Solutions _ Acids & Bases

Solute ← → solvent

Homogenous → mixture that the molecules of the solute is not visible → salty water

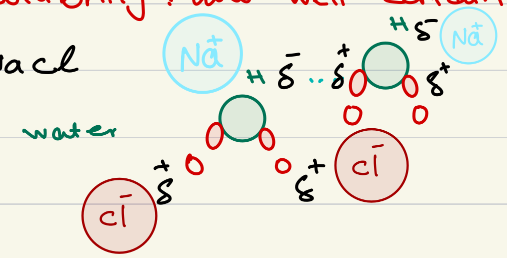
Non-homogenous → sand + water → the concentration of sand is not uniform everywhere

when water is "Solvent" → aqueous solutions (aq)

- A dilute solution contains relatively low amount of solute compared to a concentrated solution which contains a higher amount of solute in the same volume of solvent.

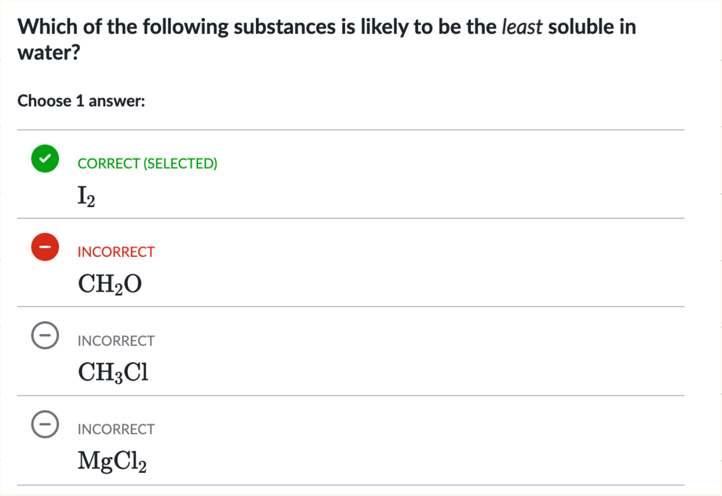
Solubility : how well certain solutes that can dissolve in certain solvents.

NaCl



since NaCl is an ionic Compound it can dissociate into its Constituent ions → Na^+, Cl^-

- If a Compound is ionic → dissolves well
- If a Compound has some polarity → dissolves well
- Molecular compounds that are non-polar → don't dissolve well



Which of the following substances is likely to be the *least* soluble in water?

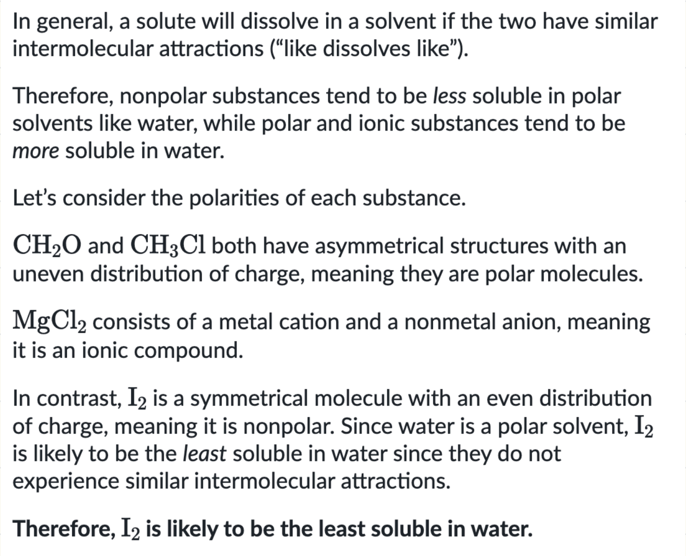
Choose 1 answer:

CORRECT (SELECTED)
I_2

INCORRECT
CH_2O

INCORRECT
CH_3Cl

INCORRECT
$MgCl_2$



In general, a solute will dissolve in a solvent if the two have similar intermolecular attractions ("like dissolves like").

Therefore, nonpolar substances tend to be *less* soluble in polar solvents like water, while polar and ionic substances tend to be *more* soluble in water.

Let's consider the polarities of each substance.

CH_2O and CH_3Cl both have asymmetrical structures with an uneven distribution of charge, meaning they are polar molecules.

$MgCl_2$ consists of a metal cation and a nonmetal anion, meaning it is an ionic compound.

In contrast, I_2 is a symmetrical molecule with an even distribution of charge, meaning it is nonpolar. Since water is a polar solvent, I_2 is likely to be the *least* soluble in water since they do not experience similar intermolecular attractions.

Therefore, I_2 is likely to be the least soluble in water.

Weak electrolytes partially dissociate into ions in solution, resulting in moderate conductivity.
Weak electrolytes partially dissociate to produce relatively few ions in solution, resulting in moderate conductivity.

→ like citric acid → $C_6H_8O_7$ → organic & Colorless chemistry 30-1 60 on Diploma 18 questions of Electrochem

Electro Chem

Reduction / Oxidation

losing e⁻ → oxidize

$$2\overset{0}{Al}_{(s)} + 3\overset{0}{Cl}_{2(g)} \rightarrow 2\overset{3+}{Al}\overset{1-}{Cl}_{3}\ {(s)}$$

gaining e⁻ → reduced

anything in its elemental form → oxidation number 0.

LEO GER

LEO ↓ losing e⁻: Ox

GER ↳ gaining e⁻: Red

losing e⁻ → oxidize

$$\overset{2+}{Cu}\overset{2-}{S}_{(s)} + \overset{0}{H}_{2(g)} \rightarrow \overset{0}{Cu}_{(s)} + \overset{1+}{H}_2\overset{2-}{S}_{(g)}$$

gaining e⁻ → reduced

5e⁻

$$① \overset{+7}{Mn}\overset{-2}{O}_4^{-} + ⑤ \overset{+1}{H}_2\overset{+3}{C}_2\overset{-2}{O}_4 \rightarrow 10\overset{+4}{C}\overset{-2}{O}_2 + 1\overset{+2}{Mn}^{2+} + ---$$

5 × e⁻

---

1 , 3 , 4 , 2

$4Fe + 3O_2 + H_2O \longrightarrow Fe(OH)_3 + \text{energy}$

20g V=?

SATP

$20g\ Fe \times \frac{1\ mol}{55.845g} \times \frac{3O_2}{4\ mFe} = 0.2686\ mol$

1 mol 24.8 L

0.2686 ↓ 6.66129 L

$2e^-$

$3Cu + 2\overset{+6}{Cr}_2\overset{-2}{O}_7^{2-} + 14H^+ \longrightarrow Cu^{2+} + 2\overset{+3}{Cr}^{3+}$

$+3e^-$ H_2O

$2O - 14 = -2$

$+14$ $+4$

12

Section 2: Identifying Oxidation and Reduction

Definition (Application-Based): In exams, you won't be asked to define oxidation or reduction in words. Instead, you'll need to *apply* the concepts to chemical reactions.

Strategy: Use Oxidation Numbers

The most reliable way to identify oxidation and reduction is by tracking changes in **oxidation numbers**.

- **Elemental Form:** Any element in its uncombined, elemental form has an oxidation state of zero.
    - **Sample:** Al(s), Cl_2(g), O_2(g), H_2(g), Na(s) all have an oxidation state of 0.
- **Ionic Compounds:** Determine oxidation states based on the known charges of the ions.
    - **Sample:** In $AlCl_3$, aluminum is Al^{3+} (oxidation state +3) and chlorine is Cl^- (oxidation state -1).
    - **Sample:** In CuS, sulfur is S^{2-} (oxidation state -2), so copper must be Cu^{2+} (oxidation state +2).
- **Change in Oxidation State:**
    - **Oxidation: Losing electrons**. The oxidation number **increases**.
        - *Mnemonic:* **L**EO - **L**osing **E**lectrons **O**xidized
        - **Sample:** Al(s) (0) → Al^{3+} (+3). Aluminum loses 3 electrons, and its oxidation number increases from 0 to +3. Aluminum is oxidized.
    - **Reduction: Gaining electrons**. The oxidation number **decreases**.
        - *Mnemonic:* G**ER** - **G**aining **E**lectrons **R**educed
        - **Sample:** Cl_2(g) (0) → $2Cl^-$ (-1). Each chlorine atom gains 1 electron, and its oxidation number decreases from 0 to -1. Chlorine is reduced.

Examples:

- **Example 1: $Al + Cl_2 \rightarrow AlCl_3$**
    - Initial states: Al (0), Cl (0)
    - Final states: Al (+3), Cl (-1)
    - **Aluminum (Al):** Goes from 0 to +3. It loses 3 electrons. **Aluminum is oxidized.**
    - **Chlorine (Cl_2):** Goes from 0 to -1. It gains 1 electron per atom. **Chlorine is reduced.**
- **Example 2: $CuS + H_2 \rightarrow Cu + H_2S$**
    - Initial states: Cu (+2) (in CuS), S (-2) (in CuS), H (0) (in H_2)
    - Final states: Cu (0), H (+1) (in H_2S), S (-2) (in H_2S)
    - **Copper (Cu):** Goes from +2 (in CuS) to 0 (elemental Cu). It gains 2 electrons. **Copper is reduced.**
    - **Hydrogen (H_2):** Goes from 0 (elemental H_2) to +1 (in H_2S). It loses 1 electron per atom. **Hydrogen is oxidized.**

   - **Sulfur (S):** Remains -2 throughout the reaction. Sulfur is neither oxidized nor reduced (no redox change).

Application to Voltaic Cells:

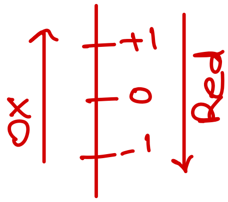
Voltaic (or Galvanic) cells are electrochemical cells that generate electrical energy from spontaneous redox reactions.

For the table, you don't need to check the electron states or the oxidation number. Higher half-raction is the best oxidizing agent; which means it will be reduced!

- **Key Principle:** Voltaic cells must involve two complete half-reactions.
- **Using the Standard Reduction Table:** This table lists various half-reactions and their standard reduction potentials.
  ↳ page 7 of the data booklet
   - The **higher** half-reaction on the table will involve the species that is the **best oxidizing agent** (it will be **reduced**).
   - The **lower** half-reaction on the table will involve the species that will be **oxidized**.
- **Example: Copper-Zinc Cell**
   - On a standard reduction table, the reduction of Cu^{2+} is typically higher than the reduction of Zn^{2+}.
   - Therefore:
      - **Cu^{2+} gets reduced** (it is the oxidizing agent).
      - **Zn gets oxidized** (it is the reducing agent).

Section 3: Deep Dive into Oxidation States, Half-Reactions, and Disproportionation

The Number Line for Oxidation States:



This visual aid helps determine if an atom is oxidized or reduced.

- **Going UP in charge (more positive):** Means electrons are **lost** → **Oxidized**
   - *Example:* From -4 to +4 (Carbon in CH_4 to CO_2)
- **Going DOWN in charge (more negative):** Means electrons are **gained** → **Reduced**
   - *Example:* From 0 to -2 (Oxygen in O_2 to H_2O)

Identifying Oxidation States in Compounds (A Guaranteed Diploma Question!):

This skill is crucial for success. You must be able to assign oxidation states correctly for various elements within compounds.

- **General Rules:**
   - **Oxygen:** Usually has an oxidation state of **-2** in compounds.
      - *Exception:* In peroxides (e.g., H_2O_2, Na_2O_2), oxygen is **-1**. (These exceptions will usually be clearly stated or identifiable.)
   - **Hydrogen:** Usually has an oxidation state of **+1** in compounds.

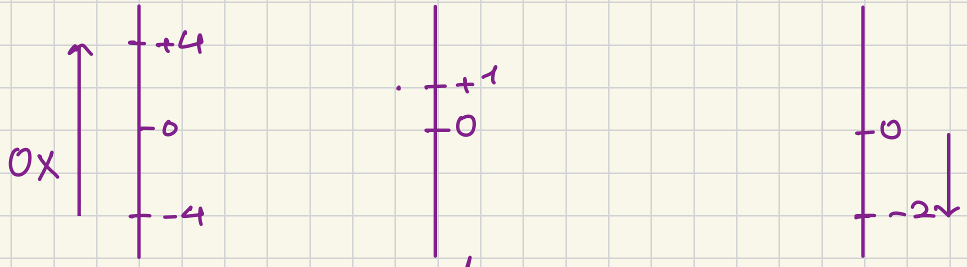
Example

OA / RA / ON / half_reactions / disproportionation

$\overset{4-}{C}\overset{1+}{H}_4(g) + 2\overset{0}{O}_2(g) \longrightarrow \overset{4+}{C}\overset{2-}{O}_2(g) + 2\overset{1+}{H}_2\overset{2-}{O}(g)$ Combustion of Methane

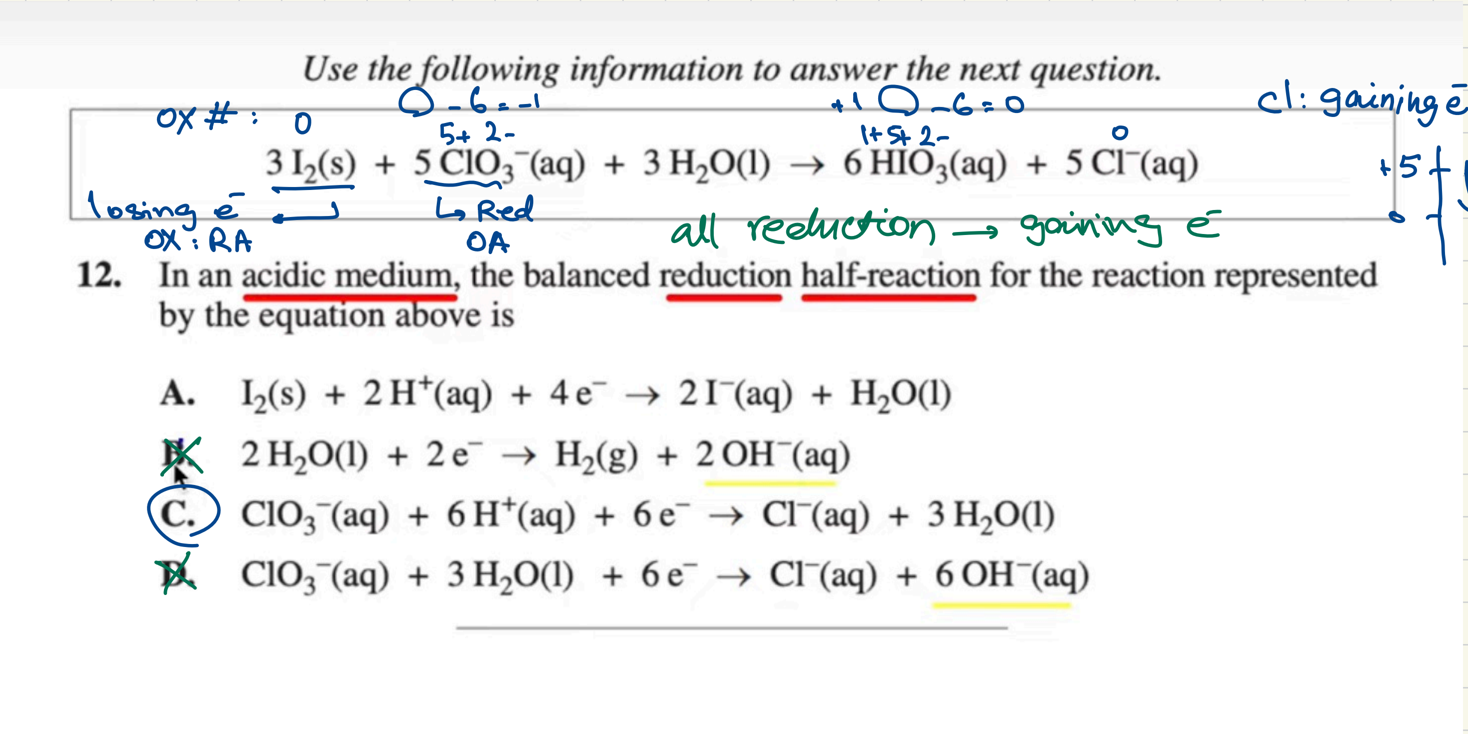
red A ← OX red : OA

draw a number line for each



Carbon : OX Hydrogen : nothing Oxygen : Red

reducing agent ↵ ↳ oxidizing agent



Use the following information to answer the next question.

OX # : $\overset{0}{}$ $\bigcirc - 6 = -1$ $+1 \bigcirc - 6 = 0$ cl: gaining $\bar{e}$

$3\,\overset{0}{I_2}(s) + 5\,\overset{5+\,2-}{ClO_3^-}(aq) + 3\,H_2O(l) \rightarrow 6\,\overset{1+\,5+\,2-}{HIO_3}(aq) + 5\,\overset{0}{Cl^-}(aq)$

losing $\bar{e}$ OX : RA ↳ Red OA all reduction → gaining $\bar{e}$

+5 0 ↓Red

12. In an acidic medium, the balanced reduction half-reaction for the reaction represented by the equation above is

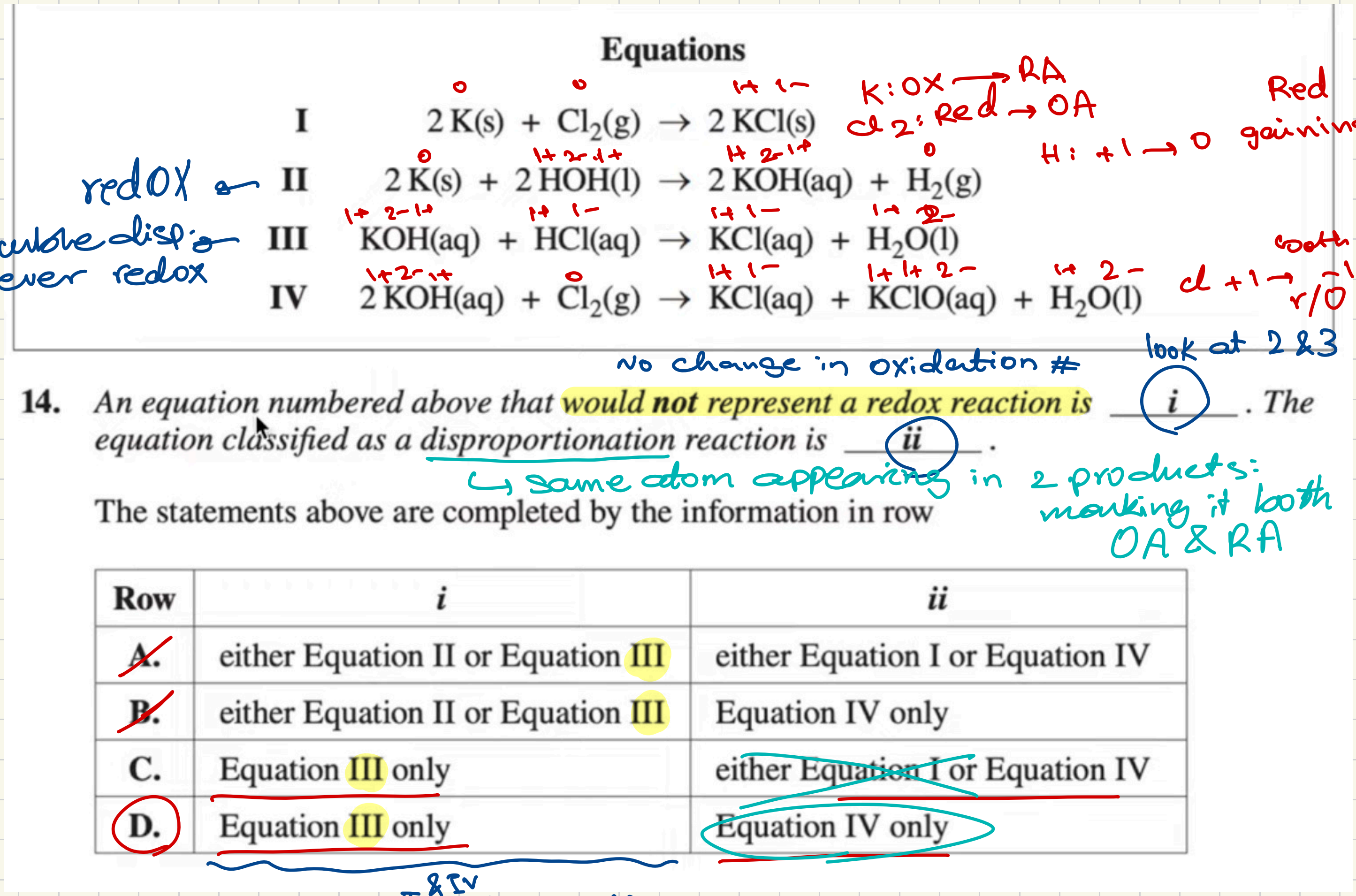
A. $I_2(s) + 2\,H^+(aq) + 4\,e^- \rightarrow 2\,I^-(aq) + H_2O(l)$

B. $2\,H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2\,OH^-(aq)$

[✓ **C.**] $ClO_3^-(aq) + 6\,H^+(aq) + 6\,e^- \rightarrow Cl^-(aq) + 3\,H_2O(l)$

D. $ClO_3^-(aq) + 3\,H_2O(l) + 6\,e^- \rightarrow Cl^-(aq) + 6\,OH^-(aq)$

- we want acidic, which means H^+ or H_3O^+ is involved. A/C
- For basic reactions, we would look for OH^-. B/D



Equations

I $2\overset{0}{K}(s) + \overset{0}{Cl_2}(g) \rightarrow 2\overset{1+\;1-}{KCl}(s)$

II $2\overset{0}{K}(s) + 2\overset{1+\,2-\,1+}{HOH}(l) \rightarrow 2\overset{1+\,2-\,1+}{KOH}(aq) + \overset{0}{H_2}(g)$

III $\overset{1+\,2-\,1+}{KOH}(aq) + \overset{1+\;1-}{HCl}(aq) \rightarrow \overset{1+\;1-}{KCl}(aq) + \overset{1+\;2-}{H_2O}(l)$

IV $2\overset{1+\,2-\,1+}{KOH}(aq) + \overset{0}{Cl_2}(g) \rightarrow \overset{1+\;1-}{KCl}(aq) + \overset{1+\,1+\,2-}{KClO}(aq) + \overset{1+\;2-}{H_2O}(l)$

K: OX → RA
Cl 2: Red → OA
Red
H: +1 → 0 gaining
redOX ←
double disp ←
never redox
both
Cl +1 → −1, 1+
r/O A
look at 2 & 3
No change in oxidation #

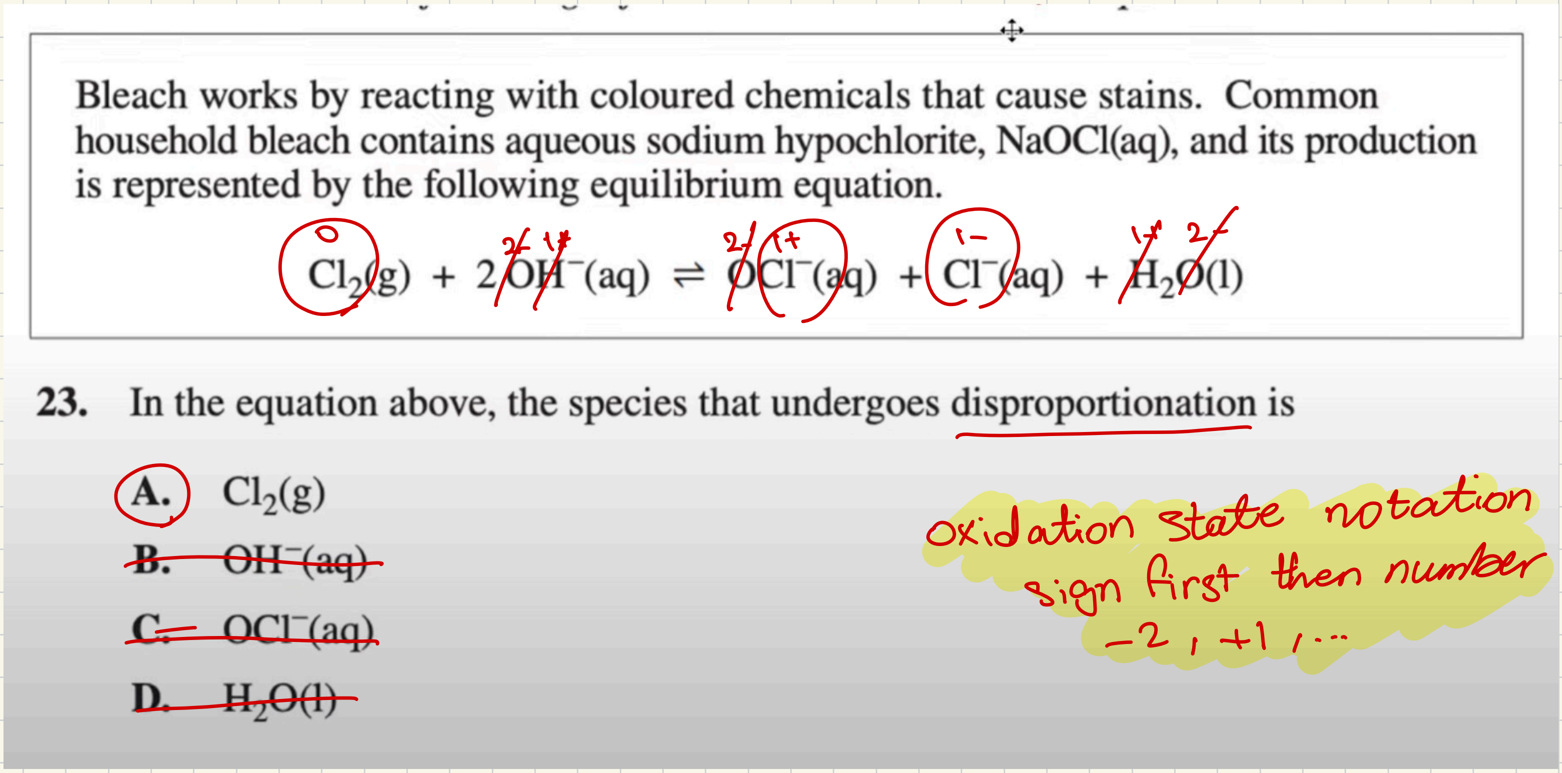
14. *An equation numbered above that would **not** represent a redox reaction is* ___i___ *. The equation classified as a disproportionation reaction is* ___ii___ .

↳ same atom appearing in 2 products: making it both OA & RA

The statements above are completed by the information in row

Row	i	ii
~~A.~~	either Equation II or Equation III	either Equation I or Equation IV
~~B.~~	either Equation II or Equation III	Equation IV only
C.	Equation III only	either Equation I or Equation IV
[✓ D.]	Equation III only	Equation IV only

since I & IV isn't mentioned,
don't waste your time.
only 2 would be enough



Bleach works by reacting with coloured chemicals that cause stains. Common household bleach contains aqueous sodium hypochlorite, NaOCl(aq), and its production is represented by the following equilibrium equation.

$$\overset{0}{Cl_2}(g) + 2\overset{2-\;1+}{OH^-}(aq) \rightleftharpoons \overset{2-\;1+}{OCl^-}(aq) + \overset{1-}{Cl^-}(aq) + \overset{1+\;2-}{H_2O}(l)$$

23. In the equation above, the species that undergoes disproportionation is

[✓ **A.**] $Cl_2(g)$
~~**B.** $OH^-(aq)$~~
~~**C.** $OCl^-(aq)$~~
~~**D.** $H_2O(l)$~~

oxidation state notation
sign first then number
−2, +1, ...

reduction →

oxidation ←

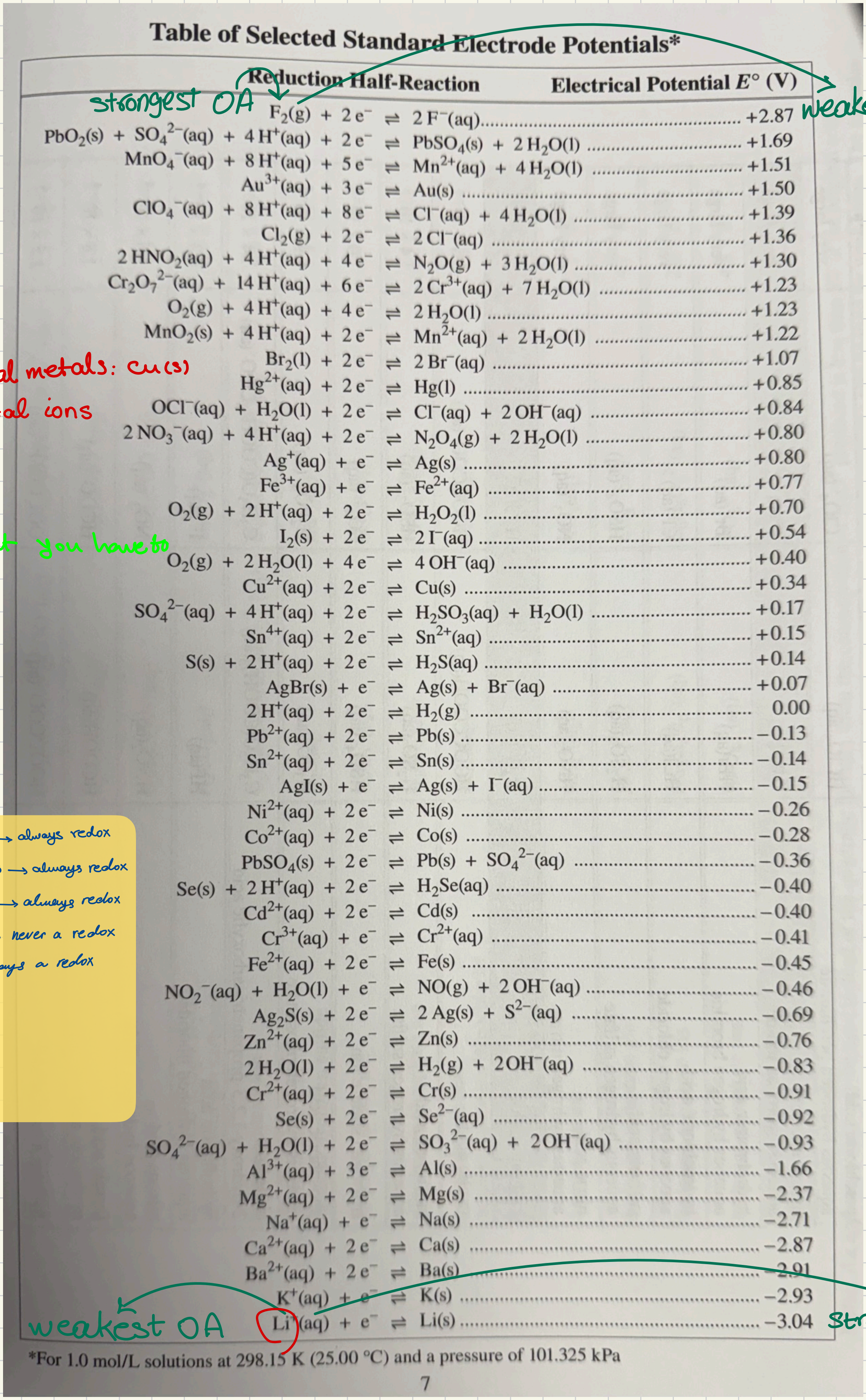


Table of Selected Standard Electrode Potentials*

strongest OA

weakest OA

Reduction Half-Reaction	Electrical Potential $E°$ (V)
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$	+1.69
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.07
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$	+0.85
$OCl^-(aq) + H_2O(l) + 2e^- \rightleftharpoons Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s)$	−2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

weakest OA

strongest RA

*For 1.0 mol/L solutions at 298.15 K (25.00 °C) and a pressure of 101.325 kPa

OA	RA
H^+	OH^-
metal ions Cu^{2+}	elemental metals: $Cu(s)$
	non-metal ions

↓ most likely but you have to check!

OA	RA
-ines Cl_2 ↓ non-metal elements	- ides Cl^-

- Formation Reactions → always redox
- Decomposition Reactions → always redox
- Single Replacement → always redox
- Double Replacement → never a redox
- Combustion → always a redox

        - *Exception:* When bonded to a metal (forming a metal hydride, e.g., NaH), hydrogen is **-1**.
    - **Sum of Oxidation States:**
        - For a **neutral compound**, the sum of all oxidation states must equal **zero**.
        - For a **polyatomic ion**, the sum of all oxidation states must equal the **charge of the ion**.
- **Examples of Sulfur Oxidation States (Practice!):**
    - **SO_2 (Sulfur Dioxide):**
        - Oxygen is -2. There are 2 Oxygen atoms: 2 * (-2) = -4.
        - The compound is neutral (sum = 0).
        - S + (-4) = 0 → **Sulfur is +4**
    - **SO_3 (Sulfur Trioxide):**
        - Oxygen is -2. There are 3 Oxygen atoms: 3 * (-2) = -6.
        - The compound is neutral (sum = 0).
        - S + (-6) = 0 → **Sulfur is +6**
    - **H_2SO_3 (Sulfurous Acid):**
        - Hydrogen is +1. There are 2 Hydrogen atoms: 2 * (+1) = +2.
        - Oxygen is -2. There are 3 Oxygen atoms: 3 * (-2) = -6.
        - The compound is neutral (sum = 0).
        - (+2) + S + (-6) = 0 → S - 4 = 0 → **Sulfur is +4**
    - **H_2SO_4 (Sulfuric Acid):**
        - Hydrogen is +1. There are 2 Hydrogen atoms: 2 * (+1) = +2.
        - Oxygen is -2. There are 4 Oxygen atoms: 4 * (-2) = -8.
        - The compound is neutral (sum = 0).
        - (+2) + S + (-8) = 0 → S - 6 = 0 → **Sulfur is +6**

Oxidizing Agents (OA) and Reducing Agents (RA):

These terms describe the *role* a substance plays in a redox reaction.

- **Oxidizing Agent (OA):**
    - **Definition:** The substance that **causes something else to be oxidized**.
    - **Itself:** The oxidizing agent is **itself reduced** (gains electrons).
    - **Sample:** In many reactions, **elemental oxygen (O_2) is always an oxidizing agent**.
- **Reducing Agent (RA):**
    - **Definition:** The substance that **causes something else to be reduced**.
    - **Itself:** The reducing agent is **itself oxidized** (loses electrons).
- **Example: CH_4 (methane) combustion with O_2**
    - $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
    - **Carbon in CH_4** goes from -4 to +4 (oxidized) → CH_4 is the **reducing agent**.
    - **Oxygen in O_2:** goes from 0 to -2 (reduced) → O_2 is the **oxidizing agent**.

Disproportionation:

A special type of redox reaction where a single element is both oxidized and reduced.

- **Definition:** A reaction in which the **same species (atom)** acts as both the **oxidizing agent** and the **reducing agent**.
- **Characteristic:** You will see the *same element* appearing in two different compounds on the product side, with two different oxidation states (one higher, one lower than the reactant).
- **Example:** Chlorine disproportionation
  - $Cl_2(0) + 2OH^- \rightarrow Cl^-(-1) + ClO^-(+1) + H_2O$
  - **Chlorine (Cl_2):** Starts with an oxidation state of 0.
  - **In Cl^-:** Chlorine's oxidation state is -1 (reduced).
  - **In ClO^-:** Chlorine's oxidation state is +1 (oxidized).
  - Since elemental chlorine (Cl_2) is both reduced to Cl^- and oxidized to ClO^-, this is a disproportionation reaction.

Section 4: Trends and Patterns in Redox Reactions

V. Trends in the Standard Reduction Table:

The standard reduction table is an essential tool for predicting the spontaneity and direction of redox reactions.

- **Direction of Reactions:**
  - **Forward Direction (Left to Right):** Represents **reduction** (gain of electrons).
  - **Backward Direction (Right to Left):** Represents **oxidation** (loss of electrons).
- **Strength of Oxidizing Agents (Left Column):**
  - **Strongest:** Found at the **top** of the left-hand column. These species are easily reduced.
  - **Weakest:** Found at the **bottom** of the left-hand column.
  - **Important Trend:** As you go *up* the left column, the oxidizing agent strength *increases*.
  - **Key Note: Acids (H^+)** are always found on the **oxidizing agent (left)** side of the table.
- **Strength of Reducing Agents (Right Column):**
  - **Strongest:** Found at the **bottom** of the right-hand column (e.g., Lithium metal, Li). These species are easily oxidized.
  - **Weakest:** Found at the **top** of the right-hand column (e.g., Fluoride ion, F^-).
  - **Important Trend:** As you go *down* the right column, the reducing agent strength *increases*.
  - **Key Note: Bases (OH^-)** are always found on the **reducing agent (right)** side of the table.
- **Elemental Metals vs. Metal Ions:**

  - **Elemental Metals (e.g., Cu, Zn, Ni):** Always act as **reducing agents** (they tend to lose electrons and become ions). There are no exceptions for common elemental metals.
  - **Metal Ions:** Mostly act as **oxidizing agents**. However, be cautious with **multivalent elements** (metals that can have multiple stable ion charges).
    - **Sample:** Copper(I) ion (Cu^+) can appear on *both* the oxidizing agent and reducing agent sides of the table, depending on the specific half-reaction. This also applies to ions like Ni^{2+} and Pb^{2+}.

Section 5: Identifying Redox vs. Non-Redox Reactions

Knowing the common reaction types that are always or never redox can save you time.

VI. Reactions That Are ALWAYS Redox:

These reaction types inherently involve changes in oxidation states.

- **Formation (Synthesis) Reactions:**
  - **Definition:** Elements combine to form a compound.
  - **Characteristic:** Reactants are always in their elemental form (oxidation state 0).
  - **Sample:** $Na(s)$ (0) + $Cl_2(g)$ (0) → $NaCl(s)$ (Na +1, Cl -1).
    - Sodium is oxidized, and Chlorine is reduced.
- **Decomposition Reactions:**
  - **Definition:** A compound breaks down into simpler substances, often elements.
  - **Characteristic:** Products often include elements in their elemental form (oxidation state 0).
  - **Sample:** $2NaCl(s)$ (Na +1, Cl -1) → $2Na(s)$ (0) + $Cl_2(g)$ (0).
    - Sodium is reduced, and Chlorine is oxidized.
- **Single Replacement Reactions:**
  - **Definition:** An element reacts with a compound, displacing another element.
  - **Characteristic:** One reactant is always an element (oxidation state 0), which will become an ion. The displaced element becomes an element (oxidation state 0).
  - **Sample:** $Ag(s)$ (0) + $CuCl_2(aq)$ (Cu +2, Cl -1) → $Cu(s)$ (0) + $2AgCl(aq)$ (Ag +1, Cl -1).
    - Silver is oxidized, and Copper is reduced.
- **Combustion Reactions:**
  - **Definition:** A rapid reaction between a substance with an oxidant, usually oxygen, producing heat and light.
  - **Characteristic:** Always involves **elemental oxygen (O_2)** as a reactant (oxidation state 0), which will always be reduced to **-2** in the products (e.g., CO_2, H_2O).
  - **Sample:** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$.
    - Oxygen is reduced (0 to -2). Carbon is oxidized (-4 to +4).

VII. Reactions That Are NEVER Redox:

These reaction types involve ion exchange without electron transfer.

- **Double Replacement Reactions (Metathesis):**
  - **Definition:** Two ionic compounds exchange ions, typically forming a precipitate, gas, or water.
  - **Characteristic:** All ions retain their original oxidation states. There is no change in oxidation numbers for any element.
  - **Sample:** $KI(aq)$ (K +1, I -1) + $Pb(NO_3)_2(aq)$ (Pb +2, N +5, O -2) → $KNO_3(aq)$ (K +1, N +5, O -2) + $PbI_2(s)$ (Pb +2, I -1).
    - No element changes its oxidation state. All ions simply switch partners.

Section 6: Redox Chemistry in Living Systems

Application of Redox Principles:

- **Key Takeaway:** The fundamental concepts of oxidation and reduction are universal and apply directly to complex biological processes.
- **Examples:**
  - **Cellular Respiration:** The process by which cells break down glucose to release energy, involving the oxidation of glucose and the reduction of oxygen.
  - **Photosynthesis:** The process by which plants convert light energy into chemical energy, involving the oxidation of water and the reduction of carbon dioxide.
- **Study Tip:** Do not try to memorize the specific redox details of these biological processes. Instead, focus on **applying the learned skills** of identifying oxidation states, oxidized/reduced species, and agents to analyze these complex reactions.
- **Interdisciplinary Overlap:** These biological redox reactions often overlap with concepts studied in the thermodynamics section of chemistry (e.g., energy changes, enthalpy).

Building Reduction Tables

Building reduction tables is a common task on tests and diploma exams. The goal is often to list oxidizing agents from strongest to weakest based on provided data or observations. You don't need a perfect reduction table; just the correct ordering is sufficient.

Trends in Reduction Tables

Important trends to remember when building reduction tables:

- **Oxidizing Agents (OA)** are on the left side.
- **Reducing Agents (RA)** are on the right side.
- Acidic half-reactions with H+ are on the OA side.
- Hydroxide (OH−) is on the RA side.
- Elemental metals are generally on the RA side.
- Metal ions are mostly on the OA side.
- When assigning OAs and RAs, remember that metals becoming metal ions is oxidation, confirming the metal as a reducing agent.

Building a Reduction Table: Example 1

1. **Assign OAs and RAs:** Based on the given data and the trends, quickly assign each substance as either an OA or RA.
2. **Analyze Spontaneous Reactions:** In a spontaneous reaction, the oxidizing agent is above the reducing agent in the table.
3. **Analyze Non-Spontaneous Reactions:** In a non-spontaneous reaction, the reducing agent is below the oxidizing agent.
4. **Incorporate Data:** Systematically incorporate each piece of data to order the chemicals in the table correctly.
5. **Complete the Table:** Construct the reduction table with the oxidizing agents on the left and reducing agents on the right. Electrons can be omitted if only ordering is needed.

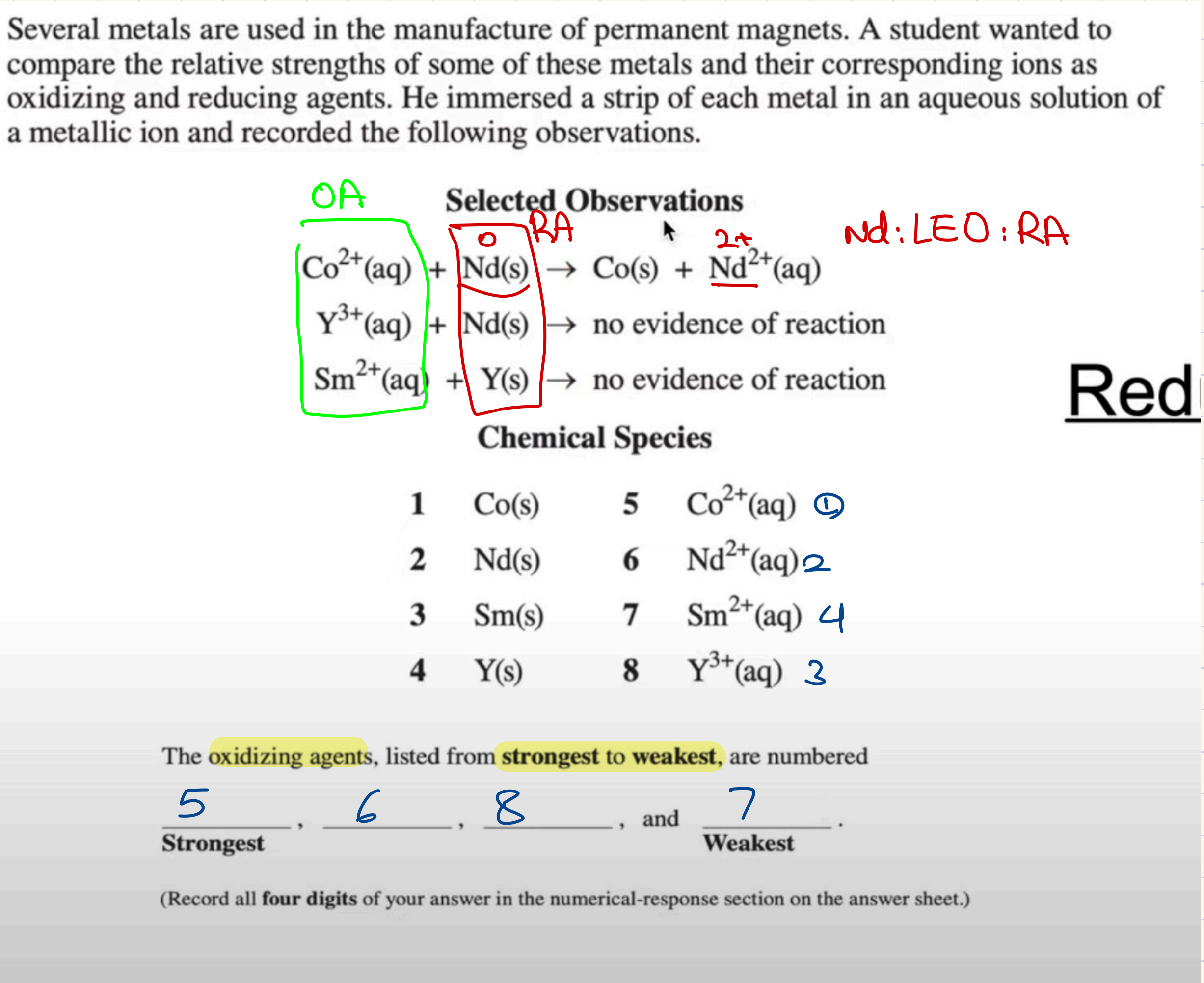
Example:

- Cobalt Co2+ reacts spontaneously with a metal (M).
- Yttrium Y3+ shows no reaction with another metal (N).
- Tin Sn2+ shows no reaction with elemental yttrium (Y).

This data helps establish the relative positions of Co2+, Y3+, and Sn2+ in the reduction table.

Building a Reduction Table: Example 2

- **Data Interpretation:** When given a table of reaction data (check marks indicating a reaction occurs), identify the best oxidizing agent and best reducing agent.
  - The best oxidizing agent is the one that reacts with everything but itself.



Several metals are used in the manufacture of permanent magnets. A student wanted to compare the relative strengths of some of these metals and their corresponding ions as oxidizing and reducing agents. He immersed a strip of each metal in an aqueous solution of a metallic ion and recorded the following observations.

OA

Selected Observations

RA

$Co^{2+}(aq) + \overset{0}{Nd(s)} \rightarrow Co(s) + \overset{2+}{Nd^{2+}}(aq)$ Nd: LEO: RA

$Y^{3+}(aq) + Nd(s) \rightarrow$ no evidence of reaction

$Sm^{2+}(aq) + Y(s) \rightarrow$ no evidence of reaction

Red

Chemical Species

1	Co(s)	5	$Co^{2+}(aq)$	①
2	Nd(s)	6	$Nd^{2+}(aq)$	2
3	Sm(s)	7	$Sm^{2+}(aq)$	4
4	Y(s)	8	$Y^{3+}(aq)$	3

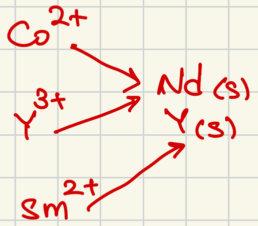
The oxidizing agents, listed from **strongest** to **weakest**, are numbered

5 (**Strongest**), 6, 8, and 7 (**Weakest**).

(Record all **four digits** of your answer in the numerical-response section on the answer sheet.)

Reduction Table: after finding RA's & OA's, we go through them to order them.

- when there's an spontaneous reaction, the OA is above the RA.
- There's no spontaneous reaction for $Y^{3+} + Nd$, which means Y^{3+} is lower than the previous OA.



OA

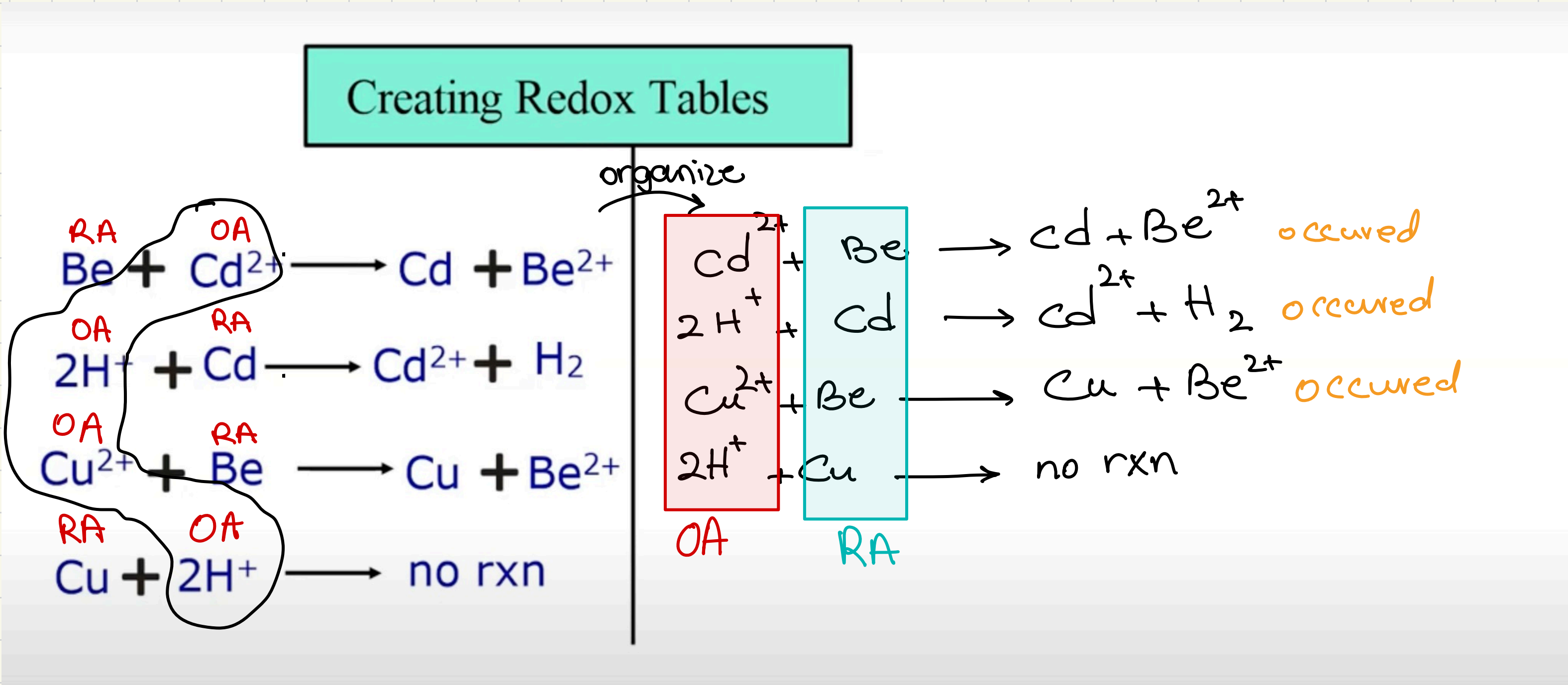
$Co^{2+} + 2e^- \rightarrow Co_{(s)}$ strongest

$Nd^{2+} + 2e^- \rightarrow Nd(s)$

$Y^{3+} + 3e^- \rightarrow Y_{(s)}$

$Sm^{2+} + 2e^- \rightarrow Sm_{(s)}$ weakest

Remember: OA's get reduced so get more negative, RA's get oxidized so get more positive!



Creating Redox Tables

$\overset{RA}{Be} + \overset{OA}{Cd^{2+}} \rightarrow Cd + Be^{2+}$

$\overset{OA}{2H^+} + \overset{RA}{Cd} \rightarrow Cd^{2+} + H_2$

$\overset{OA}{Cu^{2+}} + \overset{RA}{Be} \rightarrow Cu + Be^{2+}$

$\overset{RA}{Cu} + \overset{OA}{2H^+} \rightarrow$ no rxn

organize

$Cd^{2+} + Be \rightarrow Cd + Be^{2+}$ occured

$2H^+ + Cd \rightarrow Cd^{2+} + H_2$ occured

$Cu^{2+} + Be \rightarrow Cu + Be^{2+}$ occured

$2H^+ + Cu \rightarrow$ no rxn

OA RA

OA · reduced : gains e^- : becomes more −

RA: oxidized · loses e^-: becomes more +

→ Creating the table : if a reaction happens, OA is above RA .

Cu^{2+}	Cu
$2H^+$	H_2
Cd^{2+}	Cd
Be^{2+}	Be

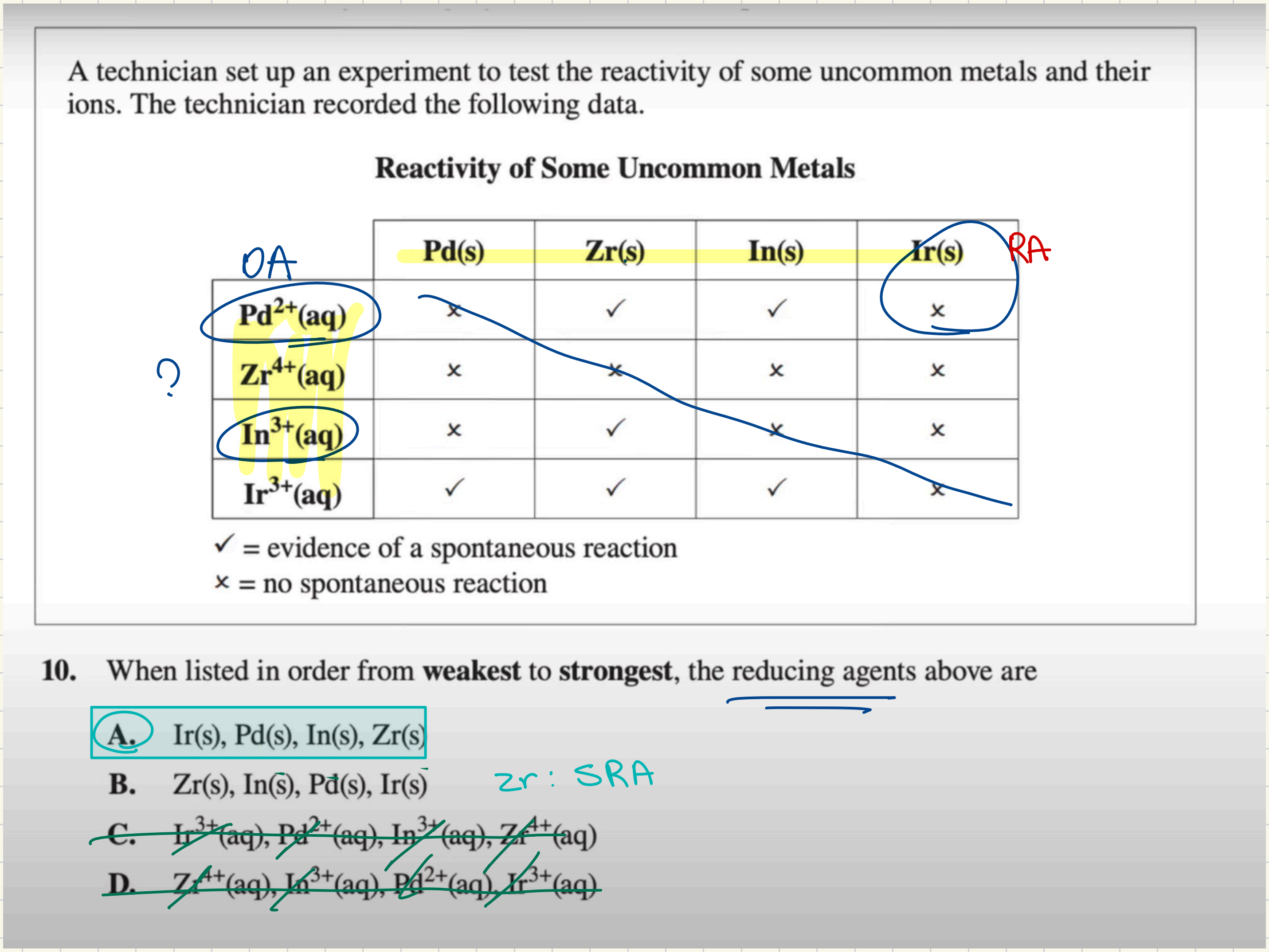
strong

$\underline{Cu^{2+}} + 2e^- \rightarrow Cu_{(s)}$

$\underline{2H^+} + 2e^- \rightarrow H_{2(g)}$

$\underline{Cd^{2+}} + 2e^- \rightarrow Cd_{(s)}$

$\underline{Be^{2+}} + 2e^- \rightarrow Be_{(s)}$



A technician set up an experiment to test the reactivity of some uncommon metals and their ions. The technician recorded the following data.

Reactivity of Some Uncommon Metals

OA

RA

	Pd(s)	Zr(s)	In(s)	Ir(s)
$Pd^{2+}(aq)$	×	✓	✓	×
$Zr^{4+}(aq)$?	×	×	×	×
$In^{3+}(aq)$	×	✓	×	×
$Ir^{3+}(aq)$	✓	✓	✓	×

✓ = evidence of a spontaneous reaction

× = no spontaneous reaction

10. When listed in order from **weakest** to **strongest**, the reducing agents above are

[✓ **A.**] Ir(s), Pd(s), In(s), Zr(s)

B. Zr(s), In(s), Pd(s), Ir(s)

Zr : SRA

~~**C.** $Ir^{3+}(aq)$, $Pd^{2+}(aq)$, $In^{3+}(aq)$, $Zr^{4+}(aq)$~~

~~**D.** $Zr^{4+}(aq)$, $In^{3+}(aq)$, $Pd^{2+}(aq)$, $Ir^{3+}(aq)$~~

more − OA

more + RA

S OA

$Ir^{3+} + 3e^- \longrightarrow Ir(s)$ WRA

$Pd^{2+} + 2e^- \longrightarrow Pd(s)$

$In^{3+} + 3e^- \longrightarrow In(s)$

$Zr^{4+} + 4e^- \longrightarrow Zr(s)$ S RA

W OA

Ir^{3+}

$Pd^{2+}(aq)$

In^{3+}

Zr^{4+}

Ir (s)

Pd (s)

~~In (s)~~

Zr (S)

- Strongest OA is the one that reacts with everything but itself → Ir^{3+}

    - Neutral metals are typically reducing agents.
- **Ordering:** Use the reactivity data to order the oxidizing and reducing agents in the table.
- **Answering Questions:** Pay close attention to what the question is asking (e.g., listing from weakest to strongest reducing agents) to avoid errors.

Spontaneity and ECell◦

not spont. [$E^\circ_{cell}{}^{-}$; OA → RA]

spont. [$E^\circ_{cell}{}^{+}$; OA → RA]

- **Spontaneous Reaction:** Oxidizing agent is above the reducing agent in the table, resulting in a positive ECell◦. This is akin to a reaction occurring "downhill" and releasing energy.
- **Non-Spontaneous Reaction:** Oxidizing agent is below the reducing agent, resulting in a negative ECell◦. This requires work to occur (like going "uphill").

Percent Error :

$$\frac{E^\circ_{measured} - E^\circ_{predicted}}{E^\circ_{predicted}} \times 100$$

→ from the table

} + high error
  - low error

Percent Error=Predicted ECell◦ ∣Experimental ECell◦−Predicted ECell◦∣×100

Percent Error: A measure of how much the experimental voltage (measured in a lab) deviates from the predicted voltage (from a reduction table).

Percent Error Calculations 🧮

When calculating percent error, you use the following formula:

Percent Error=TheoreticalExperimental−Theoretical×100

Note that some variations of this formula include absolute values. The way the formula is written above indicates whether your answer is high or low: a positive error means your answer is high, while a negative percent error indicates it's low.

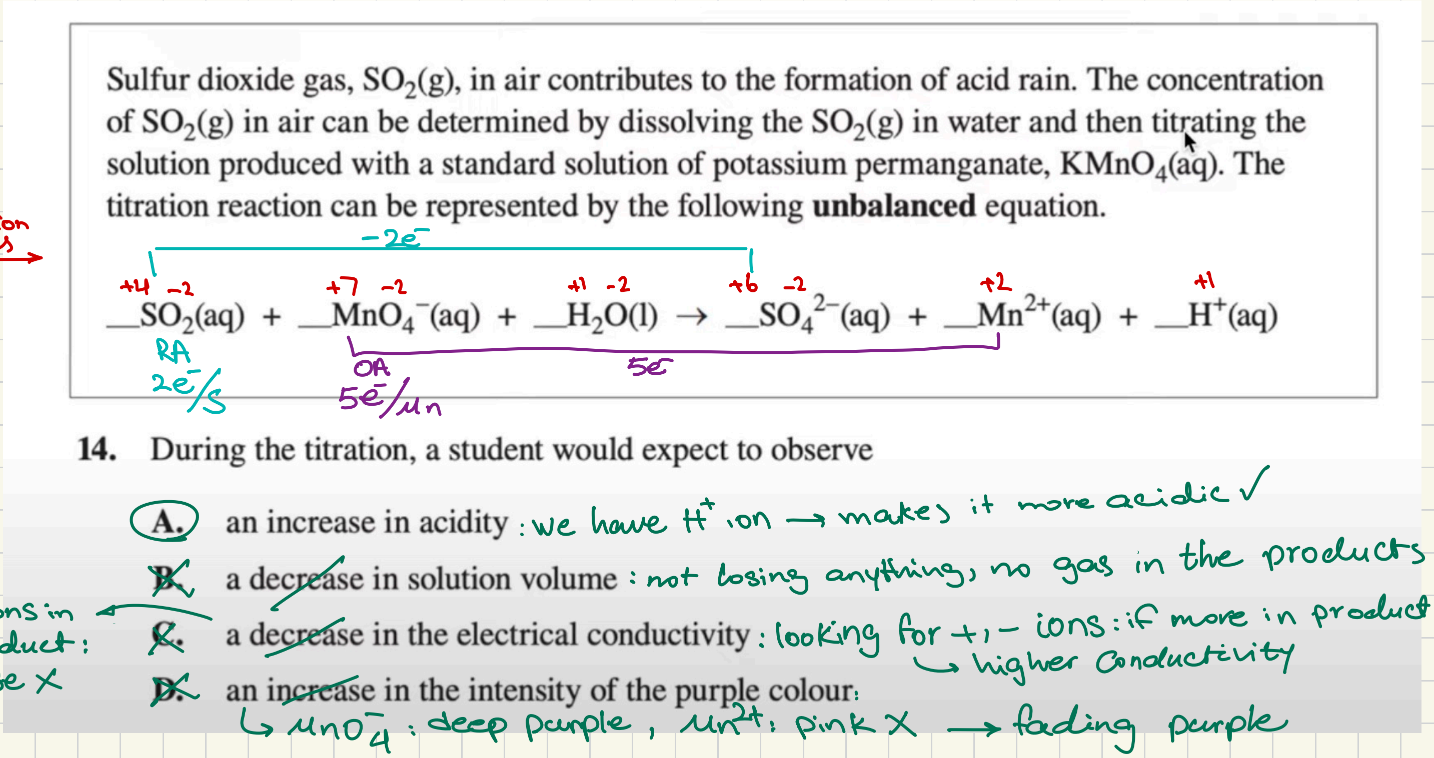
Understanding whether your error is high or low can be particularly important in fields like thermodynamics, where the implications of being high or low can differ significantly. For example, when constructing a cell, a slightly undersized salt bridge often leads to a lower-than-expected voltage.

Balancing Redox Reactions 🧪

We'll delve into balancing redox reactions in acidic and neutral solutions. Note that balancing in basic solutions will not be covered.

Balancing with Oxidation States

When balancing redox reactions, you can't simply use grade 10 balancing skills. You must balance the electron exchange first, before balancing everything else.



Sulfur dioxide gas, $SO_2(g)$, in air contributes to the formation of acid rain. The concentration of $SO_2(g)$ in air can be determined by dissolving the $SO_2(g)$ in water and then titrating the solution produced with a standard solution of potassium permanganate, $KMnO_4(aq)$. The titration reaction can be represented by the following **unbalanced** equation.

oxidation states →

$-2e^-$

$$\_\overset{+4\ -2}{SO_2}(aq) + \_\overset{+7\ -2}{MnO_4^-}(aq) + \_\overset{+1\ -2}{H_2O}(l) \rightarrow \_\overset{+6\ -2}{SO_4^{2-}}(aq) + \_\overset{+2}{Mn^{2+}}(aq) + \_\overset{+1}{H^+}(aq)$$

RA $2e^-/S$ OA $5e^-/Mn$ $5e^-$

14. During the titration, a student would expect to observe

(A.) an increase in acidity : we have H^+ ion → makes it more acidic ✓

~~B.~~ a ~~decrease~~ in solution volume : not losing anything, no gas in the products X

~~C.~~ a ~~decrease~~ in the electrical conductivity : looking for +, − ions : if more in products, ↳ higher conductivity

more ions in the product : increase X

~~D.~~ an ~~increase~~ in the intensity of the purple colour:
↳ MnO_4^- : deep purple , Mn^{2+} : pink X → fading purple

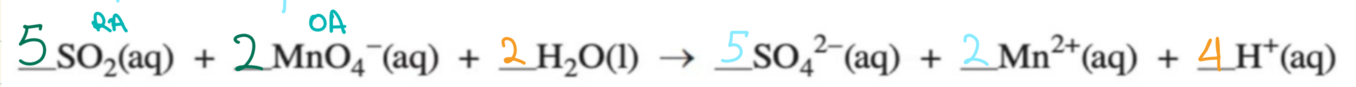
- Balancing full reactions with oxidation numbers:
it's not the same as balancing in grade 10, the electrons must be balanced too.

- After finding the oxidation states, we have to balance electron exchange.

S: +4 → +6 : what agents? RA : losing

Mn: +7 → +2 : what agent? OA : gaining

S: 5 Mn: 2



$$\underline{5}\overset{RA}{SO_2}(aq) + \underline{2}\overset{OA}{MnO_4^-}(aq) + \underline{2}H_2O(l) \rightarrow \underline{5}SO_4^{2-}(aq) + \underline{2}Mn^{2+}(aq) + \underline{4}H^+(aq)$$

$5\times$ $2e^-/S$, $2e^-/SO_2$ $2\times$ $5e^-/Mn$, $5e^-/MnO_4^-$

$10e^-$. we'll find LCM to balance the e^- ⇒ then the rest of the balancing is similar to grade 10

O: R → P 18+x → 20 H: R → P 4 → 4

OA/RA

$NO_2 + H_2O \longrightarrow HNO_3 + NO$: disportionation : write the compound twice

Balancing as a whole reaction:

$$\overset{+4}{N}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O} \longrightarrow \overset{+1}{H}\overset{\overset{+5}{\uparrow}}{N}\overset{-2}{O_3} + \overset{+2}{N}\overset{-2}{O}$$

N
- $+4 \rightarrow +5$: losing → RA
- $+4 \rightarrow +2$: gaining → OA

2e⁻

1e⁻

$$\underbrace{NO_2 + 2NO_2}_{} + 1H_2O \longrightarrow 2HNO_3 + 1NO$$

Not an acidic redox!

No H^+ in the products

$2e^-/N$ $1e^-/N$

$2e^-/NO_2$ $1e^-/NO_2 \times 2$

LCM=2

$3NO_2 + H_2O \longrightarrow 2HNO_3 + NO$

O:

R	P
6+x	7

H:

R	P
2	2

Balancing half_reactions:

OA : $NO_2 \longrightarrow NO$

$NO_2 + 2H^+ + 2e^- \longrightarrow NO + H_2O$

1 ✓ N: 1 → 1
2 O: water to the side w/ less Os
3 H: add H to the side w/ less H
4 e^-: changes

steps:

1 _ balance other than O&H
2 _ balance O : adding H_2O
3 _ balance H : adding H^+
4 _ balance e^- : changes
5 _ net 2 half_reactions

changes

R	P
0, 2+	0, 0
2+	0

: gaining $2e^-$

RA : $NO_2 + H_2O \longrightarrow HNO_3 + H^+ + 1e^-$

1 ✓ N: 1 → 1
2 O : add water
3 H : add H^+
4 e^- : write the charge

changes

R	P
0, 0	0, 1+
0	1+

→ losing e^-

2 ½ reac

$$\left\{\begin{array}{l} NO_2 + 2H^+ + 2e^- \longrightarrow NO + H_2O \\ NO_2 + H_2O \longrightarrow HNO_3 + H^+ + 1e^- \times 2 \end{array}\right.$$

⇒ balance e^-

→ next page

$$NO_2 + \cancel{2H^+} + \cancel{2e^-} \longrightarrow NO + \cancel{H_2O}$$

+

$$2NO_2 + \cancel{2}H_2O \longrightarrow 2HNO_3 + \cancel{2H^+} + \cancel{2e^-}$$

$$3NO_2 + H_2O \longrightarrow 2HNO_3 + NO$$

More Balancing ½-reactions:

$Mn^{2+} \longleftrightarrow MnO_2$: $Mn^{2+} + 2H_2O \longleftrightarrow MnO_2 + 4H^+ + 2e^-$ ①

1) Mn : 1 → 1 ✓
2) O : 0 → 2 add water to reactants : $Mn^{2+} + 2H_2O \longleftrightarrow MnO_2$
3) H : 2 → 0 add H^+ to products : $Mn^{2+} + 2H_2O \longleftrightarrow MnO_2 + 4H^+$
4) e^- : charges : 2+ → 4+ : losing $2e^-$: $Mn^{2+} + 2H_2O \longleftrightarrow MnO_2 + 4H^+ + 2e^-$

$O_2 \longleftrightarrow H_2O_2$ $O_2 + 2H^+ + 2e^- \longleftrightarrow H_2O_2$ ②

1) nothing but H & O
2) O : 2 → 2 ✓ No need to add water.
3) H : 0 → 2 : add $2H^+$ to reactants : $O_2 + 2H^+ \longleftrightarrow H_2O_2$
4) e^- : charges : 2+ → 0 gaining $2e^-$: $O_2 + 2H^+ + 2e^- \longleftrightarrow H_2O_2$

$$+\begin{cases} Mn^{2+} + 2H_2O \longleftrightarrow MnO_2 + \overset{2H^+}{4\cancel{H^+}} + \cancel{2e^-} \\ O_2 + 2\cancel{H^+} + \cancel{2e^-} \longleftrightarrow H_2O_2 \end{cases}$$

$$Mn^{2+} + 2H_2O + O_2 \longleftrightarrow MnO_2 + H_2O_2 + 2H^+$$

1. **Assign oxidation states:** Determine the oxidation states of each element in the reaction. For common elements like oxygen (usually -2 in compounds) and hydrogen (usually +1), these can often be memorized.
    - Oxygen in a compound: -2 (unless it's a peroxide)
    - Hydrogen: +1
    - Monoatomic ions: Oxidation state is equal to their charge
2. **Identify changes in oxidation states:** Look for elements that undergo a change in oxidation state during the reaction.
    - For example, in a given reaction, sulfur (S) might go from +4 to +6, while manganese (Mn) goes from +7 to +2.
3. **Calculate electron transfer:** Determine the number of electrons transferred per atom of the changing elements.
    - For sulfur going from +4 to +6, this is a loss of 2 electrons (oxidation).
    - For manganese going from +7 to +2, this is a gain of 5 electrons (reduction).
4. **Balance electron exchange:** Find the lowest common multiple of the electron transfers to balance the total number of electrons exchanged.
    - If sulfur involves 2 electrons and manganese involves 5 electrons, the lowest common multiple is 10.
    - Multiply the coefficients of the sulfur species by 5 and the manganese species by 2 to achieve 10-electron transfer for both.
5. **Balance atoms other than O and H:** Balance all elements other than oxygen and hydrogen.
6. **Balance oxygen with water:** Add water (H_2O) molecules to balance the oxygen atoms.
7. **Balance hydrogen with H+ ions:** Add hydrogen ions (H^+) to balance the hydrogen atoms.

Example:

Consider the following unbalanced redox reaction:

$SO_2+MnO_4^-+H_2O \rightarrow SO_4^{2-}+Mn^{2+}+H^+$

1. **Assign oxidation states:**
    - $O=-2$
    - $H=+1$
    - S in $SO_2=+4$
    - Mn in $MnO_4^-=+7$
    - S in $SO_4^{2-}=+6$
    - Mn in $Mn^{2+}=+2$
2. **Identify changes in oxidation states:**
    - S: +4 → +6 (loss of 2 electrons)
    - Mn: +7 → +2 (gain of 5 electrons)
3. **Balance electron exchange:**
    - Lowest common multiple of 2 and 5 is 10.
    - Multiply S by 5 (5 * 2 = 10 electrons)

    - Multiply Mn by 2 (2 * 5 = 10 electrons)
4. **Balance atoms other than O and H:**
    - 5SO2
    - 2MnO4−
    - 5SO42−
    - 2Mn2+
5. **Balance oxygen with water:**
    - Reactant side: (5 * 2) + (2 * 4) + 1 = 19 oxygen atoms
    - Product side: (5 * 4) = 20 oxygen atoms
    - Add 2H2O to the reactant side to balance oxygen.
6. **Balance hydrogen with H+ ions:**
    - Reactant side: 2 * 2 = 4 hydrogen atoms
    - Product side: 4H+

Balanced equation:

5SO2+2MnO4−+2H2O→5SO42−+2Mn2++4H+

Lab Observations

In lab settings, you might need to connect observations to the redox reactions.

For example, you need to be aware of the colors of different ions to be successful.

Permanganate (MnO4−) is purple.

Consider a redox titration where the skeleton equation is given:

SO2+MnO4−+H2O→SO42−+Mn2++H+

and the question asks about possible observations during the titration.

- (A) **Increase in acidity:** This is correct because H+ is a product.
- (B) **Decrease in volume:** This is incorrect; there is no significant volume change.
- (C) **Decrease in electrical conductivity:** This is incorrect; the number of ions increases, so conductivity increases.
- (D) **Increase in purple:** This is incorrect; the purple permanganate is a reactant and will fade as the reaction proceeds.

Disproportionation Reactions ➗

In disproportionation reactions, a single species acts as both the oxidizing agent (OA) and reducing agent (RA).

Example:

Consider the reaction:

$NO2+H2O \rightarrow HNO3+NO$

1. **Assign oxidation states:**
    - $O=-2$
    - $H=+1$
    - N in $NO2=+4$
    - N in $HNO3=+5$
    - N in $NO=+2$
2. **Identify disproportionation:**
    - Nitrogen in NO2 goes both up (+4 to +5) and down (+4 to +2) in oxidation state.
    - Therefore, NO2 acts as both the OA and RA.
3. **Separate the changes:**
    - One NO2 goes from +4 to +5 (1 electron transfer).
    - The other NO2 goes from +4 to +2 (2 electron transfer).

Balancing Redox Reactions 🧪

Balancing with Electron Exchange

When balancing redox reactions, you don't need to identify which substance is undergoing oxidation and which is undergoing reduction.

1. Determine the **electron exchange** for each substance.
2. Find the **lowest common multiple** of the electron exchanges.
3. Multiply each substance by the factor needed to achieve the lowest common multiple. This balances the electron exchange.
4. Balance the **atoms** other than oxygen and hydrogen.
5. Balance **oxygen** by adding **water (H2O)**.
6. Balance **hydrogen** by adding **hydrogen ions (H+)**.

Example:

Consider the following unbalanced reaction:

$NO2+H2O \rightarrow HNO3+NO$

1. **Balance Electron Exchange:**
    - NO2 undergoes a one-electron exchange (N from +4 to +5 in HNO3).
    - The other half undergoes a two-electron exchange (N from +4 to +2 in NO).
    - The lowest common multiple between one and two is two, thus it is a two electron exchange.
2. **Double the NO2:** $2NO2+H2O \rightarrow HNO3+NO$
3. **Balance Nitrogen:** There are three nitrogen atoms on both sides of the reaction (1 in HNO3, 1 in NO, and 2 in NO2 on the reactant side). *Correction: The coefficient of NO2*

will need to be 3 for N atoms to balance in the end, as one NO2 forms HNO3 and the other two form NO. Let's re-evaluate the electron exchange balancing by considering the products.

Let's re-approach this example using the half-reaction method first, as it's often more straightforward for disproportionation.

Balancing Half Reactions ➗

An alternative approach involves breaking the redox reaction into two **half-reactions**: one for oxidation and one for reduction.

Steps for Balancing Half Reactions:

1. Balance **atoms** other than oxygen and hydrogen.
2. Balance **oxygen** by adding **water (H2O)**.
3. Balance **hydrogen** by adding **hydrogen ions (H+)**.
4. Balance the **charge** by adding **electrons (e−)**. Add electrons to the side with the higher charge to balance it.

Steps:

1 _ balance other than O & H
2 _ balance O : adding H_2O
3 _ balance H : adding H^+
4 _ balance e^- : charges
5 _ net 2 half_reactions

Example:

Consider the half-reaction:

NO2→HNO3

1. **Balance Nitrogen:** One nitrogen atom on each side.
2. **Balance Oxygen with Water:** Add water to the reactant side: NO2+H2O→HNO3
3. **Balance Hydrogen with Hydrogen Ions:** Add a hydrogen ion to the product side: NO2+H2O→HNO3+H+
4. **Balance Charge with Electrons:**
    - Reactant side has no charge.
    - Product side has a +1 charge.
    - Add one electron to the product side: NO2+H2O→HNO3+H++e−

Now, consider the other half-reaction:

NO2→NO

1. **Balance Nitrogen:** One nitrogen atom on each side.
2. **Balance Oxygen with Water:** Add water to the product side: NO2→NO+H2O
3. **Balance Hydrogen with Hydrogen Ions:** Add two hydrogen ions to the reactant side: NO2+2H+→NO+H2O
4. **Balance Charge with Electrons:**
    - Reactant side has a +2 charge.

    - Product side has no charge.
    - Add two electrons to the reactant side: 2e−+NO2+2H+→NO+H2O

Netting Half Reactions:

To combine half-reactions into a balanced redox reaction, multiply each half-reaction by a factor so that the number of electrons is the same in both half-reactions, then add the half-reactions together. Cancel out anything that appears on both sides of the reaction to arrive at the balanced redox reaction.

Continuing the example:

1. Multiply the first half-reaction by two: 2NO2+2H2O→2HNO3+2H++2e−
2. Add the two half-reactions together:
   (2NO2+2H2O+2e−)+(NO2+2H++2e−)→(2HNO3+2H++2e−)+(NO+H2O)
   Note: The electrons should be on opposite sides to cancel. Let's rewrite the first half-reaction to show electron loss for oxidation.
   First half-reaction (oxidation): NO2+H2O→HNO3+H++e− (Multiply by 2 for 2 electrons)
   2NO2+2H2O→2HNO3+2H++2e−
   Second half-reaction (reduction): 2e−+NO2+2H+→NO+H2O
   Adding them:
   (2NO2+2H2O)+(2e−+NO2+2H+)→(2HNO3+2H++2e−)+(NO+H2O)
3. Cancel out anything that appears on both sides:
    - Cancel 2e− from both sides.
    - Cancel 2H+ from both sides.
    - Cancel 1H2O from both sides.
4. Resulting in: 3NO2+H2O→2HNO3+NO
   Correction: My previous example of "2NO2 + H2O -> HNO3 + NO" was incorrect. The correct stoichiometry for this disproportionation is 3NO2+H2O→2HNO3+NO.

Note: Both methods (oxidation states and half-reactions) yield the same balanced redox reaction.

Additional Half Reaction Balancing Examples

Example 1

MnO2→Mn2+

1. **Balance atoms other than O and H:** Already balanced with 1 Mn on each side.
2. **Balance oxygen with water:** MnO2→Mn2++2H2O
3. **Balance hydrogen with hydrogen ions:** 4H++MnO2→Mn2++2H2O
4. **Balance the electrons by balancing the charge:** 2e−+4H++MnO2→Mn2++2H2O

Example 2

$H_2O_2 \rightarrow O_2$

1. **Balance atoms other than O and H:** Nothing to do.
2. **Balance oxygen with water:** Nothing to do. (Oxygen is already balanced in terms of atoms, but not charge yet).
3. **Balance hydrogen with hydrogen ions:** $H_2O_2 \rightarrow O_2 + 2H^+$
4. **Balance the electrons by balancing the charge:** $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$

Mnemonic device:

Loss of Electrons is Oxidation

Gain of Electrons is Reduction

Oxidation is when electrons are lost. Reduction is when electrons are gained.

An **incomplete** equation indicates that components like H^+ and water are missing and the equation is unbalanced.

Balancing Redox Reactions 🧪

When balancing redox reactions, especially in acidic conditions, remember that this is a **half reaction**. Applying the correct balancing strategy, whether it's a half-reaction or full reaction, is crucial. Here's a step-by-step approach:

1. Balance all atoms other than oxygen (O) and hydrogen (H).
2. Balance oxygen by adding water (H_2O) to the side that needs more oxygen.
3. Balance hydrogen by adding hydrogen ions (H^+) to the side that needs more hydrogen.
4. Balance charge by adding electrons (e^-) to the side that is more positive to make the charges equal on both sides of the equation.

Example: Balancing with Lowest Whole Number Coefficients

Consider a problem where you need to find the lowest whole number coefficients for water and H^+ when balancing a redox reaction. Let's work through an example:

$Br^- \rightarrow BrO_3^-$

1. Balance Bromine (Br): Start by balancing the bromine atoms. You have one Br on the reactant side and one on the product side, so bromine atoms are already balanced. $Br^- \rightarrow BrO_3^-$
2. Balance Oxygen (O): Next, balance oxygen by adding water. If you have three oxygen atoms on the product side, you'll need three water molecules on the reactant side. $3H_2O + Br^- \rightarrow BrO_3^-$
3. Balance Hydrogen (H): Balance hydrogen by adding H^+ ions. If you have six hydrogen atoms on the reactant side (3 × 2), you'll need six H^+ ions on the product side. $3H_2O + Br^- \rightarrow BrO_3^- + 6H^+$

Balancing Half_Reactions

$Br_2 \longrightarrow BrO_3^-$

$Br_2(l) \longrightarrow 2BrO_3^-(aq)$ 1. start balancing Br : $2 \rightarrow 1$

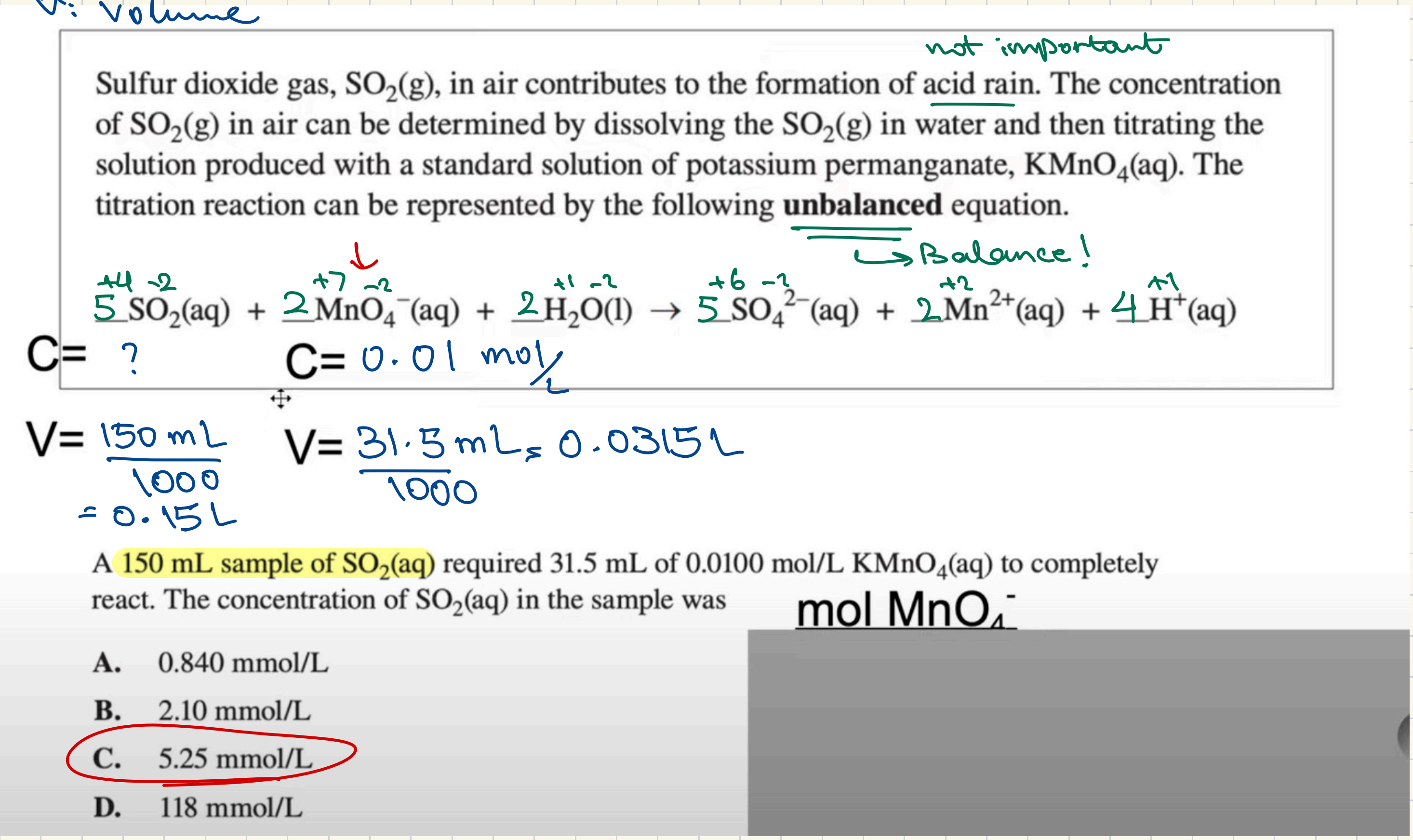
$Br_2 + 6H_2O \rightarrow 2BrO_3^-$ 2. O : $0 \rightarrow 6 \rightarrow$ add $6\ H_2O$

$Br_2 + 6H_2O \rightarrow 2BrO_3^- + 12H^+$ 3. H : $12 \rightarrow 0 \rightarrow$ add $12H^+$

$Br_2 + 6H_2O \rightarrow 2BrO_3^- + 12H^+ + 10e^-$ 4. e^- : $0 \rightarrow 10 \rightarrow$ losing $10e^-$

Redox Titration

C: Concentration of the limiting reactants
V: Volume



not important

Sulfur dioxide gas, $SO_2(g)$, in air contributes to the formation of acid rain. The concentration of $SO_2(g)$ in air can be determined by dissolving the $SO_2(g)$ in water and then titrating the solution produced with a standard solution of potassium permanganate, $KMnO_4(aq)$. The titration reaction can be represented by the following **unbalanced** equation.

$\hookrightarrow$ Balance!

$$\underline{5}\,\overset{+4\ -2}{SO_2}(aq) + \underline{2}\,\overset{+7\ -2}{MnO_4^-}(aq) + \underline{2}\,\overset{+1\ -2}{H_2O}(l) \rightarrow \underline{5}\,\overset{+6\ -2}{SO_4^{2-}}(aq) + \underline{2}\,\overset{+2}{Mn^{2+}}(aq) + \underline{4}\,\overset{+1}{H^+}(aq)$$

C= ? C= 0.01 mol/L

$V = \frac{150\ mL}{1000} = 0.15\ L$ $V = \frac{31.5\ mL}{1000} = 0.0315\ L$

A 150 mL sample of $SO_2(aq)$ required 31.5 mL of 0.0100 mol/L $KMnO_4(aq)$ to completely react. The concentration of $SO_2(aq)$ in the sample was

mol MnO_4^-

- **A.** 0.840 mmol/L
- **B.** 2.10 mmol/L
- [✓ **C.** 5.25 mmol/L]
- **D.** 118 mmol/L

Balancing: S: $+4 \rightarrow +6$: $2e^-/S \rightarrow 5 \times 2e^-/SO_2$
Mn: $+7 \rightarrow +2$: $5e^-/Mn \rightarrow 2 \times 5e^-/MnO_4^-$
LCM=10 ✓

O: $10 + 8 + x = 20 \rightarrow x = 2$
H: $4 = x$

---

mol $MnO_4 \rightarrow$ mol $SO_2 \rightarrow$ Conc mol/L SO_2

$$\text{mol/L } SO_2 = 0.01\ \text{mol/L } MnO_4^- \times 0.0315\ L \times \frac{5\ \text{mol } SO_2}{2\ \text{mol } MnO_4} \times \frac{1}{0.15\ L} =$$

$$0.00525\ \text{mol/L} \xrightarrow{\times 1000} = 5.25\ \text{mmol/L}$$

At this point, you could continue to balance the charge. However, if the question only asks for the coefficients of water and H+, you can stop here. In this case, the coefficient for water is 3, and for H+ is 6.

Balancing Charge (Continued)

If you need to balance the charge, here's how:

1. Calculate the total charge on each side.
2. Add electrons to the more positive side to balance the charge.

- Reactant side: -1 (from Br−)
- Product side: -1 (from BrO3−) + +6 (from 6H+) = +5

To balance, add 6 electrons to the product side to make it -1:

3H2O+Br−→BrO3−+6H++6e−

Redox Titration Example 🧫

Titrations, whether acid-base or redox, require you to visualize what's happening. Always determine the concentration and volume of the reactants.

Steps for Solving Titration Problems

1. **Organize Given Information:** List all known concentrations and volumes. Color-coding can be helpful.
2. **Convert:** Convert volumes to liters if necessary.
3. Find Moles of Known Reactant: Use the formula: moles=concentration×volume
4. **Use Stoichiometry:** Use the balanced equation to find the mole ratio between the known and unknown substances.
5. Find Concentration of Unknown: Use the formula: concentration=volumemoles

Example: Determining SO2 Concentration in Acid Rain

Problem: Sulfur dioxide gas causes acid rain. The SO2 concentration can be determined by titrating with potassium permanganate (KMnO4). 150 mL of SO2 requires 31.5 mL of 0.1 M KMnO4. What is the concentration of SO2?

The balanced equation is:

5SO2+2KMnO4+2H2O→K2SO4+2MnSO4+2H2SO4

Solution:

1. **Organize Information:**

Electrochemical Applications

This section focuses on electrochemical applications to both voltaic and electrolytic cells.

Changing the Reference Electrode

Why Hydrogen as a Reference? The lecture discusses what would happen if the hydrogen reference electrode were changed.

Subtracting to Re-zero: To switch the reference, subtract the E◦ reduction of the new reference from all data in the table. This re-zeros the table relative to the new reference.

- If nickel becomes the new reference, you would subtract (-0.26 V), which is equivalent to adding +0.26 V to all half-reactions.
- Anything above the new reference will have a positive voltage, and anything below will have a negative voltage.

The E° written in the table (P. 11) is when hydrogen is the reference, when changing that; we have to subtract the E° of the new ref from everything in the data table.

Example:

If the nickel standard nickel half-reaction was designated as the reference, what would happen to the iron Fe3+/Fe2+ half-reaction?

- The old value is +0.77 V.
- Add +0.26 V (because we're subtracting a negative value).
- The new voltage for the iron Fe3+/Fe2+ half-reaction is +1.03 V. Since the iron half-reaction is above the new reference (nickel), the value should be positive.

Standard Cell Potential with a New Reference

To calculate the standard cell potential of a lead-silver cell using cadmium as the reference:

1. **Adjust Potentials:** Cadmium is the new zero reference at -0.4 V. Adjust the potentials of silver and lead by adding 0.4 V to each.
2. **Calculate Ecell:** Use the formula: Ecell=Ecathode−Eanode

Whether you switch the reference or not, the cell voltage remains the same. Changing the reference is unnecessary when calculating cell voltage, but necessary if asked about a specific half-reaction.

Calculating Standard Cell Potential (Ecell)

Formula: To calculate the standard cell potential: Ecell=Ecathode−Eanode

- E_cathode: Reduction potential at the cathode (strongest oxidizing agent).
- E_anode: Reduction potential at the anode (strongest reducing agent).

Cell Notation: In standard cell notation (anode || cathode), the anode is always listed first, followed by the cathode.

Voltaic Cells: For voltaic cells, the top half-reaction in the table will be at the cathode, and the bottom half-reaction will be at the anode.

Example: Lead-Silver Cell

- Silver (Ag+) is the cathode: +0.80 V
- Lead (Pb) is the anode: -0.13 V

$Ecell=0.80-(-0.13)=+0.93$ V

Voltaic Cell Reminder: Voltaic cells *must* have a positive cell potential.

Voltaic vs. Electrolytic Cells

Steps to Calculating Ecell for Electrolytic Cells

1. **List All Species Present:** Include electrode materials (except inert electrodes) and dissociated electrolyte components.
2. **Identify Oxidizing Agents (OA) and Reducing Agents (RA):**
    - All elemental metals are reducing agents.
    - Water can act as both an oxidizing and reducing agent.
3. **Determine the Strongest OA and RA:**
    - Strongest OA reacts at the cathode.
    - Strongest RA reacts at the anode.
4. **Apply the Ecell Formula:** $Ecell=Ecathode-Eanode$

Example: Copper Nitrate Electrolytic Cell

- **Electrodes:** Silver (Ag), Carbon (C - inert)
- **Electrolyte:** Copper Nitrate (Cu(NO3)2)

1. **Species:** Ag, Cu2+, NO3−, H2O
2. **Identify OA and RA:**
    - **OA:** Cu2+, H2O
    - **RA:** Ag, H2O
3. **Determine Strongest OA and RA (using a standard reduction potential table):**
    - Cu2+ beats H2O as the strongest oxidizing agent.
    - Ag beats H2O as the strongest reducing agent.
4. **Ecell Calculation:**
    - Ecathode (Copper reduction): +0.34 V
    - Eanode (Silver oxidation): +0.80 V (from reduction potential table, Ag++e−→Ag is +0.80 V. Since Ag is oxidized, we use its reduction potential here in the formula.)
5. $Ecell=0.34-0.80=-0.46$ V

Because the calculated cell potential is negative, this reaction requires an external power source to occur.

Predicting Spontaneity 🧪

We can predict whether a reaction will be spontaneous using a table of reduction potentials. The key is to identify the **oxidizing agent (OA)** and **reducing agent (RA)** in the reaction.

Identifying Oxidizing and Reducing Agents

To determine spontaneity, identify the oxidizing and reducing agents:

- Determine what happens to each reactant (does it lose or gain electrons?)
- Remember that:
    - Losing electrons means it's being **oxidized**, and thus is the **reducing agent**.
    - Gaining electrons means it's being **reduced**, and thus is the **oxidizing agent**.
- A helpful hint to remember: Elemental metals are always reducing agents.

Using the Table

Once you've identified the OA and RA, look at their positions in the table of reduction potentials.

- If the **OA is above the RA** in the table, the reaction is **spontaneous**.
- If the **OA is below the RA** in the table, the reaction is **non-spontaneous** and requires a battery to occur.

Examples

Example 1: Lead and Iron

- **Reaction:** Elemental lead (Pb) reacting with iron(III) ion (Fe^{3+}).
- **Identify OA and RA:**
    - Lead goes from 0 to +2 oxidation state, so it loses electrons (oxidized) and is the **RA**.
    - Iron(III) must be the **OA**.
- **Check the Table:**
    - If the Fe^{3+} is above Pb in the table, the reaction is spontaneous.

Example 2: Chlorine and Fluoride

- **Reaction:** Elemental chlorine (Cl_2) and fluoride ion (F^-).
- **Identify OA and RA:**
    - Chlorine goes from 0 to -1, so it gains electrons (reduced) and is the **OA**.
    - Fluoride must be the **RA**.
- **Check the Table:**
    - If Cl_2 is below F^- in the table, the reaction is non-spontaneous and requires a battery.

Higher Mental Activity Questions 🤔

These questions may provide extra information to add to the table and then ask you to use the updated table to answer a question.

Example:

An experiment concludes that a new metal ion has a cell potential of 0.25 V. The question becomes, would the elemental form of this metal react with hydrochloric acid (HCl)?

1. **Add to the Table:**
    - Place the new metal ion (M+) and its elemental form (M) in the table based on the 0.25 V potential, locating it between the copper standard and the sulfate H+.
2. **Analyze the Reaction:**
    - HCl ionizes to H+ and Cl−. H+ is on the list and is the OA.
3. **Compare the positions of H+ and the new metal M.**
    - If H+ is below M, the reaction is non-spontaneous.

Faraday's Law ⚡

Faraday's Law is used to calculate the mass or time of a substance deposited during electrolysis.

The Pathway

The general pathway for these calculations is:

$$\text{Current (I)}\times\text{Time (t)}\rightarrow\text{Moles of Electrons}(ne-)\rightarrow\text{Moles of Substance}\rightarrow\text{Mass of Substance}$$

Alternatively, this can be expressed as a series of equations:

$$ne-=FI\times t$$

Where F is Faraday's constant.

Key Conversions

- Current and Time to Moles of Electrons:
  Use the formula: $ne-=F(I\times t)$
  Where:
    - I = Current (amps)
    - t = Time (seconds)
    - F = Faraday's constant (page 3 of data booklet)
- Moles of Electrons to Moles of Substance:
  Use the half-reaction for the substance to determine the stoichiometric ratio.

- Moles to Mass:
  Use the formula: moles=molar massmass

Example Problem 💡

Question: What is the time in seconds to plate 0.100 g of copper onto steel using a current of 15.0 amps?

Given half-reaction: Cu2++2e−→Cu(s)

Steps:

1. **Mass of Copper:**
    - Given: 0.100 g
2. **Moles of Copper:**
    - moles = molar mass*mass
    - moles=63.55 g/mol0.100 g=0.00157 mol
3. **Moles of Electrons:**
    - Use the half-reaction to find the ratio of moles of electrons to moles of copper.
    - From the half-reaction, 2 moles of electrons are required for every 1 mole of copper.
    - ne−=0.00157 mol Cu×1 mol Cu2 mol e−=0.00314 mol e−
4. **Time:**
    - Rearrange Faraday's Law to solve for time: t=I(ne−×F)
    - t=15.0 A(0.00314 mol e−×96485 C/mol)=20.2 s

General Outcome 2: Voltaic and Electrolytic Cells

This section focuses on voltaic and electrolytic cells, covering aspects like anode and cathode definitions, ion movement, the role of the salt bridge, internal and external circuits, and the similarities and differences between the two cell types.

Anode and Cathode Definitions

- **Anode:** The electrode where oxidation occurs (loss of electrons).
- **Cathode:** The electrode where reduction occurs (gain of electrons).

"LEO says GER" - Losing Electrons is Oxidation, Gaining Electrons is Reduction.

Ion Movement and the Salt Bridge 🌉

In an electrochemical cell, electrolytes are needed in both half-cells and in the salt bridge. The salt bridge is crucial for maintaining charge neutrality.

- It contains a salt, providing ions that can move to balance the charge in the half-cells.
- **Cations** (positive ions) flow towards the **cathode**.
- **Anions** (negative ions) flow towards the **anode**.
- If the salt bridge is missing or contains a non-electrolyte (e.g., methanol), the cell will not operate due to charge buildup.

Polarity

- **Voltaic Cell:**
  - Anode is negative.
  - Cathode is positive.
- **Electrolytic Cell:**
  - Anode is positive.
  - Cathode is negative.

External vs. Internal Cell

- **External Cell:** Consists of the wires and components (e.g., light bulb, voltmeter, power supply) that connect the electrodes.
- **Internal Cell:** Comprises the electrolytes and the electrodes within the cell.
- In an electrolytic cell, an external power supply is required to drive the non-spontaneous reaction.

Voltaic vs. Electrolytic Cells: Similarities and Differences

Porous Cup ☕

A porous cup is a ceramic container that allows ions and water to flow through it. It can be used in a voltaic cell to keep the different electrolyte solutions separate while still allowing ion flow to maintain charge neutrality.

Half-Reactions and Predicting Chemistry with Tables 🔎

To predict the chemistry that will occur in a cell, use a table of standard reduction potentials:

1. Identify the two half-reactions that could occur.
2. The half-reaction higher up in the table will occur as written (reduction at the cathode).
3. The half-reaction lower down in the table will be flipped (oxidation at the anode).

Example: Tin/Silver Voltaic Cell

- **Half-reactions:**
  - $Ag^+(aq)+e^- \rightarrow Ag(s)$ (occurs at cathode) gaining e^- is redoction → Cathode
  - $Sn(s) \rightarrow Sn^{2+}(aq)+2e^-$ (occurs at anode, flipped from table) losing e^- is Oxidation → Anode
- Electrons flow from the anode (tin) to the cathode (silver) in the external circuit.
- Cations move towards the cathode, and anions move towards the anode to maintain charge neutrality.

Electrolytic Cells: A Deeper Dive 🤿

Electrolytic cells require more analysis than voltaic cells:

1. List all species present in the electrolytic solution (ions, water).
2. Identify the possible oxidation and reduction half-reactions.
3. Determine the strongest oxidizing agent (best option for reduction at the cathode) and the strongest reducing agent (best option for oxidation at the anode) using the table of standard reduction potentials. Remember to flip the sign of the potential for the oxidation half-reaction.

Example: Electrolysis of Potassium Iodide (KI) Solution

- **Species present:** $K^+(aq)$, $I^-(aq)$, $H_2O(l)$.
- **Possible half-reactions:**
  - **Reduction (Cathode):**
    - $2H_2O(l)+2e^- \rightarrow H_2(g)+2OH^-(aq)$
    - $K^+(aq)+e^- \rightarrow K(s)$
  - **Oxidation (Anode):**
    - $2I^-(aq) \rightarrow I_2(s)+2e^-$
    - $2H_2O(l) \rightarrow O_2(g)+4H^+(aq)+4e^-$

- **Determine winners (using standard reduction potentials):**
    - **Strongest oxidizing agent:** Water (H_2O) will be reduced at the cathode.
    - **Strongest reducing agent:** Iodide (I^-) will be oxidized at the anode.
- **Overall reactions:**
    - **Cathode (Reduction):** $2H_2O(l)+2e^- \rightarrow H_2(g)+2OH^-(aq)$
    - **Anode (Oxidation):** $2I^-(aq) \rightarrow I_2(s)+2e^-$

In this process, hydrogen gas (H_2) is produced at the cathode, and iodine (I_2) is produced at the anode, which would appear as a brown color in the solution. Hydroxide ions (OH^-) can also be detected at the cathode.

The positive side of the power supply is connected to the anode, and the negative side is connected to the cathode, forcing the non-spontaneous reactions to occur.

Electrochemical Cells: Voltaic vs. Electrolytic

Ion Flow in Electrochemical Cells

When considering ion flow, remember that ions in solution also move to balance charge. For example, iodide ions (I^-) will move towards the anode.

Comparing Voltaic and Electrolytic Cells ⚖️

Here's a comparison of voltaic and electrolytic cells, highlighting the key differences:

Key takeaway: Electron flow is always external to the cell, while ion flow occurs within the solution.

Sample Question: Identifying Cell Characteristics 📝

Consider the following question that tests your understanding of the similarities and differences between voltaic and electrolytic cells:

Which row identifies a characteristic of each type of cell?

To solve this, remember:

- Voltaic cells have $E_{cell}>0$, while electrolytic cells have $E_{cell}<0$.
- Voltaic cells convert chemical energy to electrical energy.
- Anions move toward the anode.
- Electrolytic cells require an external power supply.

Electrolytic Cell Example: The Chloride Anomaly 🧪

Sometimes, standard tables don't accurately predict reactions, especially in electrolytic cells. Consider an electrolytic cell with aqueous sodium chloride ($NaCl$) and hydrochloric acid (HCl).

Problem: What product is produced at the anode?

Solution:

1. **List all ions:**
    - $NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$
    - $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
    - $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$
2. **Identify oxidizing agents (OA) and reducing agents (RA):**
    - **OAs:** Na^+, H^+, H_2O → gains e^-
    - **RAs:** Cl^-, H_2O → loses $\bar{e}$
3. **Apply the Chloride Anomaly:** When Cl^- and H_2O are present, Cl^- reacts instead of water, despite the standard tables suggesting otherwise.
   Chloride Anomaly: In certain electrolytic conditions, chloride ions (Cl^-) react preferentially over water, even if standard reduction potentials suggest otherwise.

In this case, Cl^- is oxidized at the anode:

$$2Cl^- \rightarrow Cl_2(g) + 2e^-$$

Therefore, the product produced at the anode is chlorine gas (Cl_2).

Scholars Of Calgary Northwest

Chemistry 30, Unit 2: Electrochemical Changes, 10 Questions

Your score was 7/10=70%. You took 23 minutes, 53 seconds to complete the test.

Click here to do this test again!

1.

$MnO_4^- + H_2C_2O_4 \longrightarrow CO_2 + Mn^{2+}$

When the above reaction that occurs in acid solution is balanced, the coefficient in front of CO_2 is

A) 4
B) 6
C) 8
D) 10

***Correct**. Your answer=D, Correct answer=D*

Explanation:

reduction half reaction: $MnO_4^- \longrightarrow Mn^{2+}$

balance O atoms with H^2O: $MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$

balance H atoms using H^+: $MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$

balance charge by adding e^-: $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$ (1)

oxidation half reaction: $H_2C_2O_4 \longrightarrow CO_2$

balance the C atoms: $H_2C_2O_4 \longrightarrow 2CO_2$

balance H atoms using H^+: $H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+$

balance charge using e^-: $H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+ + 2e^-$ (2)

Gain of e^- in reduction must equal the loss of e^- in oxidation
Multiply equation (1) by 2 and equation (2) by 5

$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$ (3)

$5H_2C_2O_4 \longrightarrow 10CO_2 + 10H^+ + 10\,e^-$ (4)

Add equations (3) and (4) and cancel any species common to both sides.
$2MnO_4^- + 5H_2C_2O_4 + 6H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

2.

1 perchloric acid
2 iron (II) chloride
3 silver nitrate
4 lithium nitrate

Order the solutions above from strongest oxidizing agent to weakest.
(Enter your response as numbers separated by commas.)

Incorrect*. Your answer=1,3,4,2*

Correct answers:

1, 3, 2, 4
1,3,2,4

Explanation:

Look up these solution components in the Table of Relative Strengths of Oxidizing and Reducing Agents, and determine the SOA by comparing their position on the left hand column.

The perchloric ion, $ClO^{-}{}_{4(aq)}$ is closest to the top, followed by $Ag^{+}{}_{(aq)}$, then $Fe^{2+}{}_{(aq)}$ and finally $Li^{+}{}_{(aq)}$.

3. One way in which electrolytic cells differ from voltaic cells is that

   A) oxidation occurs at the anode in one but at the cathode in the other
   B) cations migrate to the cathode in one but to the cathode in the other
   C) electrolytic cells produce energy but voltaic consume it
   D) the cell potential for one is positive but the other is negative

   Correct*. Your answer=D, Correct answer=D*

   Explanation: An electrolytic cell differs from a voltaic (also known as galvanic, electric or electrochemical) cell in that electricity is applied to force a non-spontaneous reaction while the reaction is spontaneous in a voltaic cell. The cell potential, ΔE°, is a numerical expression of the strength and direction of the reaction; negative in electrolytic and positive in voltaic cells. In both cell types, the cathode is the site of reduction while the anode is the site of oxidation, and anions and cations migrate to the anode and cathode respectively.

4. Corrosion of iron metal is a multi-billion dollar infrastructure cost every year. The reaction that occurs during corrosion is

   $$4Fe_{(s)} + 3O_{2(g)} + 6H_2O_{(l)} \longrightarrow 4Fe(OH)_{3(s)} + \text{energy}$$

   If 20.0 g of iron corroded, then the volume of oxygen gas consumed during the reaction at SATP (to the nearest hundredth) is __________ L. (Note: 1 mol of oxygen at SATP occupies 24.8 L)

   Correct*. Your answer=6.66, Correct answer=6.66*

   Explanation: For each 4 mol of $Fe_{(s)}$ corroded, 3 moles of oxygen gas are used.
   moles of Fe = g of Fe/molar mass of Fe = 20.0 g / 55.85 g/mol = 0.358 mol
   moles of $O_2 = \frac{3}{4} \times$ moles of Fe $= \frac{3}{4} \times 0.358$ mol = 0.268 mol
   V of O_2 at SATP = mol of O_2 × V/mol at SATP = 0.268 mol × 24.8 L/mol = 6.66 L

5. A voltaic cell is constructed using a $Cu_{(s)}$ / $Cu^{2+}{}_{(aq)}$ half–cell connected to a $C_{(s)}$ / $Cr_2O_7{}^{2-}{}_{(aq)}$, $H^{+}{}_{(aq)}$ half–cell.

   The correct net reaction of the cell is

   A) $3Cu^{2+}{}_{(aq)} + Cr_2O_7{}^{2-}{}_{(aq)} + 14H^{+}{}_{(aq)} \longrightarrow 3Cu_{(s)} + 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$

B) $3Cu_{(s)} + Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} \longrightarrow 3Cu^{2+}{}_{(aq)} + 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$

C) $Cu^{2+}{}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} \longrightarrow Cu_{(s)} + 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$

D) $Cu_{(s)} + Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} \longrightarrow Cu^{2+}{}_{(aq)} + 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$

Incorrect*. Your answer=D, Correct answer=B*

Explanation: List all possible redox participants:

$Cu_{(s)}$, $Cu^{2+}{}_{(aq)}$, $Cr_2O_7^{2-}{}_{(aq)}$, $H^+{}_{(aq)}$
note: $C_{(s)}$ is an inert electrode and will not be involved.

Next identify the OAs and RAs and chose half reactions for oxidation and reduction. Refer to the Table of Relative Strengths of Oxidizing and Reducing Agents.

cathode:
$Cr_2O_7^{2-}{}_{(aq)} + 14\ H^+{}_{(aq)} + 6e^- \longleftrightarrow 2Cr^{3+} + 7H_2O_{(l)}$
anode:
$Cu_{(s)} \longleftrightarrow Cu^{2+}{}_{(aq)} + 2e^-$

The net reaction requires multiplying the copper half reaction by 3 to balance the electrons. Choice B is the correct equation.

6.

	Properties
1	Reacts spontaneously with $Zn^{2+}{}_{(aq)}$
2	Reacts spontaneously with $Ag_{(s)}$
3	Is an oxidizing **and** a reducing agent
4	Is reduced by $Au_{(s)}$
5	Reacts spontaneously with $K^+{}_{(aq)}$
6	Reacts spontaneously with $Al^{3+}{}_{(aq)}$

Which property in the list is most appropriate for $Cr_{(s)}$?

Correct*. Your answer=1, Correct answer=1*

Explanation:

Refer to the Table of Relative Strengths of Oxidizing and Reducing Agents.

Zn^{2+}, the OA (oxidizing agent), is above $Cr_{(s)}$, the RA (reducing agent) so their combination would result in a spontaneous redox reaction, with $Cr_{(s)}$ oxidized to Cr^{2+}.

$Cr_{(s)}$ would have no reaction with $Au_{(s)}$ or $Ag_{(s)}$, both weaker RAs than $Cr_{(s)}$.

$Cr_{(s)}$ as a RA is above the OAs $K^+{}_{(aq)}$ and $Al^{3+}{}_{(aq)}$ so spontaneous reactions would not occur.

7. Refer to the information in question 6.

Which property in the list is most appropriate for $Cl_{2(g)}$?

Correct*. Your answer=2, Correct answer=2*

Explanation:

$Cl_{2(g)}$ as an OA (oxidizing agent) is near the top of the Table of Relative Strengths of Oxidizing and Reducing Agents, above RA (reducing agent) $Ag_{(s)}$, so a spontaneous reaction would occur.

No reaction for $Cl_{2(g)}$ as an OA and $Au_{(s)}$ as an RA would occur since the RA is above the OA.

No reactions would occur between $Cl_{2(g)}$ and $Zn^{2+}_{(aq)}$, $K^{+}_{(aq)}$, and $Al^{3+}_{(aq)}$, since all are OAs.

8. Which of the following equations is **not** predicted to represent a redox reaction?

   A) $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$
   B) $2Fe^{2+}_{(aq)} \longrightarrow Fe_{(s)} + Fe^{3+}_{(aq)}$
   C) $Na^{+}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow NaCl_{(s)}$
   D) $2H_2O_{(l)} + O_{2(g)} \longrightarrow 2H_2O_{2(l)}$

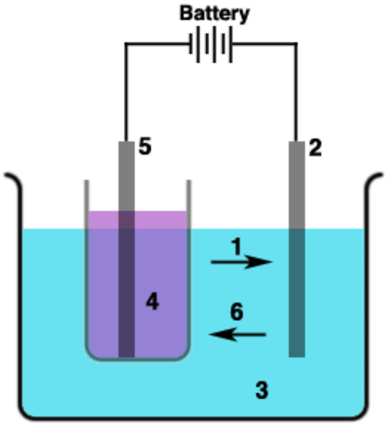
   ***Correct**. Your answer=C, Correct answer=C*

   Explanation:

   Choice C) is not a redox reaction since the oxidation numbers of the participating ions do not change.

9. The diagram below represents a cell:

   $Sn_{(s)}$ / $Sn^{2+}_{(aq)}$ // $MnO_4^{-}{}_{(aq)}$, $Mn^{2+}_{(aq)}$, $H^{+}_{(aq)}$ / $C_{(s)}$



   The anion migration is represented by

   ***Incorrect**. Your answer=6, Correct answer=1*

   Explanation: Anions are negatively charged ions. Migration is movement. The direction of ion movement is always anions towards the anode and cations towards the cathode. In this diagram of an electrolytic cell, the anode (#4) may be identified by the direction of flow of electrons towards the cathode (#5). The anions in solution identified as # 1 are moving towards the anode.

10. Which of the following chemicals is the strongest reducing agent?

    A) $Sn_{(s)}$
    B) $Cr^{2+}_{(aq)}$
    C) $Ag_{(s)}$
    D) $I^{-}_{(aq)}$

    ***Correct**. Your answer=B, Correct answer=B*

Explanation: Refer to the Table of Relative Strengths of Oxidizing and Reducing Agents.

The strongest RA will appear closest to the bottom of the table on the right hand side.

$Cr^{2+}_{(aq)}$ is below $Ag_{(s)}$, $Sn_{(s)}$ and $I^{-}_{(aq)}$, so choice B) is the strongest reducing agent.

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Chemistry 30, Unit 2: Electrochemical Changes, 10 Questions

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W

C

1. Which of the following generalizations is true of the Table of Selected Standard Electrode Potentials?

OA ✓

○ A) Metal ions and ~~non-metallic~~ elements are generally oxidizing agents.
○ B) Metal ions and non-metallic ions are generally oxidized.
● C) Metals and non-metal elements are generally oxidized.
○ D) Metals and non-metal ions are generally reduced.

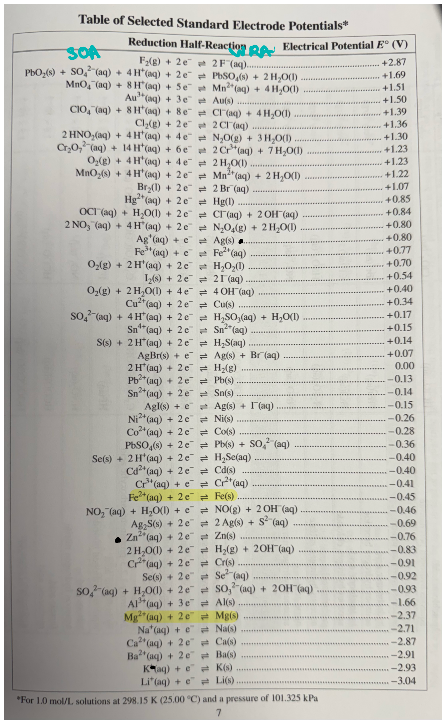

Table of Selected Standard Electrode Potentials*

SOA

WRA

Reduction Half-Reaction	Electrical Potential $E°$ (V)
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$	+1.69
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.07
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$	+0.85
$OCl^-(aq) + H_2O(l) + 2e^- \rightleftharpoons Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s)$	−2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

*For 1.0 mol/L solutions at 298.15 K (25.00 °C) and a pressure of 101.325 kPa

7

2. ✓

Properties

~~1~~ Reacts spontaneously with $Zn^{2+}_{(aq)}$
2 Reacts spontaneously with $Ag_{(s)}$ ✓
~~3~~ Is an oxidizing **and** a reducing agent
~~4~~ Is reduced by $Au_{(s)}$
~~5~~ Reacts spontaneously with $K^+_{(aq)}$
~~6~~ Reacts spontaneously with $Al^{3+}_{(aq)}$

Which property in the list is most appropriate for $Cl_{2(g)}$?

2

$-8+1=-7$

$⑤-6=-1$

3. ✓ The oxidation numbers for chlorine in $\overset{+4\ -2}{ClO_{2(g)}}$, $Cl_{2(g)}$, $\overset{+5\ -2}{ClO_3^-{}_{(aq)}}$, and $\overset{+7}{\overset{+1\ \uparrow\ -2}{HClO_{4(aq)}}}$, are:
(Write your response as four numbers separated by commas.)

4,0,5,7

RA
Mg-.OX

OA
Fe red

4. ✓ If a block of refined magnesium were selected to serve as a sacrificial anode and if it were bolted onto the iron girder of a bridge, one would expect the

○ A) iron to oxidize
● B) magnesium to oxidize before the iron
○ C) magnesium to reduce before the iron
○ D) both the iron and the magnesium to react

5.

In a breathalyzer, ethanol from the suspect's breath is oxidized by an acidic dichromate solution in the reaction ampule as in the following reaction:

$$2Cr_2O_7^{2-}{}_{(aq)} + 16H^+_{(aq)} + 3C_2H_5OH_{(aq)} \longrightarrow 4Cr^{3+}_{(aq)} + 11H_2O_{(l)} + 3CH_3COOH_{(aq)}$$

The ampule which contains the acidic dichromate testing solution would **best** be made of:

○ A) chromium.
○ B) zinc.
● C) plastic.
○ D) copper.

6.

Corrosion of iron metal is a multi-billion dollar infrastructure cost every year. The reaction that occurs during corrosion is

$$4Fe_{(s)} + 3O_{2(g)} + 6H_2O_{(l)} \longrightarrow 4Fe(OH)_{3(s)} + \text{energy}$$

If 20.0 g of iron corroded, then the volume of oxygen gas consumed during the reaction at SATP (to the nearest hundredth) is __________ L. (Note: 1 mol of oxygen at SATP occupies 24.8 L)

6.66

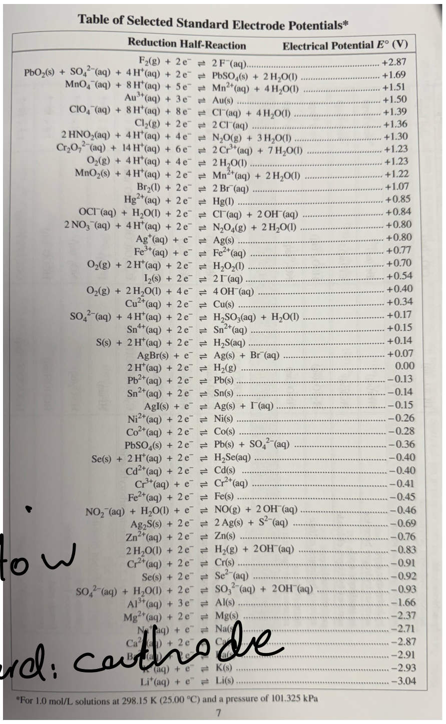

Table of Selected Standard Electrode Potentials*

Reduction Half-Reaction	Electrical Potential $E°$ (V)
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$	+1.69
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.07
$Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$	+0.85
$OCl^-(aq) + H_2O(l) + 2e^- \rightleftharpoons Cl^-(aq) + 2OH^-(aq)$	+0.84
$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	+0.17
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
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$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
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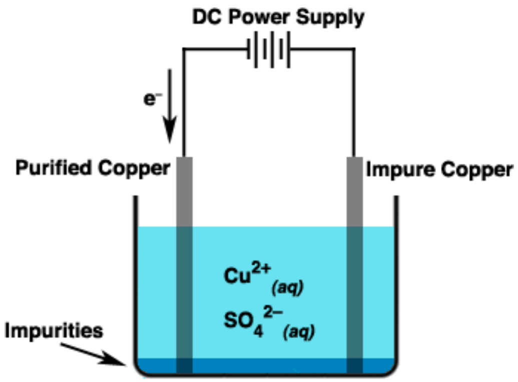
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons$ [illegible]	−2.87
[illegible]	−2.91
[illegible] $(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

*For 1.0 mol/L solutions at 298.15 K (25.00 °C) and a pressure of 101.325 kPa

7

7.

A common way to purify copper is to use an electrolytic cell like that shown below:



- ions: flow tow

ion ē flow toward: cathode

In this electrolytic cell, the impure copper acts as the

● A) anode and is the site where $Cu_{(s)}$ is oxidized ✓
○ B) anode and is the site where $Cu^{2+}{}_{(aq)}$ ions are oxidized
○ C) cathode and is the site where $Cu_{(s)}$ is reduced
○ D) cathode and is the site where $Cu^{2+}{}_{(aq)}$ ions are reduced

A: OX
C: Red

ande +
cath − SO_4^{2-}?

8.

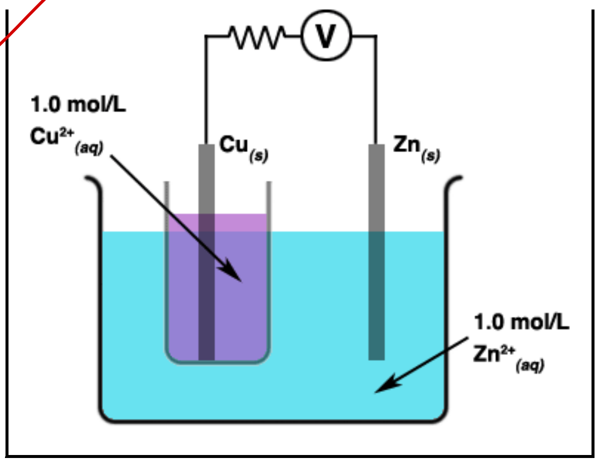
$$4H^+ + \overbrace{\overset{+4\ -2}{PbO_2} + \overset{-1}{4Cl^-} \longrightarrow \overset{+2\ -1}{PbCl_2}}^{+2e} + \overset{0}{Cl_2} + 2H_2O$$

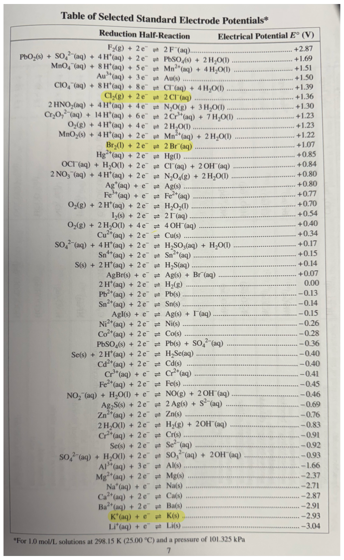
e

When the following reaction that occurs in acid solution is balanced, the coefficient in front of Cl_2 is

● A) 1
○ B) 2
○ C) 3
○ D) 4

9.





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$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
$ClO_4^-(aq) + 8H^+(aq) + 8e^- \rightleftharpoons Cl^-(aq) + 4H_2O(l)$	+1.39
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$2HNO_2(aq) + 4H^+(aq) + 4e^- \rightleftharpoons N_2O(g) + 3H_2O(l)$	+1.30
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
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$2NO_3^-(aq) + 4H^+(aq) + 2e^- \rightleftharpoons N_2O_4(g) + 2H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
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$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(l)$	+0.70
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
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$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14
$AgBr(s) + e^- \rightleftharpoons Ag(s) + Br^-(aq)$	+0.07
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$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$AgI(s) + e^- \rightleftharpoons Ag(s) + I^-(aq)$	−0.15
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.26
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	−0.36
$Se(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2Se(aq)$	−0.40
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	−0.40
$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.41
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.45
$NO_2^-(aq) + H_2O(l) + e^- \rightleftharpoons NO(g) + 2OH^-(aq)$	−0.46
$Ag_2S(s) + 2e^- \rightleftharpoons 2Ag(s) + S^{2-}(aq)$	−0.69
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$	−0.91
$Se(s) + 2e^- \rightleftharpoons Se^{2-}(aq)$	−0.92
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	−0.93
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$Ba^{2+}(aq) + 2e^- \rightleftharpoons Ba(s)$	−2.91
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

*For 1.0 mol/L solutions at 298.15 K (25.00 °C) and a pressure of 101.325 kPa

7

Which of the following is true of the above electrochemical cell?

○ A) There color of the copper electrolyte will become more intense
● B) Electrons will travel from the zinc electrode to the copper electrode.
○ C) Anions will migrate towards the copper electrode.
○ D) The reaction will not occur as there is no salt bridge.

Cu^{2+} : OA : reduced : Cath
Zn : RA : OX : ande

OA: red RA → OX

10. When chlorine gas is bubbled through a potassium bromide solution, the solution:

● A) becomes orange because bromide ions are oxidized.
○ B) becomes orange because bromide ions are reduced.
○ ~~C) becomes green because chlorine is oxidized.~~
○ ~~D) undergoes no reaction.~~

$\overset{0}{Cl_2} + \overset{+1}{K}\overset{-1}{Br} \rightarrow \overset{+1}{K}\overset{-1}{Cl} + \overset{0}{Br_2}$

R: OA ~~OX~~ : RA

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Chemistry 30, Unit 2: Electrochemical Changes, 10 Questions

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1. Which of the following generalizations is true of the Table of Selected Standard Electrode Potentials?

    A) Metal ions and non-metallic elements are generally oxidizing agents.
    B) Metal ions and non-metallic ions are generally oxidized.
    C) Metals and non-metal elements are generally oxidized.
    D) Metals and non-metal ions are generally reduced.

    You didn't answer this question.

    ***Incorrect**. Your answer= , Correct answer=A*

    Explanation: Metals in elemental form behave as RAs (reducing agents, i.e. they are oxidized, they rust). When the reaction goes in the other direction, the metal ions are OAs (oxidizing agents). Non-metal ions are usually RAs and non-metal elements are OAs. Remember that OAs reduce and RAs oxidize. Choices B, C and D are all false and only choice A is a true statement.

2.

    Properties

    1 Reacts spontaneously with $Zn^{2+}{}_{(aq)}$
    2 Reacts spontaneously with $Ag_{(s)}$
    3 Is an oxidizing **and** a reducing agent
    4 Is reduced by $Au_{(s)}$
    5 Reacts spontaneously with $K^{+}{}_{(aq)}$
    6 Reacts spontaneously with $Al^{3+}{}_{(aq)}$

    Which property in the list is most appropriate for $Cl_{2(g)}$?

    You didn't answer this question.

    ***Incorrect**. Your answer=, Correct answer=2*

    Explanation:

    $Cl_{2(g)}$ as an OA (oxidizing agent) is near the top of the Table of Relative Strengths of Oxidizing and Reducing Agents, above RA (reducing agent) $Ag_{(s)}$, so a spontaneous reaction would occur.

    No reaction for $Cl_{2(g)}$ as an OA and $Au_{(s)}$ as an RA would occur since the RA is above the OA.

    No reactions would occur between $Cl_{2(g)}$ and $Zn^{2+}{}_{(aq)}$, $K^{+}{}_{(aq)}$, and $Al^{3+}{}_{(aq)}$, since all are OAs.

3. The oxidation numbers for chlorine in $ClO_{2(g)}$, $Cl_{2(g)}$, $ClO_3{}^{-}{}_{(aq)}$, and $HClO_{4(aq)}$, are:

(Write your response as four numbers separated by commas.)

You didn't answer this question.

***Incorrect**. Your answer=*

Correct answers:

4, 0, 5, 7
4,0,5,7

Explanation:

In $ClO_2\ _{(g)}$, the 2 O ions each have a charge of 2– for a total of –4. Cl must be 4+.

$Cl_{2(g)}$ is an element and all elements have an oxidation number of 0.

In $ClO_3{}^-{}_{(aq)}$, the 3 O ions each have a charge of 2– for a total of –6. The overall charge is –1 so Cl must be 5+.

In $HClO_{4(aq)}$, the 4 O ions each have a charge of 2– for a total of –8. The H gives +1 so the Cl must be 7+.

4. If a block of refined magnesium were selected to serve as a sacrificial anode and if it were bolted onto the iron girder of a bridge, one would expect the

   A) iron to oxidize
   B) magnesium to oxidize before the iron
   C) magnesium to reduce before the iron
   D) both the iron and the magnesium to react

   You didn't answer this question.

   ***Incorrect**. Your answer= , Correct answer=B*

   Explanation: Refined magnesium is $Mg_{(s)}$. $Mg_{(s)}$ is below $Fe_{(s)}$ on the Table of Relative Strengths of Oxidizing and Reducing Agents so a spontaneous reaction would occur if $Fe_{(s)}$ is the OA and $Mg_{(s)}$ is the RA. This would result in the iron staying in its metallic form and the $Mg_{(s)}$ being oxidized as stated by choice B.

5. In a breathalyzer, ethanol from the suspect's breath is oxidized by an acidic dichromate solution in the reaction ampule as in the following reaction:

   $$2Cr_2O_7{}^{2-}{}_{(aq)} + 16H^+{}_{(aq)} + 3C_2H_5OH_{(aq)} \longrightarrow 4Cr^{3+}{}_{(aq)} + 11H_2O_{(l)} + 3CH_3COOH_{(aq)}$$

   The ampule which contains the acidic dichromate testing solution would **best** be made of:

   A) chromium.
   B) zinc.
   C) plastic.
   D) copper.

   You didn't answer this question.

   ***Incorrect**. Your answer= , Correct answer=C*

   Explanation:

   Metals like $Cr_{(s)}$, $Zn_{(s)}$, and $Cu_{(s)}$, are all reducing agents and might interfere with the redox reaction between chromium and ethanol. This could produce falsely low or erroneously high results.

   An ampule made of a non-reactive material, like plastic, would provide a more accurate test result.

6.

Corrosion of iron metal is a multi-billion dollar infrastructure cost every year. The reaction that occurs during corrosion is

$$4Fe_{(s)} + 3O_{2(g)} + 6H_2O_{(l)} \longrightarrow 4Fe(OH)_{3(s)} + \text{energy}$$

If 20.0 g of iron corroded, then the volume of oxygen gas consumed during the reaction at SATP (to the nearest hundredth) is __________ L. (Note: 1 mol of oxygen at SATP occupies 24.8 L)

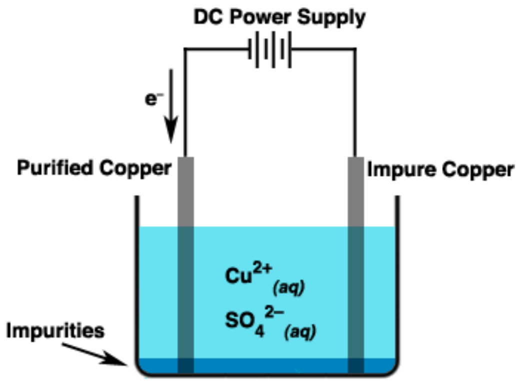
You didn't answer this question.

Incorrect*. Your answer=, Correct answer=6.66*

Explanation: For each 4 mol of $Fe_{(s)}$ corroded, 3 moles of oxygen gas are used.
moles of Fe = g of Fe/molar mass of Fe = 20.0 g / 55.85 g/mol = 0.358 mol
moles of $O_2 = \frac{3}{4} \times$ moles of Fe $= \frac{3}{4} \times 0.358$ mol = 0.268 mol
V of O_2 at SATP = mol of $O_2 \times$ V/mol at SATP = 0.268 mol × 24.8 L/mol = 6.66 L

7.

A common way to purify copper is to use an electrolytic cell like that shown below:



In this electrolytic cell, the impure copper acts as the

A) anode and is the site where $Cu_{(s)}$ is oxidized
B) anode and is the site where $Cu^{2+}{}_{(aq)}$ ions are oxidized
C) cathode and is the site where $Cu_{(s)}$ is reduced
D) cathode and is the site where $Cu^{2+}{}_{(aq)}$ ions are reduced

You didn't answer this question.

Incorrect*. Your answer= , Correct answer=A*

Explanation: The Impure Copper is the source of copper metal atoms which are oxidized to copper ions, then reduced back to pure copper metal at the cathode leaving their impurities behind. The anode is the site of oxidation, so choice A is correct.

8.

$$PbO_2 + Cl^- \longrightarrow PbCl_2 + Cl_2$$

When the following reaction that occurs in acid solution is balanced, the coefficient in front of Cl_2 is

A) 1
B) 2
C) 3
D) 4

You didn't answer this question.

***Incorrect**. Your answer= , Correct answer=A*

Explanation:

reduction half reaction: $PbO_2 \longrightarrow PbCl_2$

balance Cl atoms : $PbO_2 + Cl^- \longrightarrow PbCl_2$

balance O atoms using H_2O: $PbO_2 + 2Cl^- \longrightarrow PbCl_2 + 2H_2O$

balance H atoms using H^+: $PbO_2 + 2Cl^- + 4H^+ \longrightarrow PbCl_2 + 2H_2O$

balance charge by adding e^-: $PbO_2 + 2Cl^- + 4H^+ + 2e^- \longrightarrow PbCl_2 + 2H_2O$ (1)

oxidation half reaction: $Cl^- \longrightarrow Cl_2$

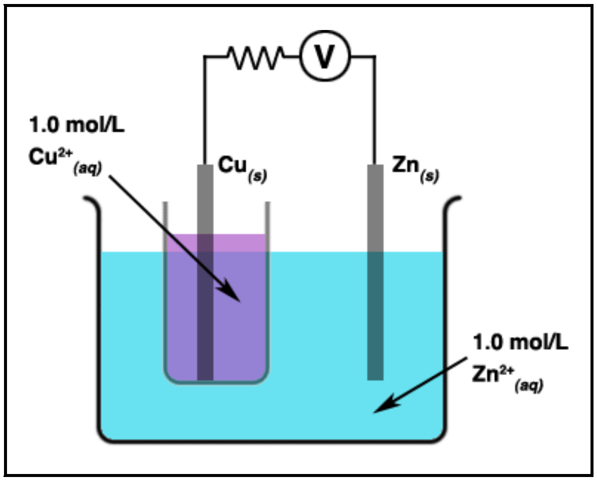
balance the Cl atoms: $2Cl^- \longrightarrow Cl_2$

balance charge using e^-: $2Cl^- \longrightarrow Cl_2 + 2e^-$ (2)

Gain of e^- in reduction must equal the loss of e^- in oxidation

Add equations (1) and (2).
$PbO_2 + 4Cl^- + 4H^+ \longrightarrow PbCl_2 + 2H_2O + Cl_2$

9.



Which of the following is true of the above electrochemical cell?

A) There color of the copper electrolyte will become more intense
B) Electrons will travel from the zinc electrode to the copper electrode.
C) Anions will migrate towards the copper electrode.
D) The reaction will not occur as there is no salt bridge.

You didn't answer this question.

***Incorrect**. Your answer= , Correct answer=B*

Explanation: In the diagram, $Cu_{(s)}$ is the cathode and $Zn_{(s)}$ is the anode. The characteristic blue colour of the copper ion will decrease with its reduction to metallic copper. Electrons always travel from anode (zinc electrode) to cathode (copper electrode) so choice B is correct. Anions move towards the anode which is the zinc electrode. A salt bridge is not necessary since the porous inner cup allows for ion movement.

10. When chlorine gas is bubbled through a potassium bromide solution, the solution:

    A) becomes orange because bromide ions are oxidized.
    B) becomes orange because bromide ions are reduced.
    C) becomes green because chlorine is oxidized.
    D) undergoes no reaction.

You didn't answer this question.

Incorrect*. Your answer= , Correct answer=A*

Explanation:

Refer to the Table of Relative Strengths of Oxidizing and Reducing Agents. $Cl_{2(g)}$ behaves as an OA and is above the RA $Br^{-}_{(aq)}$ which suggests that a spontaneous redox reaction will occur.

$2Br^{-}_{(aq)} \longleftrightarrow Br_{2(l)} + 2e^{-}$

The colour of the oxidized bromide is the characteristic orange of the liquid element bromine.

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