Chemistry 12 – Unit 5

Oxidation – Reduction

Introduction

-Demonstration of oxidation – reduction reactions

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

LEO says GER

Losing Electrons is Oxidation
Gaining Electrons is Reduction

OAR

The Oxidizing Agent is Reduced

To carry it too far...

When LEO the Lion says GER you grab your OAR and
Row Away Outa’ there!
(Reducing Agent is Oxidized)
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.) \( \text{Pb}^{2+} (\text{aq}) + \text{Mg}^{0} (\text{s}) \rightarrow \text{Pb}^{0} (\text{s}) + \text{Mg}^{2+} (\text{aq}) \)

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:_____________________
b) The species being oxidized:________________c) The reducing agent:____________________d) The species being reduced:________________e) The species gaining electrons:________________f) The species losing electrons:________________g) The product of oxidation____________________h) The product of reduction___________________

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

Half-Reactions

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

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

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

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

Electrons on the LEFT side (or GER) Means REDUCTION
Write the **oxidation** half reaction for the following redox rx.

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

Ox half rx: ____________________________

(In oxidation reactions, e⁻'s are ____ and are found on the ____ side.) (LEO)

**Note:** Half-rx'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:__________________________

Write the **reduction** half-rx:__________________________

**Do ex. 2 a-c on p. 192 SW**

**Oxidation numbers**

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

In SW, p. 193 - the charge that an atom would possess if the species containing the atom was made up of ions (even if it's not!)

**Rules to find 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.

   The oxidation # of each atom = 0

2) **In monatomic ions:** oxidation # = charge

   Eg) \( \text{In Cr}^{3+} \)-oxidation # of Cr = +3

   \( \text{S}^{2-} \)-oxidation # of S = -2
3) In **ionic compounds**
   a) the oxidation # of **Alkali Metals** is always \(+1\)
      
      eg) \(\text{NaCl} \quad \text{K}_2\text{CrO}_4\)
      
      Ox # of Na & K = +1
   
   b) the oxidation # of **Halogens** when at the end (right side) of the formula is always \(-1\)
      
      eg) \(\text{CaCl}_2 \quad \text{AlBr}_3 \quad \text{KF}\)
      
      Ox # of Cl, Br and F = -1
      
      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.) \(\text{KOH} \quad \text{CrO}_4^{2-} \quad \text{Li}_3\text{PO}_4\)
      
      Ox # of O is \(-2\)
   
   b) An exception is **Peroxides** in which ox. # of O = \(-1\)
      
      **Hydrogen Peroxide:** \(\text{H}_2\text{O}_2\)  
      
      ...Ox # of O's = -1
      
      **Alkali Peroxides:** \(\text{Na}_2\text{O}_2\)
      
      (Remember, “O” in \(\text{O}_2\) has an Ox. # of ______)

5) In **molecules** or **ions**:
   a) **Hydrogen** is almost always \(+1\)
      
      e.g.) \(\text{HNO}_3 \quad \text{H}_2\text{SO}_4 \quad \text{HPO}_4^{2-}\) Every “H” has an ox # of \(+1\)
      
      b) An exception is **metallic hydrides**, which have an ox # of \(-1\)
      
      e.g.) \(\text{NaH} \quad \text{CaH}_2\) (In each one of these Ox. # of H = \(-1\)
      
      (What is the ox # of “H” in \(\text{NH}_3\)? ______)
      
      (And remember ox # of “H” in \(\text{H}_2 = ______\)
Finding oxidation numbers of each atom in a molecule or PAI

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⁻ ← Total ionic charge (TIC) = -2

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

e.g.) Find the oxidation # of Cr in CrO₄²⁻:

(Let x = ox # of one Cr atom)

\[
\text{CrO}_4^{2-}:
\]

\[
X + 4 \text{ [# of “O”atoms]} (-2 \text{ [charge of oxygen]}) = -2 \text{ [total ionic charge]}
\]

\[
X - 8 = -2
\]

\[
X = -2 + 8
\]

\[
X = +6 \text{ So ox # of Cr here} = +6
\]

e.g.) Find ox # of Cl in HClO₄

\[
\text{HClO}_4
\]

\[
+1 + x + 4 (-2) = 0
\]

\[
1 + x - 8 = 0
\]

\[
x - 7 = 0
\]

\[
x = +7
\]

e.g.) Find Ox # of Cr in Cr₂O₇²⁻:

\[
\text{Cr}_2\text{O}_7^{2-}
\]

\[
2x + 7(-2) = -2
\]

\[
2x - 14 = -2
\]

\[
2x = +12
\]

\[
x = +6
\]
e.g.) Find ox # of P in Li₃PO₄

\[ \text{Li}_3 \text{P} \text{O}_4 \]

\[ 3(+1) + x + 4(-2) = 0 \]

\[ 3 + x - 8 = 0 \]

\[ x - 5 = 0 \]

\[ x = +5 \]

Find Ox # of the underlined element in each of the following:

a) NaH₂PO₄ _____  b) Na₂O₂ _____  c) KH ______

Find the ox # of Fe in Fe₂O₄

Find the ox # of As in As₂O₅

Read p. 193-194 of SW. Do Exercise 3 on p. 194 of SW.

**Changes in oxidation numbers**

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

e.g.) Half-rx: \( \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \)

More complex:

- When Mn⁴⁺ changes to MnO₄⁻, is Mn oxidized or reduced?

\( \text{Mn}^{3+} \rightarrow \text{MnO}_4^- \)

- What is the ox # of Mn before & after the reaction? Before _____ After _____

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

- In this process, Mn is (oxidized/reduced)__________________
**Reduction** – When an atom’s oxidation # is decreased, it is reduced.

E.g.) \( \text{Cu(NO}_3\text{)}_2 \rightarrow \text{Cu(s)} \)

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

3SO₂ + 3H₂O + ClO₃⁻ → 3SO₄²⁻ + 6H⁺ + Cl⁻

There are no O₂ molecules or peroxides, so “O” in all these has an ox # = -2

H is (+1) in both of these so it doesn’t change

Again:

\( 3\text{SO}_2 + 3\text{H}_2\text{O} + \text{ClO}_3^- \rightarrow 3\text{SO}_4^{2-} + 6\text{H}^+ + \text{Cl}^- \)

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)

\( 3\text{SO}_2 \rightarrow 3\text{SO}_4^{2-} \)

Coefficients are just for balancing.

SO₂ → SO₄²⁻

Ox # of S is +4

Ox # of S is +6

**Ox # of S increases** so S is being oxidized

Ox # of Cu = +2

Ox # of Cu = 0
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:

\[
\begin{array}{c|c}
\text{ClO}_3^- & \text{Cl}^-
\end{array}
\]

<table>
<thead>
<tr>
<th>Ox. # of Cl</th>
<th>Ox. # of Cl</th>
</tr>
</thead>
<tbody>
<tr>
<td>+5</td>
<td>-1</td>
</tr>
</tbody>
</table>

Decrease in ox # so Cl is being reduced

Therefore, the species acting as the oxidizing agent is \[\text{ClO}_3^-.\]

(They may also ask for the atom acting as the oxidizing agent
– this would be Cl in ClO₃⁻)

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 
   c) The reducing agent 
   d) The species being reduced 
   e) The oxidizing agent 
   f) The species losing electrons 
   g) 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.

\[
\begin{array}{c|c|c}
\text{C} & \text{Cr} & \text{Ox}.
\end{array}
\]

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>+2</th>
</tr>
</thead>
<tbody>
<tr>
<td>2CrO₄²⁻ + 3CH₂O + 2H₂O</td>
<td>2Cr(OH)₃ + 3HCO₂⁻ + OH⁻</td>
</tr>
</tbody>
</table>

Reduction | +3 |

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:

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

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 \( \text{NO}_2^- \rightarrow \text{NO}_3^- \) 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.

\[
\text{NO}_2^- \rightarrow \text{NO}_3^- \\
\text{Ox #} = +3 \quad \text{Ox #} = +5 \\
\Delta \text{ O.N.} = +2
\]

Since ox # increases, this is an oxidation
Find the Δ O.N. of the element in which it changes and identify each as an oxidation or reduction

a) $C_2H_5OH \rightarrow CH_3COOH$ ________________________________

b) $Fe_2O_3 \rightarrow Fe_3O_4$ ________________________________

c) $H_3PO_4 \rightarrow P_4$ ________________________________

(P₄ is the elemental form of phosphorus)

d) $CH_3COOH \rightarrow CH_3COH$ ________________________________

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 ox. # is not = 0

Eg.) Is the reaction: Zn + Cl₂ $\rightarrow$ ZnCl₂ a redox reaction?

Answer: It must be because $\Delta$ON of Zn (0 $\rightarrow$ +2 = +2) and $\Delta$ON of Cl (0 $\rightarrow$ -1 = -1)

Do Exercises 4, 5 and 6 on p. 194-195 of SW.

Half-reactions and the reduction table
- Do Experiment 21-A
- Look at “Standard Reduction Table”

<table>
<thead>
<tr>
<th>Ox agents on left + e⁻'s</th>
<th>Reducing agents on right</th>
</tr>
</thead>
<tbody>
<tr>
<td>F₂ + 2e⁻ ⇌ 2F⁻</td>
<td>Ag⁺ + e⁻ ⇌ Ag (s)</td>
</tr>
<tr>
<td>Cu²⁺ + 2e⁻ ⇌ Cu(s)</td>
<td>Zn²⁺ + 2e⁻ ⇌ Zn (s)</td>
</tr>
<tr>
<td>Li⁺ + e⁻ ⇌ Li (s)</td>
<td></td>
</tr>
</tbody>
</table>

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

**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-rx’s are written as reductions:
e.g.) F₂ + 2e⁻ ⇌ 2F⁻
     Ag⁺ + e⁻ ⇌ Ag (s)
- The double arrow implies that oxidation’s can also take place (reverse of reductions)
e.g.) reduction of Ag⁺
     Ag⁺ + e⁻ → Ag (s)
     (Same as table- single arrow)
     oxidation of Ag
     Ag (s) → Ag⁺ + e⁻
     (Reverse of that on table- single arrow)

Write half-reactions for:
- Reduction of Pb²⁺
  - Oxidation of Pb
  - Reduction of Sn²⁺
  - Oxidation of Sn⁴⁺
- Oxidation of Fe$^{2+}$ _______________________________
- Reduction of Fe$^{2+}$ _______________________________
- Oxidation of Fe$^{3+}$ _______________________________
- Reduction of acidified MnO$_4^-$ _______________________
- Oxidation of H$^+_2$ _______________________________

Which is a stronger oxidizing agent: Ni$^{2+}$ or Ag$^+$? ________
Fe$^{2+}$ or Cr$^{3+}$? ________
Sn$^{2+}$ or Sn$^{4+}$? ________

Which is a stronger reducing agent: Sn$^{2+}$ or 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 least tendency to lose electrons? ________
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>Reduced</th>
<th>Oxidized</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>F$^-$</td>
<td>F$_2$(g)</td>
<td>F$_2$(g) + 2e$^-$ $\rightleftharpoons$ 2F$^-$</td>
</tr>
<tr>
<td>SO$_4^{2-}$</td>
<td>S$_2$O$_8^{2-}$</td>
<td>S$_2$O$_8^{2-}$ + 2e$^-$ $\rightleftharpoons$ 2SO$_4^{2-}$</td>
</tr>
<tr>
<td>Li$^+$</td>
<td>Li$^+_$(s)</td>
<td>Li$^+_$(s) + e$^-$ $\rightleftharpoons$ Li$(s)$</td>
</tr>
</tbody>
</table>

Look at the reduction chart!

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

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

- An OA on the left will not react spontaneously with a RA on the right above it!

  e.g.) Au$^{3+}$ will not oxidize (or react spontaneously with) SO$_4^{2-}$. 
Some points…

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

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

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

E.g.) Li(s) will not reduce Cs(s), Rb(s), K(s), etc. because they are all on the same side only.

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

Notice: Fe²⁺ is on the left (OA) at – 0.45
       Fe²⁺ is on the right (RA) at + 0.77
       Sn²⁺ is on the left (OA) at – 0.14
       Sn²⁺ is on the right (RA) at + 0.15

A word about Cu…

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

This Cu⁺ will oxidize this Cu⁺

Since Cu⁺ oxidizes and reduces itself, any water solution of Cu⁺ is unstable – it won’t remain Cu⁺ very long!

(demo Cu in HNO₃)

Notice: Mn²⁺ is on the left at E⁰ = -1.19
       Mn²⁺ is on the right at E⁰ = +1.22

Also notice: Cr³⁺ + e⁻ = Cr²⁺ - 0.41
             Cr³⁺ + 3e⁻ = Cr(s) - 0.74
If a redox reaction is non-spontaneous, then the reverse reaction will be spontaneous!

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

But the rx: $\text{Ca}^{2+} + \text{Sr}(s) \rightarrow \text{Sr}^{2+} + \text{Ca}(s)$ is spontaneous because $\text{Sr}(s)$ is below $\text{Ca}^{2+}$ 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 $\text{Fe}^{2+}$____________

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

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

Acidified solutions

- Any reactions on the table with $\text{H}^+$ in them are acidified or acid solutions.

e.g.) Look at these: at $E^0 = +1.51$ (4th from the top)

$$\text{MnO}_4^- + 8\text{H}^+ + 5e^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$$

Called acidified permanganate solution

Note: Names of many ions can be found on the ion table!

Give the $E^0$ corresponding to each of the following:

a) acidified iodate $E^0$________________

b) acidified dichromate $E^0$________________

c) acidified manganese (IV) oxide $E^0$________________

d) acidified bromate $E^0$________________

e) acidified perchlorate $E^0$________________

f) acidified oxygen gas $E^0$________________
Nitric, Sulphuric & Phosphoric acids

- These acids are shown in **ionized form** on the table
- Nitric acid (HNO₃) is found in two places on the left side.

\[
\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{NO} + 2\text{H}_2\text{O} \quad E^0 = +0.96 \text{ v}
\]

\[
2\text{NO}_3^- + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{N}_2\text{O}_4 + 2\text{H}_2\text{O} \quad E^0 = +0.80 \text{ v}
\]

Don’t worry about coefficients yet. They are only used for balancing.

- Sulphuric acid is found at + 0.17 v

\[
\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{SO}_3 + \text{H}_2\text{O} \quad E^0 = +0.17 \text{ v}
\]

Find and write the half-reaction for the reduction of **phosphoric** acid (H₃PO₄)

Sulphurous acid (H₂SO₃)

**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 H₂O is below this at – 0.83 v but in this solution [OH⁻] = 1M so it’s **basic**, **not** neutral)

(Again H₂O 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)? _______ Cr(s)? _______ Na(s)? _______
b) Will neutral water reduce Au\(^{3+}\)? _______ Ag\(^+\)? _______
c) Will acidified permanganate oxidize SO\(_4^{2-}\)? _______ Br\(^-\)? _______ Zn? _______
d) Will nitric acid react with Ag(s)? _______ Au(s)? _______ I\(^-\)? _______ Cl\(^-\)? _______
e) Will nitric acid react with Fe\(^{2+}\)?
f) Will nitric acid react with Hg to form N\(_2\)O\(_4\)? _______
g) Will nitric acid react with Hg to form NO? _______
h) Can you safely put a gold ring in acidified dichromate solution? _______ What about acidified bromate solution? _______
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 \(\text{Sn}^{4+} + \text{H}_2\text{S} \rightarrow \) find the products
See the table at +0.15v and +0.14v
\[
\begin{align*}
\text{Sn}^{4+} + 2e^- & \rightarrow \text{Sn}^{2+} + 0.15v \\
\text{S(s)} + 2\text{H}^+ + 2e^- & \rightarrow \text{H}_2\text{S} + 0.14v \\
\end{align*}
\]
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\(^+\)
(at this point don’t worry about coefficients yet.)

Questions
a) What are the products of the reaction of acidified hydrogen peroxide (H\(_2\)O\(_2\)) and bromide (Br\(^-\))? ________________
b) What are the products of the reaction when neutral water reacts with:
\[
\begin{align*}
\text{Ca(s)} & \\
\text{Zn(s)} & \\
\text{Br}_2 & \\
\text{Acidified MnO}_2 & \\
\text{Fluorine gas} & \\
\end{align*}
\]

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