INTERNATIONAL COLLEGE OF PHARMACEUTICAL INNOVATION
国际创新药学院

Class	Pharm, BioPharm
Course	Fundamentals of Medicinal & Pharmaceutical Chemistry
Code	FUNCHEM.6
Title	Redox Reactions: Energy-Producing Reactions at the Molecular Level of Life
Lecturer	Prof. Xinchen Teng
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RECOMMENDED READING

- General Chemistry - The Essential Concepts by Chang and Goldsby (7th edition)
  - Section 4.4 – Oxidation reduction reactions
  - Section 19.1 – Redox Reaction

FUNCHEM.6 Learning Outcomes

- Define ‘oxidation’, ‘reduction’, ‘oxidising agent’, ‘reducing agent’, ‘redox reaction’ and ‘oxidation number’.
- Discuss the importance of redox reactions in the human body.
- Recall rules to assign oxidation numbers (free elements, molecules and ions). Recall exceptions to the rules.
- Demonstrate method of balancing redox equations.

Redox Reactions

Energy producing reactions at the molecular level of life.

Energy is required for 3 major purposes:

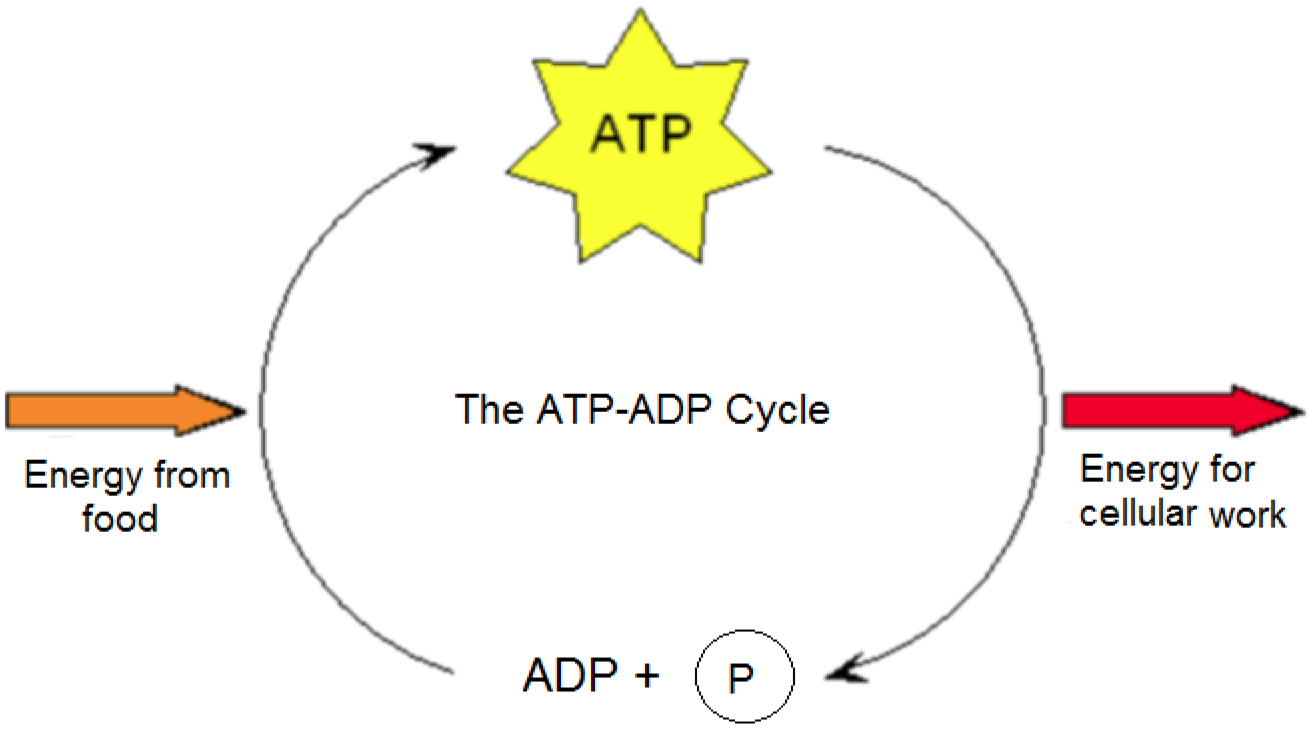
Muscle contraction and other cellular movements

Active transport of molecules and ions

Synthesis of macromolecules and other biomolecules from simple precursors

The ATP – ADP cycle

the fundamental way of energy exchange in biological systems



<u>Oxidation</u> of food

Oxidation in Redox Reaction

Oxidation occurs when a molecule:

Loses electrons

Loses hydrogen

Gains oxygen

If a molecule undergoes oxidation, it has been oxidized and it is said to be the **reducing agent**

Reduction in Redox Reaction

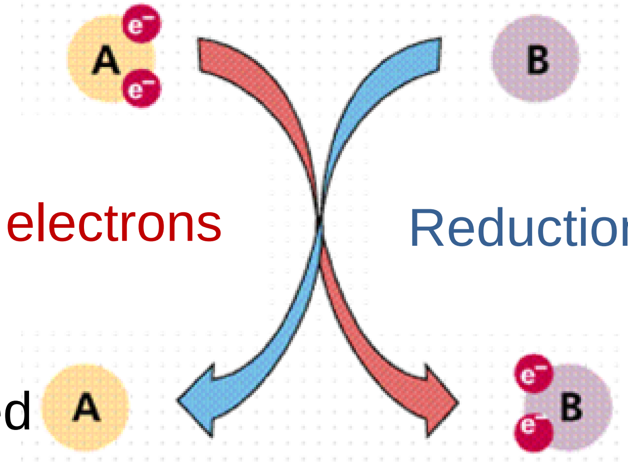
Reduction occurs when a molecule does any of the following:

Gains electrons
Gains hydrogen
Loses oxygen

If a molecule undergoes reduction, it has been reduced and it is said to be the **oxidizing agent**

Oxidation and Reduction reactions always take place simultaneously

Reducing agent **Oxidizing agent**



Oxidation: loss of electrons

Reduction: gain of electrons

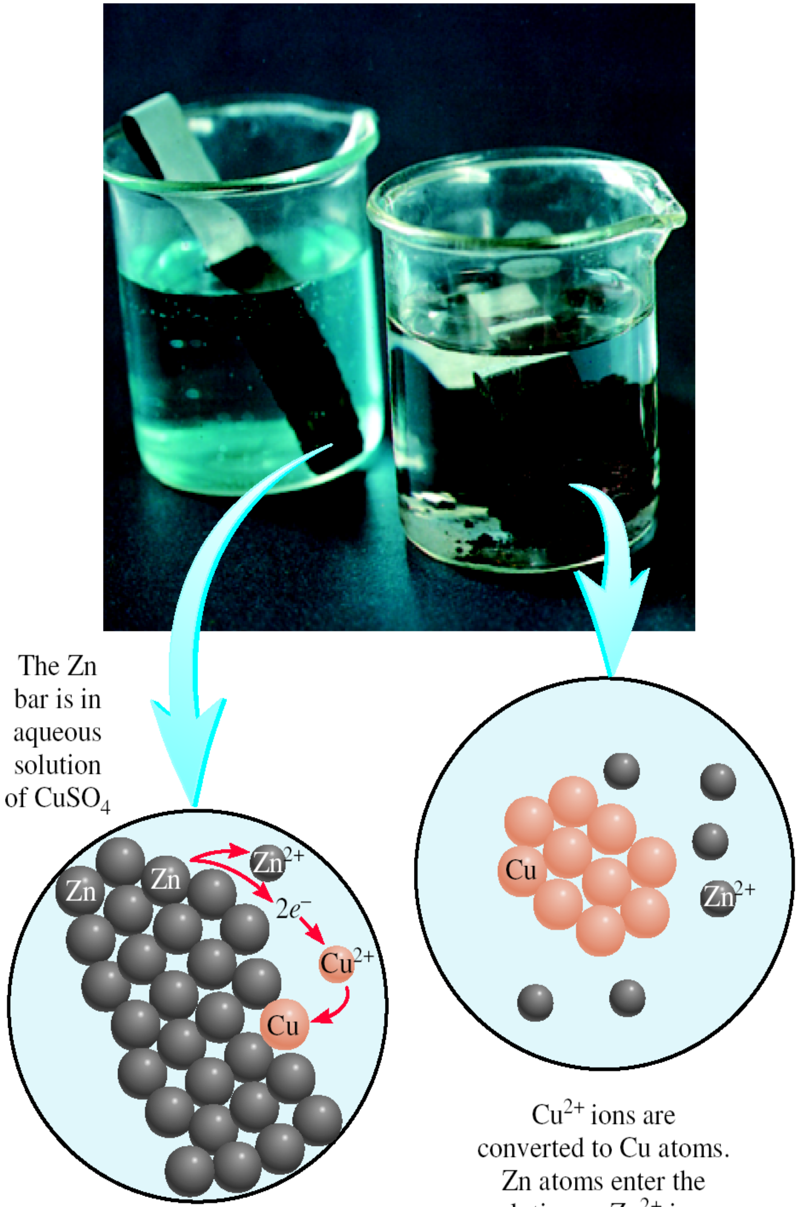
oxidized

reduced

+2
Increased
oxidation No.

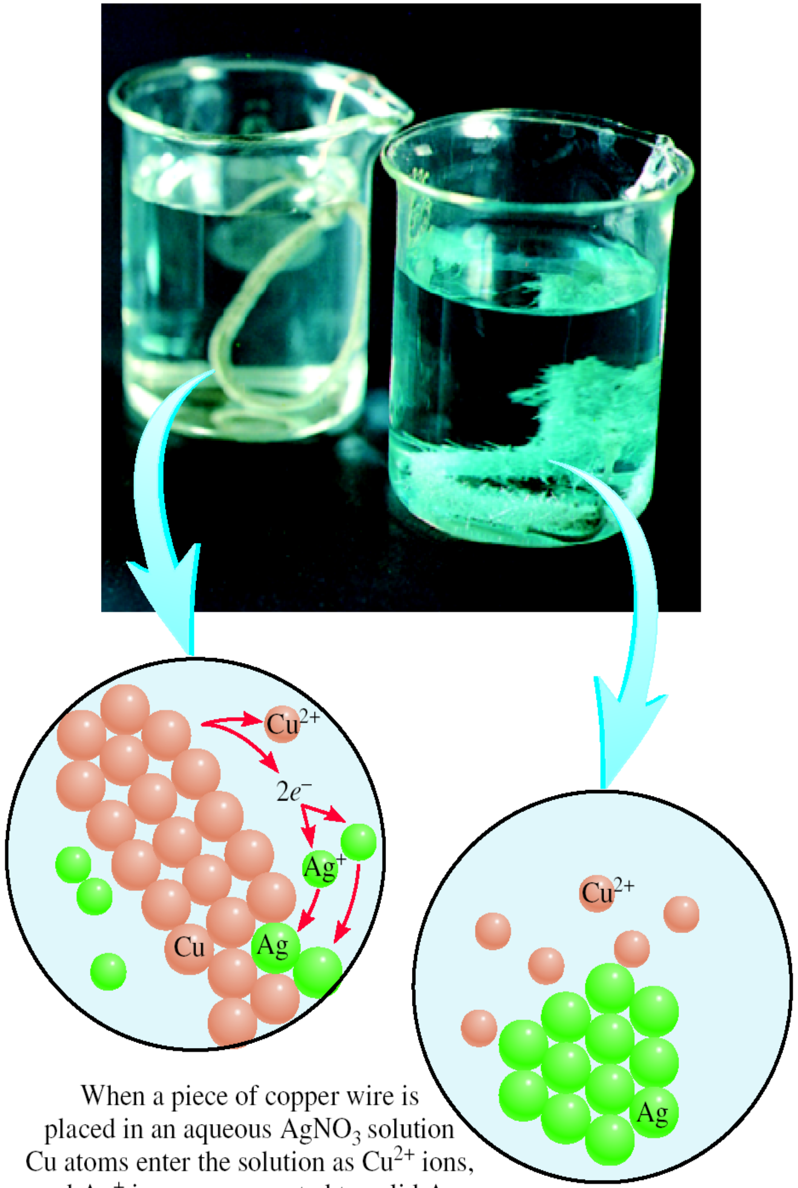
-2
Decreased
oxidation No.

Oxidation-reduction Reactions



The Zn bar is in aqueous solution of $CuSO_4$

Cu^{2+} ions are converted to Cu atoms. Zn atoms enter the solution as Zn^{2+} ions.



When a piece of copper wire is placed in an aqueous $AgNO_3$ solution Cu atoms enter the solution as Cu^{2+} ions, and Ag^+ ions are converted to solid Ag.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$

Oxidation half-reaction

$Zn \longrightarrow Zn^{2+} + 2e^-$

Zn loses 2 e^- to form Zn^{2+}

Zn is oxidized

Zn is the ***reducing agent***

Reduction half-reaction

$Cu^{2+} + 2e^- \longrightarrow Cu$

Cu^{2+} gains 2 e^- to form copper

Cu^{2+} is reduced

Cu^{2+} is the ***oxidizing agent***

The number of e^- lost by the reducing agent =
the number of electrons gained by the oxidising agent.

Watch the animation titled 'Oxidation-Reduction Reaction' on the VLE

Oxidation Numbers

To keep track of electrons in redox reactions

$$H_2(g) \quad + \quad \tfrac{1}{2}\, O_2(g) \quad \rightarrow \quad H_2O(g)$$

Oxidation number (also called **oxidation state**)
is the charge an atom would have in a molecule or ionic compound
if electrons were transferred completely.

↓

Rules to assign oxidation numbers

Rules To Assign Oxidation Number

Must Be Applied In Order!

The charge the atom would have in a molecule (or an ionic compound) if electrons were completely transferred.

1. Free elements (uncombined state) have an oxidation number of zero.

   Na, Be, K, Pb, H_2, O_2, P_4 = 0

2. In monatomic ions, the oxidation number is equal to the charge on the ion.

   Li^+, Li = +1; Fe^{3+}, Fe = +3; O^{2-}, O = -2

3. The oxidation number of hydrogen is +1 *except* when it is bonded to metals in binary compounds. In these cases, its oxidation number is –1.

4. The oxidation number of oxygen is **usually** –2. In H_2O_2 and O_2^{2-} it is –1.

5. Group IA metals are +1, IIA metals are +2 and fluorine is always –1.

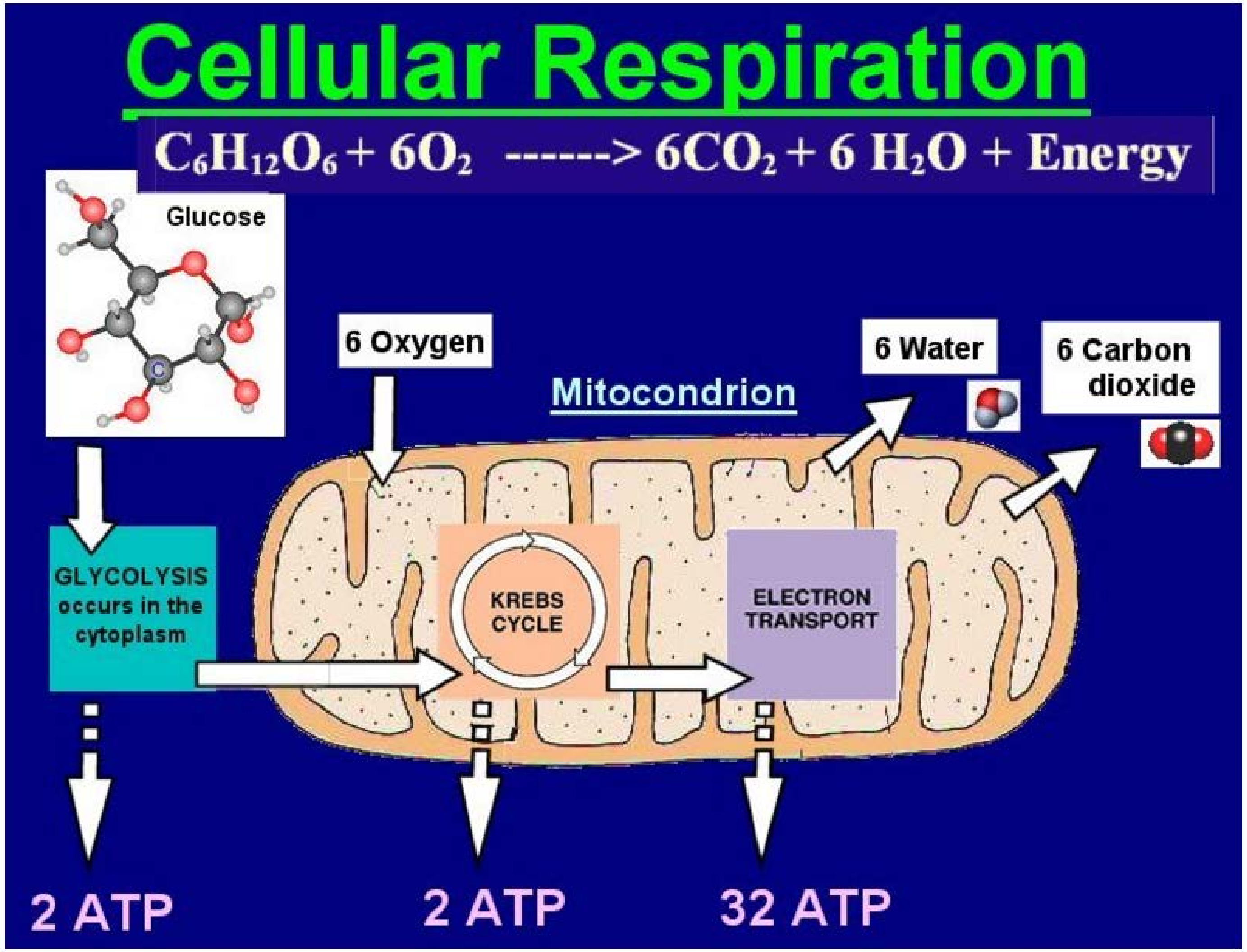
6. The sum of the oxidation numbers of all the atoms in a molecule or ion is equal to the charge on the molecule or ion.

7. Oxidation numbers do not have to be integers. The oxidation number of oxygen in the superoxide ion, O_2^-, is –½.

Some Oxidation States

<u>N</u>	NH_3	NO	NO_2^-	N_2O
	- 3	+ 2	+3	+1
<u>S</u>	H_2S	SO_3^{2-}	SO_2	H_2SO_4
	- 2	+ 4	+ 4	+ 6
<u>Fe</u>	Fe	Fe^{2+}	Fe^{3+}	
	0	+ 2	+ 3	
<u>Mn</u>	MnO_4^-	MnO_2	Mn^{2+}	Mn
	+ 7	+ 4	+ 2	0

The Breakdown Of Glucose In The Body To Produce Energy



The Breakdown Of Glucose In The Body To Produce Energy

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O + \text{energy}$$

$C_6H_{12}O_6$	$6O_2$	$6CO_2$	$6H_2O$
O.N of C: (0)	O.N of O: (0)	O.N of C: (+4)	O.N of O: (-2)

Oxidation

Reduction

Oxidation of glucose ($C_6H_{12}O_6$) to CO_2 and the reduction of oxygen to water

Balancing Redox Reactions in Acids by Half-reaction Method

e.g. $Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$

1. Determine what is oxidised and what is reduced.

$$\underset{(+2)}{Fe^{2+}} + \underset{(+6)}{Cr_2O_7^{2-}} \longrightarrow \underset{(+3)}{Fe^{3+}} + \underset{(+3)}{Cr^{3+}}$$

Remember: Increase in O.N. = oxidation

Reduction in O.N. = reduction

Fe^{2+} : oxidised

$Cr_2O_7^{2-}$: reduced

$$Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

2. Write the two half-reactions and balance all the atoms other than O and H in each half-reaction

Oxidation: $Fe^{2+} \longrightarrow Fe^{3+}$

Reduction: $Cr_2O_7^{2-} \longrightarrow 2\ Cr^{3+}$

3. For reactions in acidic medium, add H_2O to balance O atoms and H^+ to balance H atoms

Oxidation: $Fe^{2+} \longrightarrow Fe^{3+}$

Reduction:

In an acidic medium, we add $7H_2O$ molecules to the right side of the arrow to balance the O atoms

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

To balance the H atoms, we add $14H^+$ on the left side

$$Cr_2O_7^{2-} + 14\,H^+ \longrightarrow 2Cr^{3+} + 7\,H_2O$$

4. Add electrons to each half-reaction to balance the charge

Oxidation :

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

Reduction:

There are now 12 positive charges on the left side and only 6 positive charges on the right. Therefore, we add 6 electrons on the left

$$Cr_2O_7^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O$$

$$\underbrace{-2 \qquad 14+}_{12+} \qquad\qquad \underbrace{6+ \qquad 0}_{6+}$$

5. Must have same number of electrons in each half-reaction

$$Fe^{2+} \longrightarrow Fe^{3+} + e^- \qquad \boxed{X\ 6}$$

$$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \longrightarrow 2\ Cr^{3+} + 7\ H_2O$$

$$6\ Fe^{2+} \longrightarrow 6\ Fe^{3+} + \cancel{6}\ e^-$$

$$Cr_2O_7^{2-} + 14\ H^+ + \cancel{6}\ e^- \longrightarrow 2\ Cr^{3+} + 7\ H_2O$$

6. Add the two half-reactions. The e^- on both sides must cancel

$$6\,Fe^{2+} \longrightarrow 6\,Fe^{3+} + \cancel{6}\,e^-$$

$$Cr_2O_7^{2-} + 14\,H^+ + \cancel{6}\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O$$

$$6\,Fe^{2+} + Cr_2O_7^{2-} + 14\,H^+ \longrightarrow 6\,Fe^{3+} + 2\,Cr^{3+} + 7\,H_2O$$

(Check Atoms & Charges)

Try this one !

$$Fe^{2+} + MnO_4^- \longrightarrow Fe^{3+} + Mn^{2+}$$

(+2) (+7) (+3) (+2)

Oxidation

Reduction

1. Determine what is oxidised and what is reduced.

1. $Fe^{2+} \longrightarrow Fe^{3+}$ Oxidation

2. $MnO_4^- \longrightarrow Mn^{2+}$ Reduction

2. Write the two half-reactions and balance all the atoms <u>other than O and H</u> in each half-reaction

1. $Fe^{2+} \longrightarrow Fe^{3+}$ Oxidation

2. $MnO_4^- \longrightarrow Mn^{2+}$ Reduction

3. For reactions in acidic medium, add H_2O to balance O atoms and H^+ to balance H atoms

1. $Fe^{2+} \longrightarrow Fe^{3+}$

2. $MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$

4. Add electrons to each half-reaction to balance the charge

1. $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

2. $$\underbrace{\underset{-1}{MnO_4^-} + \underset{+8}{8H^+}}_{+7} \longrightarrow \underbrace{\underset{+2}{Mn^{2+}} + \underset{0}{4H_2O}}_{+2}$$

$$MnO_4^- + 8H^+ + 5\,e^- \longrightarrow Mn^{2+} + 4H_2O$$

5. Must have same number of electrons in each half-reaction, and add the two half-reaction up

1. $\boxed{\times 5}$ $5\ Fe^{2+} \longrightarrow 5\ Fe^{3+} + 5\ e^-$

2. $MnO_4^- + 8\ H^+ + 5\ e^- \longrightarrow Mn^{2+} + 4H_2O$

$$5\ Fe^{2+} + MnO_4^- + 8\ H^+ \longrightarrow 5\ Fe^{3+} + Mn^{2+} + 4\ H_2O$$

Bingo!

Balancing Redox Reactions in Bases by Half-reaction Method

Reduction

$$CN^- + \underset{(+7)}{MnO_4^-} \longrightarrow CNO^- + \underset{(+4)}{MnO_2}$$

Oxidation

1. Write the two half-reactions and balance all the atoms <u>other than O and H</u> in each half-reaction

CN^-	$\longrightarrow$	CNO^-	Oxidation
MnO_4^-	$\longrightarrow$	MnO_2	Reduction

2. Add H_2O to balance O atoms and H^+ to balance H atoms

Oxidation:

To balance the O atoms, add one H_2O molecule on the left

To balance the H atoms, add two H+ ions on the right

$$CN^- + H_2O \longrightarrow CNO^- + 2H^+$$

Reduction:

To balance the O atoms, add two H_2O molecule on the right.

To balance the H atoms, add four H+ ions on the left

$$MnO_4^- + 4H^+ \longrightarrow MnO_2 + 2H_2O$$

3. For reactions in basic medium, for every H^+ ion, add an equal number of OH^- ions to both sides of the equation

Oxidation: $CN^- + H_2O + 2OH^- \longrightarrow CNO^- + 2H^+ + 2OH^-$

Reduction: $MnO_4^- + 4H^+ + 4OH^- \longrightarrow MnO_2 + 2H_2O + 4OH^-$

Oxidation: $CN^- + \cancel{H_2O} + 2OH^- \longrightarrow CNO^- + \underbrace{2H^+ + 2OH^-}_{1\ \cancel{2}H_2O}$

Reduction: $MnO_4^- + \underbrace{4H^+ + 4OH^-}_{2\ \cancel{4}H_2O} \longrightarrow MnO_2 + \cancel{2H_2O} + 4OH^-$

4. For reactions in <u>basic medium</u>, for every H^+ ion, add an equal number of OH^- ions to both sides of the equation

Oxidation: $CN^- \ + \ + 2OH^- \longrightarrow CNO^- + H_2O$

Reduction: $MnO_4^- \ + \ 2\,H_2O \longrightarrow MnO_2 \ + \ 4OH^-$

5. Add electrons to each half-reaction to balance the charge

Oxidation:

$$\underbrace{CN^{-} + 2OH^{-}}_{-3} \longrightarrow \underbrace{CNO^{-} + H_2O}_{-1}$$

$$CN^{-} + 2OH^{-} \longrightarrow CNO^{-} + H_2O + 2e^{-}$$

Reduction:

$$\underbrace{MnO_4^{-} + 2H_2O}_{-1} \longrightarrow \underbrace{MnO_2 + 4OH^{-}}_{-4}$$

$$MnO_4 + 2H_2O + 3e^{-} \longrightarrow MnO_2 + 4OH^{-}$$

6. Must have same number of electrons in each half-reaction

$$CN^- + 2OH^- \longrightarrow CNO^- + H_2O + 2e^- \quad \times 3$$

$$MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4OH^- \quad \times 2$$

$$3CN^- + 6OH^- \longrightarrow 3CNO^- + 3H_2O + 6e^-$$

$$2MnO_4^- + 4H_2O + 6e^- \longrightarrow 2MnO_2 + 8OH^-$$

6. Add the two half-reactions

$$3\,CN^- + \cancel{6\,OH^-} \longrightarrow 3\,CNO^- + \cancel{3\,H_2O} + \cancel{6e^-}$$

$$2\,MnO_4^- + \underset{1}{\cancel{4}}\,H_2O + \cancel{6e^-} \longrightarrow 2\,MnO_2 + \underset{2}{\cancel{8}}OH^-$$

$$3\,CN^- + 2\,MnO_4^- + H_2O \longrightarrow 3\,CNO^- + 2\,MnO_2 + 2OH^-$$

BALANCING REDOX EQUATIONS

Reaction in Acidic Solution

1. Determine what is being oxidised and what is being reduced.
2. Write two half-reactions and balance all the atoms other than O and H in each half-reaction
3. Add H_2O to balance O atoms, and H^+ to balance H atoms.
4. Add e^- to each half-reaction to balance the charge.
5. Must have same number of electrons in each half-reaction.
6. Add the two half-reactions. The e^- on both sides must cancel.

Reaction in Basic Solution

1. Determine what is being oxidised and what is being reduced.
2. Write two half-reactions and balance all the atoms other than O and H in each half-reaction
3. Add H_2O to balance O atoms, and H^+ to balance H atoms.
4. For every H+ ion, add an equal number of OH^- ions to both sides of the equation.
5. Add e^- to each half-reaction to balance the charge.
6. Must have same number of electrons in each half-reaction.
7. Add the two half-reactions. The e^- on both sides must cancel.

Write a balanced ionic equation to represent the oxidation of iodide ion (I^-) by permanganate ion (MnO^-_4) in basic solution to yield molecular iodine (I_2) and manganese(IV) oxide (MnO_2).