mass is 284.

15	8
P	**O**
31.0	16.0

		P	O
mass	Assume 100g	43.7 g	56.3 g
moles (n)	n = m/M	43.7/31.0 = 1.41 mol	56.3/16.0 = 3.52 mol
divide by smallest		1.41/1.41 = 1	3.52/1.41 = 2.5
simplest whole number ratio		2	5

So empirical formula = P_2O_5 $M(P_2O_5)$ = (31.0 x 2) + (16.0 x 5) = 142. Since the molar mass of the compound is 284 then the formula of the compound is $\mathbf{P_4O_{10}}$ (since 284/142 = 2…………
So 2 x empirical formula = molecular formula)