Chemistry 30
Unit VII – Oxidation & Reduction

Oxidation & Reduction Reactions (REDOX)
- reactions where electrons are transferred
  - oxidation – electrons are lost
  - reduction – electrons are gained
- you can’t have one without the other

To help you remember: **LEO** (Losing Electrons Oxidation) says **GER** (Gaining Electrons Reduction)

oxidation number (oxidation state): how many electrons an atom has gained or has lost
- += lost electrons (oxidation)
- -= gained electrons (reduction)

Rules for assigning oxidation numbers:
1. in ionic compounds, the oxidation number = ion charge for the group.
   - group IA = +1
   - group IIA = +2
   - group VIA = -2
   - group VIIA = -1

2. in nonpolar covalent bonds (polyatomic elements) the oxidation number = 0 because electrons are shared equally
   - O\(_2\) (g) ox # = 0
   - P\(_4\) (s) ox # = 0

3. in polar covalent bonds (where electrons are NOT shared equally), the more electronegative atom in every bond is considered to be reduced while the other is oxidized (it’s sometimes necessary to draw structural formulas)

ex: H\(_2\)O

\[
\begin{array}{c}
\text{H – O – H} \\
\text{ex: H\(_2\)O} \\
\text{EN for hydrogen} = 2.1 \\
\text{EN for oxygen} = 3.5 \\
\end{array}
\]

- oxygen is more electronegative, so it is considered to “take” the electron from EACH hydrogen
- so each H is oxidized to H\(^+\), and the one oxygen is reduced to O\(^2-\)

some examples: 

F\(_2\) diatomic, so F = 0

NaCl ionic, so Na = +1, Cl = -1

Fe\(_2\)O\(_3\) ionic, so Fe = +3, O = -2

H\(_2\)S EN sulfur = 2.5
EN hydrogen = 2.1
so sulfur will be reduced to S = -2, each H = +1

CO\(_2\) (O=C=O)
EN oxygen = 3.5
EN carbon = 2.1
so carbon is oxidized, and oxygen is reduced
since there are double bonds, each bond represents 2 electrons transferred, therefore each O = -2, and the C = +4
HCN  \((H-C≡N)\)  
EN hydrogen = 2.1  
EN carbon = 2.5  
EN nitrogen = 3.0  

H will oxidize and C will reduce \((H = +1, C = -1\) for this bond\)
for other bonds, C will oxidize, N will reduce \((C = +3, N = -3\)
SO: overall, \(C = +3-1 = +2, H = +1, N = -2\) (total = 0)

\(H_2O_2\)  
\((H-O-O-H)\)  
H oxidizes, O reduces, but each oxygen shares equally, so each \(H = +1, O = -1\)

\(OF_2\)  
EN oxygen = 3.5  
EN fluorine = 4.0  
so, oxygen will oxidize and fluorine will reduce, \(F = -1\) (each)
and \(O = +2\)

*** The general rules using electronegativities can always be used when you are unsure of more than one atom, however in general, you can use the “quick rules” below to assign oxidation numbers to the atoms in compounds or ions.

**Quick rules:**

1. the oxidation numbers of:
   - elements = 0 (ex: \(P_4\), ox # = 0)
   - monatomic ions = ion charge (ex: \(Fe^{2+}\), ox # = 2+)
   - \(H = +1\) when bonded with nonmetals (ex: \(H_2O\), H = -1 when bonded with metals (ex: \(NaH\))
   - \(O = -2\) usually (ex: \(H_2O\), O = -1 in peroxides (ex: \(H_2O_2\)), O = +2 when bonded with F (ex: \(OF_2\))
   - nonmetals and transition metals are unpredictable and may have many different oxidation states – work these out last

2. the sum of the oxidation numbers:
   - in a compound = 0 (ex: \(\Sigma\) ox # in \(H_2O \rightarrow 2H + O = 2(+1) + 1(-2) = 0\))
   - in a polyatomic ion = charge of the ion (ex: \(\Sigma\) ox # in \(SO_4^{2-} \rightarrow 1(S) + 4(O) = -2\))

*** by using the combination of both of these rules, you can usually determine the oxidation number of all metals, nonmetals and transition metals in a compound or a molecule

ex:  
\(H_2SO_4\)  
\(+1\) \(-2\) \(H_2SO_4\)  
- start with what we know for sure:
  \(H = +1\)
  \(O = -2\)
- put these on top (where we’ll summarize all ox #s)

\(+1\) \(-2\) \(H_2SO_4\)  
- to figure out sulfur, we know the sum of all of the oxidation numbers is 0, work out this on the bottom (2Hs each at +1 = +2, 4Os each at -2 = -8)

\(+2\) \(-8\) \(H_2SO_4\)  
- since we know \(\Sigma\) ox # = 0 in all compounds, \(+2 +x -8 = 0\), so \(x = +6\)

\(+1\) \(-2\) \(H_2SO_4\)  
. put this on the bottom. This is the total “charge” on all of the sulfurs.
\(+2 +6 -8\) \(H_2SO_4\)  
Since there is only one sulfur, its oxidation number must be +6, fill this in on top
\(+1 +6 -2\) \(H_2SO_4\)  
- now you’re done ☺️
ex: \( \text{H}_2\text{PO}_4^- \)
- \( H = +1, O = -2; \Sigma Hs = +1, \Sigma Os = -8 \)
- this time, \( \Sigma \) of all the oxidation numbers = -1 (the ion charge)
  \[ so +2 +x -8 = -1, x = +5 \] (the oxidation number of the phosphoruses)

+1 +5 -2
\( \text{H}_2\text{PO}_4^- \)
+2 +5 -8
- since there is only one P, the oxidation number is 5

ex: \( (\text{NH}_4)_4\text{XeO}_6 \)
- \( O = -2, \) and we also know that \( \text{NH}_4 = +1, \) for totals of +4 and -12
+4 -12

+1 +8 -2
(\( \text{NH}_4)_4\text{XeO}_6 \)
+4 +8 -12
- since the sum of the oxidation numbers is zero, \( \text{Xe} \) must be +8

-3 +1 +8 -12
(\( \text{NH}_4)_4\text{XeO}_6 \)
- to solve for \( \text{N} \) and \( \text{H} \) inside the bracket, start with what you know
-3 +1 +8 -12
- \( H = +1, \) for a total of +4
  \[ \Sigma = +1 \] for the entire ion, so \( \text{N} = -3 \)

**you always have to solve the inside of the bracket (the individual atoms in the polyatomic ion)**

ex: \( \text{Cr}_2\text{O}_7^{2-} \)
- start with what you know \( O = -2, \) for a total of -14

+6 -2
\( \text{Cr}_2\text{O}_7^{2-} \)
-14

ex: \( \text{Cr}_2\text{O}_7^{2-} \)
- \( \Sigma = -2 \) (ion charge) so the \( \text{Cr} \) total must be +12 (12 -14 = -2)
+12 -14

- since there are two chromium ions, each ion has an oxidation # = +6

Balancing Equations using Oxidation Numbers

Many of the types of reactions you have worked with already in grade 10 and grade 11 have been redox reactions. All decomposition, composition, single replacement and combustion reactions are straightforward redox reactions. As well, most of the complex “other” reactions that are hard to balance the usual way are redox reactions.

- A redox reaction can be identified by changing oxidation numbers for at least two chemical species

**ex: is the following a redox reaction?**

- to determine whether or not a reaction is redox, assign numbers and see if they change (use the method in the previous section)

\[
\begin{align*}
\text{As}_2\text{O}_3(s) + \text{Cl}_2(g) + \text{H}_2\text{O}(l) & \rightarrow \text{H}_3\text{AsO}_4(aq) + \text{HCl}(aq) \\
+3 & \quad -2 \\
+1 & \quad -2 \\
+1 & \quad +5 \\
+1 & \quad -1 \\
\end{align*}
\]

Before the reaction, the oxidation number of arsenic is +3, and after the reaction, the oxidation number is +5 (gone UP), so arsenic is oxidized (LEO). Chlorine’s oxidation number changed from 0 to -1. Since it’s oxidation number went DOWN, chlorine was reduced (GER). Since the oxidation numbers change, this is a redox reaction.
Oxidation number changes can be used to balance redox reactions. Using the equation above, and the assigned oxidation numbers,

\[
\begin{align*}
+3 & \quad -2 & \quad 0 & \quad +1 & \quad -2 & \quad +1 & \quad +5 & \quad -2 & \quad +1 & \quad -1 \\
\text{As}_2\text{O}_3(s) + \text{Cl}_2(g) + \text{H}_2\text{O}(l) & \rightarrow \text{H}_3\text{AsO}_4(aq) + \text{HCl}(aq) \\
+6 & \quad -6 & \quad 0 & \quad +2 & \quad -2 & \quad +3 & \quad +5 & \quad -8 & \quad +1 & \quad -1
\end{align*}
\]

Write the two atoms that changed in oxidation number and the change that occurred. This represents the number of electrons that had to be lost (oxidation) or gained (reduction). Since chlorine and arsenic are found in sets of two on the reactant side, we have to consider that in the number of electrons that must be transferred.

\[
\begin{align*}
\text{As} & \quad 2(0 \rightarrow +5) & \quad (+2)(2) & \quad = 4 \\
\text{Cl} & \quad 2(0 \rightarrow -1) & \quad (-1)(2) & \quad = -2
\end{align*}
\]

To balance the lost electrons with the gained electrons, use the lowest common multiple (LCM) between the oxidized atoms and the reduced atoms. In this case, the LCM between 4 and 2 is 4. So for this equation, 4 electrons will be transferred. Whatever number must be multiplied by the change in oxidation number is the number of that atom required per side of the reaction.

\[
\begin{align*}
2\text{As} & \quad 2(0 \rightarrow +5) & \quad (+2)(2) & \quad = 4 \\
2\text{Cl} & \quad 2(0 \rightarrow -1) & \quad (-1)(2) & \quad = -2
\end{align*}
\]

Put these numbers back into the equation, and balance the other species by inspection. Whatever you do, you CANNOT change the coefficients of those species balanced using oxidation numbers.

\[
\text{1 As}_2\text{O}_3(s) + 2 \text{Cl}_2(g) + 5 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_3\text{AsO}_4(aq) + 4 \text{HCl}(aq)
\]

The bold numbers are derived from balancing the electrons – 4 electrons must be transferred, a total of 2 As required per side, and 4 Cl required per side. You have to take into account that As is 2 per As$_2$O$_3$ but only 1 in H$_3$AsO$_4$.

**Half Reactions**

- half reactions are chemical equations that detail EITHER the oxidation half or the reduction half of a redox reaction separately
- these ½ reactions can be:
  i) written from scratch (in neutral or acidic solutions)
  ii) written using p. 9 (the table of reduction ½ reactions)
  iii) used to balance redox reactions
  iv) written from a molecular equation

**Writing and Balancing Half Reactions**

There are just a few steps to writing half reactions from scratch:

1. write the skeletal equation
2. balance the charge by adding electrons (é)

ex: Write the oxidation half reaction for magnesium metal

\[
\text{Mg}(s) \rightarrow \text{Mg}^{2+}(aq) + 2\text{é}
\]

As magnesium metal oxidizes, it loses 2 electrons since it is in group IIA.
Writing and Balancing Half Reactions in Acidic Solution

These half reactions are slightly more complicated to write. Since the reaction takes place in acid, automatically \( \text{H}^+ (aq) \) and \( \text{H}_2\text{O}(l) \) are available to help you balance the half reaction.

As a result, atoms of hydrogen and oxygen may not appear in the skeletal equation that may appear in the final half reaction.

**ex: write the half reaction for the reduction of nitric acid to nitrogen dioxide**

1. **skeletal equation**
   \[
   \text{HNO}_3(aq) \rightarrow \text{NO}_2(g)
   \]

2. **balance the core atoms (N)** *(already balanced)*
   \[
   \text{HNO}_3(aq) \rightarrow \text{NO}_2(g)
   \]

3. **balance the oxygens** *(add \( \text{H}_2\text{O}(l) \))*
   \[
   \text{HNO}_3(aq) \rightarrow \text{NO}_2(g) + \text{H}_2\text{O}(l)
   \]

4. **balance the hydrogens** *(add \( \text{H}^+(aq) \))*
   \[
   \text{H}^+(aq) + \text{HNO}_3(aq) \rightarrow \text{NO}_2(g) + \text{H}_2\text{O}(l)
   \]

5. **balance the charge** *(add \( \text{é} \))*
   \[
   \text{é} + \text{H}^+(aq) + \text{HNO}_3(aq) \rightarrow \text{NO}_2(g) + \text{H}_2\text{O}(l)
   \]

**Rules:**
1. start by writing the skeletal equation *(get the information from the question)*
2. balance all core *(non H and O atoms)*
3. balance the oxygens by adding \( \text{H}_2\text{O}(l) \)
4. balance the hydrogens by adding \( \text{H}^+(aq) \)
5. balance the charge by adding electrons

**Balancing Using Half Reactions**

Half reactions can be used to balance redox reactions – by balancing the number of electrons lost in the oxidation reaction with the number of electrons gained in the reduction half reaction.

**ex: balance the following redox reaction taking place in acidic solution**

\[
\text{Ag(s)} + \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow \text{Ag}^+(aq) + \text{Cr}^{3+}(aq)
\]

1. **separate the net ionic into half reactions using up ALL of the reactants and products (the new skeletal equations)**
   \[
   \text{Ag(s)} \rightarrow \text{Ag}^+(aq)
   \quad \text{Cr}_2\text{O}_7^{2-}(aq) \rightarrow \text{Cr}^{3+}(aq)
   \]

2. **complete the half reactions (using \( \text{H}^+ \) and \( \text{H}_2\text{O}(l) \) if in acid), and electrons**
   \[
   6\text{é} + \text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)
   \]

3. **(x6)** *(x1)*
   \[
   6\text{Ag(s)} \rightarrow 6\text{Ag}^+(aq) + 6\text{é}
   \]

4. **cancel the electrons on both sides of the equation, cancel anything else you can, then add the two together**

\[
\text{6Ag(s)} \rightarrow 6\text{Ag}^+(aq)
\]

\[
\text{6é} + \text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 7\text{H}_2\text{O}(l)
\]
4. \[6 \text{Ag(s)} \rightarrow 6 \text{Ag}^+(aq) + 6 \text{é} \]
\[6 \text{é} + \text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) \rightarrow 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l) \]

\[6 \text{Ag(s)} + \text{Cr}_2\text{O}_7^{2-}(aq) + 14\text{H}^+(aq) \rightarrow 6 \text{Ag}^+(aq) + 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l) \]

(sometimes the half reactions can be taken directly from p.9 of the data booklet instead of you having to write them from scratch)

**Oxidizing and Reducing Agents**

**oxidizing agent (OA)** the chemical species in a redox reaction that causes reduction to happen
- the species in the reduction half reaction is the OA – in the above reaction, \(\text{Cr}_2\text{O}_7^{2-}(aq)\)

**reducing agent (RA)** the chemical species in a redox reaction that causes oxidation to happen
- the species in the oxidation half reaction is the RA – in the above reaction, \(\text{Ag(s)}\)

**Balancing Molecular Equations using Half Reactions**

Before attempting to balance an equation using half reactions, you should always rewrite the equation as a net ionic first – all species whose oxidation numbers don’t change are spectators

ex: *Balance the following redox reaction using the ion-electron (half reaction) method*

\[\text{Na}_2\text{HAsO}_3(aq) + \text{KBrO}_3(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{KBr}(aq) + \text{H}_3\text{AsO}_4(aq)\]

\[2\text{Na}^+(aq) + \text{H}^+(aq) + \text{AsO}_3^-(aq) + \text{K}^+(aq) + \text{BrO}_3^-(aq) + \text{H}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) + \text{K}^+(aq) + \text{Br}^-(aq) + 3\text{H}^+(aq) + \text{AsO}_4^3(aq)\]

\[\text{AsO}_3^3(aq) \rightarrow \text{AsO}_4^3(aq)\]

\[\text{BrO}_3^- (aq) \rightarrow \text{Br}^-(aq)\]

**skeletal equations** – now balance core, then Os and Hs, and finally electrons

\[3 [\text{H}_2\text{O}(l) + \text{AsO}_3^3(aq) \rightarrow \text{AsO}_4^3(aq) + 2\text{H}^+(aq) + 2\text{é}] \]
\[6\text{é} + 6\text{H}^+(aq) + \text{BrO}_3^-(aq) \rightarrow \text{Br}^-(aq) + 3\text{H}_2\text{O}(l) \]

\[3 \text{AsO}_3^3(aq) + \text{BrO}_3^- (aq) \rightarrow 3 \text{AsO}_4^3(aq) + \text{Br}^-(aq)\]

multiply by 3
to cancel é

(note how much work it was to balance a molecular equation this way – usually it is much easier to balance molecular equations using the oxidation number method – if you have to balance in acid, however, it will be much more straightforward to write half reactions)

**Predicting Redox Reactions and their Spontaneity**

(from the Activity Series lab)

\[2 \text{Ag}^+(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2 \text{Ag}(s)\]

\[\text{Zn}^{2+}(aq) + 2 \text{Ag}(s) \rightarrow \text{no reaction}\]

why do some redox reactions happen while others don’t?
- In order for a redox reaction to happen spontaneously, the species that is oxidizing (the reducing agent) has to willingly give up electrons to the species that is reducing (the oxidizing agent).
- For this to happen, the oxidizing agent must be “strong” enough (have a high enough affinity for electrons) to take electrons from the reducing agent.

**Strong reducing agent** – will willingly oxidize (give up electrons)
**Strong oxidizing agent** – has a high affinity for electrons, and will easily remove electrons from an RA

- The combination of a strong enough RA with a strong enough OA will lead to a spontaneous reaction
- We can predict the spontaneity of a reaction using p. 9 in the data booklet (Reduction Half Reaction Table)

### Reduction Half Reactions Table (p. 9 in the data booklet)
- This table is an activity series for metals, nonmetals, polyatomic ions and common combinations of ions.
- The reactions written are all reduction reactions.
- For an oxidation half reaction, flip the reaction as written – nothing else changes.
- The strongest OA (oxidizing agents) are at the top left of the table (F₂(g) is THE strongest OA) (this makes sense since fluorine is the most electronegative element on the periodic table, so has the highest affinity for electrons)
- The strongest RA (reducing agents) are at the bottom right of the table (Li(s)) is THE strongest RA) (which also makes sense – given it’s position on the periodic table, it will quite willingly lose an electron to obtain a full valence shell)
- Some ions will appear twice on the table – the polyvalent ions (transition metals that have more than one stable ion form, ex: Cu⁺(aq) and Cu²⁺(aq)) since its lower charged ion, in this case Cu⁺(aq) can both oxidize to Cu²⁺(aq), and can reduce to Cu(s)

Using this table, we can easily predict spontaneous reactions, from before:

\[
2 \text{Ag}⁺(aq) + \text{Zn(s)} \rightarrow \text{Zn}²⁺(aq) + 2 \text{Ag(s)}
\]
\[
\text{OA} \quad \text{RA}
\]

All spontaneous reactions will have the OA higher on the left hand side of the table and the RA lower on the right hand side of the table.

\[
\text{Zn}²⁺(aq) + 2 \rightarrow \text{no reaction}
\]
\[
\text{OA} \quad \text{RA}
\]

Since the OA (Zn²⁺) is below the RA (Ag) on the table, Zn²⁺ is not a strong enough OA, and this reaction is not spontaneous.

Always look at the shape made between the OA and the RA – if it moves up, the reaction is not spontaneous, if it moves down, then the reaction is spontaneous.
ex: Which of the following containers could hold a solution of Co(NO$_3$)$_2$?

<table>
<thead>
<tr>
<th>Ni?</th>
<th>Au</th>
<th>Pb</th>
<th>Fe?</th>
</tr>
</thead>
</table>

The solution has both Co$^{2+}$ and NO$_3^-$ ions. For each metal container, compare it as an RA with the ions in the solution as the OA.

NO$_3^-$ (aq) is a spectator ion – it doesn’t appear on the table at all (by itself).

Co$^{2+}$ (aq) + 2e$^-$

from p. 9, only Pb and Fe would react with Co$^{2+}$, so both Ni and Au would hold the solution

Would your answer change if the solution were acidified?

<table>
<thead>
<tr>
<th>Au</th>
<th>Ni</th>
<th>Pb</th>
<th>Fe</th>
</tr>
</thead>
</table>

Yes! If the solution were acidified, then NO$_3^-$ (aq) is no longer a spectator, and now Ni(s) will also react, so only the gold container could hold the solution

Predicting Redox Reactions
- From a list of reactants, the likely reaction along with its spontaneity can be written (just like B-L reactions in the Acid Base unit), and the steps are almost identical
- Just like B-L reactions where only the strongest acid will react, and react only with the strongest base, the strongest oxidizing agent (SOA) will preferentially react with the strongest reducing agent (SRA)
- In the Acid Base unit, we drew a third arrow depicting the favoured equilibrium, in this unit, we will label the reaction as either spontaneous or nonspontaneous

*** for this type of problem, the half reactions will already be constructed and can be found on p.9, so all you will have to do is pull the right half reactions, and balance the electrons lost and gained.
ex: write the net ionic equation for the reaction between sulfuric acid, potassium permanganate and tin (II) chloride. Will the reaction be spontaneous?

**Rules:** (just like BL reactions)
1. write out all species as they really appear (in this unit, we don’t write $\text{H}_3\text{O}^+(aq)$ but $\text{H}^+(aq)$ for acids)
2. label all species as OA, RA or spectators (for this purpose, anything not on the chart will be a spectator)
3. determine the SOA and SRA
4. write the half reactions for them from p.9 (remember to flip the oxidation reaction)
5. balance the electrons by multiplying to get the LCM
6. add the half reactions
7. determine spontaneity using p.9

Redox Stoichiometry & Titration
- stoichiometry can be used just as with any other type of reaction, the only complicating factor is that since we most often use half reactions to get a balanced NET IONIC EQUATION instead of molecular equations, the molar ratios are slightly more complicated.

ex: *in an experiment, 25.0 mL of an unknown concentration of CaBr$_2$(aq) reacted with 15.0 mL of 0.0200 mol/L solution of KMnO$_4$(aq) in acid. Calculate the concentration of CaBr$_2$(aq).*

**Rules:**
1. write out all species as they really appear (in this unit, we don’t write $\text{H}_3\text{O}^+(aq)$ but $\text{H}^+(aq)$ for acids)
2. label all species as OA, RA or spectators (for this purpose, anything not on the chart will be a spectator)
3. determine the SOA and SRA
4. write the half reactions for them from p.9 (remember to flip the oxidation reaction)
5. balance the electrons by multiplying to get the LCM
6. add the half reactions
7. determine spontaneity using p.9

Spontaneous?

YES!
Now that we have the balanced equation \( \text{E} \) we can do the stoich (same as always)

\[
\frac{0.200\text{molKMnO}_4}{L} \times \frac{0.0150L}{1} \times \frac{1\text{molMnO}_4^-}{1\text{molKMnO}_4} \times \frac{10\text{molBr}^-}{2\text{molMnO}_4} \times \frac{1\text{molCaBr}_2}{2\text{molBr}^-} \times \frac{1}{0.0250L} = 3.00 \times 10^{-2} \text{mol/LCaBr}_2
\]

because we used a net ionic equation for the molar ratios, we have to includes these conversions to get back to the correct units for concentration as given in the question

Along with stoichiometry, we can make qualitative observations to go along with the calculations. For this reaction, the pH will increase (since \( \text{H}^+(aq) \) is consumed), and there will be a colour change from purple (\( \text{MnO}_4^-(aq) \)) to clear (\( \text{Mn}^{2+}(aq) \)) (see the second to last page of the data booklet for the colours)

**Redox Titration**
- unlike acid base titrations where we had a list of indicators to choose from, in redox titrations, we choose a titrant that will change colour as it reduces
- two popular choices are \( \text{KMnO}_4(aq) \) (with acid) and \( \text{K}_2\text{Cr}_2\text{O}_7(aq) \) (with acid) because they:
  - have a noticeable colour change
  - are strong OAs (when acidified)

The colour change observed at the equivalence point is slightly different than what we look for in Acid Base.

\[
\text{MnO}_4^-(aq) + 8 \text{H}^+(aq) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(aq) + 4 \text{H}_2\text{O}(l)
\]

purple \( \rightarrow \) clear

\[
\text{Cr}_2\text{O}_7^{2-}(aq) + 14 \text{H}^+(aq) + 6\text{e}^- \rightarrow 2 \text{Cr}^{3+}(aq) + 7 \text{H}_2\text{O}(l)
\]

orange \( \rightarrow \) blue

For the endpoint, since the “indicator” is a true reactant in the titration, we can’t look for the initial colour change because that SHOULD be taking place throughout the titration. Instead, we look for the first drop of titrant that does not react (ie: there is no more of the RA left in the Erlenmeyer to react)

for \( \text{MnO}_4^-(aq) \) \( \rightarrow \) the endpoint is a change to light purple (one drop of unreacted \( \text{MnO}_4^-(aq) \))

for \( \text{Cr}_2\text{O}_7^{2-}(aq) \) \( \rightarrow \) the endpoint is a change to blue + orange (pukey green) (one drop of unreacted \( \text{Cr}_2\text{O}_7^{2-}(aq) \) which is orange, plus the colour already in the Erlenmeyer, blue)