DPP-2 [Redox Titration] – Solutions

Chapter: Practical Physical Chemistry

*"Sab log talent ki baat karte hain…Par asli game un logon ne jeeta hai…
Jinhone boring dinon mein bhi kaam kiya."*

TYPE 1 : Fundamentals of Redox Titration

Q.1 The reagent which forces other molecule or ion to liberate electron is

Explanation

An **oxidizing agent** is the species that forces others to lose (liberate) electrons. In doing so, the oxidizing agent itself *gains* electrons and gets reduced. Think of it as the "electron thief" – it doesn't give electrons, it **takes** them away from others, thereby forcing the victim to liberate those electrons.

Green Approach

Analogy – The Bully at school: An oxidizing agent is like a school bully who forces weaker students to hand over (liberate) their lunch money (electrons). The bully takes – so the victim *loses*. Remember: OIL RIG – **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain. The oxidizing agent causes oxidation in others, so it makes them *lose* electrons.

Answer

Option (1) – Oxidizing agent

Q.2 The reagent which forces other molecule or ion to capture electron is

Explanation

A **reducing agent** is the species that forces others to gain (capture) electrons. The reducing agent itself *gives away* electrons and gets oxidised. It is the "electron donor" that pushes electrons onto others, forcing them to capture those electrons.

Green Approach

Analogy – The Generous Uncle: A reducing agent is like a generous uncle who forces gifts (electrons) on relatives – the relatives *must* accept (capture). The uncle loses money (gets oxidized), while the relative gains (gets reduced). Reducing agent → causes reduction in others → makes them *gain* electrons.

Answer

Option (2) – Reducing agent

Q.3 Which is the self indicator used in redox titration?

Explanation

$KMnO_4$ (potassium permanganate) is a **self-indicator**. It is intensely purple/violet in colour. During titration with a reducing agent (e.g. oxalic acid or $FeSO_4$), $KMnO_4$ gets reduced to the almost colourless Mn^{2+} ion. At the equivalence point, one extra drop of $KMnO_4$ imparts a permanent faint pink/violet colour – no external indicator is needed. Hence it acts as its own indicator. Ferrous ammonium sulphate (Mohr's salt) is the *analyte*, not a self-indicator.

Green Approach

Analogy – The colour-changing shirt: Imagine $KMnO_4$ wears a dark purple shirt. As it reacts (gets reduced), it changes into colourless Mn^{2+} – shirt gone! The moment the reaction is over and one extra drop of $KMnO_4$ arrives, the solution turns faint pink again – the shirt reappears. No outside person needed to signal the endpoint; the reagent itself signals it!

Answer

Option (2) – $KMnO_4$

Q.4 Which meniscus is read in burette in case of $KMnO_4$?

Explanation

For most colourless or light-coloured solutions in a burette, the **lower (concave) meniscus** is read. However, $KMnO_4$ is dark purple and almost opaque – you *cannot* see the lower meniscus clearly. Therefore, the **upper (convex) meniscus** is read for $KMnO_4$.

Green Approach

Analogy – Dark tinted glass: Reading the lower meniscus of $KMnO_4$ is like trying to see through a bottle of dark ink – impossible! So you read the top of the liquid (upper meniscus) instead, just as you'd look at the top of an opaque dark liquid.

Answer

Option (1) – Upper meniscus

Q.5 The titration of $KMnO_4$ against oxalic acid or $FeSO_4$ is example of

Explanation

In this titration, $KMnO_4$ acts as an oxidizing agent and oxalic acid/$FeSO_4$ acts as the reducing agent. Electrons are transferred from the reducing agent to $KMnO_4$. Since the fundamental reaction involves **transfer of electrons (oxidation–reduction)**, this is classified as a **redox titration**.

Green Approach

Quick recall – classify by what changes: Acid-base = proton transfer; precipitation = insoluble salt forms; complexometric = coordination compound forms; **redox = electrons transfer**. Here electrons move from oxalic acid/Fe^{2+} to Mn^{7+} $\Rightarrow$ redox titration.

Answer

Option (4) – Redox titration

Q.6 The titration gives unsatisfactory results when carried out in the presence of HCl, because HCl…

Explanation

$KMnO_4$ is a very strong oxidizing agent. When HCl is used instead of H_2SO_4:

- $KMnO_4$ oxidizes Cl^- ions from HCl to Cl_2 gas.
- This means $KMnO_4$ is "wasted" on oxidizing Cl^- instead of just the analyte (oxalic acid).
- More $KMnO_4$ is consumed than expected ⇒ results are unsatisfactory.

In other words, HCl **reduces permanganate to Mn^{2+}** (the Cl^- donates electrons to MnO_4^-, reducing it).

Green Approach

Analogy – Two thieves, one cop: $KMnO_4$ (the cop) should arrest only the oxalic acid (criminal 1). But Cl^- from HCl is also a "criminal" (reducing agent). Now the cop is busy arresting both – it uses up more energy (more $KMnO_4$ consumed) than the actual crime. Result: the reading is inflated and inaccurate. That's why we use H_2SO_4 – it provides H^+ ions without any reducing Cl^-.

Answer

Option (2) – HCl reduces permanganate to Mn^{2+}

Q.7 Match the following:

Explanation

Titration Type	Basis of Reaction	Match
1. Neutralization	Formation of salt and water (acid+base)	R
2. Precipitation	Formation of insoluble precipitate	S
3. Redox	Oxidation–reduction (electron transfer)	Q
4. Complexometric	Formation of coordination compound	P

So: 1-R, 2-S, 3-Q, 4-P

Green Approach

Memory trick: **N**eutralization = **N**o more acid/base (salt + water); **P**recipitation = something falls down (precipitate); **R**edox = electrons **R**oam between species; **C**omplexometric = complex **C**oordination compound forms.

Answer

Option (1) – 1-R, 2-S, 3-Q, 4-P

Q.8 Assertion A & Reason R – Permanganate titrations not performed in HCl.

Explanation

Assertion (A): True – permanganate titrations are indeed not performed in HCl (as explained in Q6).

Reason (R): True – $KMnO_4$ oxidizes HCl (specifically Cl^- ions) to produce Cl_2 gas. This is oxidation of HCl, generating chlorine.

Is R the correct explanation of A? Yes – the very reason permanganate cannot be used with HCl is because Cl^- gets oxidised to Cl_2, consuming extra $KMnO_4$ and making results unreliable. R directly explains why A is true.

Green Approach

Approach for A & R questions: First check if A is true, then check if R is true, then ask – does R *logically explain why* A happens? Here: A true ✓, R true ✓, R explains A ✓ ⇒ Option (1).

Answer

Option (1) – Both A and R are true; R is the correct explanation of A.

TYPE 2 : Equivalent Weight & Normality Calculations

Q.9 The equivalent mass of potassium permanganate in alkaline medium is

Explanation

The equivalent weight of any oxidant/reductant:

$$E = \frac{\text{Molar Mass}}{n\text{-factor}}$$

where n-factor = number of electrons gained/lost per formula unit.
In **alkaline medium**, $KMnO_4$ (Mn^{7+}) is reduced to MnO_2 (Mn^{4+}):

$$MnO_4^- + {}_2H_2O + {}_3e^- \longrightarrow MnO_2 + {}_4OH^-$$

n-factor = **3** (3 electrons gained per MnO_4^-)

$$E = \frac{\text{Molar Mass}}{3}$$

Green Approach

Quick table – $KMnO_4$ n-factors to memorise:

Medium	Product	n-factor
Acidic	Mn^{2+}	5
Neutral/weakly alkaline	MnO_2	3
Strongly alkaline	MnO_4^{2-}	1

Analogy: Think of Mn as a staircase – in acid it falls 5 steps, in neutral 3 steps, in strong alkali just 1 step.

Answer

Option (2) – Molar Mass / 3

Q.10 The equivalent weight of $KMnO_4$ in alkaline medium as an oxidant will be

Explanation

Molar mass of $KMnO_4$:

$$M = 39 + 55 + 4(16) = 39 + 55 + 64 = 158 \text{ g/mol}$$

In alkaline medium, n-factor = 3 (as established in Q9):

$$E = \frac{158}{3} = 52.67 \approx 52.66 \text{ g/eq}$$

Green Approach

Memory shortcut: $158/3 \approx 52.7$. The options have 52.66 – match! Always compute molar mass step by step: K(39) + Mn(55) + O_4(64) = 158.

Answer

Option (2) – 52.66 g/eq

Q.11 The overall change in the oxidation number of manganese in the reaction (acidic medium, $KMnO_4$ + oxalic acid)

Explanation

In acidic medium, $KMnO_4$ is reduced to Mn^{2+}:

$$Mn^{7+} \rightarrow Mn^{2+}$$

Change in oxidation number of Mn:

$$\Delta\text{O.N.} = 7 - 2 = \mathbf{5}$$

This is the n-factor = 5 (5 electrons gained per Mn atom in acidic medium).

Green Approach

Remember: In acidic medium $KMnO_4$ falls from +7 to +2 – a drop of 5. The dark purple colour disappearing and solution becoming almost colourless confirms Mn^{2+} (faint pink/colourless) is formed.

Answer

Option (1) – 5

Q.12 The weight of $H_2C_2O_4 \cdot 2H_2O$ (MW = 126) required to make 500 mL of 0.2 N solution

Explanation

$H_2C_2O_4$ (oxalic acid) in redox reactions:

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$$

n-factor of oxalic acid = **2** (2 electrons lost per molecule)
Equivalent weight:

$$E = \frac{MW}{n\text{-factor}} = \frac{126}{2} = 63 \text{ g/eq}$$

Using: Weight = Normality × Volume (L) × Equivalent weight

$$W = 0.2 \times 0.5 \times 63 = 6.3 \text{ g}$$

Green Approach

Formula to memorise:

$$W = N \times V_{(L)} \times E$$

Think of it as: **N**ormality tells concentration, **V**olume tells how much, **E** tells per-gram-equivalent weight. Multiply all three to get grams needed.

Answer

Option (1) – 6.3 g

Q.13 The weight of $KMnO_4$ required to make 500 mL of 0.1 N $KMnO_4$ for titration in acidic medium

Explanation

In acidic medium, n-factor of $KMnO_4$ = 5.
Molar mass of $KMnO_4$ = 158 g/mol
Equivalent weight:

$$E = \frac{158}{5} = 31.6 \text{ g/eq}$$

Weight required:

$$W = N \times V_{(L)} \times E = 0.1 \times 0.5 \times 31.6 = 1.58 \text{ g}$$

Green Approach

Quick check – medium matters! Acidic: $n = 5$, $E = 31.6$. Alkaline: $n = 3$, $E = 52.7$. Always identify the medium first before calculating equivalent weight.

Answer

Option (2) – 1.58 g

TYPE 3 : Iodometry, Iodimetry & Reactions

Q.14 What is the colour of the starch iodine complex formed?

Explanation

When iodine (I_2) reacts with starch (amylose), it forms a characteristic **deep blue/blue-black coloured complex**. The I_3^- ions (formed when I_2 dissolves in KI solution) fit inside the helical structure of amylose, producing this intense blue colour. This colour disappears when I_2 is consumed (reduced), making starch a very sensitive indicator for iodine.

Green Approach

Just remember: Starch + Iodine = **Blue**. This is one of the most classic colour tests in chemistry. Think of a blue ink pen – starch iodine colour. It's used in iodometric titrations to detect the endpoint (when blue disappears, I_2 is gone).

Answer

Option (2) – Blue

Q.15 In the reaction $I_2 + {}_2S_2O_3^{2-} \longrightarrow {}_2I^- + S_4O_6^{2-}$; the equivalent weight of iodine will be equal to

Explanation

From the reaction, I_2 is reduced to $2I^-$:

$$I_2 + 2e^- \to 2I^-$$

Each I_2 molecule gains **2 electrons** ⇒ n-factor of I_2 = 2
Equivalent weight:

$$E(I_2) = \frac{M(I_2)}{2} = \frac{1}{2} \times \text{Molecular weight}$$

Green Approach

Analogy – Two brothers sharing: I_2 is like two brothers (two iodine atoms) who each grab one electron. Together they take 2 electrons. So per molecule, $n = 2$, and equivalent weight = mol. wt. / 2.

Answer

Option (2) – 1/2 the molecular weight

Q.16 In standardization of $Na_2S_2O_3$ using $K_2Cr_2O_7$ by iodometry, the equivalent weight of $K_2Cr_2O_7$ is

Explanation

In iodometric estimation, $K_2Cr_2O_7$ (dichromate) is reduced from Cr^{6+} to Cr^{3+}:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \to 2Cr^{3+} + 7H_2O$$

Each $Cr_2O_7^{2-}$ (one formula unit of $K_2Cr_2O_7$) gains **6 electrons** ⇒ n-factor = 6

$$E = \frac{\text{Molar Mass}}{6}$$

Green Approach

Key fact: $K_2Cr_2O_7$ always has n-factor = 6 in acidic medium (two Cr atoms, each drops from +6 to +3, so $2 \times 3 = 6$ electrons per formula unit). This is a very standard value for JEE – memorise it!

Answer

Option (2) – Molecular weight / 6

Q.17 In the iodometric estimation in laboratory, which process is involved?

Explanation

Iodometric estimation involves two steps:

1. An oxidizing agent (here $K_2Cr_2O_7$) oxidizes I^- to I_2 in acidic medium:

$$Cr_2O_7^{2-} + H^+ + I^- \longrightarrow {}_2Cr^{3+} + I_2$$

2. The liberated I_2 is then titrated against $Na_2S_2O_3$:

$$I_2 + S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + I^-$$

The reaction must be in **acidic medium with H^+**; alkaline medium would give different/incorrect products.
Option (1) matches: $Cr_2O_7^{2-} + H^+ + I^- \longrightarrow Cr^{3+} + I_2$ followed by $I_2 + S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + I^-$

Green Approach

Iodometry vs Iodimetry: Iodometry = I_2 is *liberated* (indirect – you measure indirectly via $Na_2S_2O_3$). **Iodimetry** = I_2 is used directly as the titrant. Here, $K_2Cr_2O_7$ liberates I_2 first ⇒ **iodometry**. Acidic H^+ is essential.

Answer

Option (1): $Cr_2O_7^{2-} + H^+ + I^- \longrightarrow {}_2Cr^{3+} + I_2$, then $I_2 + S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + I^-$

Q.18 Cu^{2+} salt reacts with potassium iodide to give

Explanation

When Cu^{2+} reacts with KI:

$$2Cu^{2+} + 4I^- \rightarrow Cu_2I_2 + I_2$$

Cu^{2+} is reduced to Cu^+, and I^- is oxidised to I_2. The Cu^+ precipitates as Cu_2I_2 (cuprous iodide) – a white precipitate. The I_2 liberated is then titrated with $Na_2S_2O_3$.

Green Approach

Remember the product: $Cu^{2+} \longrightarrow Cu^+$ (reduced). Cu^+ with I^- forms CuI, but as an ionic compound it is Cu_2I_2 (since Cu^+ is monovalent, two Cu^+ + two I^- = Cu_2I_2). White precipitate = Cu_2I_2. This reaction is the basis of iodometric determination of copper.

Answer

Option (1) – Cu_2I_2

Q.19 Indicator X forms blue coloured complex with compound A present in the solution. X and A are respectively

Explanation

From Q14, we know that **starch** forms a characteristic **blue colour** with **iodine (I_2)**. In this experiment studying iodide oxidation by H_2O_2, iodide ions (I^-) are oxidised to I_2. The starch indicator detects I_2 by turning blue.
So: **X = Starch** (indicator), **A = Iodine (I_2)** (compound that forms blue complex with starch).

Green Approach

Key logic: The question says "X forms blue complex with A." Only one pair gives blue: Starch (X) + Iodine (A). Eliminate: methyl orange is an acid-base indicator (red/yellow), not associated with blue. H_2O_2 doesn't form blue complexes.

Answer

Option (1) – Starch and Iodine

Q.20 When Cu^{2+} ion is treated with KI, white precipitate X appears. Solution titrated with sodium thiosulphate gives compound Y. X and Y are

Explanation

Step 1 – Formation of X:

$$2Cu^{2+} + 4I^- \rightarrow \underbrace{Cu_2I_2}_{\text{white ppt. = X}} + I_2$$

White precipitate X = Cu_2I_2

Step 2 – Titration of liberated I_2 with $Na_2S_2O_3$:

$$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$$

Product Y = $Na_2S_4O_6$ (sodium tetrathionate)

Green Approach

Classic iodometric copper determination: This is a standard experiment. Remember: X = Cu_2I_2 (white), Y = $Na_2S_4O_6$ (sodium *tetra*thionate – subscript 4 in S_4O_6). Common mistake: writing $Na_2S_4O_5$ or $Na_2S_4O_3$ – the correct product is $\mathbf{Na_2S_4O_6}$.

Answer

Option (2) – X = Cu_2I_2; Y = $Na_2S_4O_6$

Answer Key – DPP-2

Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	(1)	6	(2)	11	(1)	16	(2)
2	(2)	7	(1)	12	(1)	17	(1)
3	(2)	8	(1)	13	(2)	18	(1)
4	(1)	9	(2)	14	(2)	19	(1)
5	(4)	10	(2)	15	(2)	20	(2)