Chem Factsheet

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A FEW H's Will Produce Any Half-Equation

After studying and working your way through this Factsheet you should be confident that you can write any half-equation that is required.

Before proceeding with this, let's summarise the important background information, definitions and ideas which you will need to achieve this.

> **Definitions**
> Redox reactions involve both **red**uction and **ox**idation processes
> Oxidation is **loss** of electrons
> Reduction is **gain** of electrons
> Oxidation involves an **increase** in oxidation state (number)
> Reduction involves an **decrease** in oxidation state (number)
> An oxidising agent (or oxidant) is an **electron acceptor**
> A reducing agent (or reductant) is an **electron donor**

> **Note:** These definitions mean that, during the course of a redox reaction, the oxidant will be reduced and the reductant will be oxidised

Any redox reaction must involve a reduction (electron gain) part **and** an oxidation (electron loss) part. A **half-equation** shows either the reduction or the oxidation process separately. A reduction half-equation will contain electrons on the left-hand side (accepting) whilst an oxidation half-equation will contain electrons on the right-hand side (losing).

e.g. A reduction half-equation : $Cu^{2+} + 2e^- \rightarrow Cu$
e.g. An oxidation half-equation : $Zn \rightarrow Zn^{2+} + 2e^-$

> ***Remember**: If you reverse a reduction half equation, it becomes an oxidation half-equation, and vice-versa.*
>
> *e.g.* $2H^+ + 2e^- \rightarrow H_2$ *represents a reduction but* $H_2 \rightarrow 2H^+ + 2e^-$ *represents an oxidation*

Oxidation states (numbers) provide a quick way to work out if oxidation and reduction has taken place before attempting to write a half-equation. Each atom in a reaction is assigned an oxidation number by using a set of arithmetic rules which are shown below ; if this number increases (oxidation) or decreases (reduction) during the reaction then redox processes are easily spotted.

Rules:

1. In all uncombined ELEMENTS, an atom's oxidation number = 0 .
2. In all COMPOUNDS, the sum of all the oxidation numbers of all the atoms equals zero.
3. In all IONS, the sum of all the oxidation numbers of all the atoms equals the charge on the ion.
4. In all COMPOUNDS, group 1 elements have oxidation number +1, group 2 elements have oxidation number +2, group 3 elements have oxidation number +3 and fluorine (F) has oxidation number -1.
5. In most COMPOUNDS **hydrogen** has an oxidation state of **+1** except in MHn where the metal M will have the positive oxidation number (+n) and and H will be -1.
6. In most COMPOUNDS **oxygen** has an oxidation state of **-2** except in F_2O where F must be -1 (rule 4) causing O to be +2 and peroxides (O_2^{2-}) where the oxidation number is -1 (rule 3)

> ***Remember:** There is a hierarchy of rules, descending in priority from 1 to 6.. Hence the "exceptions" in rules 5 and 6 because earlier rules must be given priority.*

It is vital that half-equations are balanced **both** for atoms **and** overall electrical charge. The most common error is to forget the charge balance. For example,

$$NO_3^- + 4H^+ + e^- \rightarrow NO + 2H_2O$$

represents the reduction (electron gain) of nitrate (NO_3^-) to nitrogen monoxide (NO). It balances in terms of atoms but not in terms of charges. Check it. The FEW H's method for constructing half-equations will automatically balance both atoms and charges.

The FEW H's Method

This is merely a way of remembering the steps, and the order of those steps, required to form a balanced half-equation. There are **up to** four steps in every case ; these are prompted by the acronym "FEWH" where "F", "E", "W" and "H" represent the steps shown below.

F	***Write the <u>F</u>ormulas of the reduced or oxidised particle and the particle produced, on opposite sides of an equation*** e.g. $NO_3^- \rightarrow NO$
E	***Insert the appropriate number of <u>E</u>lectrons into the equation. This is given by the change in oxidation number of the particles shown above. Electrons go on the left for a reduction and on the right for an oxidation.*** $NO_3^- \rightarrow NO$ involves N(+5) changing to N(+2) oxidation state. This shows the process is a reduction and that 3 electrons must be involved in the final half-equation. e.g. $NO_3^- + 3e^- \rightarrow NO$
W	***Insert the appropriate number of <u>W</u>ater molecules into the equation so that oxygen atoms are made to balance.*** $NO_3^- + 3e^- \rightarrow NO$ shows a shortage of 2 O atoms on the right. Hence, $2H_2O$ is inserted on the right. e.g. $NO_3^- + 3e^- \rightarrow NO + 2H_2O$
H	***Insert the appropriate number of <u>H</u>ydrogen ion into the equation so that hydrogen atoms are made to balance*** $NO_3^- + 3e^- \rightarrow NO + 2H_2O$ shows a shortage of 4 H atoms on the left. Hence, $4H^+$ are inserted on the left e.g. $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ This automatically completes the fully balanced half-equation.

> Simple half-equations such as $Cu^{2+} + 2e^- \rightarrow Cu$ and $Zn \rightarrow Zn^{2+} + 2e^-$ do not need the "W"and"H" steps but "F" and "E" still apply. In fact, such examples can usually be done simply by inspection! In general, 2 or 4 of the FEWH steps will be needed but never 1 or 3.

Half-equations involving H^+ imply acidic conditions. The corresponding half-equation in alkaline conditions can be formed by adding the corresponding number of OH^- ions to each side of the "acidic" half-equation and then cancelling the water molecules.

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ becomes
$NO_3^- + 4H^+ + 4OH^- + 3e^- \rightarrow NO + 2H_2O + 4OH^-$
➔ $NO_3^- + 4H_2O + 3e^- \rightarrow NO + 2H_2O + 4OH^-$
➔ $NO_3^- + 2H_2O + 3e^- \rightarrow NO + 4OH^-$

Some Further Examples

Construct fully balanced half-equations for each of the following changes

1. Cl_2 to ClO^- in alkaline solution
2. BrO_3^- to Br^- in acid solution
3. Sn to SnO_2 in alkaline solution
4. MnO_4^- to Mn^{2+} in acid soution
5. $Cr_2O_7^{2-}$ to Cr^{3+} in acid solution.

1.	F	$Cl_2 \rightarrow 2ClO^-$ (note how 2 moles of ClO^- must be formed)
	E	$Cl_2 \rightarrow 2ClO^-$ involves 2Cl(0) changing to 2Cl(+1) oxidation state. This shows the process is an oxidation and that 2 electrons must be involved on the right in the final half-equation. $Cl_2 \rightarrow 2ClO^- + 2e^-$
	W	$Cl_2 \rightarrow 2ClO^- + 2e^-$ shows a shortage of 2 O atoms on the left. Hence, $2H_2O$ is inserted on the left. $Cl_2 + 2H_2O \rightarrow 2ClO^- + 2e^-$
	H	$Cl_2 + 2H_2O \rightarrow 2ClO^- + 2e^-$ shows a shortage of 4 H atoms on the right. Hence, $4H^+$ are inserted on the right. $Cl_2 + 2H_2O \rightarrow 2ClO^- + 4H^+ + 2e^-$ ➔ $Cl_2 + 4OH^- + 2H_2O$ ➔ $2ClO^- + 4H^+ + 4OH^- + 2e^-$ **➔ $Cl_2 + 4OH^- \rightarrow 2ClO^- + 2H_2O + 2e^-$**

2.	F	$BrO_3^- \rightarrow Br^-$
	E	$BrO_3^- \rightarrow Br^-$ involves Br(+5) changing to Br(-1) oxidation state. This shows the process is a reduction and that 6 electrons must be involved on the left in the final half-equation. $BrO_3^- + 6e^- \rightarrow Br^-$
	W	$BrO_3^- + 6e^- \rightarrow Br^-$ shows a shortage of 3 O atoms on the right. Hence, $3H_2O$ is inserted on the right. $BrO_3^- + 6e^- \rightarrow Br^- + 3H_2O$
	H	$BrO_3^- + 6e^- \rightarrow Br^- + 3H_2O$ shows a shortage of 6 H atoms on the left. Hence, $6H^+$ are inserted on the left **$BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O$**

3.	F	$Sn \rightarrow SnO_2$
	E	$Sn \rightarrow SnO_2$ involves Sn(0) changing to Sn(+4) oxidation state. This shows the process is an oxidation and that 4 electrons must be involved on the right in the final half-equation. $Sn \rightarrow SnO_2 + 4e^-$
	W	$Sn \rightarrow SnO_2 + 4e^-$ shows a shortage of 2 O atoms on the left. Hence, $2H_2O$ is inserted on the left. $Sn + 2H_2O \rightarrow SnO_2 + 4e^-$
	H	$Sn + 2H_2O \rightarrow SnO_2 + 4e^-$ shows a shortage of 4 H atoms on the rightt. Hence, $4H^+$ are inserted on the right $Sn + 2H_2O \rightarrow SnO_2 + 4H^+ + 4e^-$ ➔ $Sn + 4OH^- + 2H_2O \rightarrow SnO_2 + 4H^+ + 4OH^- + 4e^-$ **➔ $Sn + 4OH^- \rightarrow SnO_2 + 2H_2O + 4e^-$**

4.	F	$MnO_4^- \rightarrow Mn^{2+}$
	E	$MnO_4^- \rightarrow Mn^{2+}$ involves Mn(+7) changing to Mn(+2) oxidation state. This shows the process is a reduction and that 5 electrons must be involved on the left in the final half-equation. $MnO_4^- + 5e^- \rightarrow Mn^{2+}$
	W	$MnO_4^- + 6e^- \rightarrow Mn^{2+}$ shows a shortage of 4 O atoms on the right. Hence, $4H_2O$ is inserted on the right. $MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$
	H	$MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$ shows a shortage of 8 H atoms on the left. Hence, $8H^+$ are inserted on the left **$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$**

5.	F	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ (note how 2 moles of Cr^{3+} must be formed)
	E	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ involves 2Cr(+6) changing to 2Cr(+3) oxidation state. This shows the process is a reduction and that 6 electrons electrons must be involved on the left in the final half-equation. $Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+}$
	W	$Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+}$ shows a shortage of 7 O atoms on the right. Hence, $7H_2O$ is inserted on the right. $Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
	H	$Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ shows a shortage of 14 H atoms on the left. Hence, 14 H^+ are inserted on the left **$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$**

Now try a few examples for your self:

Practice Questions

1. I_2 to IO_4^- in alkaline solution
2. H_2SO_4 to H_2S in acid solution
3. NO_2 to NO_3^- in alkaline solution
4. VO_3^- to V^{2+} in acid soution
5. RCHO (an aldehyde) to RCOOH (an acid) in alkaline solution. [Hint ; assign oxidation state 0 to the R group since it remains unchanged]

Answers

1. $I_2 + 16OH^- \rightarrow 2IO_4^- + 8H_2O + 14e^-$
2. $H_2SO_4 + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$
3. $NO_2 + 2OH^- \rightarrow NO_3^- + H_2O + e^-$
4. $VO_3^- + 6H^+ + 3e^- \rightarrow V^{2+} + 3H_2O$
5. $RCHO + 2OH^- \rightarrow RCOOH + H_2O + 2e^-$

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