**Sec 5.1 - Oxidation – Reduction**

**Definitions:** (species means atom, ion or molecule)

- **Oxidation** – a species undergoing oxidation **loses electrons**
  (charge becomes more **positive**)

- **Reduction** – a species undergoing reduction **gains electrons**
  (charge becomes more **negative**)

- **Oxidizing agent** – The **species being reduced**
  (gains electrons, causes the other one to be oxidized)

- **Reducing agent** – The **species being oxidized**
  (loses electrons, causes the other one to be reduced)

\[ 2 \text{ e}^- \]

E.g.) \( \text{Cu}^{2+} (aq) + \text{Zn} (s) \rightarrow \text{Cu} (s) + \text{Zn}^{2+} (aq) \)

\[ \text{Oxidizing agent} \]

\**\text{Redox}\** – Short for **Oxidation – Reduction**

**Redox identification**

Charge on neutral atom or molecule = 0

- **Oxidation** – Charge gets more + (loses electrons)
- **Reduction** – Charge gets more − (gains electrons)

Reduction (charge decreases)

E.g.) \( \text{Pb}^{2+} (aq) + \text{Mg}^0 (s) \rightarrow \text{Pb}^0 (s) + \text{Mg}^{2+} (aq) \)

Oxidation (Charge increases)
**Question**

In the reaction: 
\[ 2\text{Fe}^{2+} + \text{Cl}_2 \rightarrow 2\text{Fe}^{3+} + 2\text{Cl}^- \]

Identify:

a) The Oxidizing Agent: \( \text{Cl}_2 \)

b) The species being oxidized: \( \text{Fe}^{2+} \)

c) The reducing agent: \( \text{Fe}^{2+} \)

d) The species being reduced: \( \text{Cl}_2 \)

e) The species gaining electrons: \( \text{Cl}_2 \)

f) The species losing electrons: \( \text{Fe}^{2+} \)

g) The product of oxidation: \( \text{Fe}^{3+} \)

h) The product of reduction: \( \text{Cl}^- \)

**Do Ex. 1 (a-e) pp. 192**

**Half-Reactions**

- Redox reactions can be broken up into oxidation & reduction half reactions.

  e.g.) Redox rxn: \( \text{Pb}^{2+}_{(aq)} + \text{Zn}_{(s)} \rightarrow \text{Pb}_{(s)} + \text{Zn}^{2+}_{(aq)} \)

  The \( \text{Pb}^{2+} \) (loses/gains) \( \text{2 electrons} \).

**Reduction Half-rx:** \( \text{Pb}^{2+}_{(aq)} + 2\text{e}^- \rightarrow \text{Pb}_{(s)} \)

Write the **oxidation** half reaction for the following redox rxn.
\( \text{Pb}^{2+}_{(aq)} + \text{Zn}_{(s)} \rightarrow \text{Pb}_{(s)} + \text{Zn}^{2+}_{(aq)} \)

Oxidation half rxn: \( \text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2\text{e}^- \)

(In oxidation reactions, e\(^{-}\)'s are _lost_ and are found on the _R_ side.) (LEO)

**Note:** Half-rxn's always have e\(^{-}\)'s, redox (oxidation-reduction) reactions never show e\(^{-}\)'s!

Given the redox reaction:
\[ \text{F}_2_{(g)} + \text{Sn}^{2+}_{(aq)} \rightarrow 2\text{F}^-_{(aq)} + \text{Sn}^{4+}_{(aq)} \]

Write the **oxidation** half-rx: \( \text{Sn}^{2+} \rightarrow \text{Sn}^{4+} + 2\text{e}^- \)

Write the **reduction** half-rx: \( \text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^- \)

**Do ex. 2 a-c on p. 192**
Sec 5.2 Oxidation numbers

- Real or apparent charge on an atom in a molecule or ion

Rules to find the oxidation number of an atom

1) In elemental form: (Single atoms of monatomic elements) or (diatomic molecules of diatomic elements)

   Oxidation number of atoms = 0     Eg.) Mn, Cr, N₂, F₂, Sn, O₂, etc.

2) In monatomic ions: oxidation # = charge ex; Na⁺ = +1

3) In ionic compounds
   a) the oxidation # of Alkali Metals is always +1   eg) NaCl and K₂CrO₄
   b) the oxidation # of Halogens when at the end (right side) of the formula is always -1   eg) CaCl₂  AlBr₃  KF

   Note: Halogens are not always -1! (Only when it is written last in formula.)

4) In molecules or polyatomic ions:
   a) Ox. # of oxygen is almost always -2  e.g.) KOH,  CrO₄²⁻, Li₃PO₄
   b) An exception is Peroxides in which ox. # of O = -1

   Hydrogen Peroxide: H₂O₂
   Alkali Peroxides: Na₂O₂

   (Remember, “O” in O₂ has an Ox. # of _ZERO_)

5) In molecules or ions:

   e.g.)  HNO₃,  H₂SO₄,  HPO₄²⁻  Every “H” has an ox # of +1

   e.g.)  NaH  CaH₂  (In each one of these Ox. # of H = -1)

   (What is the ox # of “H” in NH₃? __________)

   (And remember ox # of “H” in H₂ = ________)
Finding oxidation numbers of each atom in a molecule

In a **neutral molecule** the **total charge = 0**

- e.g.) NH₃ ← Total charge = 0 (no charge)

In a **polyatomic ion** – the **total ionic charge** is written on the **top right**

- e.g.) CrO₄²⁻

**Oxidation numbers** of all atoms **add up to total ionic charge (TIC)**

- e.g.) Find the oxidation # of Cr in CrO₄²⁻
  
  (Let x = oxidation # of one Cr atom)
  
  
  \[
  \begin{align*}
  \text{CrO}_4^{2-} & \quad \text{X} + 4 \text{ [# of “O” atoms]} (-2 \text{ [charge of oxygen]}) = -2 \text{ [total ionic charge]} \\
  \text{X} - 8 &= -2 \\
  \text{X} &= 6 \\
  \text{So ox # of Cr here} &= +6
  \end{align*}
  \]

- e.g.) Find ox # of Cl in HClO₄
  
  \[
  \begin{align*}
  \text{HClO}_4 & \quad +1 + x + 4 (-2) = 0 \\
  1 + x - 8 &= 0 \\
  x - 7 &= 0 \\
  x &= 7
  \end{align*}
  \]

- e.g.) Find Ox # of Cr in Cr₂O₇²⁻
  
  \[
  \begin{align*}
  \text{Cr}_2\text{O}_7^{2-} & \quad 2x + 7(-2) = -2 \\
  2x - 14 &= -2 \\
  2x &= +12 \\
  x &= 6
  \end{align*}
  \]

- e.g.) Find ox # of P in Li₃PO₄
  
  \[
  \begin{align*}
  \text{Li}_3\text{P}_4\text{O}_4 & \quad 3(+1) + x + 4 (-2) = 0 \\
  3 + x - 8 &= 0 \\
  x - 5 &= 0 \\
  x &= 5
  \end{align*}
  \]
Find Ox # of the underlined element in each of the following:

a) NaH$_2$PO$_4$

b) Na$_2$O$_2$

c) KH

Find the ox # of Fe in Fe$_3$O$_4$

Find the ox # of As in As$_3$O$_5$


Changes in oxidation numbers

When an atom’s oxidation # is increased, it is oxidized.

e.g.) Half-rxn: Fe$^{2+}$ $\rightarrow$ Fe$^{3+}$ + e$^{-}$

More complex:

-When Mn$^{3+}$ changes to MnO$_4^-$, is Mn oxidized or reduced?

  Mn$^{3+}$ $\rightarrow$ MnO$_4^-$  
  $^{+3}$ $\rightarrow$ $^{+7}$  
  $\times$ -8 = -1

  - What is the ox # of Mn before & after the reaction? Before $^{+3}$  
  After $^{+7}$

  - The ox # of Mn is (de/in) INcreased.

  - In this process, Mn is (oxidized/reduced) OXIDIZED

Reduction – When an atom’s oxidation # is decreased, it is reduced.

e.g.) Cu(NO$_3$)$_2$ $\rightarrow$ Cu(s)  Ox # decreases (reduction)
Redox ID using oxidation #'s

Given a more complex equation – identify atoms which do not change ox #'s
(often “O” or “H” but not always!)

e.g.) \[ 3\text{SO}_2 + 3\text{H}_2\text{O} + \text{ClO}_3^- \rightarrow 3\text{SO}_4^{2-} + 6\text{H}^+ + \text{Cl}^- \]

The only atoms left are “S” and “Cl”. Find the Ox #'s of S and Cl⁻ in species that contain them. (Ox # of 1 atom in each case)

\[ 3\text{SO}_2 \rightarrow 3\text{SO}_4^{2-} \]

\[ \text{SO}_2 \rightarrow \text{SO}_4^{2-} \]

Ox # of S is +4

Note:

- R.A.O., the reducing agent is oxidized
- The species SO₂ is acting as the reducing agent.
- The element S is being oxidized so S is losing electrons.

Look at the species with Cl: \[ \text{ClO}_3^- \rightarrow \text{Cl}^- \]

Decrease in ox # so Cl is being reduced

Therefore, the species acting as the oxidizing agent is ____________.
Eg. given the reaction:

\[ 2\text{CrO}_4^{2-} + 3\text{HCHO} + 2\text{H}_2\text{O} \rightarrow 2\text{Cr(OH)}_3 + 3\text{HCOO}^- + \text{OH}^- \]

Find:

a) The species being oxidized
b) The reducing agent
c) The species being reduced
d) The oxidizing agent
e) The species losing electrons
f) The species gaining electrons

Notes:
- For hydrocarbons it’s best to rewrite them as simple molecular formulas.
- All O’s are in molecules or ions, no O₂ & no peroxides so O remains unchanged as -2
- All H’s are in molecules or ions, no H₂ or metallic hydrides so H remains unchanged as +1
- The atoms to check for changes are C and Cr.

\[ \begin{array}{c}
0 & +2 \\
2\text{CrO}_4^{2-} + 3\text{CH}_2\text{O} + 2\text{H}_2\text{O} & \rightarrow 2\text{Cr(OH)}_3 + 3\text{HCOO}^- + \text{OH}^- \\
+6 & +3
\end{array} \]

So...
a) the species being oxidized is (CH₂O) HCHO (inc. in ox #)
b) the reducing agent is (CH₂O) HCHO (RAO)
c) The species being reduced is CrO₄²⁻ (decrease in ox #)
d) The oxidizing agent is CrO₄²⁻ (OAR)
e) The species losing e⁺s is (CH₂O) HCHO (LEO)
f) The species gaining e⁺s is CrO₄²⁻ (GER)

Given the redox reaction:

\[ 2\text{MnO}_4^- + 3\text{C}_2\text{O}_4^{2-} + 4\text{H}_2\text{O} \rightarrow 2\text{MnO}_2 + 6\text{CO}_2 + 8\text{OH}^- \]

Find:

a) The species being reduced: ________________.
b) The species undergoing oxidation: ________________.
c) The oxidizing agent: ________________.
d) The reducing agent: ________________.
e) The species gaining electrons: ________________.
f) The species losing electrons: ________________.
Given the balanced redox reaction:

\[ 3S + 4\text{HNO}_3 \rightarrow 3\text{SO}_2 + 4\text{NO} + 2\text{H}_2\text{O} \]

Find:

a) The oxidizing agent: ______________.
b) The reducing agent: ______________.
c) The species being reduced: ____________.
d) The species being oxidized: ____________.
e) The species losing electrons: ____________.
f) The species gaining electrons: ______________.
g) The product of oxidation: ______________.
h) The product of reduction: ______________.

Given the following:

\[ 6\text{Br}_2 + 12\text{KOH} \rightarrow 10\text{KBr} + 2\text{KBrO}_3 + 6\text{H}_2\text{O} \]

Find:

a) The oxidizing agent: ______________.
b) The reducing agent: ______________.
c) The species undergoing oxidation: ______________.
d) The species being reduced: ______________.
e) The product of oxidation: ______________.
f) The product of reduction: ______________.
Using oxidation numbers to identify half-reactions

They don’t have to be balanced

e.g.) If NO₂⁻ → NO₃⁻ is an example of (oxidation or reduction?) ________________.

(“O” does not change it’s ox # (no O₂ or peroxides)) so find ox # of N on both sides.

\[ \text{NO}_2^- \rightarrow \text{NO}_3^- \]

e.g.) \( \text{H}_2\text{O}_2 \rightarrow \text{H}_2\text{O} \)

Find the \( \Delta \text{Ox.#} \) of the element in which it changes and identify each as an oxidation or reduction

a) \( \text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOH} \) ____________________________

b) \( \text{Fe}_2\text{O}_3 \rightarrow \text{Fe}_3\text{O}_4 \) ____________________________

c) \( \text{H}_3\text{PO}_4 \rightarrow \text{P}_4 \) ____________________________

(P₄ is the elemental form of phosphorus)

d) \( \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COH} \) ____________________________

NOTE: When asked if a given reaction is a redox or not:

Look for a change from an element \( \rightarrow \) compound or compound \( \rightarrow \) an element

These will always be redox, because in elemental form ox. # = 0 and in compounds usually the ox. # is not = 0

Eg.) Is the reaction: \( \text{Zn} + \text{Cl}_2 \rightarrow \text{ZnCl}_2 \) a redox reaction? YES

Answer: It must be because \( \Delta \text{Ox#} \) of Zn (0 \( \rightarrow \) +2 = +2) and \( \Delta \text{Ox#} \) of Cl (0 \( \rightarrow \) -1 = -1)

Do Exercises 4, 5 and 6 on p. 194-195.
**Sec 5.3 - Half-reactions and the reduction table**

- Look at "Standard Reduction Table"

<table>
<thead>
<tr>
<th>Oxidizing agents on left + e⁻'s ⇌ Reducing agents on right</th>
</tr>
</thead>
<tbody>
<tr>
<td>Stronger ox agents (More tendency to be reduced) (To gain e⁻'s) (low tendency to lose e⁻'s)</td>
</tr>
<tr>
<td>F₂ + 2e⁻ ⇌ 2F⁻</td>
</tr>
<tr>
<td>Ag⁺ + e⁻ ⇌ Ag (s)</td>
</tr>
<tr>
<td>Cu²⁺ + 2e⁻ ⇌ Cu (s)</td>
</tr>
<tr>
<td>Zn²⁺ + 2e⁻ ⇌ Zn (s)</td>
</tr>
<tr>
<td>Li⁺ + e⁻ ⇌ Li (s)</td>
</tr>
<tr>
<td>Stronger reducing agents (More tendency to be oxidized RAO) (To lose e⁻'s) (low tendency to gain e⁻'s)</td>
</tr>
</tbody>
</table>

- So F₂ is a stronger ox. agent than Ag⁺, etc.
- The strongest reducing agent on your chart is: __Li⁺__.

### Help in Hunting

- Solid metals mostly on bottom right (less active ones Ag, Au, farther up on the right side)
- Halogens (e.g. Cl₂) and oxyanions e.g. BrO₃⁻, MnO₄⁻, IO₃⁻ found near top left
- Some metal ions found on **both sides** e.g. Fe²⁺, Sn²⁺, Cu⁺, Mn²⁺ **can act as OA's or RA's**

All the half-rxn's are written as **reductions:**

\[ \text{e.g.) } F₂ + 2e⁻ \rightleftharpoons 2F⁻ \]
\[ \text{Ag⁺ + e⁻} \rightleftharpoons \text{Ag(s)} \]

- The double arrow implies that **oxidation's** can also take place (reverse of reductions)

\[ \text{e.g.) reduction of Ag⁺ (Same as table- single arrow)} \]
\[ \text{Ag⁺ + e⁻ } \rightarrow \text{Ag(s)} \]

\[ \text{oxidation of Ag(Reverse of that on table- single arrow)} \]
\[ \text{Ag(s) } \rightarrow \text{Ag⁺ + e⁻} \]
Write half-reactions for:

- Reduction of Pb
  \[ \text{Pb} \rightarrow \text{Pb}^{2+} + 2e^- \]

- Oxidation of Pb
  \[ \text{Sn}^{2+} + 2e^- \rightarrow \text{Sn(s)} \]

- Reduction of Sn
  \[ \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \]

- Oxidation of Fe
  \[ \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \]

Which is a stronger oxidizing agent:

- Ni\(^{2+}\) or Ag\(^+\)?
  \[ \text{Ag}^+ \]

- Fe\(^{2+}\) or Cr\(^{3+}\)?
  \[ \text{Cr}^{3+} \]

- Sn\(^{2+}\) or Sn\(^{4+}\)?
  \[ \text{Sn}^{4+} \]

Which is a stronger reducing agent:

- Sn\(^{2+}\) or Fe\(^{2+}\)?
  \[ \text{Sn}^{2+} \]

- Zn or Ba?
  \[ \text{Ba} \]

- Cl\(^-\) or Br\(^-\)?
  \[ \text{Br}^- \]

- Fe\(^{2+}\) or Au?
  \[ \text{Fe}^{2+} \]

Which has a greater tendency to lose electrons, Ni or Zn?

Which has a greater tendency to gain electrons, Fe\(^{3+}\) or Cr\(^{3+}\)?

Which solid metal has the greatest tendency to lose e\(^-\)s?

Give the formula for an ion that is a stronger oxidizing agent that Ni\(^{2+}\), but is weaker than Pb\(^{2+}\)?
Using the reduction table to predict which reactions are spontaneous

An **oxidizing agent** will react **spontaneously** with (oxidize) a reducing agent below it on the right.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Reduction</th>
<th>Oxidation</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{F}_2(g) + 2e^- \rightleftharpoons 2\text{F}^- )</td>
<td>( \text{F}_2 ), the strongest OA, oxidize (react spontaneously with) all species below it on the right side from ( \text{SO}_4^{2-} ) all the way down to ( \text{Li(s)} )</td>
<td></td>
</tr>
<tr>
<td>( \text{S}_2\text{O}_8^{2-} + 2e^- \rightleftharpoons 2\text{SO}_4^{2-} )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>( \text{Li}^+ + e^- \rightleftharpoons \text{Li(s)} )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Look at your reduction chart!**

A **reducing agent** on the right will react **spontaneously** with (reduce) any oxidizing agent on the left above it.

e.g.) \( \text{Li(s)} \) (bottom right) will **reduce all species** on the left side except \( \text{Li}^+ \). \( \text{SO}_4^{2-} \) (near top right) will reduce **only** \( \text{F}_2 \).

- An OA on the left will **not** react spontaneously with a RA on the right above it! e.g.) \( \text{Au}^{3+} \) will **not** oxidize (or react spontaneously with) \( \text{SO}_4^{2-} \).

**Some points...**

1) Be very careful with charges e.g. \( \text{Li}^+ \) is a totally different thing than \( \text{Li(s)} \).

2) Things don’t react with species which are on the same side (these are **impossible** – not just non-spontaneous.)

E.g.) \( \text{K}^+ \) (4th from bottom on the left) will **not** oxidize \( \text{Rb}^+ \) or \( \text{Cs}^+ \) \( \text{Li}^+ \) etc. –because they are on the same side. (Impossible)

E.g.) \( \text{Li(s)} \) will **not** reduce \( \text{Cs(s)} \), \( \text{Rb(s)} \), \( \text{K(s)} \), etc. because they are all on the same side.

3) Some elements with **multiple oxidization** numbers e.g.) \( \text{Sn, Cu, Mn, Fe} \) have ions on both sides of the chart! – **Look carefully at your table to find these.**

**Note** – Don’t worry what \( E^0 \) means yet, I will just use it to let you locate half-reactions.

Notice:  
\( \text{Fe}^{2+} \) is on the left (OA) at \(-0.45 \)  
\( \text{Fe}^{2+} \) is on the right (RA) at \(+0.77 \)  
\( \text{Sn}^{2+} \) is on the left (OA) at \(-0.14 \)  
\( \text{Sn}^{2+} \) is on the right (RA) at \(+0.15 \)
A word about Cu...

Notice: Cu⁺ is on the left at + 0.52
Cu²⁺ is on the right at + 0.15
- recall that anything on the left will oxidize a species below it on the right.

\[
\begin{align*}
\text{Cu}^+ + e^- &= \text{Cu}^0, \quad E^o = 0.52 \\
\text{Cu}^{2+} + e^- &= \text{Cu}^+, \quad E^o = 0.15
\end{align*}
\]

Since Cu⁺ oxidizes and reduces itself, any water solution of Cu⁺ is unstable

Notice: Mn²⁺ is on the left at \(E^o = -1.19\)  
Mn²⁺ is on the right at \(E^o = +1.22\)  

Also notice: \(\text{Cr}^{3+} + e^- = \text{Cr}^{2+} \quad -0.41\)  
\(\text{Cr}^{3+} + 3e^- = \text{Cr}^{(s)} \quad -0.74\)

If a redox reaction is non-spontaneous, then the reverse reaction will be spontaneous!

\(\text{e.g.})\) The reaction \(\text{Sr}^{2+} + \text{Ca}^{(s)} \rightarrow \text{Ca}^{2+} + \text{Sr}^{(s)}\) is non-spontaneous because Ca is above Sr²⁺ on the right side.

But the rx: \(\text{Ca}^{2+} + \text{Sr}^{(s)} \rightarrow \text{Sr}^{2+} + \text{Ca}^{(s)}\) is spontaneous because Sr(s) is below Ca²⁺ on the right side

Use the reduction table to answer the following questions:

a) Will \(\text{Br}_2\) oxidize \(\text{Au}^{(s)}\) ?

b) Will \(\text{Pb}^{(s)}\) reduce \(\text{Fe}^{2+}\) ?

c) Will \(\text{Zn}^{2+}\) react with \(\text{Cr}^{3+}\) ?

d) Will \(\text{Mg}^{2+}\) react with \(\text{Cr}^{3+}\) ?

e) Give the symbol of an ion that will oxidize \(\text{Mn}^{(s)}\) but not \(\text{Cr}^{(s)}\).......

f) Give the formula for a compound which will reduce \(\text{Co}^{2+}\) but will not reduce \(\text{Fe}^{2+}\) .................

\[\text{H}_2\text{Se}\]

g) Which is a stronger reducing agent, \(\text{Sn}^{2+}\) or \(\text{Fe}^{2+}\)? (Hint – you must look for both on the \_R\_ side).......

\[\text{Sn}^{2+}\]

h) Which is a stronger oxidizing agent, \(\text{Cu}^+\) or \(\text{Sn}^{2+}\)? (Hint – you must look for both on the \_L\_ side)..............

\[\text{Cu}^+\]
**Acidified solutions**

- Any reactions on the table with H\(^+\) in them are **acidified** or **acid solutions**.

  e.g.) Look at these: at E\(^o\) = +1.51V (4\(^{th}\) from the top)

\[
\text{MnO}_4^- + 8H^+ + 5e^- \leftrightarrow \text{Mn}^{2+} + 4H_2O
\]

**Give the E\(^o\) corresponding to each of the following:**

a) acidified iodate ........... E\(^o\) = +1.20
b) acidified dichromate........... E\(^o\) = +1.23
c) acidified manganese (IV) oxide... E\(^o\) = +1.22
d) acidified bromate.............. E\(^o\) = +1.48
e) acidified perchlorate........... E\(^o\) = +1.39
f) acidified oxygen gas............ E\(^o\) = +1.23

**Nitric, Sulphuric & Phosphoric acids**

- These acids are shown in **ionized form** on the table
- Nitric acid (HNO\(_3\)) is found in two places on the left side.

\[
\text{HNO}_3 \leftrightarrow \text{NO}_3^- + 4H^+ + 3e^- \quad \text{NO} + 2H_2O \quad \text{E}^o = +0.96 \text{ V}
\]

\[
2\text{NO}_3^- + 4H^+ + 2e^- \leftrightarrow \text{N}_2\text{O}_4 + 2H_2O \quad \text{E}^o = +0.80 \text{ V}
\]

- Sulphuric acid is found at + 0.17 V

\[
\text{SO}_4^{2-} + 4H^+ + 2e^- \leftrightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O} \quad \text{E}^o = +0.17 \text{ V}
\]

Find and write the half-reaction for the reduction of **phosphoric** acid (H\(_3\)PO\(_4\))

\[
\text{H}_3\text{PO}_4 + 2H^+ + 2e^- \rightarrow \text{H}_3\text{PO}_3 + \text{H}_2\text{O} \quad -0.28 \text{ V}
\]

Sulphurous acid (H\(_2\)SO\(_3\))

\[
\text{H}_2\text{SO}_3 + 4H^+ + 4e^- \rightarrow \text{S} + 3\text{H}_2\text{O} + 0.45 \text{ V}
\]
A note about water

- On the top of the table it says “ionic concentrations are at 1M”
- This includes [H\(^+\)] = 1M with two exceptions:
  - Neutral water is found on the shaded lines at + 0.82 V and – 0.41 V
  - Neutral water as a reducing agent is on the right side at + 0.82 V
  - Neutral water as an oxidizing agent is on the left side at – 0.41 V

(Notice H\(_2\)O is below this at – 0.83 V, but in this solution [OH\(^-\)] = 1M (so it’s basic, not neutral)

(Again H\(_2\)O is also found at + 1.23 V but here [H\(^+\)] = 1M so it’s acidic, not neutral)

Questions

a) Will neutral water oxidize Fe(s)? ___Y____ Cr(s)? ___Y___ Na(s)? ___Y___
b) Will neutral water reduce Au\(^{3+}\)? ___Y___ Ag\(^+\)? ___N___
c) Will acidified permanganate oxidize SO\(_4^{2-}\)? ___N___ Br\(^-\)? ___Y___ Zn? ___Y___
d) Will nitric acid react with Ag(s)? ___Y___ Au(s)? ___N___ I\(^-\)? ___Y___ Cl\(^-\)? ___N___
e) Will nitric acid react with Fe\(^{2+}\)? ___Y___
f) Will nitric acid react with Hg to form N\(_2\)O\(_4\)? ___Y___
g) Will nitric acid react with Hg to form NO? ___Y___
h) Can you safely put a gold ring in acidified dichromate solution? ___Y___ What about acidified bromate solution? ___Y___
i) If Cl\(_2\) gas is bubbled into water, will it all remain as Cl\(_2\), or will some be converted to ___Cl\(^-\)___

Finding products of spontaneous reactions

eg) Given Sn\(^{4+}\) + H\(_2\)S, find the products

See the table at +0.15V and +0.14V

\[
\begin{align*}
\text{Sn}^{4+} + 2e^- & \rightleftharpoons \text{Sn}^{2+} + 0.15V \\
\text{S(s)} + 2\text{H}^+ + 2e^- & \rightleftharpoons \text{H}_2\text{S} + 0.14V
\end{align*}
\]

The higher reaction will be reduction (\(\rightarrow\)), the lower reaction will proceed to the left (\(\leftarrow\)) and be an oxidation.

\[
\begin{align*}
\text{Sn}^{4+} + 2e^- & \rightarrow \text{Sn}^{2+} \\
\text{S(s)} + 2\text{H}^+ + 2e^- & \leftarrow \text{H}_2\text{S (reversed! Lower one is reversed- is an oxidation)}
\end{align*}
\]

-So the products are Sn\(^{2+}\), S, and H\(^+\)

\[
\text{Sn}^{4+} + \text{H}_2\text{S} \rightarrow \text{S(s)} + 2\text{H}^+ + \text{Sn}^{2+}
\]
Questions

a) What are the products of the reaction of acidified hydrogen peroxide (H₂O₂) and bromide (Br⁻)? _____H₂O , Br₂_________

b) What are the products of the reaction when neutral water reacts with:
   Ca(s) ___Ca²⁺, H₂, OH⁻ (10⁻⁷)
   Zn(s) ___Zn²⁺, H₂, OH⁻ (10⁻⁷)
   Br₂ ___½ O₂, 2H⁺, 2Br⁻
   Acidified MnO₂ ___Mn²⁺, H₂O, ½O₂, H⁺
   Fluorine gas ___F⁻, ½ O₂, 2H⁺

Read text p. 195-199
Do Ex 7-12 p, 199-200