Chem Factsheet

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Why Students Lose Marks: AS Redox Questions

This Factsheet analyses students' real answers to exam questions on reduction and oxidation. By the end of this Factsheet, you should be more confident about:

- What the examiners want
- The kinds of things you are likely to be asked
- Common mistakes and misunderstandings

As you read the students' answers to the questions and the comments, try to work out what the student should have done - using the hints and comments if necessary - before you read the markscheme.

What do you have to know?

In this type of question, the examiner is trying to assess whether you can:

- recognise examples of redox reactions, oxidising and reducing agents
- write half-equations for standard redox conversions
- combine half equations
- calculate oxidation numbers usingthe standard rules
- write half-equations for new reactions, given appropriate information

(a) *Oxidation* used to be defined as 'combination with oxygen'.
Explain why, even though the definition has now been broadened, it is still generally true to say that combination with oxygen is oxidation.

Because oxygen is an effective oxidising agent - it takes electrons away from other elements easily

Although this answer is not on the markscheme, 1 mark awarded for the idea that "oxygen takes electrons away from other elements easily", which demonstrates an understanding of oxidation

..... [2]

(b) When oxygen reacts with fluorine, converting it into oxygen difluoride, OF_2, does the fluorine become oxidised? Explain your answer.

✓
No, fluorine always has an oxidation state of -1 in its compounds

× Although the rule quoted for fluorine is true, a rule is not an explanation - the student should have referred to fluorine being the most electronegative element - or to it being more electronegative than oxygen ..... [2]

Hints and Comments

- *Although it's vital to know the rules for assigning oxidation numbers, they are not explanations - you need to understand **why** fluorine always has an oxidation number of -1 in compounds, oxygen always -2 except with fluorine or in peroxides/superoxides etc.*
- *Unless a previous part of the question has asked for a definition of oxidation - or told you it! - there are likely to be marks available for showing that you understand oxidation is loss of electrons or increase in oxidation number.*

Markscheme

(a) Oxygen is a highly electronegative element (1) so the other element generally loses electrons and is oxidised (1)

(b) Fluorine is not oxidised (1) because it is more electronegative than oxygen/ is the most electronegative element (1)

TiO_2, known in industry as 'titanium dioxide', dissolves in concentrated sulphuric acid to give an orange solution of 'titanyl sulphate', $TiO^{2+}\ SO_4^{2-}$ It also reacts with fused potassium hydroxide to give potassium titanate, $(K^+)_2TiO_3^{2-}$.

(a) Write down the oxidation number of titanium in each of the following :

✓ *correct working and answer*

TiO_2 *? + 2 x -2 = 0 ? = 4* [1]

✓ *correct working and answer*

TiO^{2+} *? - 2 = 2 ? = 4* [1]

× *the working is correct, but the candidate has got into problems with negative numbers*

TiO_3^{2-} *? + 3 x -2 = -2 ? = -8* [1]

(b) Hence state whether each of these conversions involves 'oxidation', 'reduction' or 'neither'.

$TiO_2 \rightarrow TiO^{2+}$ *neither -ON doesn't change* [1]
✓ *correct working and answer*

$TiO_2 \rightarrow TiO_3^{2-}$ *ON goes down - reduction* [1]
ecf✓ *candidate has drawn conclusions correctly using previous incorrect answer*

Hints and Comments

- *You generally should put + signs on oxidation numbers where appropriate*
- *Make sure your answer is plausible - you aren't likely to find oxidation numbers of -8*
- *It was sensible for the candidate to continue with the question - despite the "weird" answer - as "follow through" marks were awarded*

Now try marking this student answer before you look at the markscheme and comments

(a) Define disproportionation *One element is oxidised and reduced*

..... [2]

(b) Copper(I) compounds (except for complexes) are unstable in aqueous solution and decompose by disproportionation. Suggest what you would expect to observe when the white soluble salt, copper(I) sulphate, dissolves in water and write a chemical equation for the reaction.

You get copper (II) sulphate and copper

$Cu_2SO_4 \rightarrow CuSO_4 + Cu$

..... [3]

Here's the markscheme:

(a) *The **simultaneous** oxidation and reduction (1) of a single substance / species (1)*

(b) *A blue solution (1) and a red / brown precipitate (1)*
$Cu_2SO_4(s) \rightarrow CuSO_4(aq) + Cu(s)$ (1)

So how did the student score?

(a) ***One element is oxidised and reduced*** *1/2 - the word "simultaneously" (or "in the same reaction") needed*
The student clearly understood the term, but lost a mark through not being precise enough in the definition

(b) ***You get copper (II) sulphate and copper*** *0/2 -these are not observations - observations refer to what you actually **see** (or smell!) - colours of solutions/solids/gases, effervescence etc*

*$Cu_2SO_4 \rightarrow CuSO_4 + Cu$ 0/1 - although the equation is basically correct, **state symbols** are extremely important here - copper (I) sulphate only disproportionates when it's in solution, and it's necessary to show the copper (II) sulphate is formed as a solution, and that there is a precipitate of copper metal*

Note that you did not have to be familiar with copper (I) compounds or properties to answer this question - although you did need to know what copper (II) sulphate and copper metal looked like.

Arsenic is a toxic element and there are many cases of it being used by murderers. It is regularly monitored in European public drinking water supplies. Modern analytical methods used in water quality assurance begin with complete oxidation of arsenic in the sample, with sodium persulphate and concentrated sulphuric acid, to form H_3AsO_4. This is then reduced in two stages; first, using potassium iodide and acid, to form $HAsO_2$ and, second, using aluminium powder and acid, to form arsenic(III) hydride, AsH_3.

Write balanced ionic half-equations for the following reactions.

(i) H_3AsO_4 to $HAsO_2$ *$H_3AsO_4 \rightarrow HAsO_2 + H_2O + \frac{1}{2}O_2$* *×× no marks awarded - the candidate has introduced an additional product - oxygen gas - in an attempt to balance the equation! The information given about the use of acid has not been included* ***[2]***

(ii) Iodide ion to iodine *$2I^- \rightarrow I_2 + 2e^-$* *✓ full marks here - the candidate has recalled this common redox half equation correctly, with balanced charges* ***[1]***

Hints and Comments

- *Examiners may give you questions about substances you're not familiar with, like arsenic. This is to test your understanding of the general principles and how to apply them - in this sort of question they know you are not relying on memory.*
- *If you are asked to write half-equations, it's always a good idea to work out the oxidation numbers of the element being reduced or oxidised - that tells you how many electrons to put in the half-equation - eg if the ON decreases by 5, you need 5 electrons on the left; if it goes up by 3, you need 3 electrons on the right.*
- *The only substances you should generally be adding to half equations are H^+ (if acid conditions are mentioned) and H_2O.*
- *The iodide - iodine conversion is a standard (and easy) one - you should be able to write the equation down without thinking twice*

Markscheme

(i) Calculating oxidation numbers of As in H_3AsO_4 and $HAsO_2$ as +5 and +3 respectively 1
Correct equation: $H_3AsO_4 + 2H^+ + 2e^- \rightarrow HAsO_2 + 2H_2O$ 1

(ii) $2I^- \rightarrow I_2 + 2e^-$ 1

(a) Write ionic half-equations for the following changes. State symbols may be omitted.

(i) sulphite ions to sulphate ions in acid solution

$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$ ✓ *completely correct equation* [1]

(ii) Manganate(VII) ions to manganese(II) ions in acidic solution.

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ ✓ *completely correct equation* [1]

(b) Combine the ionic half-equations from above to obtain the complete ionic equation for the redox reaction between manganate(VII) ions and sulphite ions in acidic solution.

$2(SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+) + 5(MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O)$

$2SO_3^{2-} + 5MnO_4^- + 2H_2O + 40H^+ \rightarrow 2SO_4^{2-} + 4H^+ + 5Mn^{2+} + 20H_2O$

$2SO_3^{2-} + 5MnO_4^- + 36H^+ \rightarrow 2SO_4^{2-} + 5Mn^{2+} + 18H_2O$ [3]

××× *No marks awarded. The candidate has clearly mis-remembered a method for combining half equations (the correct approach would be to multiply by the number of electrons in the* ***other*** *half-equation).*
If s/he had included the electrons in the working (rather than assuming they would cancel out) or had thought to check the ***charges*** *balanced at the end, s/he might have realised the mistake.*

Hints and Comments

- *It makes life simpler in the exam if you make sure you learn the commoner half-equations*
- *When you are combining half equations, there are two methods you can use - balancing* ***electrons*** *(as mentioned above) or balancing* ***oxidation number change*** *- in the above equation, the oxidation number of the sulphur increases by 2, while that of the manganese goes down by 5 - so 5 sulphurs must be oxidised for every 2 manganese that are reduced. Whichever method you choose, however - always make sure the* ***charges*** *balance at the end*

Markscheme

(a) (i) $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$ (1)

(ii) $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (1)

(b) (i)×5 + (ii)×2 (1)

$2MnO_4^- + 16H^+ + 5SO_3^{2-} + 5H_2O \rightarrow 2Mn^{2+} + 8H_2O + 5SO_4^{2-} + 10H^+$ (1)

$2MnO_4^- + 6H^+ + 5SO_3^{2-} \rightarrow 2Mn^{2+} + 3H_2O + 5SO_4^{2-}$ (1)

Questions

1. (a) Define *oxidation* and *reduction* in terms of electrons. [2]
   (b) Explain why it is that oxidation and reduction always occur together. [1]

2. Calculate the oxidation number of the stated element in each of the following species
   (i) Mn in MnO_4^{2-} (ii) N in HNO_3
   (iii) S in $S_2O_3^{2-}$ (iv) Cr in $Cr_2O_7^{2-}$
   (v) V in VO^{2+} (vi) O in F_2O

3. Titanium dissolves in concentrated hydrochloric acid to give titanium(III) chloride and hydrogen. Construct an ionic equation for this reaction by writing down two ionic half-equations and then combining them. [3]

Answers

1. (a) Oxidation is electron loss (1)
   Reduction is electron gain (1)
   (b) When a substance is oxidised it loses electrons; these must be accepted by another substance, which becomes reduced (1)

2. (i) + 6 (ii) + 5 (iii) + 2 (iv) + 6 (v) + 4 (vi) + 2

3. $Ti \rightarrow Ti^{3+} + 3e^-$ (1)
   $2H^+ + 2e^- \rightarrow H_2$ (1)
   $2Ti + 6H^+ \rightarrow 2Ti^{3+} + 3H_2$ (1)

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