



Chemistry

Redox Reactions

Grade 12 advanced

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Prepared by

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Student`s name :.....

Class:.....

هذه المذكرة لاتعني عن الكتاب المدرسي وليست بهدف البيع او الربح

تمنيتي لجميع الطلبة والطالبات بالتفوق والنجاح

Redox Reactions

Section 1 : Oxidation and Reduction

Main idea : oxidation and reduction are complementary – as a substance is oxidized , another substance is reduced

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- ✓ the light produced by a glow stick is result of a chemical reaction
- ✓ when you snap the glass capsule inside the plastic case
- ✓ two chemicals are mixed and electron transfer occurs
- ✓ as the electrons are transferred , chemical energy is converted into light energy

items	Definition
Oxidation – reduction reaction Or (redox reaction)	A reaction in which electrons are transferred from one atom to another
Oxidation (LEO)	Is defined as the loss of electrons from atoms of a substance.
Reduction (GER)	Is the gain of electrons by atoms of a substance.
Oxidizing agent	The substance that oxidizes another substance by accepting its electrons
Reducing agent	The substance that reduces another substance by losing its electrons
The oxidation number of an atom	Is the number of electrons lost or gained by the atom when it forms ions

➤ Electron transfer and redox reactions :

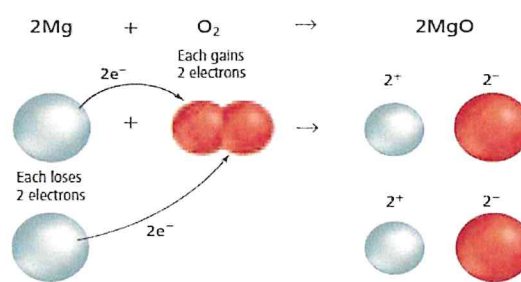
- ✓ A defining characteristic of combustion and single replacement reactions is that they always involve the transfer of electrons from one atom to another

Example :

combustion reaction of **magnesium in air**, which involves the transfer of electrons

complete chemical equation : $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$

net ionic equation : $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{Mg}^{2+}\text{(aq)} + 2\text{O}^{2-}\text{(aq)}$



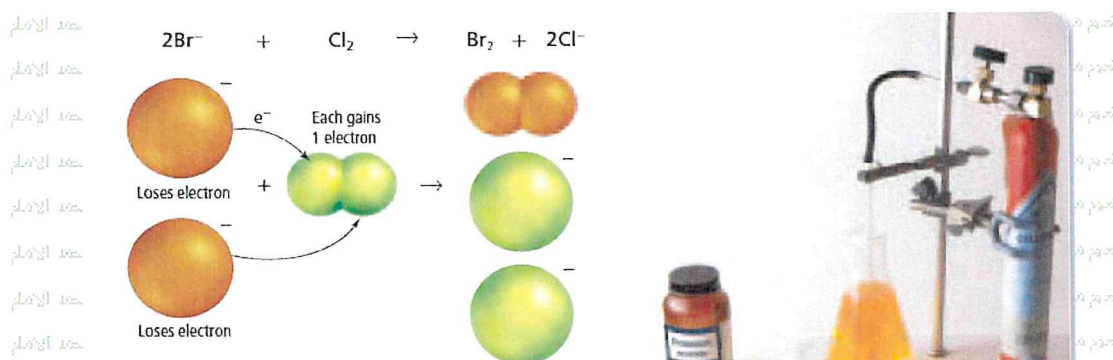
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✓ **The single – replacement reaction** in which chlorine in an aqueous solution reacts with bromide ions from an aqueous solution of potassium bromide ,
complete chemical equation : $2\text{KBr}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{KCl}(\text{aq}) + \text{Br}_2(\text{g})$
net ionic equation : $2\text{K}^+(\text{aq}) + 2\text{Br}^-(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{K}^+(\text{aq}) + 2\text{Cl}^-(\text{aq}) + \text{Br}_2(\text{g})$
Spectators ions: $2\text{K}^+(\text{aq})$

✚ note that Cl_2 takes electrons from bromide ions to become chloride ions Cl^- and two bromide ions lose electrons the two bromine atoms form a covalent bond with each other to produce bromine Br_2 molecules
 ✚ the formation of the covalent bond by sharing of electrons is also an oxidation – reduction reactions .



✓ **Double – replacement reactions are not redox reactions (why ?)**
 Because it has not transfer electrons between atoms

✓ As do many synthesis and decomposition reactions are redox reactions

Example : synthesis reaction between chlorine Cl_2 and sodium Na to form sodium chloride NaCl
complete chemical equation : $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NaCl}(\text{s})$
net ionic equation : $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{Na}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$

➤ Chemical facts :

- ✓ The glow of an activated light stick can be made brighter by warming it , although the glow not last as long
- ✓ Light generated by redox reactions doesn't generally result in the formation of heat
- ✓ About 90% of marine life uses some form of bioluminescence generating light through redox reactions

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➤ oxidation and reduction :

Example:

✓ **oxidation** : referred to a substance combined with oxygen. Or loss of electrons from atoms of a substance



sodium is oxidized because it loses an electrons

✓ **reduction** : is the gain of electrons by a toms of asubstance



Chlorine is reduced because it gain an electrons

✓ oxidation and reduction are complementary processes that mean : oxidation can not occur unless reduction also occurs

✓ **LEO** : Loss of Electrons is Oxidation

✓ **GER** : Gain of Electrons is Reduction

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➤ Change in oxidation number :

✓ Oxidation numbers are tools that scientists use in written chemical equations to help them keep track of the movment of electrons in a redox reaction

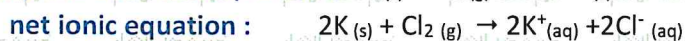
✓ Group (1) element : that tends to lose 1 electron in reaction because of its low electronegativity is assigned an oxidation number of +1

✓ Group (17) element that tends to gain one electron in reactions because of its high electronegativity is assigned an oxidation number of -1

✓ Oxidation numbers are written with the positive or negative sign before the numbers (+3 , +2) whereas ionic charge is written with the sign after the number (3+ , 2+)

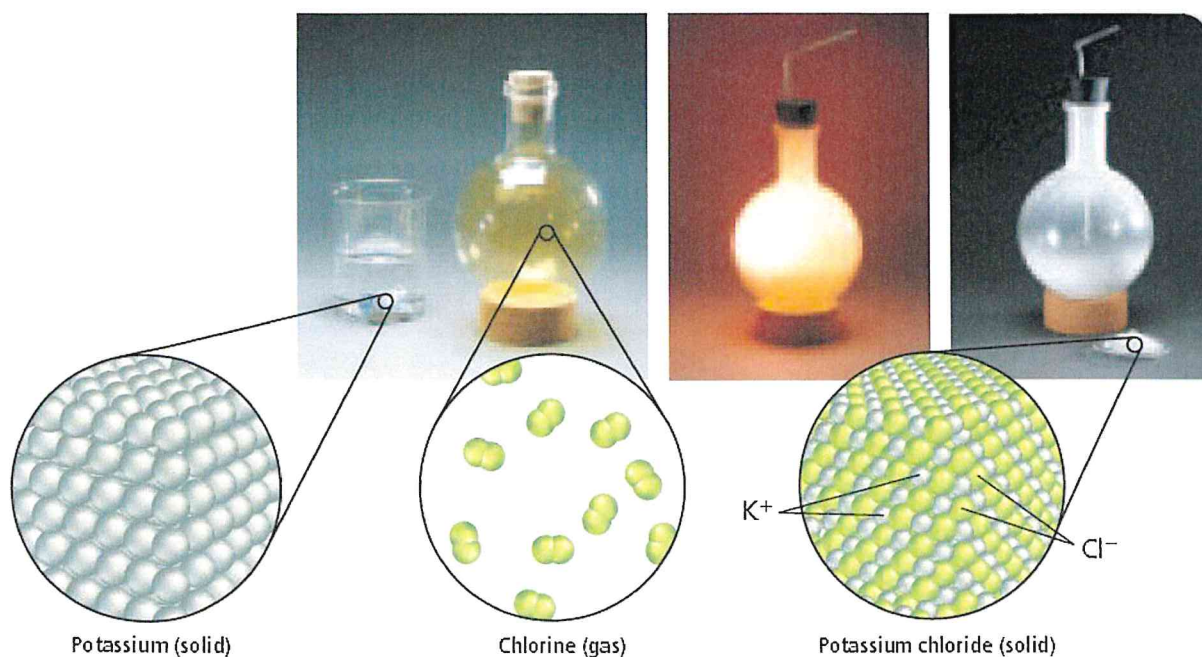
Example:

The equation for the reaction of potassium metal with chloride vapor is as follows :



✓ potassium atoms are oxidized from 0 to +1 state because each atom loses an electron

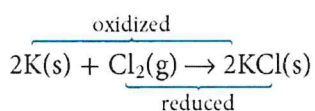
✓ chlorine atoms are reduced from 0 to the -1state because each atom gains an electron



	Oxidation	Reduction
Definition	Is the loss of electrons from atoms of a substance.	Is the gain of electrons by atoms of a substance.
The numerical value of oxidation number	Increase	Decrease
Example	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$
(oxidizing / reducing) Agent	Na : is the reducing agent	Cl_2 : is the oxidizing agent

➤ Oxidizing and reducing agent :

- ✓ The reducing agent is oxidized because it loses electrons so it supplies electron to substance being reduced (gaining electrons)



Oxidizing agent: Cl_2

Reducing agent: K

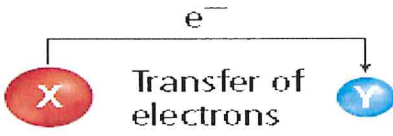
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➤ **A common application of redox chemistry :**

- 1- Remove **tarnish** from metal objects
- 2- **Whiten clothes** : by using chlorine bleach (sodium hypochlorite NaOCl as **oxidizing agent**)and add it to your laundry . it oxidizes dyes , stain , and other materials that discolor clothes

Process	
Oxidation • A reactant loses an electron. • Reducing agent is oxidized. • Oxidation number increases.	• X loses an electron. • X is the reducing agent and becomes oxidized. • The oxidation number of X increases.
Reduction • Other reactant gains an electron. • Oxidizing agent is reduced. • Oxidation number decreases.	• Y gains an electron. • Y is the oxidizing agent and becomes reduced. • The oxidation number of Y decreases.

Notes :

Rust : when moist air comes in contact with iron , the iron will oxidizes iron oxide (Fe_2O_3) pure iron is uncommon in nature . steel a mixture that contain iron is a commonly used of iron several protective methods such as :

✓ Plating

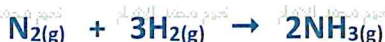
✓ Painting

✓ Applying an enamel or plastic coating , can inhibit the production of iron oxide .

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➤ Redox and electronegativity :

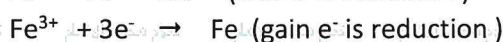
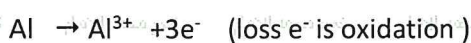
- ✓ The chemistry of oxidation – reduction reactions is not limited to atoms of an element changing to ions or the reverse
- ✓ Some redox reactions involve changes in molecular substances or **polyatomic ions** in which atoms are covalently bonded to other atoms
- ✓ When atom lost electrons and was oxidized while the other atom gained electrons and was reduced
- ✓ **Electronegativity increases left to right across a period** and generally decreases down a group
- ✓ **Which atom is more electronegativity** (attract electrons more strongly) like (Group 17 and oxygen in Group 16) are **strong oxidizing agents**
- ✓ Which atom is less electronegativity (Group 1 and 2) are strong reducing agents



- ✓ **Nitrogen N_2 :** is the **oxidizing agent** , **H_2 :** is **reducing agent**
- ✓ Nitrogen is more electronegativity so it will reduced by gaining electrons from the other atom (hydrogen)
- ✓ Conversely, the less electronegative atom (hydrogen) is treated as if it had been oxidized by losing electrons to other atom (nitrogen)

Example:

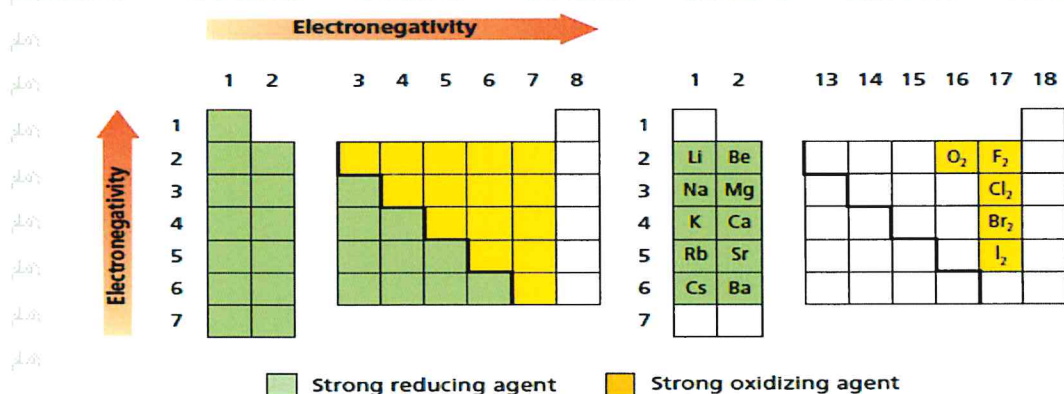
Identify oxidation – reduction reactions



Aluminum is oxidized and is therefore the reducing agent

Iron is reduced and therefore the oxidizing agent

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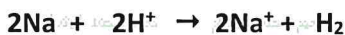
➤ Practice problems :

1- Identify each of the following changes as either oxidation or reduction .

Recall that e^- is the symbol for an electron .



2-identify what is oxidized and what is reduced in the following :



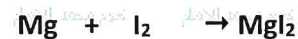
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3-identify the oxidizing agent and the reducing agent in the following equation .

explain your answer :



4-challenge : identify the oxidizing agent and the reducing agent in each reaction :



➤ Determining oxidation numbers :

- ✓ Iron has different oxidation numbers indicate by the different colors which mineral is also present

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Rules for determining oxidation numbers			
Rule		Example	n element
1	The oxidation number of an uncombined atom (element is Zero)	$\text{Cl}_2, \text{F}_2, \text{Br}_2, \text{O}_2, \text{H}_2, \text{Na}$	0
2	The oxidation number of a monoatomic ion is equal to charge of the ion	Ca^{2+}	+2
		Br^-	-1
3	The oxidation number of the more electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion	N in NH_3	-3
		O in NO	-2
4	The oxidation number of the most electronegative element, F is always -1 when it is bonded to another element	F in LiF	-1
5	✓ The oxidation number of oxygen in compounds is always -2 ✓ The oxidation number of oxygen in hydrogen peroxide H_2O_2 is -1 ✓ The oxidation number of oxygen in OF_2 is +1	O in NO_2	-2
		O in H_2O_2	-1
		O in OF_2	+2
6	✓ The oxidation number of hydrogen in most of its compounds is +1 ✓ The oxidation number of hydrogen in hydrides is -1	H in H_2O	+1
		H in CaH_2	-1
7	The sum of oxidation numbers of the atoms in a polyatomic ions is equal to the charge of the ions	SO_3^{2-}	$(+4)+3(-2) = -2$
8	The oxidation number of Group 1 is +1	K in KCl	+1
9	The oxidation number of Group 2 is +2	Ca in CaBr_2	+2
10	The oxidation number of Group 13 is +3	Al in AlCl_3	+3

Notes :

- 1- The oxidation number of Group 1, 2, 13 are positive and equal to their number of valence electrons
- 2- The sum of the oxidation numbers in a neutral compound is zero

Example:

Calculate the oxidation number of each element in potassium chlorate KClO_3 and a sulfite ion (SO_3^{2-})

Solution :

$$\text{KClO}_3 : +1 + X + (-2) = 0$$

$$\text{Cl} = +5$$

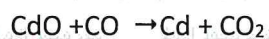
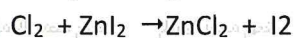
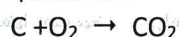
$$(\text{SO}_3^{2-}) = X + 3(-2) = -2$$

$$X = +4$$

Determine the oxidation number of boldface element in the formulas for compounds :

Oxidation number	ion	Oxidation number	compound
	SO_3^{2-}		KClO_3
	NH_4^+		NaClO_4
	AsO_4^{3-}		HNO_2
	CrO_4^{2-}		AlPO_4
	$\text{S}_2\text{O}_4^{2-}$		NH_3
	ClO_3		KCN
	NO_2^-		N_2H_4
	IO_4^-		HNO_3
	$\text{B}_4\text{O}_7^{2-}$		Sb_2O_5
	MnO_4^-		Ca_3N_2
	NH_2^-		CuWO_4

Challenge : determine the net change of oxidation number of each of the element in these redox equations :



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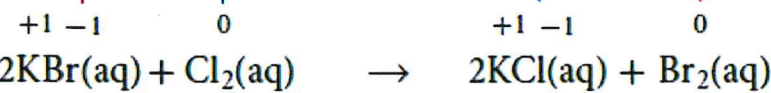
➤ Oxidation numbers in redox reactions :

- ✓ When an atom is **oxidized** , its oxidation number increases
- ✓ When an atom is **reduced** , its oxidation number decreases
- ✓ A **spectator ion** : that ion takes no part in the reaction

Example:

change: -1 reduction

change: +1 oxidation



no change in oxidation number

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أهدي هذا العمل المتواضع الى موجهي الكيمياء

أ. عادل الجيار - أ. محمد ابو عبيد

داعيا المولي عز وجل لهما بالتوفيق والسداد لما قدموه لنا من علم وخلق

أ. نعيم الامام عقل

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Section2: balancing redox equations

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Main idea: redox equations are balanced when the total increase in oxidation numbers equals the total decrease in oxidation numbers of the atoms involved in the reaction.

CHEM 4 YOU :

When fatty substances in foods spoil they are referred to as rancid. Large molecules are broken down through redox reactions that result in foul-smelling products be balanced using the same rules for simpler equations.

The oxidation – number method :

Is the balancing technique based on the principle of conservation of charge (The number of electrons transferred from atoms must equal the number of electrons accepted by other atoms).

Oxidation number procedures :

- 1- Assign oxidation numbers to all atoms in the equation
- 2- Identify the atoms that are oxidized and the atoms that are reduced
- 3- Determine the change in oxidation number for the atoms that are oxidized and for atoms that are reduced
- 4- Make the change in oxidation numbers equal in magnitude by adjusting coefficients in the equation .
- 5- If necessary use the conventional method to balance the remainder of the equation

Example:

study the following unbalanced equation for the reaction that occurs when copper metal placed in concentrated nitric acid

Cu oxidized

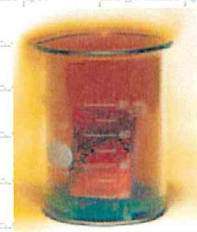
N reduced

H,O not change

N in NO_3^- not change



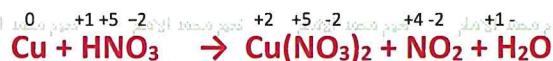
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The brown gas that is produced is NO_2 from the reduction of nitrate ions (NO_3^-) and the blue solution is the result of oxidation of copper (Cu) to copper (II) ions (Cu^{2+})



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Make the change in oxidation numbers equal in magnitude by adjusting coefficients



Use the conventional method to balance the reminder of equation.

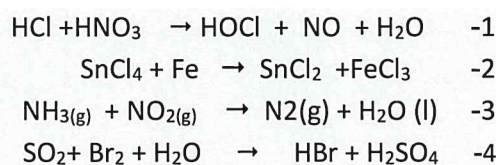


Notes :

- ✓ This reaction can be difficult to balance because elements might appear more than once on each side of the equation
- ✓ The number of electrons transferred from atoms must equal the number of electrons accepted by other atoms
- ✓ Therefore , the total increase in oxidation numbers (oxidation) must equal the total decrease in oxidation numbers (reduction) of atoms involved in the reaction

Practice problems :

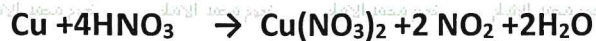
Use the oxidation- number method to balance these redox equations :



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➤ Balancing net ionic redox equations :

- ✓ Sometimes chemists prefer to express redox reactions in the simplest possible terms as an equation showing only the oxidation and reduction processes



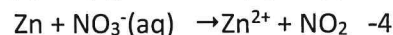
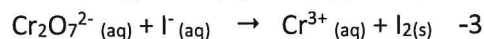
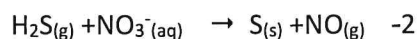
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Notes :

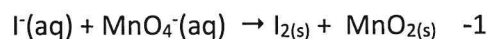
- ✓ in this case , the hydrogen ion and the water molecules are eliminated because neither is oxidized nor reduced
- ✓ in acidic solution : hydrogen ion (H^+) and water molecules are abundant and free to participate in redox reactions as either reactants and products
- ✓ some redox reactions can occur only **in basic reduction**
- ✓ when you balance equations for these reactions you can **added (OH^-) and water** molecules to either side of the equation

Practice problems :

Use the oxidation – number method to balance the following net ionic redox equations :



Balance the following equation in basic solution :



Look at the following reaction and answer the questions :



- 1- Use the oxidation number method to balance that equation ?
- 2- Which elements are reduced and which are oxidized ?

➤ Bioluminescence :



- ✓ is the conversion of potential energy in chemical bonds into light during a redox reaction
- ✓ **some organisms emit light like** : many deep-sea fishes and fireflies have in common the bacterium , *xenorhabdus luminescens*
- ✓ depending on : the species , bioluminescence is produced by different chemicals and by different means
- ✓ **in the fireflies** : light results from the oxidation of the molecule luciferin.
- ✓ Scientists are still unraveling the mystery of bioluminescence
- ✓ **Some luminescent organisms emit light constantly** , whereas other emit light when they are disturbed
- ✓ **Deep-sea fishes** and some jellyfish appear to be able to control the light they emit and one species of mushroom is known to emit light of two different colors
- ✓ **Zoologists** have also determined that some light-emitting organisms do not produce light themselves , they produce light by harboring bioluminescent bacteria .
- ✓ Organisms appear to use bioluminescence for different purposes some purposes might include attracting a mate and defence against prey in the ocean depths , bioluminescence probably aids vision and recognition

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➤ Balancing redox equations using half-reactions:

- ✓ **A species** is any kind of chemical unit involved in a process
- ✓ **A half-reaction** : is one of the two parts of a redox reaction

Example:



There are 4 species

Oxidation – reduction reactions occur whenever a species that can give up electrons (reducing agent) comes in contact with another species that can accept them (oxidizing agent)

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Example:

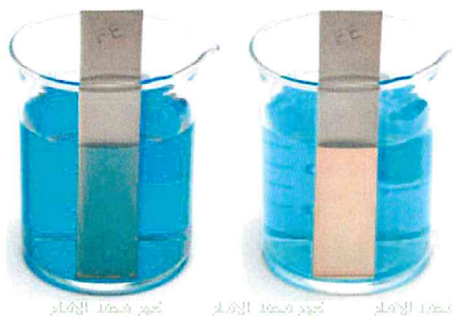
iron can reduce many species that are oxidizing agents (avariety of reduction half reactions that involve the oxidation of Fe to Fe³⁺)

Redox reactions that oxidize iron		
Reduction half- reaction	Oxidation half-reaction	Overall reaction (unbalanced)
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$		$\text{Fe} + \text{Cl}_2 \rightarrow \text{FeCl}_3$
$\text{O}_2 + 4\text{e}^- \rightarrow 2\text{O}^{2-}$		$\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{Fe} + \text{F}_2 \rightarrow \text{FeF}_3$
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$		$\text{Fe} + \text{HBr} \rightarrow \text{H}_2 + \text{FeBr}_3$
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$		$\text{Fe} + \text{AgNO}_3 \rightarrow \text{Ag} + \text{Fe(NO}_3)_3$
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	اعداد أ.نعيم عقل	$\text{Fe(s)} + \text{CuSO}_4(\text{aq}) \rightarrow \text{Cu(s)} + \text{Fe}_2(\text{SO}_4)_3(\text{aq})$ When you put an iron nail into a solution of copper (II) sulfate Solid copper metal is deposited on the iron

➤ **The half – reaction method:**

- 1- Write the net ionic equation for the reaction , omitting spectator ions
- 2- Write the oxidation and reduction half reactions for the net ionic equation
- 3- Balance the atoms and charge in each half- reaction
- 4- Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction
- 5- Added the balanced half- reactions and return spectator ions

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Example: balance a redox equation by using half-reaction :



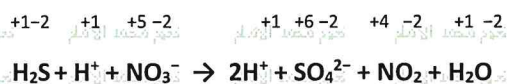
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Solution :

1- write the net ionic equation for the reaction



2- eliminate coefficient , spectator ions and state symbols :

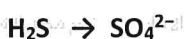
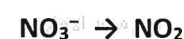


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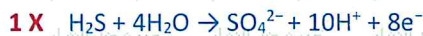
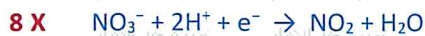


3- write the oxidation and reduction half-reactions for the net ionic equation

including oxidation numbers.:



4- balance the atoms and charges in the half reactions:



Addition two half equations



Add the balanced half- reactions and simply by canceling or reducing like terms on both sides of the equation :



5-In acidic solution , H_2O molecules are available in abundance and can be used to balance oxygen atoms in the half-reactions

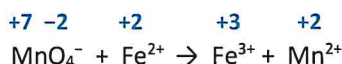
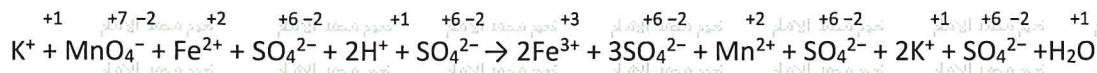
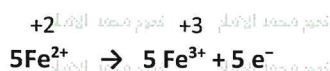
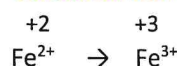
6-In acidic solution , balance hydrogen atom by added H^+ to lower side of hydrogen atom

7-In acidic solution , balance charge by added electrons to higher positive charge side

8- balance the number of electrons in two sides (conservation of charge)

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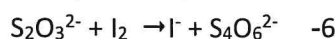
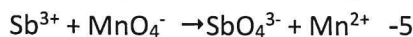
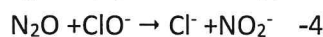
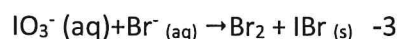
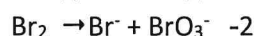
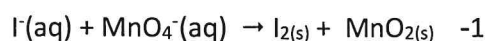
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Example:**balance a redox equation by using half-reactions :****oxidation half- reaction****Balance the following chemical equation :**

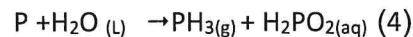
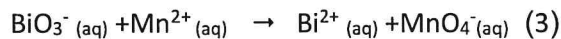
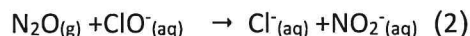
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0508194296**Application 1:**

Balance the following chemical reactions by half reactions in basic medium :

**Application 2:**

Balance the following chemical reaction by half reactions in acidic medium :



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