Redox Reaction

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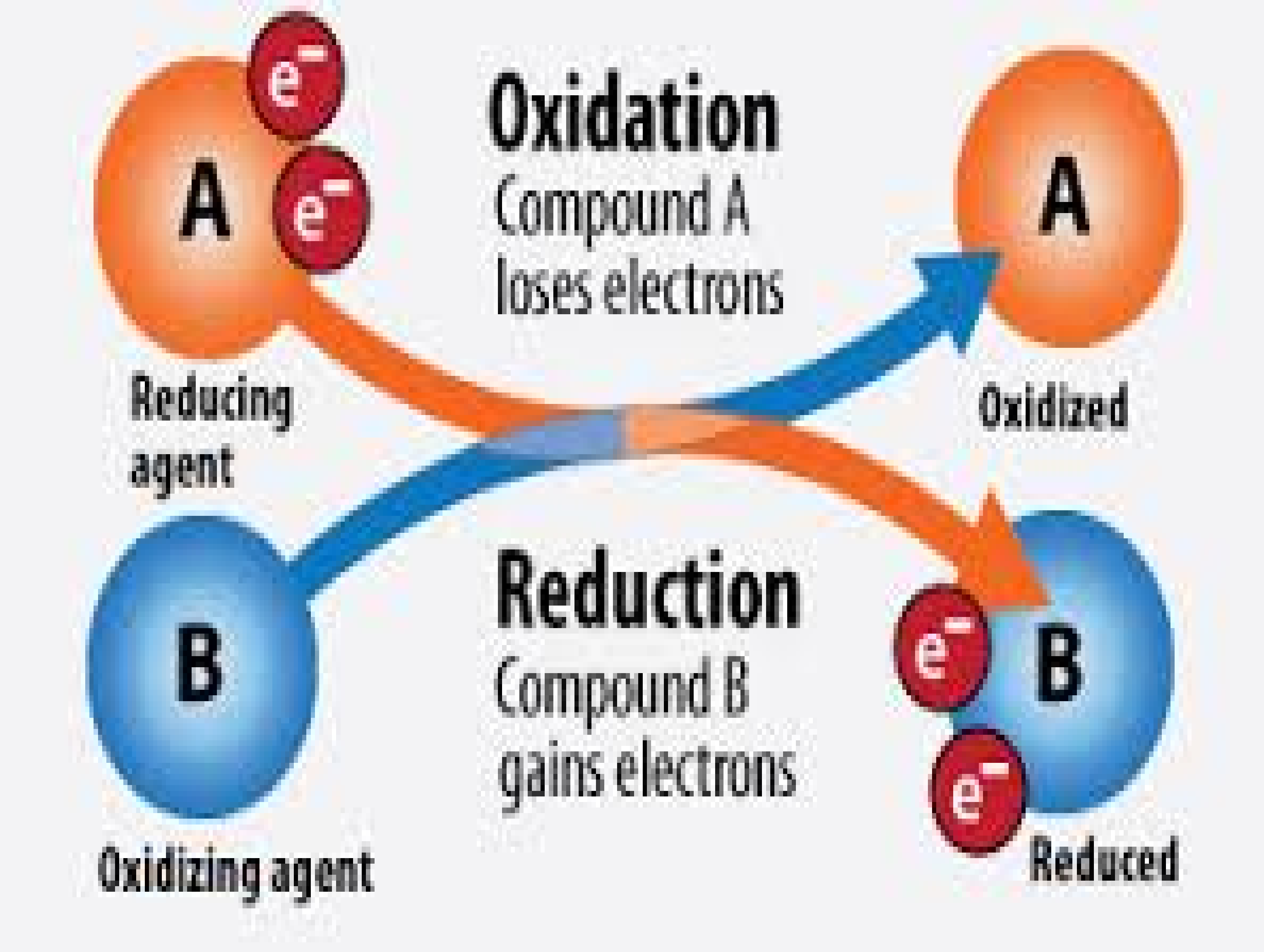

e−
e−
A
Oxidation
Compound A
loses electrons
A
Reducing
agent
Oxidized
B
Reduction
Compound B
gains electrons
e−
B
e−
Oxidizing agent
Reduced

Redox reaction: A redox reaction is a reaction that contains both an oxidation process and a reduction process. Actually, this pretty much describes any electrochemical process, because you can't reduce something (i.e. give it electrons) unless something has given these electrons to it by being oxidized. As a result, in redox reactions, the thing that gets reduced is called an "oxidizing agent" because it takes electrons from something else, and the thing that gets oxidized is called a "reducing agent" because it give electrons to something else.

ELECTRODE POTENTIAL

A metal (M) consists of metal ions (M^{n+}) with valence electrons. When the metal (M) is placed in a solution of its own salt, any one of the following reactions will occur.

(a) Positive metal ions may pass into the solution.

$M \longrightarrow M^{n+} + ne^-$ (oxidation)

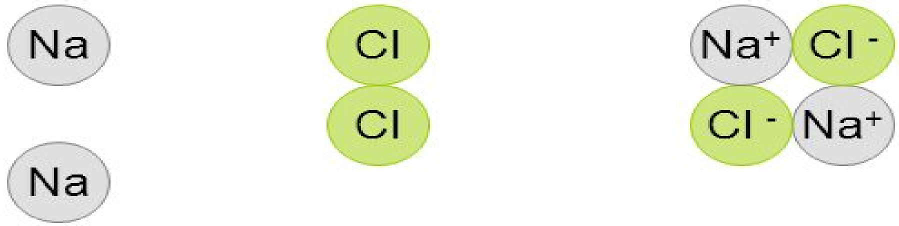
(b) Positive metal ions from the solution may deposit over the metal.

$M^{n+} + ne^- \longrightarrow M$ (reduction)

Synthesis Reaction

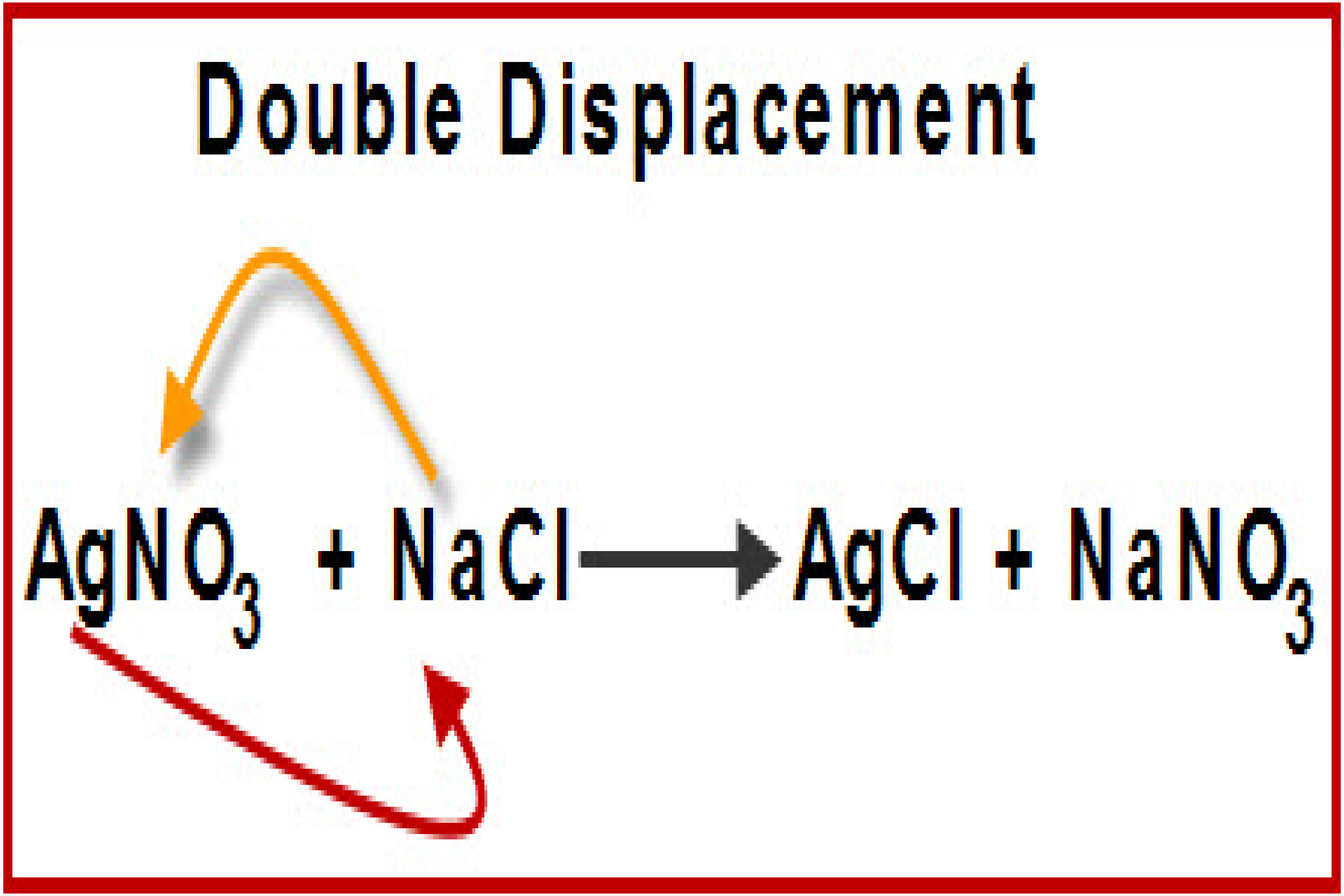
Direct combination reaction (Synthesis)

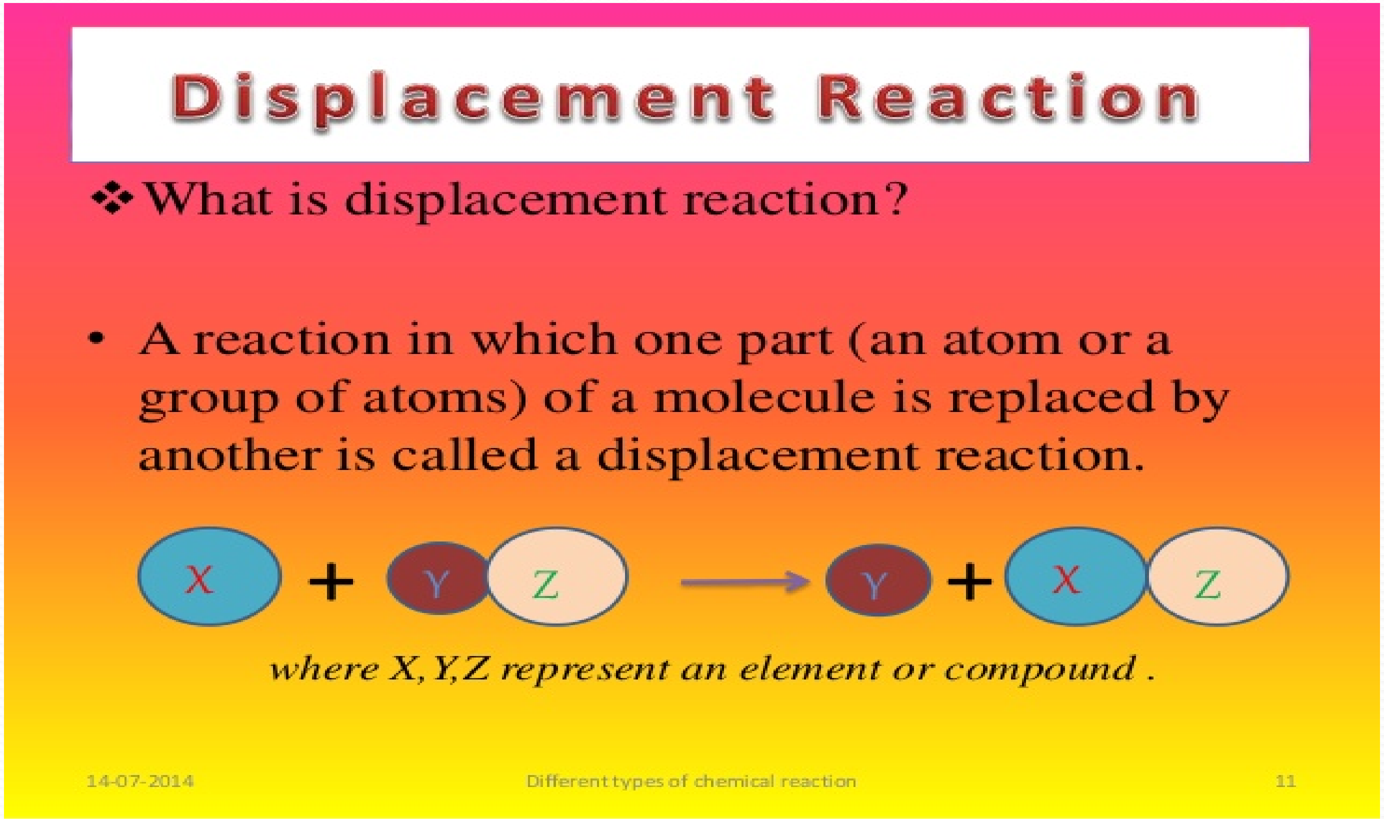
$$2\,Na + Cl_2 \rightarrow 2\,NaCl$$



General form: **A** + **B** → **AB**

A	+	B	→	AB
element or compound		element or compound		compound


Double Displacement
AgNO3 + NaCl → AgCl + NaNO3


Displacement Reaction
❖ What is displacement reaction?
• A reaction in which one part (an atom or a group of atoms) of a molecule is replaced by another is called a displacement reaction.
X
+
Y
Z
Y
+
X
Z
where X, Y, Z represent an element or compound .
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Different types of chemical reaction
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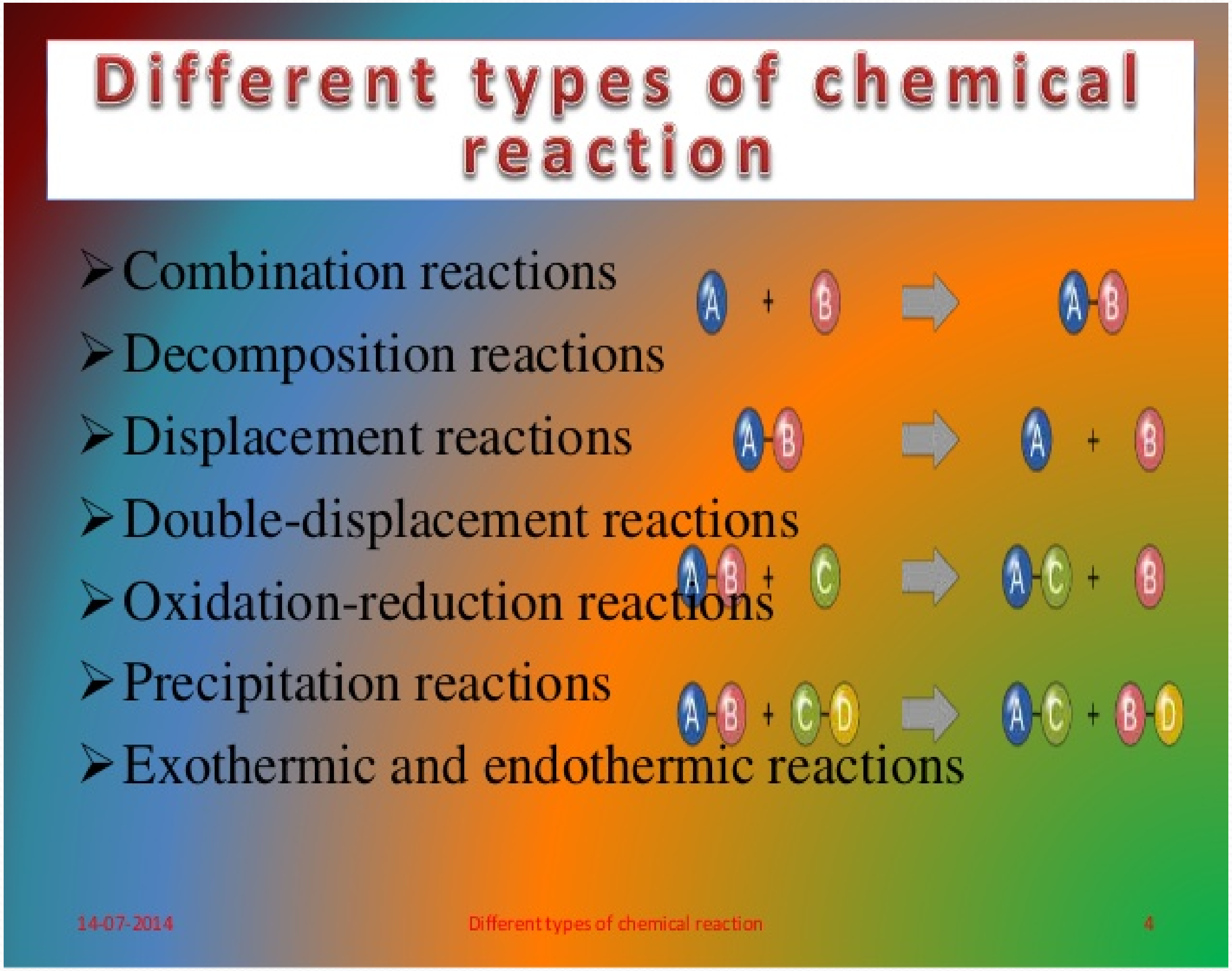
Different types of chemical reaction
➢Combination reactions
➢Decomposition reactions
➢Displacement reactions
➢Double-displacement reactions
➢Oxidation-reduction reactions
➢Precipitation reactions
➢Exothermic and endothermic reactions
A + B
A-B
A-B
A + B
A-B + C
A-C + B
A-B + C-D
A-C + B-D
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Different types of chemical reaction
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- Combustion Reactions
- **Combustion** reactions almost always involve oxygen in the form of O_2, and are almost always exothermic, meaning they produce heat. Chemical reactions that give off light and heat and light are colloquially referred to as "burning."
- CxHy+O2→CO2+H2O(1.12)(1.12)CxHy+O2→CO2+H2O
- Although combustion reactions typically involve redox reactions with a chemical being oxidized by oxygen, many chemicals "burn" in other environments. For example, both titanium and magnesium burn in nitrogen as well:
- 2Ti(s)+N2(g)→2TiN(s)(1.13)(1.13)2Ti(s)+N2(g)→2TiN(s)
- 3Mg(s)+N2(g)→Mg3N2(s)(1.14)(1.14)3Mg(s)+N2(g)→Mg3N2(s)
- Moreover, chemicals can be oxidized by other chemicals than oxygen, such as Cl_2 or F_2; these processes are also considered combustion reactions

- Disproportionation Reactions
- **Disproportionation Reactions:** In some redox reactions a single substance can be both oxidized and reduced. These are known as disproportionation reactions, with the following general equation:
- 2A→A+n+A−n(1.15)(1.15)2A→A+n+A−n
- Where n is the number of electrons transferred. Disproportionation reactions do not need begin with neutral molecules, and can involve more than two species with differing oxidation states (but rarely).

- Disproportionation reactions have some practical significance in everyday life, including the reaction of hydrogen peroxide, H_2O_2 poured over a cut. This a decomposition reaction of hydrogen peroxide, which produces oxygen and water. Oxygen is present in all parts of the chemical equation and as a result it is both oxidized and reduced. The reaction is as follows:
- 2H2O2(aq)→2H2O(l)+O2(g)(1.16)(1.16)2H2O2(aq)→2H2O(l)+O2(g)
- **DISCUSSION**On the reactant side, H has an oxidation state of +1 and O has an oxidation state of -1, which changes to -2 for the product H_2O (oxygen is reduced), and 0 in the product O2O2 (oxygen is oxidized).

Disproportionation Reaction

- Example:
  Is this a disproportionation reaction?

$$NH_4NO_3 \longrightarrow N_2O + 2H_2O$$

-3 +5 +1

This is NOT a disproportionation reaction

- Disproportionation requires that the <u>same</u> atom is both oxidised and reduced simultaneously.
- In this case, <u>different</u> atoms (of nitrogen) are oxidised and reduced.

- It is often useful to follow chemical reactions by looking at changes in the oxidation numbers of the atoms in each compound during the reaction. Oxidation numbers also play an important role in the systematic nomenclature of chemical compounds. By definition, the **oxidation number** of an atom is the charge that atom would have if the compound was composed of ions.
- 1. The oxidation number of an atom is zero in a neutral substance that contains atoms of only one element. Thus, the atoms in O_2, O_3, P_4, S_8, and aluminum metal all have an oxidation number of 0.
- 2. The oxidation number of simple ions is equal to the charge on the ion. The oxidation number of sodium in the Na^+ ion is +1, for example, and the oxidation number of chlorine in the Cl^- ion is -1.
- 3. The oxidation number of hydrogen is +1 when it is combined with a *nonmetal* as in CH_4, NH_3, H_2O, and HCl.
- 4. The oxidation number of hydrogen is -1 when it is combined with a *metal* as in. LiH, NaH, CaH_2, and $LiAlH_4$.
- 5. The metals in Group IA form compounds (such as Li_3N and Na_2S) in which the metal atom has an oxidation number of +1.
- 6. The elements in Group IIA form compounds (such as Mg_3N_2 and $CaCO_3$) in which the metal atom has a +2 oxidation number.
- 7. Oxygen usually has an oxidation number of -2. Exceptions include molecules and polyatomic ions that contain O-O bonds, such as O_2, O_3, H_2O_2, and the O_2^{2-} ion.
- 8. The elements in Group VIIA often form compounds (such as AlF_3, HCl, and $ZnBr_2$) in which the nonmetal has a -1 oxidation number.
- 9. The sum of the oxidation numbers in a neutral compound is zero.
- H_2O: 2(+1) + (-2) = 0
- 10. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion. The oxidation number of the sulfur atom in the SO_4^{2-} ion must be +6, for example, because the sum of the oxidation numbers of the atoms in this ion must equal -2.
- SO_4^{2-}: (+6) + 4(-2) = -2
- 11. Elements toward the bottom left corner of the periodic table are more likely to have positive oxidation numbers than those toward the upper right corner of the table. Sulfur has a positive oxidation number in SO_2, for example, because it is below oxygen in the periodic table.
- SO_2: (+4) + 2(-2) = 0

Oxidation Numbers

- The oxidation numbers of atoms in an ion add up to the charge on the the ion.

Oxidation state of S in SO_4^{2-}?

? – 8 = -2

? = +6

F	-1
O	-2
H	+1
Cl	-1

- The **Standard hydrogen electrode** (abbreviated SHE), is a redox **electrode** which forms the basis of the thermodynamic scale of oxidation-reduction potentials. … **Hydrogen electrode** is based on the redox half cell: $2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$ This redox reaction occurs at a platinized platinum**electrode**.

- In electrochemistry, the **Nernst equation** is an **equation** that relates the reduction potential of an electrochemical reaction (half-cell or full cell reaction) to the standard electrode potential, temperature, and activities (often approximated by concentrations) of the chemical species undergoing reduction and oxidation …

- **Standard Hydrogen Electrode** (SHE) is an **electrode** that scientists use for reference on all half-cell potential reactions. The value of the **standard electrode**potential is zero, which forms the basis one needs to calculate cell potentials using different **electrodes** or different concentrations.

- In the most common forms of potentiometry, two different types of **electrodes** are used. The potential of the **indicator electrode** varies, depending on the concentration of the analyte, while the potential of the reference **electrode** is constant. Potentiometry is probably the most frequently used…

- **Reference electrode**. A **reference electrode** is an **electrode** which has a stable and well-known **electrode** potential. The high stability of the **electrode** potential is usually reached by employing a redox system with constant (buffered or saturated) concentrations of each participants of the redox reaction.

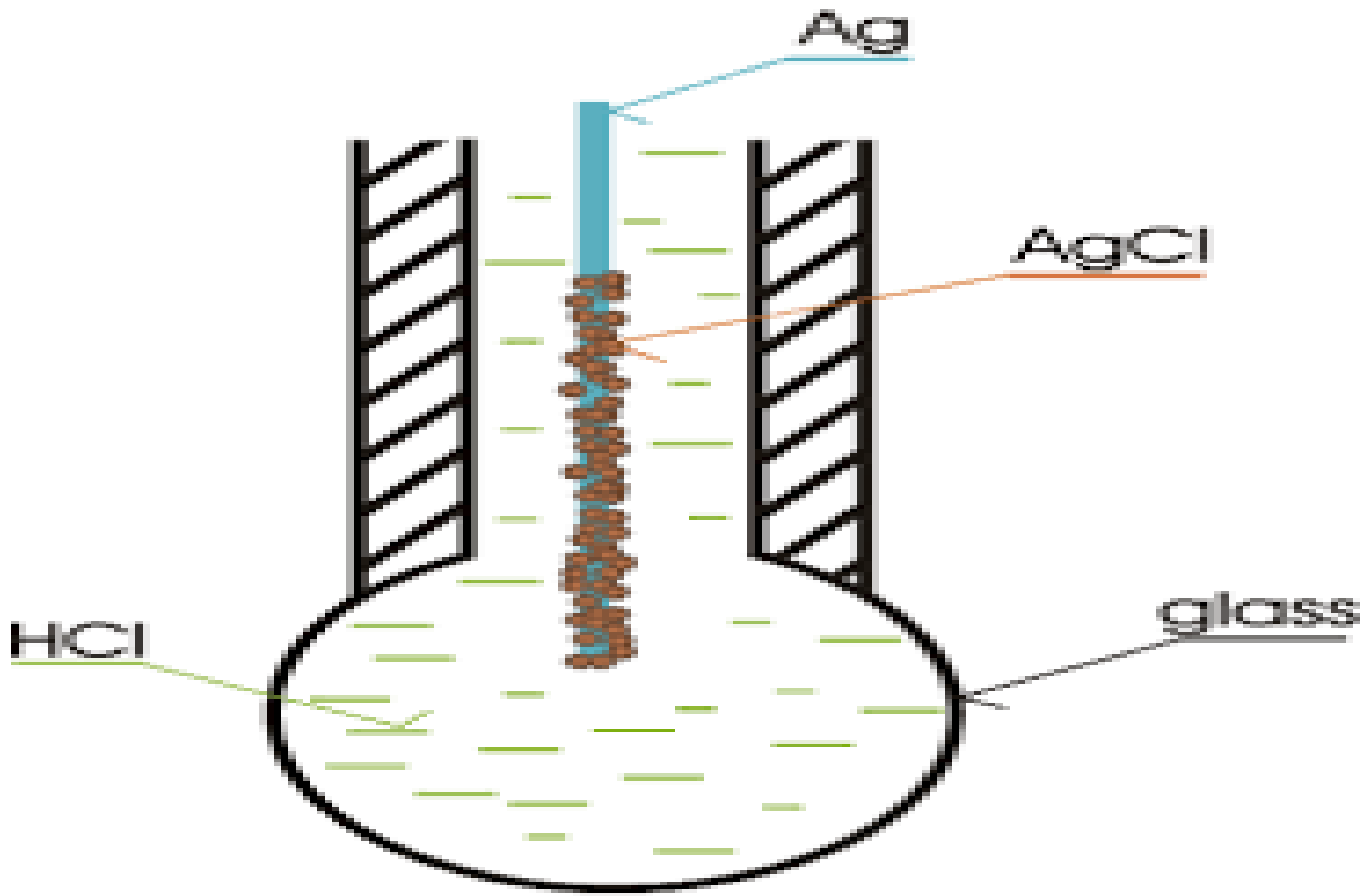

FIG. 1. Schematics of a glass electrode.

- The **Saturated calomel electrode** (SCE) is a reference **electrode** based on the reaction between elemental mercury and mercury(I) chloride. ... The aqueous phase in contact with the mercury and the mercury(I) chloride (Hg_2Cl_2, "**calomel**") is a**saturated** solution of potassium chloride in water.

Saturated calomel electrode

It is a commonly used reference electrode, it consists of a glass tube, that contains Hg at the bottom covered with solid Hg_2Cl_2 and above this the tube is filled with KCl solution. A platinum wire is in touch with Hg and it is used for electrical contact. The KCl solution inside the tube can have ionic contact with solution outside and acts as a salt bridge.

The electrode potential of the calomel electrode is +0.2422V.