Chem Factsheet

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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 $2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g)$;
$\{2Na^+(aq) + 2OH^-(aq)$ 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 → atoms balance.

In e.g. A, on both sides the sum of the charges is zero → 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, $Na(s) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + \frac{1}{2}H_2(g)$
is an alternative to e.g. A.

Note. The number of molecules (atoms, ions)
= Avogadro's number × 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 ∴ Mg^{2+} and Cl is in group 7 ∴ 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 2. Count and balance atoms of other elements. See notes in 1 e.gs. 1, 3 and 4. Note If charges are included they will automatically be balanced if steps 1 and 2 are correct.	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. If unequal, make the number of electrons in each half-equation equal by appropriate multiplication. 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

Balance: $Al(OH)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$

Step 1. Left: 1 SO_4; right: 3 SO_4. ∴ $Al(OH)_3 + \mathbf{3}H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$

Step 2. (a) Left: 1 Al in $Al(OH)_3$; right: 2 Al in $Al_2(SO_4)_3$. ∴ $\mathbf{2}Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$

Note The SO_4 groups have been balanced. ∴ balance Al before O's or H's since Al is in $Al_2(SO_4)_3$ and its coefficient will stay at 1.

(b) Left: 6 O in $2Al(OH)_3$; right: 1 O in H_2O. ∴ $2Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + \mathbf{6}H_2O$

This also balances the hydrogen atoms but check it! [(2x3) + (3x2)] = (6x2).

Balanced equation is: $\mathbf{2Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6H_2O}$

e.g. 2. Precipitation

Balance: $Pb(NO_3)_2(aq) + KI(s) \rightarrow PbI_2(s) + KNO_3(aq)$ or $Pb^{2+}(aq) + I^-(aq) \rightarrow PbI_2(s)$

Step 1. Left: 2 NO_3, right: 1 NO_3. ∴ $Pb(NO_3)_2(aq) + KI(s) \rightarrow PbI_2(s) + \mathbf{2}KNO_3(aq)$

Step 2. Left: 1 K; right: 2 K. ∴ $Pb(NO_3)_2(aq) + \mathbf{2}KI(s) \rightarrow PbI_2(s) + 2KNO_3(aq)$

or Left: 1 I; right: 2 I. ∴ $Pb(NO_3)_2(aq) + \mathbf{2}KI(s) \rightarrow PbI_2(s) + 2KNO_3(aq)$
or $Pb^{2+}(aq) + \mathbf{2}I^-(aq) \rightarrow PbI_2(s)$

Note. $Pb^{2+}(aq) + I^-(aq) \rightarrow PbI_2(s)$ may also be balanced by counting the ionic charges.

Balanced equation is: $\mathbf{Pb(NO_3)_2(aq) + 2KI(s) \rightarrow PbI_2(s) + 2KNO_3(aq)}$
or $\mathbf{Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)}$

e.g. 3. Combustion

Balance: $CH_3CH_2CH_2OH + O_2 \rightarrow CO_2 + H_2O$

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

Step 2. (a) Left: 3 C; right: 1 C. ∴ $CH_3CH_2CH_2OH + O_2 \rightarrow \mathbf{3}CO_2 + H_2O$
(b) Left: 8 H; right: 2 H. ∴ $CH_3CH_2CH_2OH + O_2 \rightarrow 3CO_2 + \mathbf{4}H_2O$
(c) Left: 3 O; right: 10 O. ∴ $\mathbf{CH_3CH_2CH_2OH + 4\tfrac{1}{2}O_2 \rightarrow 3CO_2 + 4H_2O}$
or $\mathbf{2CH_3CH_2CH_2OH + 9O_2 \rightarrow 6CO_2 + 8H_2O}$

e.g. 4. Simple Redox Reactions

(a) Balance: $Al + Cl_2 \rightarrow AlCl_3$ Step 2. Left: 2 Cl ; right: 3 Cl. ∴ $\mathbf{Al + 1\tfrac{1}{2}Cl_2 \rightarrow AlCl_3}$ or double: $\mathbf{2Al + 3Cl_2 \rightarrow 2AlCl_3}$
(b) Balance: $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$ Step 2. Left: 1 O ; right: 2 O. Note Do O before H as O's are in only two substances while H is in three. ∴ $\mathbf{Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2}$

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.

Given: Reduction : $MnO_4^- + H^+ \rightarrow Mn^{2+} + 4H_2O$ and Oxidation: $SO_3^{2-} \rightarrow SO_4^{2-}$

Step 1. Mn(+7) → (+2) ∴ 5 e- on left.
4 O's make $4H_2O$; ∴ $8H^+$.
∴ Half equation is $MnO_4^- + \mathbf{8}H^+ + \mathbf{5}e^- \rightarrow Mn^{2+} + \mathbf{4}H_2O$ - - - (1)

S(+4) → (+6) ∴ 2 e- on right.
Extra O on right → $1H_2O$ on left since this provides O
∴ $2H^+$ on right
∴ Half equation is $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + \mathbf{2}H^+ + \mathbf{2}e^-$ - - - (2)

Step 2. Half-equation (1) × 2: $2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$ --- (3)

Half-equation (2) × 5: $5SO_3^{2-} + 5H_2O \rightarrow 10e^- + 5SO_4^{2-} + 10H^+$ --- (4)

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

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

Cancel H^+ and H_2O.

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

Note 1 Equation (5) can be obtained directly from the half-equations (1) and (2).
(5) = (1) × 2 add (2) × 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.
e.g. $MnO_4^- + 8H^+ + ?e^- \rightarrow Mn^{2+} + 4H_2O$.
$5e^-$ 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+}.

Given: Reduction : $Cr_2O_7^{2-} + H^+ \rightarrow Cr^{3+} + H_2O$ and Oxidation: $CH_3CH_2OH \rightarrow CH_3CHO$

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

Left: 2 Cr; right: 1 Cr. ∴ **2**Cr^{3+}

$Cr_2O_7^{2-} + H^+ \rightarrow \mathbf{2}Cr^{3+} + H_2O$

2Cr(+6) → 2(+3) ∴ 6 e- on left.
7 O's make $7H_2O$; ∴ $14H^+$.
∴ Half equation is $Cr_2O_7^{2-} + \mathbf{14}H^+ + \mathbf{6}e^- \rightarrow \mathbf{2}Cr^{3+} + 7H_2O$ --- (1)

2C(-2) → 2(-1) ∴ 2 e- on right.
2 H lost
∴ $2H^+$ on right
∴ Half equation is $CH_3CH_2OH \rightarrow CH_3CHO + \mathbf{2}H^+ + \mathbf{2}e^-$ --- (2)

Step 2. Half-equation (1) × 1: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ --- (3)

Half-equation (2) × 3: $3CH_3CH_2OH \rightarrow 3CH_3CHO + 6H^+ + 6e^-$ --- (4)

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

$Cr_2O_7^{2-} + 14H^+ + 3CH_3CH_2OH \rightarrow 2Cr^{3+} + 7H_2O + 3CH_3CHO + 6H^+$ --- (5)

Cancel H^+ and H_2O.

$Cr_2O_7^{2-} + 8H^+ + 3CH_3CH_2OH \rightarrow 2Cr^{3+} + 7H_2O + 3CH_3CHO$

(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.
Given: Oxidation : $Cr(OH)_3 + OH^- \rightarrow CrO_4^{2-} + H_2O$ and Reduction: $H_2O_2 \rightarrow OH^-$

Step 1. Cr(+3) → (+6) ∴ 3 e- on right.
∴ $5OH^-$ to balance charge.
∴ $4H_2O$ to balance 8H.
∴ Half equation is $Cr(OH)_3 + \mathbf{5}OH^- \rightarrow CrO_4^{2-} + \mathbf{4}H_2O + \mathbf{3}e^-$ --- (1)

2O(-1) → 2(-2) ∴ 2 e- on left.
2O → $2OH^-$
∴ Half equation is $H_2O_2 + \mathbf{2}e^- \rightarrow \mathbf{2}OH^-$ --- (2)

Step 2. Half-equation (1) x 2: $2Cr(OH)_3 + 10OH^- \rightarrow 2CrO_4^{2-} + 8H_2O + 6e-$ --- (3)

Half-equation (2) x 3: $3H_2O_2 + 6e- \rightarrow 6OH^-$ --- (4)

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

$2Cr(OH)_3 + 10OH^- + 3H_2O_2 \rightarrow 2CrO_4^{2-} + 8H_2O + 6OH^-$ --- (5)

Cancel OH^- and H_2O.

$2Cr(OH)_3 + 4OH^- + 3H_2O_2 \rightarrow 2CrO_4^{2-} + 8H_2O$ (Balanced equation)

(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 $E^\ominus$ values.

$MnO_4^-(aq) + 2H_2O(l) + 3e^- \rightarrow MnO_2(s) + 4OH^-(aq)$; $E^\ominus = +0.59$ V (1)

$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$; $E^\ominus = +1.23$ V (2)

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 $E^\ominus$.
Oxidising agents are on the left. The better oxidant has the more positive $E^\ominus$.

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

A. $MnO_2(s) + 4OH^-(aq)$ is the better reducing agent.

Hence, 2 × reverse of (1) is combined with 3 × (2) giving:

$5MnO_2(s) + 4H^+(aq) \rightarrow 2MnO_4^-(aq) + 3Mn^{2+}(aq) + 2H_2O(l)$

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 M $S_2O_3^{2-}$ which is converted to $S_4O_6^{2-}$, then by inspection there are twice as many moles $S_2O_3^{2-}$ than I_2.

Also, since $2S_2O_3^{2-} \rightarrow S_4O_6^{2-}$, two negative charges are needed on the right. These must be provided by I_2 converting to $2I^-$.

Hence, the ionic equation is: **$I_2 + 2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2I^-$**.

4. Writing Half-Equations from the Balanced Equation

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

$8H^+ + SO_4^{2-} + 6I^- \rightarrow 3I_2 + S + 4H_2O$

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

$\therefore 6I^- \rightarrow 3I_2$, $\therefore 6I^- \rightarrow 3I_2 + 6e^-$, $\therefore$ **$2I^- \rightarrow I_2 + 2e^-$**

$\therefore 8H^+ + SO_4^{2-} \rightarrow S + 4H_2O$, $\therefore$ **$8H^+ + SO_4^{2-} + 6e^- \rightarrow S + 4H_2O$**

Practice Questions

1. Write equations for the reaction between phosphoric(V) acid (H_3PO_4) and magnesium hydroxide.
2. Write equations for the combustion of pent-2,4-dione
3. Balance: $Cr_2O_7^{2-} + H^+ + Fe^{2+} \rightarrow Fe^{3+} + Cr^{3+} + H_2O$
4. Balance: $Fe(s) + HCl(aq) \rightarrow HFeCl_4(aq) + H_2(g)$
5. Use the $E^\varnothing$ date to write the reaction that is feasible in a lead-acid battery.
   $PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$; $E^\ominus = -0.36$ V
   $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$; $E^\ominus = +1.69$ V
6. The following is the net reaction that occurs in a fuel cell with potassium hydroxide as the electrolyte and ethanol as fuel: $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$.
   Write the half-equations for the reaction that occur at the each electrode. (<u>Note</u> Both half-equation require OH^-)

Answers

1. $2H_3PO_4 + 3Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6H_2O$
2. $CH_3COCH_2COCH_3 + 6O_2 \rightarrow 5CO_2 + 4H_2O$
3. $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$
4. $Fe(s) + 4HCl(aq) \rightarrow HFeCl_4(aq) + 1½H_2(g)$ (or double)
5. $Pb(s) + 2SO_4^{2-}(aq) + PbO_2(s) + 4H^+(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$
6. Negative electrode : $C_2H_5OH + 12OH^- \rightarrow 2CO_2 + 9H_2O + 12e^-$
   Positive electrode : $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

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