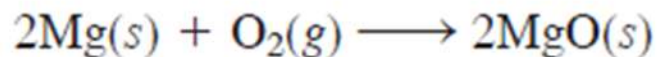


# Chapter 10

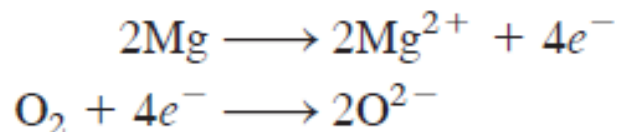
**Redox reactions in aqueous solutions**

# Oxidation-Reduction Reactions

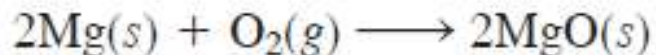
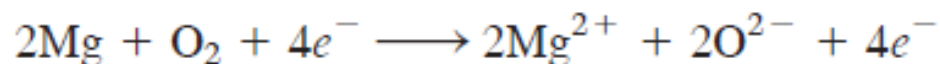
- The class of reactions called oxidation-reduction, or redox, reactions are considered electron transfer reactions.
- Consider the formation of magnesium oxide (MgO) from magnesium and oxygen.



- In this reaction, two Mg atoms give up or transfer four electrons to two O atoms (in O<sub>2</sub>). Each of these steps is called a half-reaction.

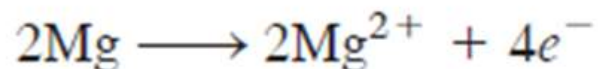


- The sum of the half-reactions gives the overall reaction:



# Oxidation-Reduction Reactions

- An **oxidation reaction** is a half-reaction that involves loss of electrons. In the formation of magnesium oxide, magnesium is oxidized. It is said to act as a **reducing agent or reductant** because it donates electrons to oxygen and causes oxygen to be reduced.



- A **reduction reaction** is a half-reaction that involves gain of electrons. Oxygen is reduced and acts as an **oxidizing agent or oxidant** because it accepts electrons from magnesium, causing magnesium to be oxidized.



# Oxidation Number

- **An atom's oxidation number, also called oxidation state**, signifies the number of charges the atom would have in a molecule (or an ionic compound) if electrons were transferred completely.
- **We use the following rules to assign oxidation numbers:**
  - 1. In free elements** (that is, in the uncombined state), each atom has an oxidation number of zero. Thus, each atom in  $\text{H}_2$ ,  $\text{Br}_2$ , Na, Be, K,  $\text{O}_2$ , and  $\text{P}_4$  has the same oxidation number: zero.
  - 2. For ions composed of only one atom** (that is, monatomic ions), the oxidation number is equal to the charge on the ion. Thus,  $\text{Li}^{+1}$  ion has an oxidation number of +1;  $\text{Ba}^{+2}$  ion, +2;  $\text{Fe}^{+3}$  ion, +3; I ion, -1;  $\text{O}^{-2}$  ion, -2; and so on.
  - 3- All alkali metals have an oxidation number of +1 and all alkaline earth metals have an oxidation number of +2 in their compounds. Aluminum has an oxidation number of +3 in all its compounds.**
  - 4. The oxidation number of oxygen** in most compounds (for example,  $\text{MgO}$  and  $\text{H}_2\text{O}$ ) is -2, but in peroxide ion ( $\text{O}_2^{-2}$ ) as in  $\text{H}_2\text{O}_2$  it is -1.

**5- The oxidation number of hydrogen** is +1, except when it is bonded to metals in binary compounds. In these cases (for example, LiH, NaH, CaH<sub>2</sub>), its oxidation number is -1.

**6. Fluorine has an oxidation number** of -1 in all its compounds. Other halogens (Cl, Br, and I) have negative oxidation numbers when they occur as halide ions in their compounds (as HCl). When combined with O as (ClO<sub>4</sub><sup>-</sup>) they have positive oxidation numbers.

**7- In a neutral molecule**, the sum of the oxidation numbers of all the atoms must be zero. **In a polyatomic ion**, the sum of oxidation numbers of all the elements in the ion must be equal to the net charge of the ion.

**8- Oxidation numbers do not have to be integers.** For example, the oxidation number of O in the superoxide ion, O<sub>2</sub><sup>-</sup>, is -1/2 (as in KO<sub>2</sub>).

Give the oxidation numbers of the underlined atoms in the following molecules and ions: (a)  $\text{Mg}_3\underline{\text{N}}_2$ , (b)  $\text{Cs}\underline{\text{O}}_2$ , (c)  $\text{Ca}\underline{\text{C}}_2$ , (d)  $\underline{\text{C}}\text{O}_3^{2-}$ , (e)  $\underline{\text{C}}_2\underline{\text{O}}_4^{2-}$ , (f)  $\text{Zn}\underline{\text{O}}_2^{2-}$ , (g)  $\text{Na}\underline{\text{B}}\text{H}_4$ , (h)  $\underline{\text{W}}\text{O}_4^{2-}$ .

• تناقش في المحاضرة

# Balancing Redox Equations

- **There are two methods for balancing redox reactions:**
  - 1- Oxidation number method.
  - 2- Half reaction method.
- **We consider the balancing of equations for redox reactions that occur in two situations:**
  - 1- When the redox equation written as molecular equation.
  - 2- When the redox equation written as a net ionic equation.

## Oxidation number method for molecular equation

- **Step 1:** assign oxidation numbers for all atoms.
- **Step 2:** note which atom appear to lose and which atom appear to gain electrons and determine how many electrons are lost or gained for each atom.
- **Step 3:** determine the total loss or gain of electrons per formula unit.
- **Step 4:** make the total gain of electrons equals the total loss of electrons by multiplying a coefficient on the left side of equation.
- **Step 5:** Balance the atoms that gain or lose electrons then balance other atoms except O and H. Then, balance O atoms then hydrogen atoms.





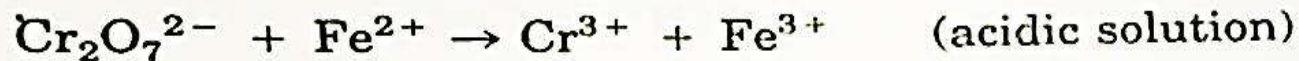


# Oxidation number method for net ionic equation

- **Steps 1-4: as former**
- **Step 5 :** Balance the atoms that gain or lose electrons.
- **Step 6:** Balance other atoms except O and H.
- **Step 7: Balance the charges:**
  - A- If the reaction takes place in **acidic** medium, balance the hydrogen atoms by add  $H^+$  in the site deficient in positive charge.
  - B- If the reaction takes place in **basic** medium, balance the charge by adding  $OH^-$  in the site deficient in negative charge.
- **Step 8:** add water molecules to balance oxygen atoms. So, if your balance is correct, hydrogen atom would be already balanced.

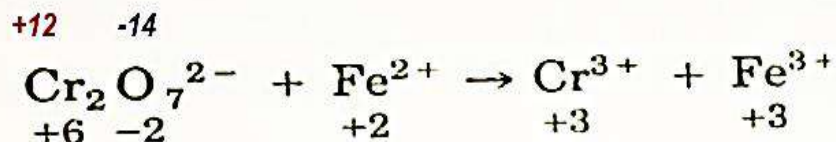
## Oxidation number method for net ionic equation in acidic medium

Complete and balance the following equation for a reaction which takes in acidic solution:

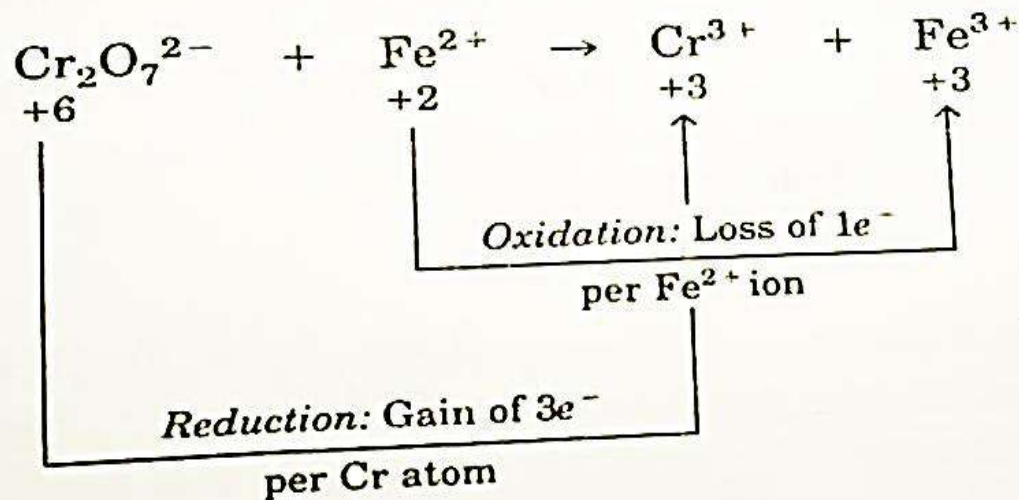


### Solution

Step 1:



Step 2:





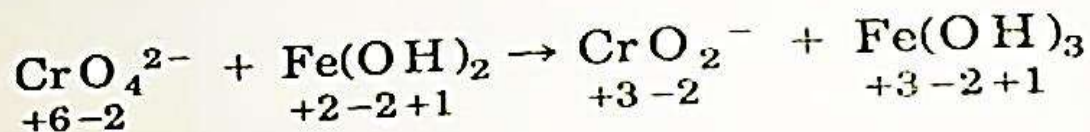
# Oxidation number method for net ionic equation in basic medium

Complete and balance the following equation for a reaction in basic solution:

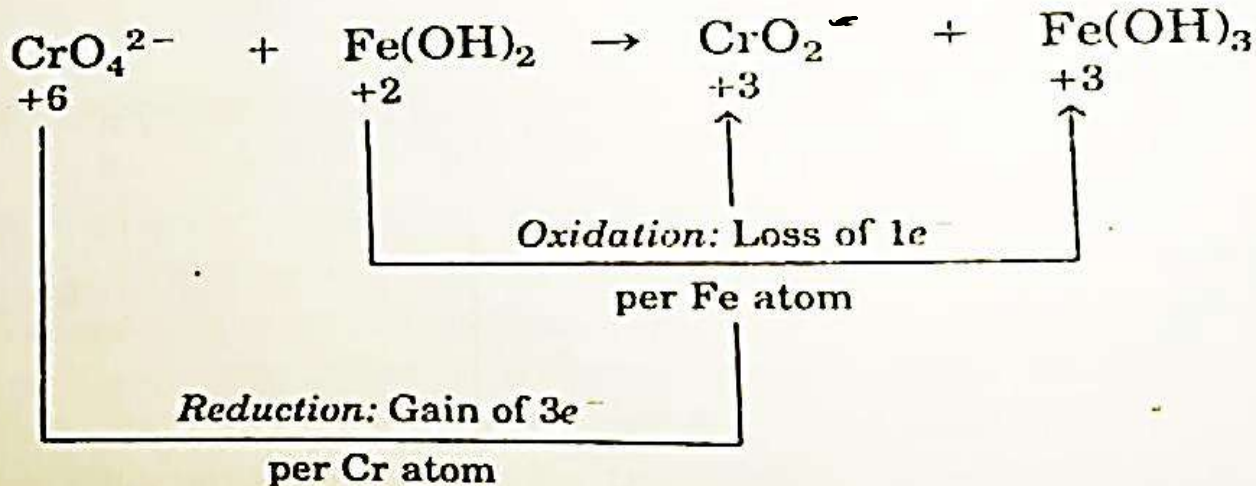
$$\overset{+6}{\text{Cr}}\overset{-2}{\text{O}_4}^{2-} + \overset{+2}{\text{Fe}}\overset{-2}{\text{(OH)}}_2 \rightarrow \overset{+3}{\text{Cr}}\overset{-2}{\text{O}_2}^{-} + \overset{+3}{\text{Fe}}\overset{-2}{\text{(OH)}}_3 \quad (\text{basic solution})$$

**Solution**

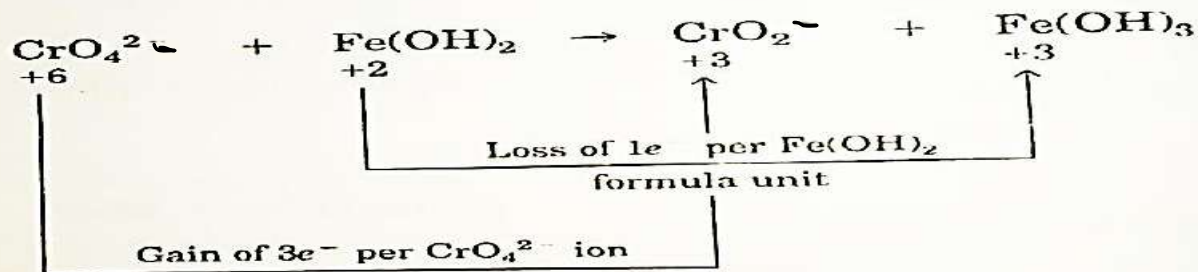
*Step 1:*



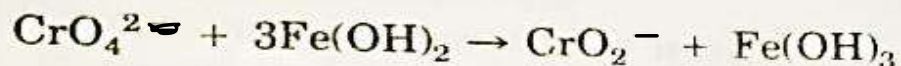
*Step 2:*



Step 3:



Step 4:



Step 5:



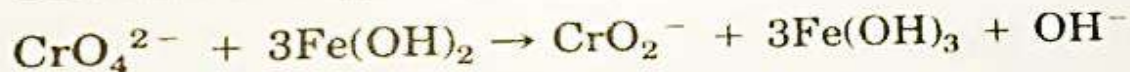
Step 6: Done!

Step 7:

$$\text{Total charge on left} = -2 + 3(0) = -2$$

$$\text{Total charge on right} = -1 + 3(0) = -1$$

$$\text{Additional (negative) charge needed on right} = -1$$



Step 8:

