

How to Balance Chemical Equations

Chemistry, like all other sciences, is the study of the behaviour of materials. What especially sets chemistry apart from other sciences are chemical reactions. A chemical reaction is a process that leads to the change of one set of chemical species into a different set of chemical species without the change in the nuclei of any species. A chemical equation is a written shorthand symbolic representation of such a chemical reaction.

e.g. A $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{Na}^+(\text{aq}) + 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$;
{ $2\text{Na}^+(\text{aq}) + 2\text{OH}^-(\text{aq})$ may be written as 2NaOH(aq) }

The balanced chemical equation always shows:

- the symbols / formulas of the species reacting on the left (reactants) and also those of the species formed on the right (products),
- the coefficients in front of each species showing in what ratio the species react and are produced,
- an arrow from reactants to products which means "produces".

State (phase) symbols may also be included; e.g. (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution. A reversible chemical reaction would have two arrows (\rightleftharpoons).

Ignoring the state symbols, the equation in example A "reads": two (moles of) sodium atoms react with two (moles of) water molecules to produce two (moles of) sodium ions, two (moles of) hydroxide ions and one (mole of) hydrogen molecule(s).

An equation is **balanced** when:

- the number of (moles of) atoms of each element in the reactants and the products are equal,
- the sum of the ionic charges on the reactant particles and the product particles are equal.

In e.g. A, on the left and on the right there are 2Na, 4H and 2O \rightarrow atoms balance.

In e.g. A, on both sides the sum of the charges is zero \rightarrow charges balance.

The equation can also be balanced using fractions where the coefficients represent "moles" since a fraction of a molecule, atom or formula does not exist but a fraction of a mole is feasible.

Hence, $\text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) + \frac{1}{2}\text{H}_2(\text{g})$
is an alternative to e.g. A.

Note. The number of molecules (atoms, ions)
= Avogadro's number \times number of moles (n)
= $6.023 \times 10^{23} \times n$ particles

The law of conservation of mass applies to a balanced chemical equation so that the total mass of the reactants is always equal to the total mass of products and also the total number and mass of each type of atom is unchanged.

The Chemical Formulas Must Be Correct

Before an equation can be balanced the chemical formulas must be correct. Chemical formulas may either be provided, remembered or worked out in some way. For example, from the position of the elements in the Periodic Table, from oxidation states or by remembering the formulas, especially for ions such as NH_4^+ , SO_4^{2-} , etc.

Four common errors are:

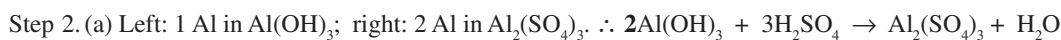
- writing the symbol for an atom of an element instead of the formula for the molecule of the element. Thus H_2 , N_2 , O_2 , F_2 , Cl_2 , Br_2 and I_2 must be used rather than H, N, O etc. P_4 and S_8 may also be used but on most occasions P and S are allowable. For all other elements, the chemical symbols are used e.g. Na, C, Fe, etc,
- writing the wrong chemical formula because no thought is given to the positions of the elements in the Periodic Table. e.g. magnesium chloride is written as MgCl because of the two words in the name! But Mg is in group 2 \therefore Mg^{2+} and Cl is in group 7 \therefore Cl^- so the formula is MgCl_2 .
- writing the wrong chemical formula because the formulas of ions such as carbonate (CO_3^{2-}) have not been learned.
- not using given oxidation states to deduce the formula. For example, sodium chlorate(I) - Na is +1 (group 1), O = -2 (group 6-8), Cl = +1 (given) - hence the formula is NaClO as the sum of the oxidation states must be zero.

How to Decide Which Method to Use

Is it a REDOX REACTION?	
No	Yes
Types of reaction: e.g. acid-base, precipitation	If the reaction is combustion or "simple" redox use the counting method as for a non-redox reaction
Method: Counting	Method: Half-equations / oxidation numbers
Step 1. If they are present, count and balance groups of atoms, (e.g. SO_4 , CO_3 , NH_4).	Step 1. Write the two half-equations. Reduction: decrease in ox. no. = electrons gained. Oxidation: increase in ox. no. = electrons lost. • Check charges balance!
Step 2. Count and balance atoms of other elements. See notes in 1 e.g.s. 1, 3 and 4.	Step 2. If unequal, make the number of electrons in each half-equation equal by appropriate multiplication.
Note If charges are included they will automatically be balanced if steps 1 and 2 are correct.	Step 3. Add these and cancel like terms

1. Examples of Equations Balanced by Counting Groups of Atoms and Individual Atoms

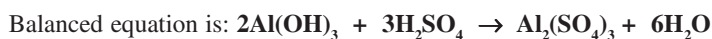
e.g. 1. Acid-Base



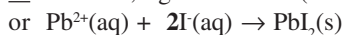
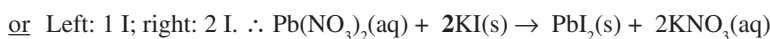
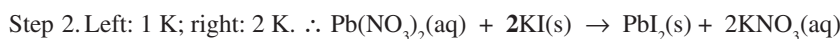
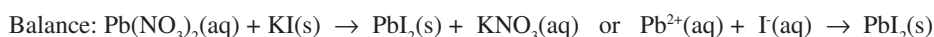
Note The SO_4 groups have been balanced. \therefore balance Al before O's or H's since Al is in $\text{Al}_2(\text{SO}_4)_3$ and its coefficient will stay at 1.



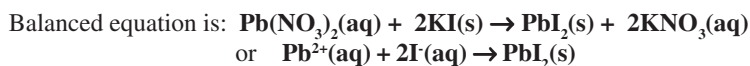
This also balances the hydrogen atoms but check it! $[(2 \times 3) + (3 \times 2)] = (6 \times 2)$.



e.g. 2. Precipitation



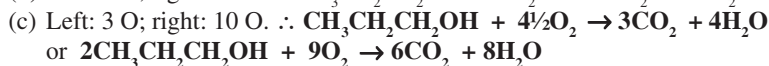
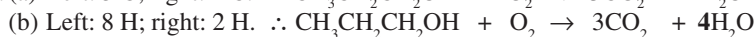
Note. $\text{Pb}^{2+}(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$ may also be balanced by counting the ionic charges.



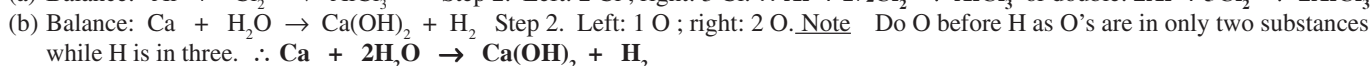
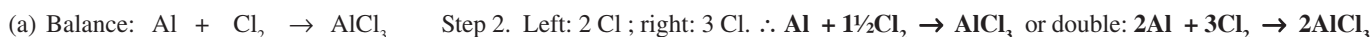
e.g. 3. Combustion



Note. Do C and H's before O since all C's got to CO_2 and all H's got to H_2O .



e.g. 4. Simple Redox Reactions

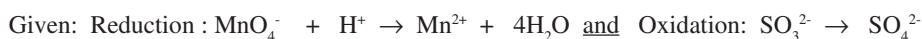


2. Examples of Balancing Redox Equations

(i) Balancing Equations Under Acidic Conditions.

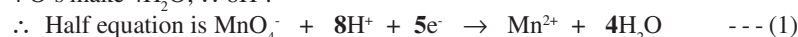
Acid conditions means H^+ ions are included. These ions will usually result in water as a product.

e.g. 1. Balance the equation for the reaction under acid conditions between manganate(VII) ions and sulfate(IV) ions if these ions are converted to manganese(II) ions and sulfate(VI) ions respectively.



Step 1. $\text{Mn}(+7) \rightarrow (+2) \therefore 5 \text{ e}^-$ on left.

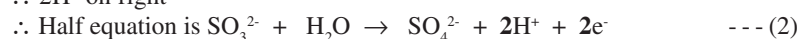
4 O's make $4\text{H}_2\text{O}$; $\therefore 8\text{H}^+$.

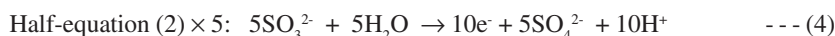


$\text{S}(+4) \rightarrow (+6) \therefore 2 \text{ e}^-$ on right.

Extra O on right $\rightarrow 1\text{H}_2\text{O}$ on left since this provides O

$\therefore 2\text{H}^+$ on right

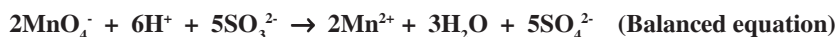




Step 3. Add (3) and (4).

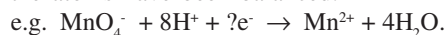


Cancel H^+ and H_2O .



Note 1 Equation (5) can be obtained directly from the half-equations (1) and (2).
(5) = (1) $\times 2$ add (2) $\times 5$.

Note 2 An alternative method for finding the number of electrons for the half-equations is to balance the charge **after** all the atoms have been balanced.



5e^- are needed to make the total charge +2 on both sides.

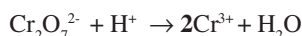
Using the change in ox. nos. the number of electrons can be worked out before all the atoms are balanced; but the atom whose oxidation number changes **must** be balanced first.

e.g. 2. An organic example. Write the equation for the oxidation of ethanol to ethanal using potassium dichromate(VI) in sulfuric acid where chromium forms Cr^{3+} .



Step 1. **Immediately balance atoms that have a change in oxidation number.**

Left: 2 Cr; right: 1 Cr. $\therefore 2\text{Cr}^{3+}$



$2\text{Cr}(+6) \rightarrow 2(+3) \therefore 6\text{e}^-$ on left.

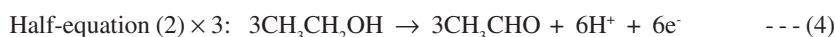
7 O's make $7\text{H}_2\text{O}$; $\therefore 14\text{H}^+$.



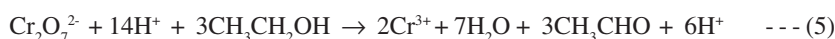
$2\text{C}(-2) \rightarrow 2(-1) \therefore 2\text{e}^-$ on right.

2 H lost

$\therefore 2\text{H}^+$ on right



Step 3. Add (3) and (4).



Cancel H^+ and H_2O .

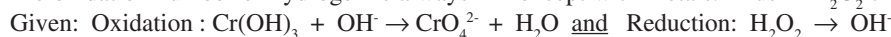


(ii) Balancing Equations Under Alkaline Conditions

Alkaline conditions means OH^- ions are needed. These ions will usually result in water as a product.

e.g. Write the equation for the reaction for the oxidation of a chromium(III) hydroxide to chromate(VI) ions by hydrogen peroxide under alkaline conditions.

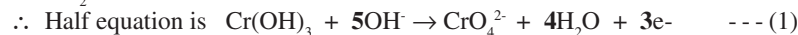
Note. The oxidation number of hydrogen is always +1 except with metals. Thus in H_2O_2 the ox. no. of oxygen is -1.



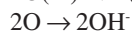
Step 1. $\text{Cr}(+3) \rightarrow (+6) \therefore 3\text{e}^-$ on right.

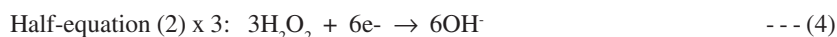
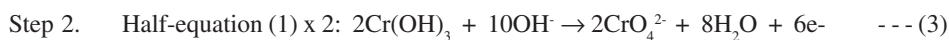
$\therefore 5\text{OH}^-$ to balance charge.

$\therefore 4\text{H}_2\text{O}$ to balance 8H.

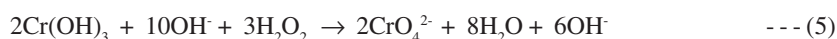


$2\text{O}(-1) \rightarrow 2(-2) \therefore 2\text{e}^-$ on left.





Step 3. Add (3) and (4).



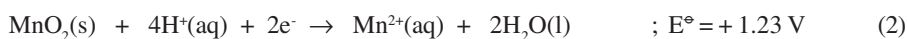
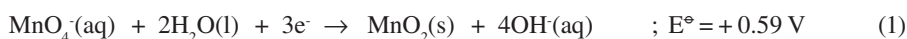
Cancel OH^- and H_2O .



(iii) Balancing an Equation from Standard Redox Potential Values

The E° values are used to decide which direction is favoured by any given half equation. The half equations are then combined as in part (ii).

e.g. Balance the reaction that is feasible from the following E° values.



A. Decide which is the better reducing agent and which is the better oxidising agent. These react together.

Reducing agents are on the right. The better reductant has the more negative E° .

Oxidising agents are on the left. The better oxidant has the more positive E° .

B. Reverse the half-equation for the better reducing agent.

A. $\text{MnO}_2(\text{s}) + 4\text{OH}^-(\text{aq})$ is the better reducing agent.

Hence, $2 \times$ reverse of (1) is combined with $3 \times$ (2) giving:



3. Writing a Balanced Equation from Experimental Data

This is possible when the moles of the each reactant can be experimentally determined and the products are known or can be deduced.

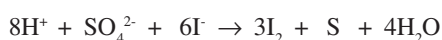
If 10.0 cm^3 0.005 M I_2 just reacts with 10.0 cm^3 $0.100 \text{ M S}_2\text{O}_3^{2-}$ which is converted to $\text{S}_4\text{O}_6^{2-}$, then by inspection there are twice as many moles $\text{S}_2\text{O}_3^{2-}$ than I_2 .

Also, since $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-}$, two negative charges are needed on the right. These must be provided by I_2 converting to 2I^- .

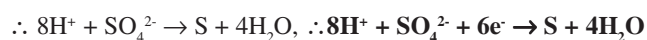
Hence, the ionic equation is: $\text{I}_2 + 2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{I}^-$.

4. Writing Half-Equations from the Balanced Equation

e.g. Sulfate(VI) ions can be reduced by iodide ions to sulfur:



To obtain the half-equations select all "connected species" and add the missing electrons appropriately before simplifying.



Practice Questions

- Write equations for the reaction between phosphoric(V) acid (H_3PO_4) and magnesium hydroxide.
- Write equations for the combustion of pent-2,4-dione
- Balance: $\text{Cr}_2\text{O}_7^{2-} + \text{H}^+ + \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{Cr}^{3+} + \text{H}_2\text{O}$
- Balance: $\text{Fe}(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{HFeCl}_4(\text{aq}) + \text{H}_2(\text{g})$
- Use the E° data to write the reaction that is feasible in a lead-acid battery.
 $\text{PbSO}_4(\text{s}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq}) \quad ; E^\circ = -0.36 \text{ V}$
 $\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^- \rightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \quad ; E^\circ = +1.69 \text{ V}$
- The following is the net reaction that occurs in a fuel cell with potassium hydroxide as the electrolyte and ethanol as fuel: $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$.
Write the half-equations for the reaction that occur at the each electrode. (Note Both half-equation require OH^-)

Answers

- $2\text{H}_3\text{PO}_4 + 3\text{Mg}(\text{OH})_2 \rightarrow \text{Mg}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}$
- $\text{CH}_3\text{COCH}_2\text{COCH}_3 + 6\text{O}_2 \rightarrow 5\text{CO}_2 + 4\text{H}_2\text{O}$
- $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{Fe}^{2+} \rightarrow 6\text{Fe}^{3+} + 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
- $\text{Fe}(\text{s}) + 4\text{HCl}(\text{aq}) \rightarrow \text{HFeCl}_4(\text{aq}) + 1\frac{1}{2}\text{H}_2(\text{g})$ (or double)
- $\text{Pb}(\text{s}) + 2\text{SO}_4^{2-}(\text{aq}) + \text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) \rightarrow 2\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- Negative electrode: $\text{C}_2\text{H}_5\text{OH} + 12\text{OH}^- \rightarrow 2\text{CO}_2 + 9\text{H}_2\text{O} + 12\text{e}^-$
Positive electrode: $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$

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