4.7 Types of Chemical Reactions

Essential knowledge statements from the AP Chemistry CED:

- **Acid-base** reactions involve transfer of one or more protons between chemical species.
- **Oxidation-reduction** reactions (also known as **redox** reactions) involve transfer of one or more electrons between chemical species, as indicated by changes in oxidation numbers of the involved species. Combustion is an important subclass of oxidation-reduction reactions, in which a species reacts with oxygen gas. In the case of hydrocarbons, carbon dioxide and water are products of complete combustion.
- In a redox reaction, electrons are transferred from the species that is oxidized to the species that is reduced.
- Oxidation numbers may be assigned to each of the atoms in the reactants and products; this is often an effective way to identify the oxidized and reduced species in a redox reaction.
- **Precipitation** reactions frequently involve mixing ions in aqueous solution to produce an insoluble or sparingly soluble ionic compound. All sodium, potassium, ammonium, and nitrate salts are soluble in water.

In a previous packet you were studying Topic 4.2 (Net Ionic Equations), and you learned the following information about precipitation reactions and how to write equations to represent them.

- A **precipitation** reaction occurs when two different aqueous solutions containing ions are combined, resulting in the formation of an insoluble (or slightly soluble) solid ionic compound. The solid product is called the **precipitate**.
- There are three different ways to write a balanced equation for a precipitation reaction: a molecular equation, a complete ionic equation, and a net ionic equation
  - **molecular equation** – each reactant and product is written as a neutral compound
  - **complete ionic equation** – any substance that ionizes completely is broken up into separate ions; any substance that does not ionize (or is only partially ionized) is written as a neutral compound
  - **spectator ions** – ions that appear in identical forms on both sides of a complete ionic equation
  - **net ionic equation** – the balanced equation that describes the actual reaction that occurs in aqueous solution; it is obtained after the spectator ions are eliminated from the complete ionic equation

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1. For each of the following precipitation reactions, write the molecular equation and the net ionic equation.

Solutions of $KOH(aq)$ and $Mg(NO_3)_2(aq)$ are combined.
Molecular Equation:
Net Ionic Equation:

1. (continued)

Solutions of $AgNO_3(aq)$ and $Na_2CO_3(aq)$ are combined.
Molecular Equation:
Net Ionic Equation:

Solutions of $K_3PO_4(aq)$ and $Ca(NO_3)_2(aq)$ are combined.
Molecular Equation:
Net Ionic Equation:

Acid-base reactions are described in Topic 4.8, as shown below.

4.8 Introduction to Acid-Base Reactions

Essential knowledge statements from the AP Chemistry CED:

- By definition, a Brønsted-Lowry acid is a proton donor and a Brønsted-Lowry base is a proton acceptor.
- Only in aqueous solutions, water plays an important role in many acid-base reactions, as its molecular structure allows it to accept protons from and donate protons to dissolved species.
- When an acid or base ionizes in water, the conjugate acid-base pairs can be identified and their relative strengths compared.

When referring to acid-base reactions, the term *proton* and H^+ are used interchangeably.

When a strong acid is dissolved into water, it completely ionizes into positive and negative ions. The following substances are classified as strong acids.

hydrochloric acid	HCl	perchloric acid	$HClO_4$
hydrobromic acid	HBr	nitric acid	HNO_3
hydroiodic acid	HI	sulfuric acid	H_2SO_4

According to the Brønsted-Lowry definition of acids and bases,

- An acid is a substance that donates H^+ ions. Acid = H^+ donor.
- A base is a substance that accepts H^+ ions. Base = H^+ acceptor.

2. Each of the following reactions is classified as an acid-base reaction because it involves proton transfer. Classify each reactant as either an acid (i.e., H^+ donor) or a base (i.e., H^+ acceptor).

(a) $HBr + NaOH \rightarrow H_2O + NaBr$

Acid or Base? ________ ________

(b) $H_2SO_4 + Ca(OH)_2 \rightarrow 2\ H_2O + CaSO_4$

Acid or Base? ________ ________

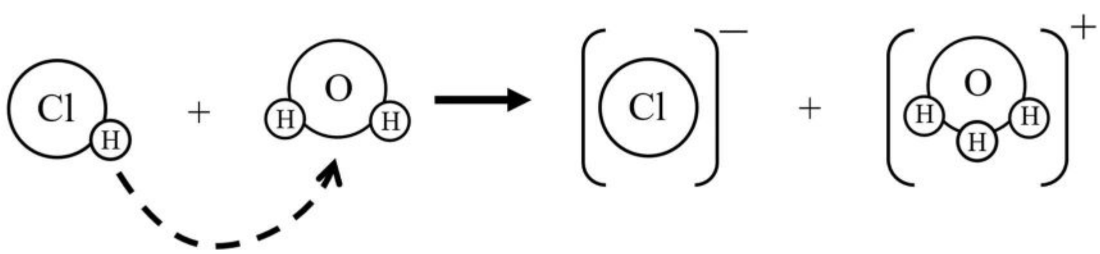
(c) $HNO_3 + KOH \rightarrow H_2O + KNO_3$

Acid or Base? ________ ________

From the three examples shown above, we can make some generalizations.

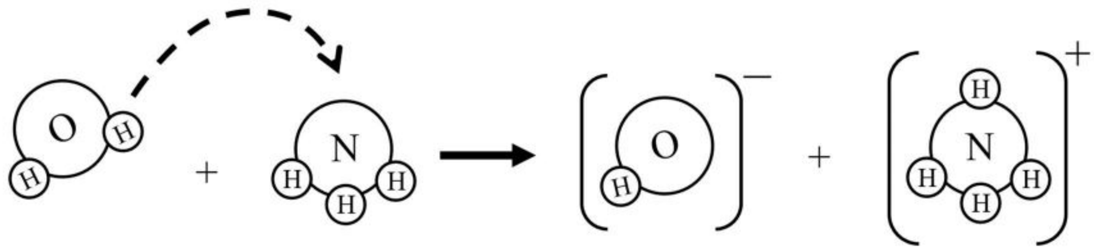
- The chemical formula of an acid usually begins with the element hydrogen (H).
- The formula of an acid can often be recognized as a partnership between H^+ ions and negative ions such as chloride (Cl^-), bromide (Br^-), nitrate (NO_3^-), sulfate (SO_4^{2-}), etc.
- Metal hydroxides contain metal cations and hydroxide ions. These substances are classified as bases because the hydroxide ion (OH^-) acts as a proton acceptor.

Hydrochloric acid (HCl) is a compound that is classified as an acid. When HCl dissolves in water, HCl molecules react with water molecules, forming chloride ions and hydronium ions, as represented by the following particle diagram and chemical equation.



$$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

Ammonia (NH_3) is a compound that is classified as a base. When NH_3 dissolves in water, NH_3 molecules react with water molecules, forming hydroxide ions and ammonium ions, as represented by the following particle diagram and chemical equation.



$$H_2O \quad + \quad NH_3 \quad \rightarrow \quad OH^- \quad + \quad NH_4^+$$

3. Notice that water has the ability to act either as an acid or as a base.

   (a) When HCl reacts with water, H_2O acts as (an acid a base)

   because it (donates accepts) a proton.

   (b) When NH_3 reacts with water, H_2O acts as (an acid a base)

   because it (donates accepts) a proton.

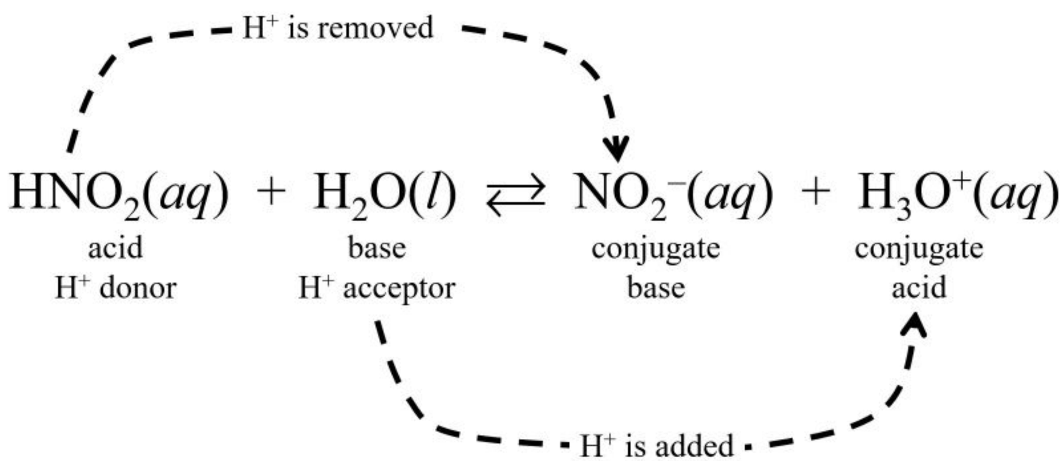
- Acid-base reactions involve proton transfer from an acid to a base.
- Some types of acid-base reactions are reversible and can occur in both directions. This is indicated by a pair of double arrows ($\rightleftarrows$).
- A **conjugate acid-base pair** is an acid and a base that differ by only a single H^+.

4. Each of the following substances can act as an acid. Write the correct formula and charge of the conjugate base for each acid. If there is more than one H atom in the formula, remove **only one H^+ ion** to form the conjugate base.

Acid	**Conjugate Base**
HF	
HNO_3	
H_2CO_3	
H_2SO_4	
H_3O^+	
NH_4^+	
HCO_3^-	
HPO_4^{2-}	

5. Each of the following substances can act as a base. Fill in the missing information in the table below by writing the correct formula and charge of the conjugate acid. Add **only one H^+ ion** to form the conjugate acid.

Base	**Conjugate Acid**
NH_3	
ClO^-	
NO_2^-	
OH^-	
HCO_3^-	
S^{2-}	
SO_3^{2-}	
PO_4^{3-}	



In the acid-base reaction shown above, there are two conjugate acid-base pairs. A conjugate acid-base pair is an acid and a base that differ by only a single H^+.

- HNO_2 and NO_2^- represent a conjugate acid-base pair.
- H_3O^+ and H_2O represent a conjugate acid-base pair.

$$CH_3CH_2COOH(aq) \; + \; H_2O(l) \rightleftarrows CH_3CH_2COO^-(aq) \; + \; H_3O^+(aq)$$

6. Identify both sets of conjugate acid-base pairs in the reaction above.

and	
and	

$$CH_3CH_2NH_2(aq) \; + \; H_2O(l) \rightleftarrows CH_3CH_2NH_3^{+}(aq) \; + \; OH^{-}(aq)$$

7. Identify both sets of conjugate acid-base pairs in the reaction above.

and
and

Redox reactions are described in Topic 4.9, as shown below.

4.9 Oxidation-Reduction (Redox) Reactions

Essential knowledge statement from the AP Chemistry CED:
- Balanced chemical equations for redox reactions can be constructed from half-reactions.

When a piece of magnesium (Mg) is burned in air, a bright light is released. Magnesium reacts with oxygen (O_2), forming the ionic solid magnesium oxide (MgO). You can watch a video of this experiment by scanning the QR code shown at right.

The balanced chemical equation for the reaction between magnesium and oxygen is shown below.

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

Magnesium oxide (MgO) contains Mg^{2+} ions and O^{2-} ions. The equation shown above can be split up into two separate processes.

An atom of Mg changes into a Mg^{2+} ion	A molecule of O_2 changes into two O^{2-} ions
$Mg \rightarrow Mg^{2+}$	$O_2 \rightarrow 2\ O^{2-}$

oxidation: a substance loses electrons

reduction: a substance gains electrons

You can use the mnemonic device **OIL RIG** to help you remember these definitions.

OIL: Oxidation is the Loss of electrons

RIG: Reduction is the Gain of electrons

Equations that show either oxidation or reduction separately are known as **half-reactions**.

- In an oxidation half-reaction, electrons are lost and appear on the right side of the equation.
- In a reduction half-reaction, electrons are gained and appear on the left side of the equation.
- A half-reaction should be balanced in terms of both atoms and charge.

Oxidation Half-Reaction	Reduction Half-Reaction
$Mg \rightarrow Mg^{2+} + 2\,e^{-}$	$4\,e^{-} + O_2 \rightarrow 2\,O^{2-}$

- The number of electrons lost in the oxidation half-reaction must be equal to the number of electrons gained in the reduction half-reaction.
- You may need to multiply a half-reaction by a certain number so that the electrons can be cancelled out on both sides of the equation.
- When you add the two half-reactions together, the electrons are cancelled out on both sides. The result is the overall balanced oxidation-reduction (redox) equation.
- A redox equation should be balanced in terms of both atoms and charge.

$$2\,Mg \rightarrow 2\,Mg^{2+} + 4\,e^{-}$$ This half-reaction was multiplied by 2.

$$4\,e^{-} + O_2 \rightarrow 2\,O^{2-}$$

$$2\,Mg + O_2 \rightarrow 2\,MgO$$ Electrons are cancelled out on both sides of the equation and do not appear in the overall redox equation.

$$Al(s) + Cu^{2+}(aq) \rightarrow Al^{3+}(aq) + Cu(s)$$

8. Explain why the equation above is not completely balanced.
   Split this equation up into two separate half-reactions. Then add the two half-reactions together so that the electrons are cancelled out on both sides, producing a balanced redox equation.

9. In each of the following examples, add the two half-reactions together to produce a balanced redox equation in which the electrons are cancelled out on both sides.

(a) Oxidation half-reaction: $Fe^{2+} \rightarrow Fe^{3+} + e^-$

Reduction half-reaction: $5\,e^- + MnO_4^- + 8\,H^+ \rightarrow Mn^{2+} + 4\,H_2O$

(b) Oxidation half-reaction: $2\,Cl^- \rightarrow Cl_2 + 2\,e^-$

Reduction half-reaction: $6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$

(c) Oxidation half-reaction: $Ni \rightarrow Ni^{2+} + 2\,e^-$

Reduction half-reaction: $3\,e^- + 4\,H^+ + NO_3^- \rightarrow NO + 2\,H_2O$

Rules for Assigning Oxidation Numbers

In a redox reaction, electrons are transferred from the substance that is oxidized to the substance that is reduced. One of the ways to recognize a redox reaction is to **look for changes in oxidation numbers**. Oxidation numbers are assigned to elements according to the following rules.

Rule	Examples
#1: If an atom is in its elemental form, the oxidation number is zero.	H_2 = zero C = zero Li = zero Cl_2 = zero N_2 = zero Fe = zero
#2: The oxidation number of a monoatomic ion is the charge on that ion.	Na^+ = +1 Cl^- = –1 Mg^{2+} = +2 O^{2-} = –2 Al^{3+} = +3 N^{3-} = –3
#3: The oxidation number of a metal in Group 1 is +1 in all compounds. The oxidation number of a metal in Group 2 is +2 in all compounds.	Na = +1 in compounds such as NaCl, $NaNO_3$, and Na_2CO_3. Mg = +2 in compounds such as $MgBr_2$, $MgSO_4$, and MgO.
#4: The oxidation number of F is –1 in all compounds.	F = –1 in compounds such as NaF, CH_3F, and SF_2.
#5: The oxidation number of H is usually +1 when H is bonded to nonmetals and –1 when H is bonded to metals.	H = +1 in compounds such as H_2O, CH_3OH, NH_4Cl, and H_2S. H = –1 in compounds such as LiH, NaH, and CaH_2.
#6: The oxidation number of O is usually –2 in most ionic and molecular compounds. Two important exceptions to this rule are hydrogen peroxide, H_2O_2, and oxygen difluoride, OF_2.	O = –2 in compounds such as Na_3PO_4, $Mg(NO_3)_2$, CO_2, H_2SO_3, N_2O, and P_2O_5. O = –1 in H_2O_2 O = +2 in OF_2
#7: The oxidation number of the halogens Cl, Br, and I is –1 in most binary compounds that contain only two elements. When atoms Cl, Br, and I are bonded with oxygen or fluorine, they will have a positive oxidation number.	The halogens Cl, Br, or I have an oxidation number of –1 in compounds such as KCl, NaBr, MgI_2, CCl_4, SBr_2, and HI. Cl = +1 in NaClO and HClO Cl = +3 in $NaClO_2$ and $HClO_2$ Cl = +5 in $NaClO_3$ and $HClO_3$ Cl = +7 in $NaClO_4$ and $HClO_4$

- The sum of the oxidation numbers of all atoms in a neutral compound is equal to zero.
- The sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the overall charge on the polyatomic ion.

The oxidation number is assigned to an individual atom in a chemical formula.
The oxidation number is NOT assigned to the entire group of atoms in a chemical formula.

Chemical Formula	**Correct Oxidation Numbers**	Incorrect Oxidation Numbers Be careful NOT to do this!
CH_4	**C = –4** **H = +1**	H = +4
SO_2	**S = +4** **O = –2**	O = –4
Na_2CO_3	**Na = +1** **C = +4** **O = –2**	Na = +2 O = –6

10. Assign oxidation numbers to each atom.

Substance	Oxidation Numbers		
Fe	Fe =		
Cl_2	Cl =		
P_4	P =		
NaBr	Na =	Br =	
Ca_3N_2	Ca =	N =	
Fe_2O_3	Fe =	O =	
MgS	Mg =	S =	
SO_2	S =	O =	
CH_3F	C =	H =	F =
CHF_3	C =	H =	F =
CO	C =	O =	
OF_2	O =	F =	
H_2O_2	H =	O =	
KH	K =	H =	
CaH_2	Ca =	H =	

Substance	Oxidation Numbers		
HBr	H =	Br =	
HBrO	H =	Br =	O =
$HBrO_2$	H =	Br =	O =
$HBrO_3$	H =	Br =	O =
Na_2CO_3	Na =	C =	O =
$BaCrO_4$	Ba =	Cr =	O =
NH_4^+	N =	H =	
OH^-	O =	H =	
NO_3^-	N =	O =	
ClO_4^-	Cl =	O =	
SO_3^{2-}	S =	O =	
SO_4^{2-}	S =	O =	
$Cr_2O_7^{2-}$	Cr =	O =	
PO_4^{3-}	P =	O =	
H_3O^+	H =	O =	

11. Each of the following equations represents an oxidation-reduction reaction.
    - Assign oxidation numbers to each element on each side of the equation.
    - Determine which element is oxidized and which element is reduced in the reaction.
    - If the oxidation number increases, the element is oxidized.
    - If the oxidation number decreases, the element is reduced.

Balanced Equation		Element that is Oxidized	Element that is Reduced
$3\ Fe(NO_3)_2 + 2\ Al \rightarrow 3\ Fe + 2\ Al(NO_3)_3$			
Fe = N = O = Al =	Fe = N = O = Al =		

Balanced Equation		Element that is Oxidized	Element that is Reduced
$2\ BrO_3^{-} + 3\ N_2H_4 \rightarrow 2\ Br^{-} + 6\ H_2O + 3\ N_2$			
Br = O = N = H =	Br = O = N = H =		

Balanced Equation		Element that is Oxidized	Element that is Reduced
$P_4 + 10\ HClO + 6\ H_2O \rightarrow 4\ H_3PO_4 + 10\ HCl$			
P = H = Cl = O =	P = H = Cl = O =		

Balanced Equation		Element that is Oxidized	Element that is Reduced
$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$			
C = H = O =	C = H = O =		

Balanced Equation		Element that is Oxidized	Element that is Reduced
$Mn + H_2SO_4 \rightarrow MnSO_4 + H_2$			
Mn = H = S = O =	Mn = H = S = O =		

12. Classify each of the following reactions as precipitation, acid-base, or redox.

Balanced Equation for the Reaction	Reaction Type
$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	
$Ca(OH)_2(aq) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$	
$AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)$	
$2\ KOH(aq) + H_2SO_4(aq) \rightarrow 2\ H_2O(l) + K_2SO_4(aq)$	
$Pb(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow 2\ KNO_3(aq) + PbSO_4(s)$	
$PbS(s) + 4\ H_2O_2(aq) \rightarrow 4\ H_2O(l) + PbSO_4(s)$	

13. For each of the following redox reactions, write the balanced net ionic equation.

(a) $Ca(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

(b) $2\ Fe(s) + 6\ HBr(aq) \rightarrow 2\ FeBr_3(aq) + 3\ H_2(g)$

(c) $Zn(s) + Ni(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Ni(s)$

(d) $Al(s) + 3\ AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3\ Ag(s)$

(e) $Cu(s) + 4\ HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\ H_2O(l) + 2\ NO_2(g)$

(f) $Br_2(aq) + 2\ KI(aq) \rightarrow 2\ KBr(aq) + I_2(aq)$