BioMedical Admissions Test (BMAT)

Section 2: Chemistry

Topic C3: Chemical Reactions, Formulae and Equations

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Chemical reactions

- Remember that in a reaction, new arrangements of molecules (the products) are formed but no matter is created or destroyed.
- This means that the mass at the start will be the same as the mass at the end.
    - In an experiment it may seem as if mass has been lost, but this could be due to mass escaping in the form of gas.

Essential chemical formulae and symbols

There are some key formulae which are required knowledge for BMAT section 2.

Exam tip:
These need to be learnt by rote - it would be a good idea to make flashcards for this (either digitally or handwritten) and test yourself on them until they are in your memory.

Covalent Elements

Name	Formulae
Bromine	Br_2
Buckminsterfullerene	C_{60}
Chlorine	Cl_2
Fluorine	F_2
Hydrogen	H_2
Iodine	I_2
Nitrogen	N_2
Oxygen	O_2

Covalent Compounds

Name	Formulae
Ammonia	NH_3
Carbon dioxide	CO_2
Carbon monoxide	CO
Methane	CH_4
Nitrogen dioxide	NO_2
Nitrogen monoxide	NO
Sulfur dioxide	SO_2
Sulfur trioxide	SO_3
Water	H_2O

Positive ions

Name	Formulae
Lithium	Li^+
Sodium	Na^+
Potassium	K^+
Magnesium	Mg^{2+}
Calcium	Ca^{2+}
Barium	Ba^{2+}
Aluminium	Al^{3+}
Ammonium	NH_4^+
Copper II	Cu^{2+}
Hydrogen	H^+

Iron II	Fe^{2+}
Iron III	Fe^{3+}
Silver	Ag^{+}
Zinc	Zn^{2+}

Negative ions

Name	**Formulae**
Oxide	O^{2-}
Sulfide	S^{2-}
Fluoride	F^{-}
Chloride	Cl^{-}
Bromide	Br^{-}
Iodide	I^{-}
Carbonate	CO_3^{2-}
Hydroxide	OH^{-}
Nitrate	NO_3^{-}
Sulfate	SO_4^{2-}

Acids

Name	**Formulae**
Hydrochloric acid	HCl
Nitric Acid	HNO_3
Sulfuric acid	H_2SO_4
Ethanoic acid	CH_3COOH

Chemical equations

In a **balanced equation**, each element appears the same number of times on the right hand side as on the left hand side.

For example when Barium reacts with Oxygen to form barium oxide:

$$2Ba_{(s)} + O_{2(g)} \rightarrow 2BaO_{(s)}$$

- ➔ There are two barium on the left hand side and two on the right hand side.
- ➔ There are two oxygen on each side.

We can also see that the **state** is given for each element/compound.

For BMAT, you need to know:

State	Symbol
Solid	s
Liquid	l
Aqueous	aq
Gas	g

Some ions are always found in certain states when they are in a compound in the aqueous medium.

Soluble ions

Ion	Formula	Exceptions
Potassium	K^+	
Ammonium	NH_4^+	
Sodium	Na^+	
Halides (except Fluoride)	C^-, B^-, I^-	Halides of Ag^+, Hg^{2+} and Pb2+
Fluoride	F^-	Fluorides of Mg^{2+} , Ca^{2+} , Sr^{2+} Ba^{2+} , Pb^{2+}
Nitrate	NO_3^-	
Sulfate	SO_4^{2-}	Sulfates of Sr^{2+} , Ba^{2+} Pb^{2+}

Insoluble ions

Ion	Formula	Exceptions
Carbonate	CO_3^{2-}	NH_4^+ salts, Alkali metal salts (group 1)
Phosphate	PO_4^{3-}	NH_4^+ salts, Alkali metal salts (group 1)
Most Hydroxides	OH^-	NH_4^+ salts, Alkali metal salts (group 1)

Ionic Equations

- Ionic equations are equations which only include the **species** (ions) which have changed state.
    - For example an ion changes from being aqueous to being solid.

Worked example: when an acid reacts with sodium carbonate.

The full equation is this:

$$H_2SO_{4\,(aq)} + Na_2CO_{3\,(aq)} \rightarrow Na_2SO_{4\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$$

Next we can split the ionic species into the ions which make them up:

$$2H^+_{\,(aq)} + SO_4^{2-}{}_{(aq)} + 2Na^+_{\,(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)} + 2Na^+_{\,(aq)} + SO_4^{2-}{}_{(aq)}$$

We can then find the ionic species which **don't change state** in the course of the reaction. They have been marked in ***bold italics*** here.

$$2H^+_{\,(aq)} + \boldsymbol{SO_4^{2-}{}_{(aq)}} + \boldsymbol{2Na^+_{\,(aq)}} + CO_3^{2-}{}_{(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)} + \boldsymbol{2Na^+_{\,(aq)}} + \boldsymbol{SO_4^{2-}{}_{(aq)}}$$

To form the ionic equation these species which do not change are removed. The ionic equation for this reaction is as follows:

$$2H^+_{\,(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow H_2O_{\,(l)} + CO_{2\,(g)}$$

There are other reactions of acids which you need to know:

➔ **Neutralisation reactions**: $H^+_{\,(aq)} + OH^-_{\,(aq)} \rightarrow H_2O_{\,(l)}$

➔ **Ammonia reactions**: $H^+_{\,(aq)} + NH_{3\,(aq)} \rightarrow NH_4^+{}_{(aq)}$

➔ **Reactions with metals**: $2H^+_{\,(aq)} + Mg_{\,(s)} \rightarrow Mg^{2+}_{\,(aq)} + H_{2\,(g)}$ (for group 2 metals)

There is another situation in which an ionic equation is needed – **precipitation reactions** forming solids from aqueous species.

Worked example: the reaction to form silver halides in the test for the presence of halide ions. *There is more on this in Topic C16 - Chemical Tests.*

- Silver nitrate is added to an unknown sample which is suspected to contain halide ions such as bromide ions.
- As the only species which would change state (from aqueous to solid), the silver and bromide ions are found in the ionic equation, but the others are not.

$$Ag^{+}{}_{(aq)} + Br^{-}{}_{(aq)} \rightarrow AgBr_{(s)}$$

Half equations

A final type of chemical reaction is **half equations**.

- ➔ One major situation in which these can be used is in **Redox reactions** *(more on this can be found in Topic C5 - Oxidation, Reduction, and Redox)*.
- ➔ These could be either **displacement reactions** or for **electrolysis** reactions.

Worked example: an example of electrolysis is obtaining magnesium from magnesium oxide.

Each magnesium ion gains two electrons to form magnesium metal. The equation is as follows:

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

Each oxide ion loses two electrons to form oxygen gas. As oxygen is a diatomic gas, two oxide ions each lose two electrons as in the equation below:

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

For displacement reactions, one example is magnesium displacing the copper in copper sulfate. The two half-equations are as follows:

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

The overall equation would be: $Cu^{2+}{}_{(aq)} + Mg_{(s)} \rightarrow Cu_{(s)} + Mg^{2+}{}_{(aq)}$

Equilibrium

Reversible reactions in a closed system will reach a state of equilibrium:

- A **reversible reaction is one where the reactants form products, but the products can also react to reform the reactants**. I.e. there is a forward and backward reaction, e.g. **A+B ⇔ C+D.**
- A **closed system** means once the reaction has started **nothing is being added or taken out**, for example products are not being removed or new reactants are not being added.
- The term **'equilibrium'** means that the **forward reaction is occuring at the same rate as the backwards reaction**, i.e. A+B → C+D = C+D → A+B. Because both the forward reaction and the backward reaction are o**ccuring at the same rate there is no overall effect, the amount of reactants and products stay constant. This is known as a dynamic equilibrium.**

Exam tip:
A common way of picturing the concept of dynamic equilibrium is to think of the revolving doors found at the entrances to big shops.

When there are 5 people pushing their way around the revolving door out of the shop and 5 people pushing their way around the revolving door into the shop at exactly the same time, this is a dynamic equilibrium.

- ➔ The speed at which people are leaving the shop is the same as the speed at which people are entering the shop and so overall there is no net difference in the number of people in the shop.
- ➔ Although 5 people leave, another 5 are coming in.
- ➔ And so although it looks like no one is going through the revolving doors as the total number of people inside the shop is not changing, people are still moving through the revolving door, but the speed at which people leave is equal to the speed at which people come in.

Changing the equilibrium position:
The equilibrium position of a reversible reaction is not always in the 'dynamic' equilibrium position where the rate of the forward reaction = the rate of the backward reaction.

By **changing the concentration, temperature and pressure of the reversible reaction, the position of the equilibrium can be changed**; it can be pushed **either to the left side** of the reaction **or the right side**.

For example, in the reaction $H_{2(g)} + I_{2(g)} \Leftrightarrow 2HI_{(g)}$

If the equilibrium position moves to the **right**, this means the **rate of the forward reaction > rate of the backward reaction**

- ➢ i.e. the favoured reaction is $H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$ and so **more products are formed** than reactants.

If the equilibrium position moves to the **left**, this means the **rate of the backward reaction > rate of the forward reaction**,

- ➢ i.e. the favoured reaction is $2HI_{(g)} \rightarrow H_{2(g)} + I_{2(g)}$ and so **more reactants** are formed than products.

Factors that affect the position of equilibrium:

- **Concentration of reactants or product**
- **Pressure of gaseous reactants or products**
- **Temperature**

Increasing the concentration of a reactant → more products are made:

$$A+B \Leftrightarrow C+D$$

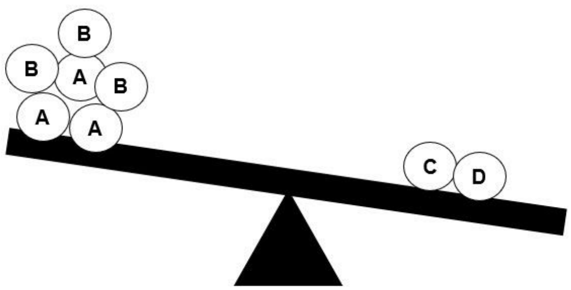
- Increasing the concentration of A or B will cause the **equilibrium position to shift to the right** and so **more products** C+D will be made.
- The **equilibrium position tries to counteract the change and restore balance to the system.**

Increasing the concentration of the product → more reactants are made:

$$A+B \Leftrightarrow C+D$$

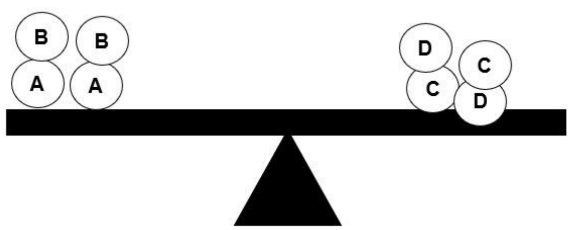
- Increasing the concentration of C or D will cause the **equilibrium position to shift to the left** and so **more reactants** A+B will be made.
- The **equilibrium position tries to counteract the change and restore balance to the system.**

A diagram may help you understand this concept:



Adding more reactant A and B causes the reaction to become unbalanced, it is no longer in the dynamic equilibrium state.

In order to counteract this change of extra reactants and restore balance to the system, the equilibrium position moves to the right to make more product C+D from the extra A+B.



Balance is now restored to the system.

Increasing pressure of the system (only affects gases) → equilibrium position moves to the side with least gaseous molecules:

$$2A_{(g)} + B_{(g)} \Leftrightarrow C_{(g)} + D_{(g)}$$

- Increasing the pressure of the above reaction will cause the **equilibrium position to shift to right**, i.e. to the side of the products C and D and so **more products** are made.
- This is because **when the pressure is increased, the equilibrium position acts to counteract this increase in pressure by decreasing pressure** and therefore **moves to the side of lower pressure which is the side with fewer gaseous molecules.**
- In the above reaction there are 3 moles of gas on the left but only 2 on the right, and so increasing the pressure shifts the equilibrium to the right in this reaction.

Decreasing the pressure of the system (only affects gases) → equilibrium position moves to the side with more gaseous molecules:

$$2A_{(g)} + B_{(g)} \Leftrightarrow C_{(g)} + D_{(g)}$$

- Decreasing the pressure of the above reaction will cause the equilibrium position to **shift to the left**, i.e. to the side of the products 2A + B and so **more reactants** are made.
- This is because **when the pressure is decreased, the equilibrium position acts to counteract this decrease in pressure by increasing the pressure of the system** and t**herefore moves to the side of greater pressure which is the side with a greater number of gaseous molecules.**

Increasing the temperature → equilibrium position shifts in the direction of the endothermic reaction:

In a reversible reaction, one reaction will be endothermic and the other will be exothermic

$$A+B \Leftrightarrow C+D$$

- For example in the above reaction A+B → C+D is the endothermic reaction and C+D → A+B is the exothermic reaction.
- When you increase the temperature by adding heat, the **equilibrium position shifts to decrease the temperature by absorbing the extra heat,** and t**herefore shifts in the direction of the endothermic reaction**. In this example, it means the equilibrium position favours the forward reaction and so more product C+D is made.

Decreasing the temperature → equilibrium position shifts in the direction of the exothermic reaction

$$A+B \Leftrightarrow C+D$$

- For example in the above reaction, C+D → A+B is the exothermic reaction.
- When you **decrease the temperature of the reaction, heat is removed. The equilibrium position shifts to counteract this change** and **add heat to increase the temperature and so the exothermic reaction is favoured** as this releases heat.
- In this example, this means the backward reaction C+D → A+B is favoured and so more reactant is made.

It is important to understand that **catalysts have no effect on the position of equilibrium.** However, catalysts provide an alternative reaction pathway that has a lower activation energy and so a **catalyst acts to increase the rate of both the forward and backward reaction which means the equilibrium position is reached faster**. Adding a catalyst **will not affect the yield.**

The position of equilibrium affects how economical a reaction is. If the equilibrium gives a very low yield under the chosen conditions, it will not be economically viable. A compromise between the concentrations, temperature and pressure to find the best conditions needs to be reached.

Exam tip:
If you remember that the position of equilibrium always acts to counteract any changes made to the reaction system, you will be able to work out whether a certain change to the reaction favours the production of reactants or products.

Take the reaction $N_2O_{4(g)} \Leftrightarrow 2NO_{2(g)}$
If the pressure is increased in this reaction, the system will try to counteract this change of increased pressure and decrease the pressure by moving the equilibrium position to the side of the fewer gaseous molecules, which is the left hand side. And so the formation of reactants will be favoured.