Diploma Focused Unit Package

This package gets you ready for the Unit B questions on the diploma exam. Unit B is about 30% of the exam. You can work through it on your own.

Part 1 is redox reactions (Cards 1 to 6). **Part 2** is cells (Cards 7 to 12).

How to use this package

1. **Do the start check.** It shows which cards you need.
2. **Fill in the checklist.** Tick “can do” or “not yet” for each card.
3. **Start with a “not yet” card.** Read the steps and the worked example.
4. **Do the questions on paper.** Use your Data Booklet, as you will on the exam.
5. **Check your answers.** Every answer is worked. Find where your work first differs.
6. **Do the card check and show Mr F.** Then tick the card off.
7. **Finish with the practice unit check.** It tells you if you are ready for the unit final.
8. **Then do the exam-style question sets.** They ask questions the way the exam does.

Stuck on a word? Look in Words to know. Stuck on an idea? Look in Big ideas.

How diploma questions are asked

The diploma asks questions in a few set ways. Know them before exam day. You can practise all of them in the exam-style question sets at the end.

“Use the following information to answer questions …”
A box of information comes first: a real situation, an equation, a table or a diagram. Read the box before the questions. If the box gives a half-reaction with its E° value, use it together with the Data Booklet table.

A sentence with blanks *i* and *ii*
The question is a sentence with two blanks, then “The statement above is completed by the information in row”, then a table of four rows. Decide blank *i* first. That removes two rows. Then decide blank *ii*.

Numbered statements
Statements are numbered I, II, III, IV. The choices are groups, such as “I and III”. Mark each statement true or false first. Then find the choice that matches.

Numerical response: four digits
You record four digits. They may be four numbered particles in order, four coefficients, four oxidation numbers, or the numbers of four true statements. Record them in the order asked. “In any order” is only allowed when the question says so.

Numerical response: +/− and a value

The question shows “+/−” and a blank. Record only the size of the value, such as 0.74. You do not record the sign.

Numerical response: scientific notation
The answer is written as a.bc × 10^{-d}. You record the four digits a, b, c and d. For 3.07 × 10^{-4} you record 3074.

The word “nonspontaneous”
The diploma writes “nonspontaneous” for a reaction that is not spontaneous. It also says “chemical species” for particles.

Start check

Twelve quick questions, one for each card. Use your Data Booklet. If you get a question wrong, or you guessed, do that card. If you get it right and it felt easy, you may skip the card’s lesson part and go straight to its questions.

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1. What is the oxidation number of nitrogen in $HNO_3(aq)$?

A. +3

B. +5

C. −3

D. +7

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2. Nitrate ions change to nitrogen dioxide gas in an acidic solution: $NO_3^-(aq) \rightarrow NO_2(g)$. How many electrons are in the balanced half-reaction?

A. 1

B. 2

C. 3

D. 4

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3. Balance $Al(s) + Ni^{2+}(aq) \rightarrow Al^{3+}(aq) + Ni(s)$. What is the coefficient of Al(s)?

A. 1

B. 2

C. 3

D. 6

4. $A^{2+}(aq)$ reacts with metal B. $A^{2+}(aq)$ does not react with metal C. Which ion is the strongest oxidizing agent?

A. $A^{2+}(aq)$

B. $B^{2+}(aq)$

C. $C^{2+}(aq)$

D. There is not enough information.

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5. Which metal reacts spontaneously with $Cu^{2+}(aq)$?

A. Ag(s)

B. Au(s)

C. Hg(l)

D. Pb(s)

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6. In a titration, $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$. How many moles of Fe^{2+} react with 2.0×10^{-4} mol of MnO_4^-?

A. 4.0×10^{-5} mol

B. 2.0×10^{-4} mol

C. 1.0×10^{-3} mol

D. 1.6×10^{-3} mol

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7. In the cell $Mg(s) \mid Mg^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$, which way do electrons move in the wire?

A. From Mg(s) to Cu(s)

B. From Cu(s) to Mg(s)

C. From $Mg^{2+}(aq)$ to $Cu^{2+}(aq)$

D. Through the salt bridge

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8. What is $E°_{cell}$ for $Pb(s) \mid Pb^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$?

A. +0.93 V

B. +0.67 V

C. +1.73 V

D. −0.93 V

9. A solution of KBr(aq) is electrolyzed with inert electrodes. What forms at the cathode?

A. K(s)

B. $Br_2(l)$

C. $H_2(g)$

D. $O_2(g)$

10. How many moles of electrons pass when 2.00 A runs for 965 s?

A. 0.0100 mol

B. 0.0200 mol

C. 0.0400 mol

D. 2.00 mol

11. Why does a zinc coating protect iron even when it is scratched?

A. Zinc is a stronger oxidizing agent than iron.

B. Zinc is a stronger reducing agent than iron.

C. Zinc does not react with oxygen.

D. Zinc is the cathode.

12. In an experiment, which variable must be something you measure directly?

A. The manipulated variable

B. The responding variable

C. The controlled variable

D. The calculated variable

▼ **Check your answers for the start check**

Question	Answer	Why	If wrong, do
1	**B**	(+1) + x + 3(−2) = 0, so x = +5.	Card 1
2	**A**	$NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$	Card 2
3	**B**	$2Al(s) + 3Ni^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Ni(s)$	Card 3
4	**C**	A^{2+} is above B and below C. From the top: C^{2+}, A^{2+}, B^{2+}.	Card 4
5	**D**	Cu^{2+} (+0.34 V) reacts with a metal below it on the right. Only Pb(s) (−0.13 V) is below it.	Card 5
6	**C**	$5 \times 2.0 \times 10^{-4}$ mol = 1.0×10^{-3} mol.	Card 6

7	**A**	The anode is on the far left. Electrons move from the anode, Mg(s), to the cathode, Cu(s).	Card 7
8	**A**	+0.80 V − (−0.13 V) = +0.93 V.	Card 8
9	**C**	Water (−0.83 V) is a stronger oxidizing agent than K^+ (−2.93 V). Water is reduced to $H_2(g)$ and $OH^-(aq)$.	Card 9
10	**B**	q = 2.00 A × 965 s = 1930 C. n(e^-) = 1930 C ÷ 9.65 × 10^4 C/mol = 0.0200 mol.	Card 10
11	**B**	Zinc is below iron on the right side of the table, so zinc is oxidized in place of iron.	Card 11
12	**B**	The responding variable is what you measure to see the effect.	Card 12

Checklist

Card	I can ...	Can do	Not yet
Part 1: redox reactions			
1	I can find oxidation numbers. I can use them to say if a reaction is redox, what is oxidized, what is reduced, and which reactant is the oxidizing agent and the reducing agent.	☐	☐
2	I can write and balance one half-reaction in an acidic or neutral solution, when it is not in the Data Booklet table.	☐	☐
3	I can balance a redox equation from the table, from a skeleton equation in acid, in a basic solution when the half-reactions are given, and for a disproportionation.	☐	☐
4	I can use the results of an experiment to rank oxidizing agents and reducing agents, and write them as a small table of reduction half-reactions.	☐	☐
5	I can list what is in the beaker, pick the strongest oxidizing agent and the strongest reducing agent, write the equation, and say if the reaction is spontaneous.	☐	☐
6	I can find a concentration or a mass from a redox titration, and I can choose the right equipment.	☐	☐
Part 2: cells			
7	I can find the anode and the cathode of a voltaic cell, the half-reaction at each, the way electrons and ions move, the cell notation, and what I would see.	☐	☐

8	I can calculate a standard cell potential. I can handle a new reference half-cell, or a half-reaction that is missing from the table.	☐	☐
9	I can tell an electrolytic cell from a voltaic cell. I can predict the half-reaction and product at each electrode, find the minimum voltage, and say what I would see.	☐	☐
10	I can use current and time to find a mass or an amount at an electrode, and I can work backwards to find a current or a time.	☐	☐
11	I can explain how corrosion is stopped, and how plating, refining, batteries and fuel cells work, using what I know about redox.	☐	☐
12	I can name the variables in an experiment, tell an observation from an inference, judge a design, and explain why a measured voltage differs from the predicted one.	☐	☐

Your Data Booklet for Unit B

You have the Chemistry Data Booklet for the whole exam. These are the parts Unit B uses.

Part of the booklet	You need it for	How to use it
Table of Selected Standard Electrode Potentials	Every Unit B prediction and voltage.	Oxidizing agents on the left, strongest at the top. Reducing agents on the right, strongest at the bottom. Values are reduction potentials: never change the sign, never multiply.
Faraday constant, $F = 9.65 \times 10^4$ C/mol e^-	Faraday's law (Card 10).	It is with the other constants near the front. Use it in $n(e^-) = q \div F$.
Periodic table	Molar masses for titration and Faraday's law questions.	Use the molar mass as printed. Round only your final answer.
Table of colours of common aqueous ions	Predicting what you would see in a cell (Cards 7 and 9).	Example: $Cu^{2+}(aq)$ is blue, $Cr_2O_7^{2-}(aq)$ is orange, $MnO_4^-(aq)$ is purple.
Solubility table	Deciding if an ionic compound is dissolved, so you can write it as ions (Card 5).	A dissolved ionic compound is written as its separate ions.

The Data Booklet does **not** have: the oxidation number rules, AN OX and RED CAT, $E°cell = E°cathode - E°anode$, $q = I \times t$, or $n(e^-) = q \div F$. You must know these. If a question needs a value that is not in the booklet, the question gives it.

Words to know

Redox words

oxidation
A particle loses electrons. Its oxidation number goes up.

reduction
A particle gains electrons. Its oxidation number goes down.

OIL RIG
A memory aid: Oxidation Is Loss of electrons, Reduction Is Gain of electrons.

redox reaction
A reaction where electrons move from one particle to another. Oxidation and reduction always happen together.

electron transfer
Electrons moving from one particle to another.

oxidizing agent
The reactant that takes electrons. It is reduced.

reducing agent
The reactant that gives electrons. It is oxidized.

strongest oxidizing agent
Of the oxidizing agents in the beaker, the one highest on the left side of the table.

strongest reducing agent
Of the reducing agents in the beaker, the one lowest on the right side of the table.

oxidation number
A number given to each atom to keep track of electrons. It is the charge the atom would have if every shared electron went to the atom that pulls harder.

peroxide
A compound with an O–O bond, such as H_2O_2 or BaO_2. Oxygen is −1 in a peroxide, not −2.

hydride
A compound of hydrogen with a metal from Group 1 or 2, such as NaH or CaH_2. Hydrogen is −1 in it, not +1.

half-reaction
An equation that shows only the oxidation, or only the reduction, with the electrons written in.

net ionic equation
An equation that shows only the particles that change.

spectator ion
An ion that is in the beaker but does not change. It is left out of the net ionic equation.

disproportionation
A redox reaction where the same kind of particle is both oxidized and reduced.

spontaneous
The reaction happens by itself when the reactants are mixed. Some questions say “favoured”.

not spontaneous
The reaction does not happen by itself.

acidified
Acid was added, so $H^+(aq)$ is in the beaker.

operational definition
A definition by what you can see or measure. Example: oxidation is a metal gaining mass as it reacts with oxygen.

theoretical definition
A definition by what the particles do. Example: oxidation is a loss of electrons.

Table words

table of standard electrode potentials
The Data Booklet table of reduction half-reactions. Oxidizing agents are on the left, strongest at the top. Reducing agents are on the right, strongest at the bottom.

standard reduction potential (E°)
The table value, in volts, beside each half-reaction.

table value
The E° number read from the table.

Cell words

electrode
A solid that carries electrons into or out of a solution. It is often a metal strip. Reactions happen on its surface.

anode
The electrode where oxidation happens. Memory aid: AN OX.

cathode
The electrode where reduction happens. Memory aid: RED CAT.

anion
A negative ion. Anions move toward the anode.

cation
A positive ion. Cations move toward the cathode.

electrolyte
A solution, or a melted compound, that has ions free to move, so it carries charge.

salt bridge
A tube of ion solution that joins two half-cells. Ions move through it, so each half-cell stays neutral.

porous cup or porous barrier
A wall with tiny holes. It does the same job as a salt bridge.

external circuit
The wire, and anything joined into it, outside the solutions. Electrons move through it.

inert electrode

An electrode that carries electrons but does not react. Carbon, C(s), and platinum, Pt(s), are used.

half-cell
One electrode in its solution. A voltaic cell has two.

voltaic cell
A cell where a spontaneous reaction makes electrical energy. It is also called a galvanic cell. A battery is one or more voltaic cells.

electrolytic cell
A cell where electrical energy from a power source forces a reaction that is not spontaneous.

power source
A battery or other supply of electrical energy that pushes electrons through an electrolytic cell. Some books call it a power supply.

voltmeter
A meter that measures the voltage of a cell, in volts.

electrolysis
Using electrical energy to make a reaction happen in an electrolytic cell.

molten
Melted. A molten salt has ions that can move, and no water.

cell notation
A one-line way to write a cell, with the anode on the far left and the cathode on the far right. A single line | separates a solid from a solution. A double line || is the salt bridge.

standard hydrogen half-cell
The reference half-cell. Its table value is set at 0.00 V. Every other table value is measured against it.

standard conditions
Solutions at 1.0 mol/L and 25 °C. Table values are for these conditions.

standard cell potential (E°cell)
The voltage of a cell at standard conditions. E°cell = E°cathode − E°anode.

minimum voltage
The smallest voltage a power source needs to make an electrolytic cell run. It is the size of the negative E°cell.

overvoltage
The extra voltage, above the minimum, that is needed before a reaction really happens at an electrode.

chloride anomaly
In the electrolysis of a chloride solution, chlorine gas forms at the anode, even though the table says water should be oxidized first.

Titration words

titration
Adding one solution to another a little at a time, until the reaction is just complete.

titrant
The solution in the burette. Its concentration is known.

sample
The solution in the flask. Its concentration is what you want to find.

endpoint
The point where the colour changes and you stop adding titrant.

equivalence point
The point where the amounts match the mole ratio in the balanced equation.

volumetric pipette
Glassware that delivers one exact volume, such as 10.00 mL. It measures the sample.

burette
A long tube with a tap and a scale. It measures the volume of titrant added. It is also spelled “buret”.

Calculation words

current (I)
How much charge passes each second. It is measured in amperes, A. One ampere is one coulomb each second.

charge (q)
The amount of electricity, in coulombs, C. $q = I \times t$, with t in seconds.

mole of electrons
An amount of electrons, counted in moles, as you count atoms.

Faraday constant (F)
The charge of one mole of electrons: 9.65×10^4 C/mol e^-. It is in the Data Booklet.

Faraday’s law
The amount of a substance made or used at an electrode depends on the charge that passes. $n(e^-) = q \div F$.

Uses

corrosion
A metal being oxidized by what is around it. Rusting of iron is corrosion. It needs oxygen and water.

cathodic protection
Stopping corrosion by making the metal the cathode, so it cannot be oxidized.

sacrificial anode
A block of a stronger reducing agent, such as magnesium or zinc, joined to the metal you want to protect. The block is oxidized instead.

galvanizing
Coating iron or steel with zinc.

electroplating
Using an electrolytic cell to coat an object with a thin layer of metal. The object is the cathode.

refining
Making a metal pure. In electrolytic refining the impure metal is the anode and pure metal forms on the cathode.

electrowinning

Getting a metal out of a solution of its ions by electrolysis.

battery
One or more voltaic cells joined together.

fuel cell
A voltaic cell that is fed fuel all the time, such as hydrogen and oxygen. It does not run down while fuel is supplied.

Lab words

manipulated variable
The one thing the experimenter changes on purpose. Some books call it the independent variable.

responding variable
What is measured to see the effect. It must be something you measure directly. Some books call it the dependent variable.

controlled variable
Something that could change, but that the experimenter keeps the same.

observation
What you see or measure: bubbles, a colour, a mass, a meter reading.

inference
What you work out from the observations: which gas it is, which ion formed.

Big ideas

Idea 1: Redox is electron transfer

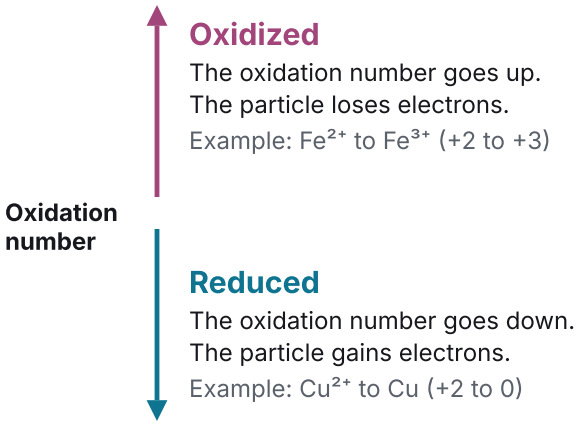
In a redox reaction, electrons move from one particle to another.

The particle that **loses** electrons is **oxidized**. The particle that **gains** electrons is **reduced**. Remember OIL RIG.

The reactant that takes the electrons is the **oxidizing agent**. It is reduced. The reactant that gives the electrons is the **reducing agent**. It is oxidized.

Example: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$. Zinc gives two electrons, so Zn(s) is oxidized and is the reducing agent. $Cu^{2+}(aq)$ takes them, so it is reduced and is the oxidizing agent.

There are two kinds of definition. An **operational** definition says what you see: a metal reacts with oxygen and gains mass. A **theoretical** definition says what the particles do: the metal loses electrons.



Idea 2: Oxidation numbers

An oxidation number keeps track of electrons, one atom at a time. Use these rules in order.

1. An element by itself is 0. Examples: Zn(s), O_2(g), Cl_2(g).

2. A simple ion is its charge. Example: Fe^{3+} is +3.

3. Hydrogen is +1. **Exception:** it is −1 in a hydride of a Group 1 or 2 metal, such as NaH or CaH_2.

4. Oxygen is −2. **Exception:** it is −1 in a peroxide, such as H_2O_2 or BaO_2.

5. The numbers add up to 0 for a compound. They add up to the charge for an ion.

Carbon can be a fraction. For a compound with several carbon atoms, find one average number. In propane, C_3H_8, the eight hydrogens add to +8, so the three carbons add to −8. Each carbon is −8/3.

Idea 3: The Data Booklet table, read by position

Every line in the table is a reduction half-reaction.

Left side: oxidizing agents. The strongest is at the **top**.

Right side: reducing agents. The strongest is at the **bottom**.

An oxidizing agent reacts spontaneously with a reducing agent that is **below** it on the right. If the reducing agent is above it, the reaction is not spontaneous.

Oxidizing agents strongest at the top		Reducing agents strongest at the bottom	E° (V)
$Ag^+(aq) + e^-$	⇌	$Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^-$	⇌	$Cu(s)$	+0.34
$2H^+(aq) + 2e^-$	⇌	$H_2(g)$	0.00
$Ni^{2+}(aq) + 2e^-$	⇌	$Ni(s)$	−0.26
$Zn^{2+}(aq) + 2e^-$	⇌	$Zn(s)$	−0.76

$Ag^+(aq)$ is above $Zn(s)$, so they react spontaneously.
$Zn^{2+}(aq)$ is below $Ag(s)$, so they do not react.

Idea 4: What is in the beaker

Before you predict a reaction, list every particle in the beaker.

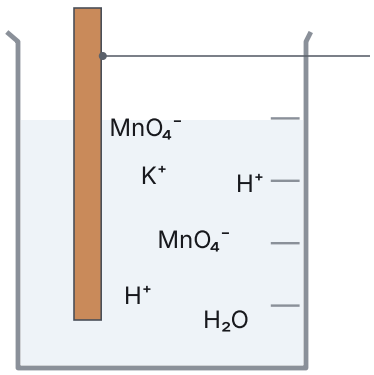
A dissolved ionic compound is its ions. Potassium permanganate solution is $K^+(aq)$ and $MnO_4^-(aq)$.

Water is always there in a solution: $H_2O(l)$.

"Acidified" means $H^+(aq)$ is there too.

"Basic" means $OH^-(aq)$ is there too.

A solid metal is written as the metal: $Cu(s)$.



Cu(s): a solid metal

Oxidizing agents
$MnO_4^-(aq)$ with $H^+(aq)$
$H^+(aq)$, $K^+(aq)$, $H_2O(l)$

Reducing agents
$Cu(s)$, $H_2O(l)$

Acidified potassium permanganate

Idea 5: The voltaic cell

A voltaic cell makes electrical energy from a spontaneous redox reaction. The two half-reactions happen in two separate half-cells, so the electrons must go through the wire.

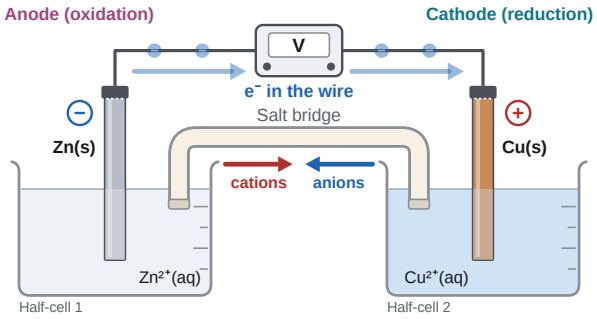
AN OX: oxidation happens at the **an**ode. **RED CAT: red**uction happens at the **cat**hode. This is true in every cell.

Electrons move through the wire from the anode to the cathode.

Ions move through the salt bridge. Anions move toward the anode. Cations move toward the cathode.

In a voltaic cell the anode is the negative electrode and the cathode is the positive electrode.

This picture is from lesson B6. It shows a zinc–copper cell.

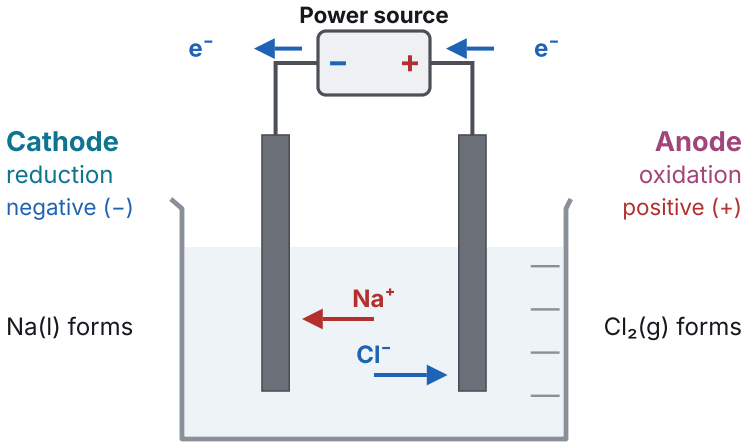


Idea 6: The electrolytic cell

An electrolytic cell uses electrical energy to force a reaction that is not spontaneous. Look for a **power source**.

AN OX and RED CAT still hold. The power source pushes electrons onto the cathode, so the cathode is **negative** here. It pulls electrons from the anode, so the anode is **positive**.

	Voltaic cell	Electrolytic cell
In the wire	a voltmeter or a device that uses energy	a power source
Reaction	spontaneous	not spontaneous
E°cell	positive	negative
Energy	chemical to electrical	electrical to chemical
Anode	oxidation; negative	oxidation; positive
Cathode	reduction; positive	reduction; negative
Usual set-up	two half-cells and a salt bridge	one container, no salt bridge



Molten sodium chloride: only ions, no water

Idea 7: E° and the hydrogen reference

A table value, E°, tells how strongly a particle pulls electrons, compared with the **standard hydrogen half-cell**. That half-cell is given the value 0.00 V.

Every table value is a reduction potential. Do not change its sign when you write the half-reaction backwards. Do not multiply it when you multiply the half-reaction.

E°cell = E°cathode − E°anode. A positive answer means spontaneous. A negative answer means not spontaneous.

If a different half-cell is chosen as the reference, every value shifts by the same amount. The **difference** between two values does not change, so E°cell does not change.

Table values are for 1.0 mol/L and 25 °C. A higher concentration of a reactant gives a larger voltage. As a cell runs, its reactants are used up and its voltage falls. When the voltage reaches zero the cell is “dead”.

Some questions are about a cell that is **not** at standard conditions, such as one inside a living thing. The question then gives its own E° values. Use the values in the question, not the Data Booklet ones.

Half-reaction	**Hydrogen** is 0.00 V	**Copper** is 0.00 V
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80	+0.46
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34	**0.00**
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	**0.00**	−0.34
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76	−1.10

The gap from silver to zinc is 1.56 V in both columns.
A new reference shifts every value. It does not change E°cell.

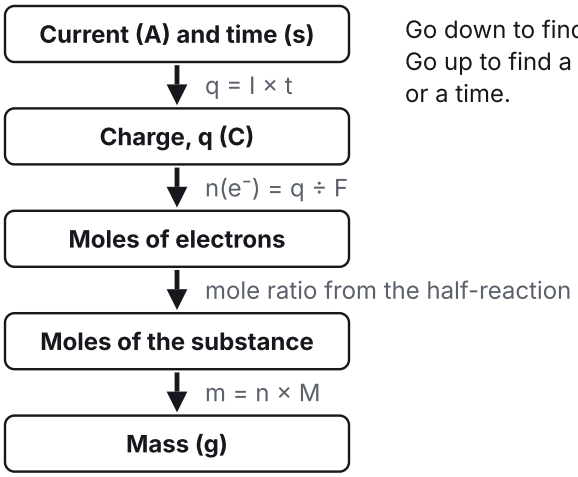
Idea 8: Charge, moles of electrons and Faraday’s law

A current is a flow of charge. **q = I × t**: charge in coulombs is current in amperes times time in **seconds**.

One mole of electrons carries 9.65×10^4 C. This is the Faraday constant, F, in the Data Booklet. So **$n(e^-) = q \div F$**.

The half-reaction is the bridge to the substance. In $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$, 2 mol of electrons make 1 mol of copper.

Then m = n × M, with M from the periodic table in the Data Booklet.



Idea 9: Corrosion and how it is stopped; cells people use

Corrosion is a spontaneous redox reaction. Iron is oxidized: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$. Oxygen, with water, is reduced. So rusting needs both oxygen and water.

A coating (paint, plastic, grease) keeps oxygen and water away.

Galvanizing coats iron with zinc. Zinc is a stronger reducing agent than iron, so the zinc is oxidized first, even if the coating is scratched.

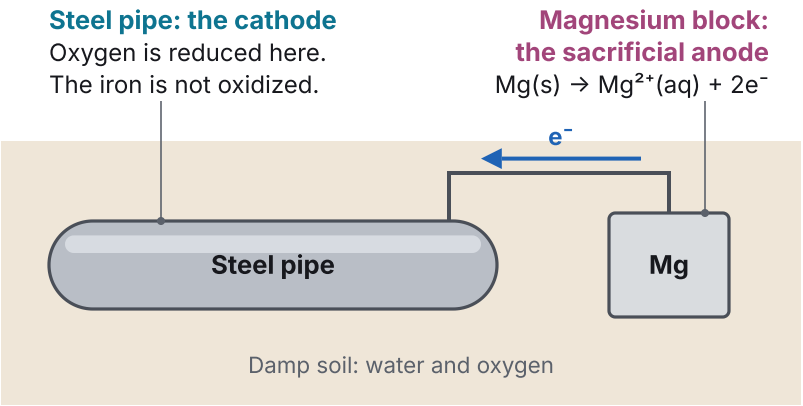
A sacrificial anode is a block of magnesium or zinc joined to the iron. The block is oxidized and the iron becomes the cathode. This is **cathodic protection**. A power source can do the same job by pushing electrons onto the iron.

Electroplating and **refining** use electrolytic cells. The object to be coated, or the pure metal, is the cathode.

Batteries and **fuel cells** are voltaic cells. Recharging a battery runs it as an electrolytic cell. Used batteries hold metals that should be recycled, not sent to a landfill.

Redox in industry. Chlorine compounds are oxidizing agents: they kill bacteria in drinking water and pools, and they bleach wood pulp and cloth. Food goes bad partly by oxidation, so some foods are packed without air. Electrolytic cells also recover metals for recycling and can clean gases from a smokestack.

Long ago, people found by trial and error that heating some rocks (ores) with charcoal gives a metal. The carbon is the reducing agent: it takes the oxygen from the metal oxide.



Card 1: Oxidation numbers, and spotting a redox reaction

I can find oxidation numbers. I can use them to say if a reaction is redox, what is oxidized, what is reduced, and which reactant is the oxidizing agent and the reducing agent.

How to spot it on the diploma

- "The oxidation number of chromium in $K_2Cr_2O_7$ is ..."
- "Which of the following equations represents a redox reaction?"
- "In this reaction, the oxidizing agent is ..."
- A new compound you have never seen. The rules still work.

The steps

1. **Write the numbers you know.** An element alone is 0. A simple ion is its charge. Hydrogen is +1 (−1 in a metal hydride). Oxygen is −2 (−1 in a peroxide).
2. **Use the sum rule for the atom you do not know.** The numbers add to 0 for a compound, or to the charge for an ion.
3. **Compare each element before and after the arrow.** Write the oxidation number above each atom on both sides.
4. **Decide.** The number went up: that atom is oxidized. The number went down: that atom is reduced. No number changed: the reaction is not redox.
5. **Name the agents.** The **reactant** that holds the oxidized atom is the reducing agent. The **reactant** that holds the reduced atom is the oxidizing agent.

Worked example

Acidified dichromate ions react with iron(II) ions: $Cr_2O_7^{2-}(aq) + 6Fe^{2+}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$. Identify the oxidizing agent and the reducing agent.

> **Step 1**
> Oxygen is −2. Hydrogen in H^+ is +1. Fe^{2+} is +2, Fe^{3+} is +3, Cr^{3+} is +3.
>
> **Step 2**
> Chromium in $Cr_2O_7^{2-}$: $2x + 7(-2) = -2$, so $2x = +12$ and $x = +6$.
>
> **Step 3**
> Chromium: +6 before, +3 after. Iron: +2 before, +3 after. Oxygen and hydrogen do not change.
>
> **Step 4**
> Chromium went down, so it is reduced. Iron went up, so it is oxidized.
>
> **Step 5**
> The oxidizing agent is $Cr_2O_7^{2-}(aq)$. The reducing agent is $Fe^{2+}(aq)$.

The common mistake

> Naming an atom as the agent. The agent is the whole reactant particle: $Cr_2O_7^{2-}(aq)$, not “Cr”. Also watch for peroxides: oxygen is −1 in H_2O_2.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

What is the oxidation number of sulfur in $H_2SO_4(aq)$?

A. +4

B. +6

C. −2

D. +8

QUESTION 2 · EASY · NUMERICAL RESPONSE

The oxidation number of chlorine in ClO_4^- is $+i$. The oxidation number of nitrogen in NO_2 is $+j$. The oxidation number of carbon in CO is $+k$. Record the three digits i, j, k in that order.

Answer: ☐

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

Which of the following equations represents a redox reaction?

A. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

B. $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

C. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

D. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

QUESTION 4 · MEDIUM · MULTIPLE CHOICE

In barium peroxide, $BaO_2(s)$, the oxidation number of oxygen is *i*. In calcium hydride, $CaH_2(s)$, the oxidation number of hydrogen is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	−2	+1
B.	−2	−1
C.	−1	+1
D.	−1	−1

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

Chlorine gas reacts with a basic solution: $3Cl_2(g) + 6OH^-(aq) \rightarrow 5Cl^-(aq) + ClO_3^-(aq) + 3H_2O(l)$. Which statement describes this reaction?

A. Chlorine is oxidized only, and $OH^-(aq)$ is the oxidizing agent.

B. Chlorine is reduced only, and $OH^-(aq)$ is the reducing agent.

C. Chlorine is both oxidized and reduced, so this is a disproportionation.

D. No oxidation number changes, so this is not a redox reaction.

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

Cells get energy from glucose in cellular respiration: $C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$.
In cellular respiration, the oxidizing agent is *i*, and the oxidation number of carbon changes from *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	$O_2(g)$	0 to +4
B.	$O_2(g)$	+4 to 0
C.	$C_6H_{12}O_6(aq)$	0 to +4
D.	$C_6H_{12}O_6(aq)$	+4 to 0

CHALLENGE

Propane burns in a barbecue: $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$. How many moles of electrons are transferred when 1 mol of propane burns?

▼ **Hint**

Find the average oxidation number of carbon in propane. Then find how much each carbon atom changes. There are three carbon atoms.

▼ **Check your answers for Card 1**

Question 1: B. Hydrogen is +1 and oxygen is −2. 2(+1) + x + 4(−2) = 0, so x = +6.

Question 2: 742. ClO_4^-: x + 4(−2) = −1, so x = +7. NO_2: x + 2(−2) = 0, so x = +4. CO: x + (−2) = 0, so x = +2.

Question 3: C. In C, zinc goes from 0 to +2 and hydrogen goes from +1 to 0. In A, B and D no oxidation number changes.

Question 4: D. BaO_2 is a peroxide, so oxygen is −1: (+2) + 2(−1) = 0. CaH_2 is a hydride of a Group 2 metal, so hydrogen is −1: (+2) + 2(−1) = 0.

Question 5: C. Chlorine is 0 in Cl_2. It is −1 in Cl^- (reduced) and +5 in ClO_3^- (oxidized). The same kind of particle is both oxidized and reduced. Oxygen stays −2 and hydrogen stays +1.

Question 6: A. In glucose: 6x + 12(+1) + 6(−2) = 0, so x = 0. In CO_2 carbon is +4. Carbon is oxidized, so glucose is the reducing agent. Oxygen goes from 0 to −2. It is reduced, so $O_2(g)$ is the oxidizing agent.

Challenge. Carbon in C_3H_8: 3x + 8(+1) = 0, so x = −8/3. Carbon in CO_2 is +4. Each carbon atom goes up by 4 − (−8/3) = 20/3. Three carbon atoms: 3 × 20/3 = 20. So 20 mol of electrons are transferred. Check with oxygen: 10 oxygen atoms each go from 0 to −2, which is also 20.

Card 1 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. What is the oxidation number of phosphorus in PO_4^{3-}?
2. In $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$, which reactant is the reducing agent?
3. Is $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$ a disproportionation? Give the oxidation numbers of oxygen.

Need more? Go back to the lesson: Lesson B1 · Lesson B2 · Lesson NI. Videos for each idea: Videos by concept.

Card 2: Write a half-reaction that is not in the table

I can write and balance one half-reaction in an acidic or neutral solution, when it is not in the Data Booklet table.

How to spot it on the diploma

- You are given one reactant and one product, such as $NO_3^-(aq) \rightarrow NO(g)$, and the words “in an acidic solution”.
- “The balanced half-reaction is …” or “The number of electrons transferred is …”
- You looked in the table and the half-reaction is not there.

The steps

1. **Write the reactant and the product.** Leave space on both sides.
2. **Balance the main atom.** This is every atom that is not oxygen or hydrogen.
3. **Balance oxygen with $H_2O(l)$.** Add water to the side that needs oxygen.
4. **Balance hydrogen with $H^+(aq)$.** Add H^+ to the side that needs hydrogen.
5. **Balance charge with electrons.** Add e^- to the side that is more positive, until both sides have the same total charge.
6. **Check.** Count each atom and the total charge on both sides. Electrons on the left: it is a reduction. Electrons on the right: it is an oxidation.

Worked example

Nitrate ions change to nitrogen monoxide gas in an acidic solution: $NO_3^-(aq) \rightarrow NO(g)$. Write the balanced half-reaction.

> **Step 2**
> Nitrogen: 1 on each side. It is balanced.
>
> **Step 3**
> Oxygen: 3 on the left, 1 on the right. Add $2H_2O(l)$ to the right.
>
> **Step 4**
> Hydrogen: 4 on the right now. Add $4H^+(aq)$ to the left.
>
> **Step 5**
> Charge: left is $-1 + 4 = +3$. Right is 0. Add $3e^-$ to the left.
>
> **Step 6**
> $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$. Charge is 0 on both sides. Electrons are on the left, so this is a reduction.

The common mistake

Putting the electrons on the wrong side. Work out the total charge of each side first. Electrons always go on the more positive side.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

Hydrazine, $N_2H_4(aq)$, is a rocket fuel. It changes to nitrogen gas in an acidic solution: $N_2H_4(aq) \rightarrow N_2(g)$. Which half-reaction is balanced?

A. $N_2H_4(aq) \rightarrow N_2(g) + 4H^+(aq) + 4e^-$

B. $N_2H_4(aq) + 4e^- \rightarrow N_2(g) + 4H^+(aq)$

C. $N_2H_4(aq) \rightarrow N_2(g) + 2H^+(aq) + 2e^-$

D. $N_2H_4(aq) + 4H^+(aq) \rightarrow N_2(g) + 4e^-$

QUESTION 2 · MEDIUM · NUMERICAL RESPONSE

Sulfur dioxide changes to sulfate ions in an acidic solution: $a\ SO_2(g) + b\ H_2O(l) \rightarrow c\ SO_4^{2-}(aq) + d\ H^+(aq) + e\ e^-$. With $a = 1$, record the three digits b, d, e in that order.

Answer: ☐

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

Chlorate ions change to chloride ions in an acidic solution: $ClO_3^-(aq) \rightarrow Cl^-(aq)$. How many electrons are in the balanced half-reaction, and is it an oxidation or a reduction?

A. 4 electrons; oxidation

B. 4 electrons; reduction

C. 6 electrons; oxidation

D. 6 electrons; reduction

QUESTION 4 · MEDIUM · NUMERICAL RESPONSE

Iodate ions change to iodine in an acidic solution: $IO_3^-(aq) \rightarrow I_2(s)$. When the half-reaction is balanced with the smallest whole numbers, how many electrons are in it? Record your answer as a whole number.

Answer: ☐

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

Some breath testers use the change of ethanol to ethanoic acid in an acidic solution: $C_2H_5OH(aq) \rightarrow CH_3COOH(aq)$. Which half-reaction is balanced?

A. $C_2H_5OH(aq) \rightarrow CH_3COOH(aq) + 2H^+(aq) + 2e^-$

B. $C_2H_5OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + 4H^+(aq) + 4e^-$

C. $C_2H_5OH(aq) + H_2O(l) + 4e^- \rightarrow CH_3COOH(aq) + 4H^+(aq)$

D. $C_2H_5OH(aq) + 2H_2O(l) \rightarrow CH_3COOH(aq) + 6H^+(aq) + 6e^-$

CHALLENGE

Arsenic(III) oxide changes to arsenic acid in an acidic solution: $As_2O_3(s) \rightarrow H_3AsO_4(aq)$. Write the balanced half-reaction. Then use oxidation numbers to check the number of electrons.

▼ **Hint**

Balance arsenic first. There are two arsenic atoms on the left.

▼ **Check your answers for Card 2**

Question 1: A. Nitrogen is balanced. There is no oxygen. Add $4H^+$ to the right for hydrogen. Left charge is 0 and right is +4, so add $4e^-$ to the right.

Question 2: 242. $SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$. Oxygen needs $2H_2O$ on the left. Hydrogen needs $4H^+$ on the right. Right charge is $-2 + 4 = +2$, so add $2e^-$ to the right.

Question 3: D. $ClO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow Cl^-(aq) + 3H_2O(l)$. Left charge before electrons is $-1 + 6 = +5$. Right is -1. Add $6e^-$ to the left. Electrons on the left: reduction.

Question 4: 10. Balance iodine first: $2IO_3^-$. Then $2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(s) + 6H_2O(l)$. Left charge before electrons is $-2 + 12 = +10$. Right is 0.

Question 5: B. Carbon: 2 on each side. Oxygen: 1 on the left, 2 on the right, so add 1 H_2O to the left. Hydrogen: 8 on the left, 4 on the right, so add $4H^+$ to the right. Right charge is +4, so add $4e^-$ to the right.

Challenge. $As_2O_3(s) + 5H_2O(l) \rightarrow 2H_3AsO_4(aq) + 4H^+(aq) + 4e^-$. Arsenic: 2 on each side. Oxygen: $3 + 5 = 8$ on the left, 8 on the right. Hydrogen: 10 on the left, $6 + 4 = 10$ on the right. Charge: 0 on the left, $+4 - 4 = 0$ on the right. Check: arsenic is +3 in As_2O_3 and +5 in H_3AsO_4. Two atoms each lose 2 electrons, which is 4 electrons.

Card 2 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. Balance in an acidic solution: $BrO_3^-(aq) \rightarrow Br_2(l)$.
2. Balance in an acidic solution: $H_2C_2O_4(aq) \rightarrow CO_2(g)$.

3. A balanced half-reaction has its electrons on the left side. Is it an oxidation or a reduction?

Need more? Go back to the lesson: Lesson B3. Videos for each idea: Videos by concept.

Card 3: Balance a whole redox equation

I can balance a redox equation from the table, from a skeleton equation in acid, in a basic solution when the half-reactions are given, and for a disproportionation.

How to spot it on the diploma

- "Balance the equation" or "The coefficient of $H^+(aq)$ in the balanced equation is …"
- A **skeleton equation**: only the main reactants and products, with no $H^+(aq)$ or $H_2O(l)$ shown. You add them.
- A basic solution: the diploma gives you the half-reactions. You only join them.

The steps

1. **Find the two half-reactions.** Use oxidation numbers to see what is oxidized and what is reduced. Copy each half from the table, or write it with the Card 2 steps.
2. **Make the electrons equal.** Multiply each half-reaction so the electrons lost equal the electrons gained.
3. **Add the two halves.** Write everything on the left of both, an arrow, then everything on the right of both.
4. **Cancel.** Remove the electrons. Remove any $H^+(aq)$, $H_2O(l)$ or $OH^-(aq)$ that is on both sides.
5. **Check.** Count each atom and the total charge on both sides.

Worked example

Balance this skeleton equation for an acidic solution: $MnO_4^-(aq) + H_2C_2O_4(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$.

> **Step 1**
> Reduction, from the table of standard electrode potentials in the Data Booklet: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$. Oxidation, by the Card 2 steps: $H_2C_2O_4(aq) \rightarrow 2CO_2(g) + 2H^+(aq) + 2e^-$.
>
> **Step 2**
> Multiply the reduction by 2 and the oxidation by 5. Each now has 10 electrons.
>
> **Step 3**
> $2MnO_4^- + 16H^+ + 10e^- + 5H_2C_2O_4 \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2 + 10H^+ + 10e^-$
>
> **Step 4**
> Cancel $10e^-$. Cancel $10H^+$ from both sides, which leaves $6H^+$ on the left.
>
> **Step 5**
> $2MnO_4^-(aq) + 5H_2C_2O_4(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$. Charge: $-2 + 6 = +4$ on the left, $+4$ on the right.

The common mistake

Forgetting to multiply before adding. The electrons must cancel completely. If any electrons are left in your final equation, go back to step 2.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

Aluminium metal reacts with copper(II) ions: Al(s) + Cu^{2+}(aq) → Al^{3+}(aq) + Cu(s). What is the coefficient of Cu^{2+}(aq) in the balanced equation?

A. 1

B. 2

C. 3

D. 6

QUESTION 2 · MEDIUM · NUMERICAL RESPONSE

Use the table of standard electrode potentials in the Data Booklet to balance the reaction of acidified dichromate ions, $Cr_2O_7^{2-}$(aq), with iron(II) ions, Fe^{2+}(aq), which form Fe^{3+}(aq). What is the coefficient of H^+(aq) in the balanced equation? Record your answer as a whole number.

Answer: ☐

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

Copper metal reacts with nitrate ions in an acidic solution. The skeleton equation is Cu(s) + NO_3^-(aq) → Cu^{2+}(aq) + NO(g). Which equation is balanced?

A. Cu(s) + NO_3^-(aq) + 4H^+(aq) → Cu^{2+}(aq) + NO(g) + 2H_2O(l)

B. 3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) → 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(l)

C. 3Cu(s) + 2NO_3^-(aq) + 4H^+(aq) → 3Cu^{2+}(aq) + 2NO(g) + 2H_2O(l)

D. 2Cu(s) + 3NO_3^-(aq) + 12H^+(aq) → 2Cu^{2+}(aq) + 3NO(g) + 6H_2O(l)

QUESTION 4 · EXAM LEVEL · NUMERICAL RESPONSE

In a basic solution, permanganate ions react with sulfite ions. The half-reactions are given:
MnO_4^-(aq) + 2H_2O(l) + 3e^- → MnO_2(s) + 4OH^-(aq)
SO_3^{2-}(aq) + 2OH^-(aq) → SO_4^{2-}(aq) + H_2O(l) + 2e^-
In the balanced equation with the smallest whole numbers, the coefficient of SO_3^{2-}(aq) is *i*, the coefficient of H_2O(l) is *j* and the coefficient of OH^-(aq) is *k*. Record the three digits *i*, *j*, *k* in that order.

Answer: ☐

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

Nitrogen dioxide reacts with water to make nitric acid and nitrogen monoxide. Nitrogen is both oxidized and reduced. The skeleton equation is $NO_2(g) \rightarrow NO_3^-(aq) + NO(g)$ in an acidic solution. What is the coefficient of $NO_2(g)$ in the balanced equation?

A. 1

B. 2

C. 3

D. 4

CHALLENGE

Manganate ions, $MnO_4^{2-}(aq)$, are green. In an acidic solution they change to purple permanganate ions, $MnO_4^-(aq)$, and solid $MnO_2(s)$. Write the balanced equation for this disproportionation.

▼ **Hint**

Write two half-reactions that both start with $MnO_4^{2-}(aq)$. One makes $MnO_4^-(aq)$. The other makes $MnO_2(s)$.

▼ **Check your answers for Card 3**

Question 1: C. Al loses 3 electrons and Cu^{2+} gains 2. Multiply by 2 and by 3: $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$.

Question 2: 14. The dichromate half-reaction has 6 electrons. Multiply $Fe^{2+} \rightarrow Fe^{3+} + e^-$ by 6: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 6Fe^{3+}(aq)$.

Question 3: B. Halves: $Cu \rightarrow Cu^{2+} + 2e^-$, and $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$. Multiply by 3 and by 2 to get 6 electrons each. Charge: −2 + 8 = +6 on the left, +6 on the right.

Question 4: 312. Multiply the first by 2 and the second by 3 (6 electrons each). Add: $2MnO_4^- + 4H_2O + 3SO_3^{2-} + 6OH^- \rightarrow 2MnO_2 + 8OH^- + 3SO_4^{2-} + 3H_2O$. Cancel $3H_2O$ and $6OH^-$: $2MnO_4^-(aq) + 3SO_3^{2-}(aq) + H_2O(l) \rightarrow 2MnO_2(s) + 3SO_4^{2-}(aq) + 2OH^-(aq)$.

Question 5: C. Oxidation: $NO_2(g) + H_2O(l) \rightarrow NO_3^-(aq) + 2H^+(aq) + e^-$. Reduction: $NO_2(g) + 2H^+(aq) + 2e^- \rightarrow NO(g) + H_2O(l)$. Multiply the oxidation by 2 and add: $3NO_2(g) + H_2O(l) \rightarrow 2NO_3^-(aq) + 2H^+(aq) + NO(g)$.

Challenge. Oxidation: $MnO_4^{2-}(aq) \rightarrow MnO_4^-(aq) + e^-$. Reduction: $MnO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow MnO_2(s) + 2H_2O(l)$. Multiply the oxidation by 2 and add: $3MnO_4^{2-}(aq) + 4H^+(aq) \rightarrow 2MnO_4^-(aq) + MnO_2(s) + 2H_2O(l)$. Charge: −6 + 4 = −2 on the left, −2 on the right.

Card 3 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. Balance: $Zn(s) + Ag^+(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$.

2. Balance: $Sn^{2+}(aq) + Fe^{3+}(aq) \rightarrow Sn^{4+}(aq) + Fe^{2+}(aq)$.
3. A reduction half-reaction has 5 electrons. An oxidation half-reaction has 2 electrons. What do you multiply each by?

Need more? Go back to the lesson: Lesson B3 · Lesson B5 · Lesson NI. Videos for each idea: Videos by concept.

Card 4: Build a table of oxidizing and reducing agents from lab results

I can use the results of an experiment to rank oxidizing agents and reducing agents, and write them as a small table of reduction half-reactions.

How to spot it on the diploma

- A table of results with “reaction” and “no reaction”, or ticks and crosses.
- Made-up metals with letters such as X, Y and Z.
- “The strongest oxidizing agent is …” or “List the oxidizing agents from strongest to weakest.”

The steps

1. **Read each result as a pair.** The ion in the solution is the oxidizing agent. The metal strip is the reducing agent.
2. **Use each reaction.** A reaction happened: the ion is **above** the metal in the table.
3. **Use each “no reaction”.** No reaction: the ion is **below** the metal in the table.
4. **Rank the ions.** The ion that reacts with the most metals is the strongest oxidizing agent. It goes at the top.
5. **Write the table.** Write each line as a reduction half-reaction, strongest oxidizing agent first. The strongest reducing agent is then at the bottom on the right.

Worked example

Three metals and their ions were tested. $X^{2+}(aq)$ reacted with Y(s) and with Z(s). $Z^{2+}(aq)$ reacted with Y(s) only. $Y^{2+}(aq)$ reacted with no metal. Build the table.

> **Step 1**
> The ions X^{2+}, Y^{2+} and Z^{2+} are the oxidizing agents. The metals are the reducing agents.
>
> **Steps 2 and 3**
> X^{2+} is above Y and above Z. Z^{2+} is above Y but below X. Y^{2+} is below X and below Z.
>
> **Step 4**
> X^{2+} reacted with 2 metals, Z^{2+} with 1, Y^{2+} with 0. Strongest to weakest: X^{2+}, Z^{2+}, Y^{2+}.
>
> **Step 5**
> $X^{2+}(aq) + 2e^- \rightarrow X(s)$
> $Z^{2+}(aq) + 2e^- \rightarrow Z(s)$
> $Y^{2+}(aq) + 2e^- \rightarrow Y(s)$
> The strongest reducing agent is Y(s), at the bottom on the right.

The common mistake

Reading the table upside down for reducing agents. The strongest oxidizing agent is at the top left. The strongest reducing agent is at the bottom right.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

A student puts a strip of each metal into a solution of each ion. A tick (✓) means a reaction was seen. A cross (×) means no reaction was seen.

	$W^{2+}(aq)$	$X^{2+}(aq)$	$Y^{2+}(aq)$	$Z^{2+}(aq)$
W(s)	—	×	×	✓
X(s)	✓	—	✓	✓
Y(s)	✓	×	—	✓
Z(s)	×	×	×	—

The letters W, X, Y and Z are not real symbols.
Which ion is the strongest oxidizing agent?

A. $W^{2+}(aq)$

B. $X^{2+}(aq)$

C. $Y^{2+}(aq)$

D. $Z^{2+}(aq)$

QUESTION 2 · EASY · MULTIPLE CHOICE

Use the results table in question 1. Which metal is the strongest reducing agent?

A. W(s)

B. X(s)

C. Y(s)

D. Z(s)

QUESTION 3 · MEDIUM · NUMERICAL RESPONSE

Use the results table in question 1. The ions are numbered: 1 is $W^{2+}(aq)$, 2 is $X^{2+}(aq)$, 3 is $Y^{2+}(aq)$, 4 is $Z^{2+}(aq)$. List the oxidizing agents from strongest to weakest. Record the four digits.

Answer: ____

QUESTION 4 · MEDIUM · MULTIPLE CHOICE

Metal Q is put into a solution of $R^+(aq)$, and solid R forms. Metal T is put into a solution of $R^+(aq)$, and nothing changes. Which table lists the half-reactions with the strongest oxidizing agent at the top?

A. $Q^+(aq) + e^- \rightarrow Q(s)$
$R^+(aq) + e^- \rightarrow R(s)$
$T^+(aq) + e^- \rightarrow T(s)$

B. $R^+(aq) + e^- \rightarrow R(s)$
$T^+(aq) + e^- \rightarrow T(s)$
$Q^+(aq) + e^- \rightarrow Q(s)$

C. $T^+(aq) + e^- \rightarrow T(s)$
$R^+(aq) + e^- \rightarrow R(s)$
$Q^+(aq) + e^- \rightarrow Q(s)$

D. $T^+(aq) + e^- \rightarrow T(s)$
$Q^+(aq) + e^- \rightarrow Q(s)$
$R^+(aq) + e^- \rightarrow R(s)$

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

In the experiment in question 1, which of the following is a controlled variable?

A. Whether a reaction is seen

B. The metal put into each solution

C. The concentration of each solution

D. The ion in each solution

CHALLENGE

Metal M reacts with hydrochloric acid and bubbles of gas form. Metal N does not react with hydrochloric acid. Metal P reacts with a solution of $M^{2+}(aq)$. (a) List $N^{2+}(aq)$, $H^+(aq)$, $M^{2+}(aq)$ and $P^{2+}(aq)$ from strongest to weakest oxidizing agent. (b) A solution of $M^{2+}(aq)$ must be stored in a metal container. Which of the metals N or P could the container be made of?

▼ **Hint**

Hydrogen ions are an oxidizing agent too. Put $H^+(aq)$ in your table like any other ion. A container must **not** react with the solution inside it.

▼ **Check your answers for Card 4**

Question 1: D. Read down each column. Z^{2+} reacted with 3 metals, W^{2+} with 2, Y^{2+} with 1 and X^{2+} with 0. Z^{2+} takes electrons from the most metals.

Question 2: B. Read across each row. X(s) reacted with all 3 ions, so it gives electrons most easily. It is also the metal of the weakest oxidizing agent, X^{2+}.

Question 3: 4132. Number of metals each ion reacted with: Z^{2+} 3, W^{2+} 2, Y^{2+} 1, X^{2+} 0. So Z^{2+}, W^{2+}, Y^{2+}, X^{2+}.

Question 4: C. R^{+} reacted with Q, so R^{+} is above Q. R^{+} did not react with T, so R^{+} is below T. Top to bottom: T^{+}, R^{+}, Q^{+}.

Question 5: C. A controlled variable is one the student could change but keeps the same. The metal and the ion are changed on purpose (manipulated). Whether a reaction is seen is what is observed (responding).

Challenge. (a) H^{+} reacts with M, so H^{+} is above M. H^{+} does not react with N, so H^{+} is below N. M^{2+} reacts with P, so M^{2+} is above P. Strongest to weakest: $N^{2+}(aq)$, $H^{+}(aq)$, $M^{2+}(aq)$, $P^{2+}(aq)$. (b) $M^{2+}(aq)$ reacts with a metal below it. P is below it, so a container of P would react. N is above it, so there is no reaction. The container could be made of N.

Card 4 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. $A^{2+}(aq)$ reacts with metal B. Which is the stronger oxidizing agent: $A^{2+}(aq)$ or $B^{2+}(aq)$?
2. In a table of reduction half-reactions, where is the strongest reducing agent?
3. $C^{2+}(aq)$ does not react with metal D. Will $D^{2+}(aq)$ react with metal C?

Need more? Go back to the lesson: Lesson B4. Videos for each idea: Videos by concept.

Card 5: Predict a reaction and say if it is spontaneous

I can list what is in the beaker, pick the strongest oxidizing agent and the strongest reducing agent, write the equation, and say if the reaction is spontaneous.

How to spot it on the diploma

- Two things are mixed and you are asked for the products, the equation, or “will a reaction happen?”
- The word “acidified”. It tells you $H^+(aq)$ is in the beaker.
- “Which metal could be used to make a container for this solution?”

The steps

1. **List every particle in the beaker.** Write the ions of each dissolved compound, any solid metal, and $H_2O(l)$. Add $H^+(aq)$ if it is acidified. Add $OH^-(aq)$ if it is basic.
2. **Label each particle.** Find it in the table of standard electrode potentials in the Data Booklet. On the left side it is an oxidizing agent. On the right side it is a reducing agent. Some particles are both.
3. **Pick the strongest of each.** The strongest oxidizing agent is the one highest on the left. The strongest reducing agent is the one lowest on the right.
4. **Decide.** Strongest oxidizing agent **above** the strongest reducing agent: spontaneous. Below it: not spontaneous, so no reaction.
5. **Write the equation.** Copy the reduction half-reaction as it is. Write the other half-reaction backwards, as an oxidation. Make the electrons equal and add (Card 3).

Worked example

A strip of copper is put into acidified potassium permanganate solution. Predict the reaction.

> **Step 1**
> In the beaker: $Cu(s)$, $K^+(aq)$, $MnO_4^-(aq)$, $H^+(aq)$, $H_2O(l)$.
>
> **Step 2**
> Oxidizing agents: $MnO_4^-(aq)$ with $H^+(aq)$ (+1.51 V), $H^+(aq)$ (0.00 V), $H_2O(l)$ (−0.83 V), $K^+(aq)$ (−2.93 V).
> Reducing agents: $Cu(s)$ (+0.34 V), $H_2O(l)$ (+1.23 V).
>
> **Step 3**
> Strongest oxidizing agent: $MnO_4^-(aq)$ with $H^+(aq)$, the highest on the left. Strongest reducing agent: $Cu(s)$, the lowest on the right.
>
> **Step 4**
> $MnO_4^-(aq)$ is above $Cu(s)$. The reaction is spontaneous.
>
> **Step 5**
> $2MnO_4^-(aq) + 16H^+(aq) + 5Cu(s) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Cu^{2+}(aq)$

The common mistake

> Leaving water and $H^+(aq)$ off the list. Also, do not pick the strongest reducing agent from the top of the table. The strongest reducing agent in the beaker is the **lowest** one on the right.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

Which metal reacts spontaneously with $Ni^{2+}(aq)$? Use the table of standard electrode potentials in the Data Booklet.

A. Ag(s)

B. Cu(s)

C. Pb(s)

D. Zn(s)

QUESTION 2 · MEDIUM · MULTIPLE CHOICE

A solution of iron(III) chloride is mixed with a solution of tin(II) chloride. Which row gives the strongest oxidizing agent and the strongest reducing agent in the beaker?

A. $Fe^{3+}(aq)$ and $Cl^-(aq)$

B. $Fe^{3+}(aq)$ and $Sn^{2+}(aq)$

C. $Sn^{2+}(aq)$ and $Cl^-(aq)$

D. $Sn^{2+}(aq)$ and $H_2O(l)$

QUESTION 3 · MEDIUM · NUMERICAL RESPONSE

A 1.0 mol/L solution that contains $Cu^{2+}(aq)$ must be stored. The negative ion in the solution is a spectator ion. Four metals are numbered: 1 is Ag(s), 2 is Zn(s), 3 is Au(s), 4 is Fe(s). Which metals will **not** react with the solution? Record the digits from smallest to largest.

Answer: ☐

QUESTION 4 · EXAM LEVEL · MULTIPLE CHOICE

Acidified potassium dichromate solution is added to potassium iodide solution. What are the products of the spontaneous reaction?

A. $Cr^{3+}(aq)$ and $I_2(s)$

B. $Cr^{3+}(aq)$ and $K(s)$

C. $H_2(g)$ and $I_2(s)$

D. $Cr^{3+}(aq)$ and $O_2(g)$

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

A strip of copper is put into hydrochloric acid. What happens, and why?

A. Hydrogen gas forms, because $H^+(aq)$ is above Cu(s) in the table.

B. Chlorine gas forms, because $Cl^-(aq)$ is the strongest reducing agent.

C. No reaction, because the strongest oxidizing agent is below the strongest reducing agent.

D. No reaction, because copper is not in the table.

CHALLENGE

A strip of nickel is put into one solution that contains $Ag^+(aq)$, $Cu^{2+}(aq)$ and $Zn^{2+}(aq)$, each at 1.0 mol/L. The negative ion is a spectator ion. (a) Which ion reacts first? Write the equation. (b) Which ion reacts next? (c) Which ion never reacts with the nickel?

▼ **Hint**

The strongest oxidizing agent in the beaker reacts first. When it is used up, make the list again without it.

▼ **Check your answers for Card 5**

Question 1: D. $Ni^{2+}(aq)$ is at −0.26 V. It reacts with a metal below it on the right. Zn(s) is at −0.76 V, below it. Ag(s), Cu(s) and Pb(s) (−0.13 V) are all above it.

Question 2: B. In the beaker: Fe^{3+}, Sn^{2+}, Cl^-, H_2O. Oxidizing agents: Fe^{3+} (+0.77 V), Sn^{2+} (−0.14 V), H_2O (−0.83 V). Reducing agents: Sn^{2+} (+0.15 V), H_2O (+1.23 V), Cl^- (+1.36 V). Highest on the left is Fe^{3+}. Lowest on the right is Sn^{2+}. The reaction is $2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)$.

Question 3: 13. $Cu^{2+}(aq)$ is at +0.34 V. It reacts with a metal below it: Zn(s) (−0.76 V) and Fe(s) (−0.45 V). Ag(s) (+0.80 V) and Au(s) (+1.50 V) are above it, so they do not react.

Question 4: A. In the beaker: K^+, $Cr_2O_7^{2-}$, H^+, I^-, H_2O. Strongest oxidizing agent: $Cr_2O_7^{2-}$ with H^+ (+1.23 V). Reducing agents: I^- (+0.54 V) and H_2O (+1.23 V). The lowest on the right is I^-. $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(s)$.

Question 5: C. In the beaker: Cu(s), H^+, Cl^-, H_2O. Strongest oxidizing agent: H^+ (0.00 V). Reducing agents: Cu (+0.34 V), H_2O (+1.23 V), Cl^- (+1.36 V). The lowest on the right is Cu(s), and H^+ is below it. Not

spontaneous.

Challenge. Oxidizing agents: Ag^+ (+0.80 V), Cu^{2+} (+0.34 V), Zn^{2+} (−0.76 V), H_2O (−0.83 V). The strongest reducing agent is Ni(s) (−0.26 V). (a) Ag^+ is the highest on the left, so it reacts first: Ni(s) + $2Ag^+$(aq) → Ni^{2+}(aq) + 2Ag(s). (b) When Ag^+ is used up, Cu^{2+} is the strongest oxidizing agent left. It is above Ni(s), so it reacts next. (c) Zn^{2+} is below Ni(s), so it never reacts.

Card 5 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. List every particle in a beaker of acidified sodium dichromate solution.
2. Will Pb(s) react spontaneously with Zn^{2+}(aq)? Why?
3. In the table, where do you find the strongest oxidizing agent in the beaker?

Need more? Go back to the lesson: Lesson B4 · Lesson RT · Lesson B7. Videos for each idea: Videos by concept.

Card 6: Redox titration: the calculation and the equipment

I can find a concentration or a mass from a redox titration, and I can choose the right equipment.

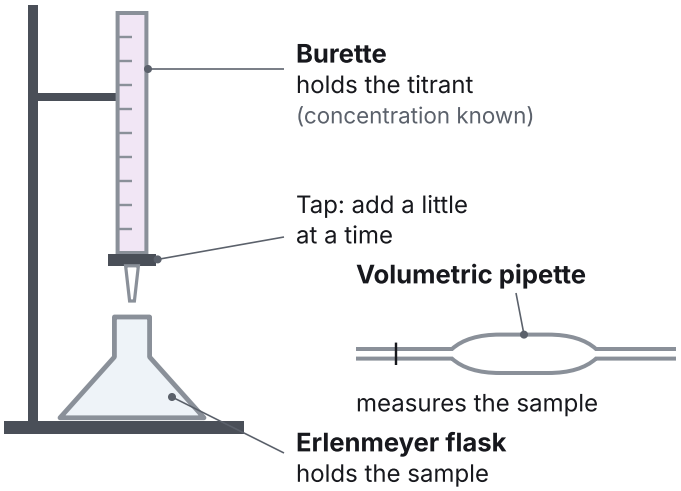
How to spot it on the diploma

- A volume of sample, and a volume and concentration of titrant such as $KMnO_4(aq)$ or $K_2Cr_2O_7(aq)$.
- A table of burette readings with several trials.
- “Which piece of equipment should be used to measure the sample?”
- Sometimes the equation is given. Sometimes you must build it first (Cards 3 and 5).

The steps

1. **Get the balanced equation.** If it is not given, build it with Card 5 or Card 3.
2. **Find the amount of titrant.** n = c × V. Change mL to L first. If there are several trials, average the ones that agree.
3. **Use the mole ratio.** Use the coefficients in the balanced equation to change moles of titrant into moles of the particle in the sample.
4. **Find what the question asks.** Concentration: c = n ÷ V of the sample, in L. Mass: m = n × M, with M from the periodic table in the Data Booklet.
5. **Round once, at the end.** Give the number of significant digits the question asks for, with the unit.

The equipment



The **sample** is measured with a **volumetric pipette** and put in an **Erlenmeyer flask**.

The **titrant** goes in the **burette**. Read the burette before and after. The difference is the volume added.

Acidified permanganate, $MnO_4^-(aq)$, is purple. In the flask it reacts and loses its colour. At the **endpoint** one more drop stays pale pink, because nothing is left for it to react with.

The first trial is often a rough one that goes past the endpoint. Average only the trials that agree.

Worked example

A 10.00 mL sample of acidified $Fe^{2+}(aq)$ is titrated with 0.0200 mol/L $KMnO_4(aq)$. The endpoint is reached at 15.50 mL. The equation is $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$. Find the concentration of $Fe^{2+}(aq)$.

> **Givens**
> Sample 10.00 mL = 0.01000 L. Titrant 15.50 mL = 0.01550 L at 0.0200 mol/L. All given in the question.
>
> **Step 2**
> $n(MnO_4^-) = 0.0200\ mol/L \times 0.01550\ L = 3.100 \times 10^{-4}\ mol$
>
> **Step 3**
> 5 mol of Fe^{2+} react with 1 mol of MnO_4^-. $n(Fe^{2+}) = 5 \times 3.100 \times 10^{-4}\ mol = 1.550 \times 10^{-3}\ mol$
>
> **Step 4**
> $c(Fe^{2+}) = 1.550 \times 10^{-3}\ mol \div 0.01000\ L = 0.155\ mol/L$

The common mistake

Skipping the mole ratio. One MnO_4^- ion reacts with five Fe^{2+} ions, not one. Students also lose marks by measuring the sample with a graduated cylinder. Use a volumetric pipette.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · NUMERICAL RESPONSE

A 25.00 mL sample that contains $Sn^{2+}(aq)$ is titrated with 0.0150 mol/L acidified $K_2Cr_2O_7(aq)$. The endpoint is reached at 18.40 mL. The equation is $Cr_2O_7^{2-}(aq) + 3Sn^{2+}(aq) + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3Sn^{4+}(aq) + 7H_2O(l)$. What is the concentration of $Sn^{2+}(aq)$, in mmol/L? Record your answer to three significant digits.

Answer: _____

QUESTION 2 · EASY · MULTIPLE CHOICE

A student must put exactly 10.00 mL of a sample into a flask for a titration. Which piece of equipment is best practice?

A. A 10 mL graduated cylinder

B. A 10.00 mL volumetric pipette

C. A 50 mL beaker

D. A 50.00 mL burette

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

Purple $KMnO_4(aq)$ is added from a burette to a colourless acidified sample of $Fe^{2+}(aq)$. What is seen at the endpoint?

A. The solution in the flask changes from purple to colourless.

B. The solution in the flask changes from colourless to pale pink.

C. A purple solid forms in the flask.

D. Bubbles of gas form in the flask.

QUESTION 4 · MEDIUM · NUMERICAL RESPONSE

A 10.00 mL sample of hydrogen peroxide solution is acidified and titrated with 0.0250 mol/L $KMnO_4(aq)$. The endpoint is reached at 14.30 mL. The equation is $2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(l)$. What is the concentration of $H_2O_2(aq)$, in mmol/L? Record your answer to three significant digits.

Answer: [ ]

QUESTION 5 · EXAM LEVEL · NUMERICAL RESPONSE

Samples of acidified $Fe^{2+}(aq)$, each 20.00 mL, are titrated with 0.0100 mol/L $K_2Cr_2O_7(aq)$. The equation is not given. Use the table of standard electrode potentials in the Data Booklet.

Trial	**1**	**2**	**3**	**4**
Volume of $K_2Cr_2O_7(aq)$ added (mL)	12.95	12.40	12.44	12.42

What is the concentration of $Fe^{2+}(aq)$, in mmol/L? Record your answer to three significant digits.

Answer: [ ]

CHALLENGE

An iron tablet has a mass of 0.500 g. It is dissolved in acid so that all its iron becomes $Fe^{2+}(aq)$. The whole solution is titrated with 0.0200 mol/L $KMnO_4(aq)$, and the endpoint is reached at 17.25 mL. The equation is not given. (a) What mass of iron, in mg, is in the tablet? (b) What percent of the tablet's mass is iron?

▼ **Hint**

Build the equation first: acidified $MnO_4^-(aq)$ with $Fe^{2+}(aq)$. You need the mole ratio. The molar mass of iron is in the periodic table in the Data Booklet.

▼ **Check your answers for Card 6**

Question 1: 33.1. $n(Cr_2O_7^{2-}) = 0.0150\ mol/L \times 0.01840\ L = 2.760 \times 10^{-4}$ mol. $n(Sn^{2+}) = 3 \times$ that $= 8.280 \times 10^{-4}$ mol. $c = 8.280 \times 10^{-4}\ mol \div 0.02500\ L = 0.0331\ mol/L = 33.1$ mmol/L.

Question 2: B. A volumetric pipette delivers one exact volume. A graduated cylinder or a burette in its place is a shortcut, and is not best practice.

Question 3: B. Before the endpoint, each drop of purple MnO_4^- reacts with Fe^{2+} and its colour goes. At the endpoint the Fe^{2+} is used up, so one more drop stays and the solution is pale pink.

Question 4: 89.4. $n(MnO_4^-) = 0.0250\ mol/L \times 0.01430\ L = 3.575 \times 10^{-4}$ mol. The ratio is 5 mol H_2O_2 to 2 mol MnO_4^-. $n(H_2O_2) = 3.575 \times 10^{-4}\ mol \times 5/2 = 8.938 \times 10^{-4}$ mol. $c =$ that $\div 0.01000\ L = 0.0894\ mol/L = 89.4$ mmol/L.

Question 5: 37.3. Trial 1 does not agree with the others, so leave it out. Average of trials 2 to 4: 12.42 mL. Build the equation: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 6Fe^{3+}(aq)$. $n(Cr_2O_7^{2-}) = 0.0100\ mol/L \times 0.01242\ L = 1.242 \times 10^{-4}$ mol. $n(Fe^{2+}) = 6 \times$ that $= 7.452 \times 10^{-4}$ mol. $c =$ that $\div 0.02000\ L = 0.0373\ mol/L = 37.3$ mmol/L.

Challenge. Equation: $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$. $n(MnO_4^-) = 0.0200\ mol/L \times 0.01725\ L = 3.450 \times 10^{-4}$ mol. $n(Fe^{2+}) = 5 \times$ that $= 1.725 \times 10^{-3}$ mol. (a) $m = 1.725 \times 10^{-3}\ mol \times 55.85\ g/mol = 0.09634\ g = 96.3$ mg. (b) $0.09634\ g \div 0.500\ g \times 100 = 19.3\%$.

Card 6 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. In $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$, how many moles of Fe^{2+} react with 1 mol of MnO_4^-?
2. How many moles of MnO_4^- are in 10.00 mL of 0.0200 mol/L $KMnO_4(aq)$?
3. Why is acid added to the flask in a permanganate titration?

Need more? Go back to the lesson: Lesson RT. Videos for each idea: Videos by concept.

Card 7: Read and label a voltaic cell

I can find the anode and the cathode of a voltaic cell, the half-reaction at each, the way electrons and ions move, the cell notation, and what I would see.

How to spot it on the diploma

- A diagram of two beakers joined by a salt bridge, with a voltmeter in the wire. Parts are numbered and you match the labels.
- A line of cell notation, such as Zn(s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu(s).
- "The electrons flow from …", "The anode is …", "The mass of the electrode will …"

The steps

1. **Find both half-reactions in the table.** Each half-cell is one line of the table of standard electrode potentials in the Data Booklet.
2. **Pick the cathode and the anode.** The line that is **higher** in the table runs forward as the reduction: that is the cathode (RED CAT). The lower line runs backwards as the oxidation: that is the anode (AN OX).
3. **Show the electrons.** Electrons move through the wire from the anode to the cathode.
4. **Show the ions.** Anions move toward the anode. Cations move toward the cathode.
5. **Predict what you would see.** A metal anode loses mass. A cathode gains mass if a metal forms on it. A coloured ion that is used up fades.
6. **Write the cell notation.** anode | anode solution || cathode solution | cathode

Worked example

A voltaic cell has a nickel electrode in Ni^{2+}(aq) and a silver electrode in Ag^{+}(aq). Describe the cell.

> **Step 1**
> Ag^{+}(aq) + e^{-} → Ag(s) is at +0.80 V. Ni^{2+}(aq) + 2e^{-} → Ni(s) is at −0.26 V.
>
> **Step 2**
> The silver line is higher. Cathode: Ag^{+}(aq) + e^{-} → Ag(s). Anode: Ni(s) → Ni^{2+}(aq) + 2e^{-}.
>
> **Step 3**
> Electrons move through the wire from the nickel electrode to the silver electrode.
>
> **Step 4**
> Anions move toward the nickel half-cell. Cations move toward the silver half-cell.
>
> **Step 5**
> The nickel electrode loses mass. The silver electrode gains mass.
>
> **Step 6**
> Ni(s) | Ni^{2+}(aq) || Ag^{+}(aq) | Ag(s). Net equation: Ni(s) + 2Ag^{+}(aq) → Ni^{2+}(aq) + 2Ag(s)

The common mistake

> Deciding the anode by which metal “sounds” more reactive. Use the table every time: the higher line is the cathode. Also, electrons never go through the salt bridge. Only ions do.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

A voltaic cell has a zinc electrode in $Zn^{2+}(aq)$ and a copper electrode in $Cu^{2+}(aq)$.
The anode is *i*, and electrons move through the wire *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	zinc	from zinc to copper
B.	zinc	from copper to zinc
C.	copper	from zinc to copper
D.	copper	from copper to zinc

QUESTION 2 · EASY · MULTIPLE CHOICE

What is the job of the salt bridge in a voltaic cell?

A. It lets electrons move from one half-cell to the other.

B. It lets ions move between the half-cells, so each solution stays neutral.

C. It measures the voltage of the cell.

D. It supplies the energy that makes the reaction happen.

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

A cell is written as $Pb(s) \mid Pb^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$. Which statement is true?

A. Pb(s) is the cathode, and $Pb^{2+}(aq)$ is reduced.

B. Pb(s) is the anode, and Pb(s) is oxidized.

C. Cu(s) is the anode, and Cu(s) is oxidized.

D. Cu(s) is the cathode, and Cu(s) is reduced.

QUESTION 4 · MEDIUM · MULTIPLE CHOICE

A cell is written as Co(s) | Co^{2+}(aq) || Cu^{2+}(aq) | Cu(s). The Data Booklet table of ion colours says Co^{2+}(aq) is pink and Cu^{2+}(aq) is blue. Which row gives two things seen as the cell runs?

A. The cobalt electrode gains mass, and the blue colour fades.

B. The cobalt electrode loses mass, and the blue colour fades.

C. The cobalt electrode loses mass, and the blue colour gets deeper.

D. The copper electrode loses mass, and the pink colour fades.

QUESTION 5 · EXAM LEVEL · NUMERICAL RESPONSE

A voltaic cell is built from an aluminium half-cell and a nickel half-cell. Four particles are numbered: 1 is Al(s), 2 is Al^{3+}(aq), 3 is Ni^{2+}(aq), 4 is Ni(s). The anode is *i*. The oxidizing agent is *j*. The particle that forms at the cathode is *k*. Cations move toward electrode *l*. Record the four digits *i*, *j*, *k*, *l* in that order.

Answer: ☐

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

A cell is written as Zn(s) | Zn^{2+}(aq) || Fe^{3+}(aq), Fe^{2+}(aq) | C(s). Which row gives the half-reaction at the cathode and what happens to the mass of the carbon electrode?

A. $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$; the mass does not change

B. $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$; the mass goes up

C. $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$; the mass does not change

D. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$; the mass goes up

CHALLENGE

A voltaic cell has a cadmium electrode in Cd^{2+}(aq). Its other half-cell is a carbon electrode in an acidified solution of $Cr_2O_7^{2-}$(aq). The table of ion colours says $Cr_2O_7^{2-}$(aq) is orange and Cr^{3+}(aq) is green. (a) Write the half-reaction at each electrode and the net equation. (b) Write the cell notation. (c) Give one thing you would see in each half-cell. (d) What happens to the pH in the cathode half-cell?

▼ **Hint**

Find both lines in the table. The higher one is the cathode. Look at what is used up at the cathode.

▼ **Check your answers for Card 7**

Question 1: A. Cu^{2+}/Cu (+0.34 V) is higher than Zn^{2+}/Zn (−0.76 V). Copper is the cathode and zinc is the anode. Electrons move from the anode to the cathode.

Question 2: B. Electrons move through the wire. Ions move through the salt bridge. Without it, charge would build up in each half-cell and the cell would stop.

Question 3: B. In cell notation the anode is on the far left. Oxidation happens at the anode: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$. At the cathode $Cu^{2+}(aq)$ is reduced, not Cu(s).

Question 4: B. Anode: $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$, so the cobalt electrode loses mass. Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$, so blue $Cu^{2+}(aq)$ is used up and its colour fades.

Question 5: 1344. Ni^{2+}/Ni (−0.26 V) is higher than Al^{3+}/Al (−1.66 V). Anode: Al(s), number 1. The oxidizing agent is $Ni^{2+}(aq)$, number 3. Ni(s), number 4, forms at the cathode. Cations move toward the cathode, number 4.

Question 6: A. The cathode is on the far right: the inert carbon electrode. $Fe^{3+}(aq)$ is reduced to $Fe^{2+}(aq)$ on its surface. Both ions stay in the solution and carbon does not react, so its mass does not change. Net: $Zn(s) + 2Fe^{3+}(aq) \rightarrow Zn^{2+}(aq) + 2Fe^{2+}(aq)$.

Challenge. (a) Cathode (+1.23 V): $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$. Anode (−0.40 V): $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$. Net: $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 3Cd(s) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Cd^{2+}(aq)$. (b) $Cd(s) \mid Cd^{2+}(aq) \parallel Cr_2O_7^{2-}(aq), H^+(aq), Cr^{3+}(aq) \mid C(s)$. (c) The cadmium electrode loses mass. The orange solution turns green. (d) $H^+(aq)$ is used up, so the pH goes up.

Card 7 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. In any cell, what happens at the anode?
2. In a cell with magnesium and lead half-cells, which metal is the cathode? Use the table.
3. Write the cell notation for a zinc–silver voltaic cell.

Need more? Go back to the lesson: Lesson B6. Videos for each idea: Videos by concept.

Card 8: Cell potential and the reference half-cell

I can calculate a standard cell potential. I can handle a new reference half-cell, or a half-reaction that is missing from the table.

How to spot it on the diploma

- “The standard cell potential is …” or “The voltage of this cell is …”
- “If the reference half-cell were changed to …”
- A cell voltage is given and you are asked for the reduction potential of one half-reaction.

The steps

1. **Read both table values.** Copy each E° with its sign from the table of standard electrode potentials in the Data Booklet.
2. **Decide which is the cathode and which is the anode.** The reduction happens at the cathode. The oxidation happens at the anode.
3. **Subtract. E°cell = E°cathode − E°anode.** Do not change a sign first. Do not multiply a value by a coefficient.
4. **Read the sign of the answer.** Positive: spontaneous, a voltaic cell. Negative: not spontaneous.
5. **For a new reference:** subtract the new reference’s old value from every table value. E°cell stays the same.
6. **For a missing value:** put what you know into E°cell = E°cathode − E°anode and solve for the one you do not know.

Worked example

Find the standard cell potential of the cell Ni(s) | Ni^{2+}(aq) || Ag^{+}(aq) | Ag(s).

> **Givens**
> E°(Ag^{+}/Ag) = +0.80 V and E°(Ni^{2+}/Ni) = −0.26 V, both from the table of standard electrode potentials in the Data Booklet.
>
> **Step 2**
> Silver is the cathode (far right). Nickel is the anode (far left).
>
> **Step 3**
> E°cell = +0.80 V − (−0.26 V) = +1.06 V
>
> **Step 4**
> The answer is positive, so the reaction is spontaneous. Silver’s half-reaction is used twice in the net equation, but +0.80 V is **not** doubled.

The common mistake

Changing the sign of the anode's value and then subtracting as well. That changes it twice. Use the table values exactly as printed in E°cell = E°cathode − E°anode.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · NUMERICAL RESPONSE

What is the standard cell potential, in volts, of the cell Zn(s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu(s)? Use the table of standard electrode potentials in the Data Booklet. Record your answer to two decimal places.

Answer: ____

QUESTION 2 · EASY · MULTIPLE CHOICE

The net equation for a cell is Al(s) + 3Ag^{+}(aq) → Al^{3+}(aq) + 3Ag(s). What is the standard cell potential?

A. +0.74 V

B. +2.46 V

C. +4.06 V

D. −2.46 V

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

A student predicts this reaction: Cu(s) + 2H^{+}(aq) → Cu^{2+}(aq) + H_2(g).
E° for this reaction is *i*, so the reaction is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	−0.34 V	nonspontaneous
B.	−0.34 V	spontaneous
C.	+0.34 V	spontaneous
D.	+0.34 V	nonspontaneous

QUESTION 4 · MEDIUM · NUMERICAL RESPONSE

A metal X is not in the Data Booklet. The cell X(s) | X^{2+}(aq) || Cu^{2+}(aq) | Cu(s) has a standard cell potential of +0.91 V. The reduction potential of $X^{2+}(aq) + 2e^- \rightarrow X(s)$ is a negative number. What is its size, in volts, without the sign? Record your answer to two decimal places.

Answer: ☐

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

Suppose the $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ half-reaction were chosen as the reference, with a value of 0.00 V. Which row gives the new values for Zn^{2+}/Zn and for Ag^+/Ag?

A. −1.10 V and +0.46 V

B. −0.42 V and +1.14 V

C. −0.76 V and +0.80 V

D. −1.10 V and +1.14 V

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

A zinc–copper voltaic cell is built with 2.0 mol/L Cu^{2+}(aq) in place of 1.0 mol/L Cu^{2+}(aq). Cu^{2+}(aq) is a reactant. Which row describes the voltage?

A. It starts above 1.10 V and falls as the cell runs.

B. It starts above 1.10 V and rises as the cell runs.

C. It starts below 1.10 V and falls as the cell runs.

D. It stays at exactly 1.10 V until the cell stops.

CHALLENGE

A different table uses $Ag^+(aq) + e^- \rightarrow Ag(s)$ as its reference, at 0.00 V. On that table, $M^{3+}(aq) + 3e^- \rightarrow M(s)$ is listed at −1.30 V. (a) What is the value for M^{3+}/M on the Data Booklet table? (b) A cell is made from an M half-cell and a copper half-cell. Which metal is the anode, and what is E°cell? (c) Write the net equation.

▼ **Hint**

On the Data Booklet table, silver is at +0.80 V, not 0.00 V. Every value on the other table is lower by that amount.

▼ **Check your answers for Card 8**

Question 1: 1.10. E°cell = +0.34 V − (−0.76 V) = +1.10 V.

Question 2: B. E°cell = +0.80 V − (−1.66 V) = +2.46 V. The value for silver is not multiplied by 3.

Question 3: A. $H^+(aq)$ is reduced (cathode) and Cu(s) is oxidized (anode). E° = 0.00 V − (+0.34 V) = −0.34 V. A negative value means not spontaneous. The diploma writes this as one word: nonspontaneous.

Question 4: 0.57. X is the anode. E°cell = E°cathode − E°anode, so +0.91 V = +0.34 V − E°(X). E°(X) = +0.34 V − 0.91 V = −0.57 V.

Question 5: A. Subtract copper's old value, +0.34 V, from every value. Zinc: −0.76 − 0.34 = −1.10 V. Silver: +0.80 − 0.34 = +0.46 V. The gap between any two values is the same as before.

Question 6: A. The table value is for 1.0 mol/L. A higher concentration of a reactant gives a larger voltage. As the cell runs, reactants are used up, so the voltage falls.

Challenge. (a) Add silver's Data Booklet value: −1.30 V + 0.80 V = −0.50 V. (b) Cu^{2+}/Cu (+0.34 V) is higher, so copper is the cathode and M is the anode. E°cell = +0.34 V − (−0.50 V) = +0.84 V. (c) Multiply to get 6 electrons: $2M(s) + 3Cu^{2+}(aq) \rightarrow 2M^{3+}(aq) + 3Cu(s)$.

Card 8 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. What is E°cell for $Pb(s) \mid Pb^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$?
2. A half-reaction is used three times in a net equation. Do you multiply its E° by 3?
3. E° for a reaction is −0.50 V. Is the reaction spontaneous?

Need more? Go back to the lesson: Lesson B7. Videos for each idea: Videos by concept.

Card 9: Read and label an electrolytic cell

I can tell an electrolytic cell from a voltaic cell. I can predict the half-reaction and product at each electrode, find the minimum voltage, and say what I would see.

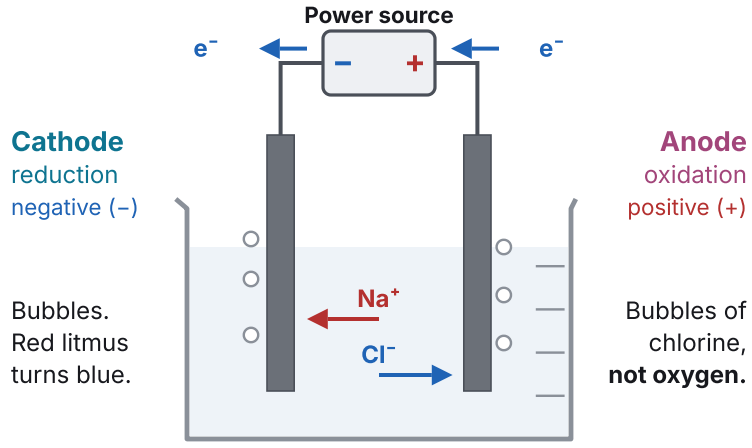
How to spot it on the diploma

- A **power source** or battery in the wire, in place of a voltmeter.
- One container, often with inert electrodes of C(s) or Pt(s), and no salt bridge.
- The words electrolysis, molten, electroplating or "minimum voltage".
- Cell notation with no double line, such as C(s) | Co^{2+}(aq), NO_3^-(aq) | C(s).

The steps

1. **List what is in the cell.** Molten compound: only its ions. Solution in water: its ions **and** H_2O(l).
2. **Find the reduction at the cathode.** The strongest oxidizing agent is reduced (RED CAT). The cathode is joined to the **negative** terminal of the power source.
3. **Find the oxidation at the anode.** The strongest reducing agent is oxidized (AN OX). The anode is joined to the **positive** terminal. In a **basic** solution, OH^-(aq) is in the cell too, and it is a reducing agent: find it on the right side of the line $O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$ (+0.40 V).
4. **Find the minimum voltage.** E°cell = E°cathode − E°anode. It is negative. The minimum voltage is its size.
5. **Predict what you would see.** A metal coats the cathode. Bubbles mean a gas. If water is reduced, OH^-(aq) forms and red litmus turns blue at the cathode. If water is oxidized, H^+(aq) forms and blue litmus turns red at the anode.
6. **Check for chloride.** With Cl^-(aq) and water, chlorine gas forms at the anode, not oxygen. This is the chloride anomaly.

The chloride anomaly



Sodium chloride solution: ions and water

In a chloride solution the reducing agents are $H_2O(l)$ (+1.23 V) and $Cl^-(aq)$ (+1.36 V). Water is lower on the right, so the table says water should be oxidized.

In a real cell, chlorine gas forms. Oxidizing water needs a large extra voltage, called an **overvoltage**. Oxidizing $Cl^-(aq)$ needs less. So as the voltage is turned up, $Cl^-(aq)$ is oxidized first.

It is **not** true that setting the supply to water's minimum voltage makes water react first.

The table predicts well for voltaic cells. For electrolytic cells, use it with care.

Worked example

A solution of potassium iodide, KI(aq), is electrolyzed with inert carbon electrodes. Predict the half-reactions, the minimum voltage and what is seen.

> **Step 1**
> In the cell: $K^+(aq)$, $I^-(aq)$, $H_2O(l)$.
>
> **Step 2**
> Oxidizing agents: $K^+(aq)$ (−2.93 V) and $H_2O(l)$ (−0.83 V). Water is higher. Cathode: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$.
>
> **Step 3**
> Reducing agents: $I^-(aq)$ (+0.54 V) and $H_2O(l)$ (+1.23 V). $I^-(aq)$ is lower. Anode: $2I^-(aq) \rightarrow I_2(s) + 2e^-$.
>
> **Step 4**
> $E°cell = -0.83\ V - (+0.54\ V) = -1.37\ V$. Minimum voltage: 1.37 V.
>
> **Step 5**
> Cathode: bubbles of gas, and red litmus turns blue. Anode: a brown colour forms. Net: $2H_2O(l) + 2I^-(aq) \rightarrow H_2(g) + 2OH^-(aq) + I_2(s)$

The common mistake

Forgetting water. In a solution, water is often the particle that reacts. Sodium and potassium metal never form in water: $H_2O(l)$ is a stronger oxidizing agent than $Na^+(aq)$ or $K^+(aq)$.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

Which feature of a diagram shows that a cell is electrolytic and not voltaic?

A. It has two electrodes.

B. It has a power source in the wire.

C. It has a solution of ions.

D. It has an anode and a cathode.

QUESTION 2 · EASY · MULTIPLE CHOICE

Molten sodium chloride, NaCl(l), is electrolyzed with inert electrodes.
The product at the cathode is *i*, and the product at the anode is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	Na(l)	$Cl_2(g)$
B.	$Cl_2(g)$	Na(l)
C.	$H_2(g)$	$Cl_2(g)$
D.	Na(l)	$O_2(g)$

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

A solution of copper(II) sulfate, $CuSO_4(aq)$, is electrolyzed with inert electrodes. Which row gives the product at each electrode?

A. Cathode: Cu(s). Anode: $O_2(g)$.

B. Cathode: $H_2(g)$. Anode: $O_2(g)$.

C. Cathode: Cu(s). Anode: $H_2(g)$.

D. Cathode: $O_2(g)$. Anode: Cu(s).

QUESTION 4 · MEDIUM · NUMERICAL RESPONSE

A solution of nickel(II) bromide, $NiBr_2(aq)$, is electrolyzed with inert electrodes. What is the minimum voltage needed, in volts? Use the table of standard electrode potentials in the Data Booklet. Record your answer to two decimal places.

Answer: ____

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

A solution of sodium chloride, NaCl(aq), is electrolyzed with inert electrodes. Which row gives the gas that really forms at the anode, and the reason?

A. $O_2(g)$, because water is the strongest reducing agent in the table.

B. $Cl_2(g)$, because oxidizing water needs a larger overvoltage than oxidizing $Cl^-(aq)$.

C. $Cl_2(g)$, because $Cl^-(aq)$ is below water on the right side of the table.

D. $H_2(g)$, because water is the strongest oxidizing agent present.

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

In an electrolytic cell, one electrode is joined to the negative terminal of the power source. This electrode is the *i*, and *ii* move toward it through the solution.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	anode	anions
B.	anode	cations
C.	cathode	cations
D.	cathode	anions

CHALLENGE

A solution of sodium sulfate, $Na_2SO_4(aq)$, is electrolyzed with inert electrodes. A piece of red litmus paper and a piece of blue litmus paper are held near each electrode. (a) Write the half-reaction at each electrode. (b) Write the net equation. (c) What is the minimum voltage? (d) Give one observation and one inference for the cathode.

▼ **Hint**

List the particles: $Na^+(aq)$, $SO_4^{2-}(aq)$, $H_2O(l)$. Look for each on the left side and on the right side of the table. In a neutral solution, $SO_4^{2-}(aq)$ is found only on the left.

▼ **Check your answers for Card 9**

Question 1: B. Both kinds of cell have electrodes, ions, an anode and a cathode. Only an electrolytic cell needs a power source, because its reaction is not spontaneous.

Question 2: A. A molten salt has no water. Only Na^+ and Cl^- are in the cell. Na^+ is reduced at the cathode. Cl^- is oxidized at the anode. Net: $2Na^+(l) + 2Cl^-(l) \rightarrow 2Na(l) + Cl_2(g)$.

Question 3: A. In the cell: Cu^{2+}, SO_4^{2-}, H_2O. Strongest oxidizing agent: Cu^{2+} (+0.34 V), above H_2O (−0.83 V). The only reducing agent is H_2O (+1.23 V). Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. Anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$. Net: $2Cu^{2+}(aq) + 2H_2O(l) \rightarrow 2Cu(s) + O_2(g) + 4H^+(aq)$.

Question 4: 1.33. Strongest oxidizing agent: Ni^{2+} (−0.26 V), above H_2O (−0.83 V). Strongest reducing agent: Br^- (+1.07 V), below H_2O (+1.23 V). $E°cell = -0.26\ V - (+1.07\ V) = -1.33\ V$. Minimum voltage: 1.33 V.

Question 5: B. The table predicts that water (+1.23 V) is oxidized before Cl^- (+1.36 V). In a real cell chlorine forms, because of the overvoltage for water. This is the chloride anomaly. $H_2(g)$ forms too, but at the cathode.

Question 6: C. The power source pushes electrons onto this electrode. Particles gain those electrons there, which is reduction, so it is the cathode. Positive ions (cations) are attracted to it.

Challenge. (a) Oxidizing agents: H_2O (−0.83 V), SO_4^{2-} in a neutral solution (−0.93 V, from the Data Booklet table) and Na^+ (−2.71 V). Water is the highest. Cathode: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$. The only reducing agent is H_2O (+1.23 V). Anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$. (b) Double the cathode half-reaction and add. The $4H^+$ and $4OH^-$ make $4H_2O$, which cancel: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$. (c) $E°cell = -0.83\ V - (+1.23\ V) = -2.06\ V$, so 2.06 V. (d) Observation: bubbles form at the cathode, and red litmus turns blue. Inference: the gas is hydrogen, and $OH^-(aq)$ formed.

Card 9 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. In an electrolytic cell, is the cathode positive or negative?
2. Molten magnesium chloride, $MgCl_2(l)$, is electrolyzed. What forms at the cathode?
3. During the electrolysis of a solution, red litmus turns blue at the cathode. Which ion formed there?

Need more? Go back to the lesson: Lesson B8 · Lesson B9. Videos for each idea: Videos by concept.

Card 10: Faraday’s law

I can use current and time to find a mass or an amount at an electrode, and I can work backwards to find a current or a time.

How to spot it on the diploma

- A current in amperes (A) and a time, with a metal being plated or a gas being made.
- “What mass of metal is deposited?”, “How long will it take?”, “What current is needed?”
- Watch the half-reaction. If it has 2 or 3 electrons, the mole ratio is not 1 to 1.

The steps

1. **Write the half-reaction.** It tells you how many moles of electrons go with one mole of the substance.
2. **Change the time to seconds.** Minutes × 60. Hours × 3600.
3. **Find the moles of electrons.** $q = I \times t$, then $n(e^-) = q \div F$. $F = 9.65 \times 10^4$ C/mol e^-, from the Data Booklet.
4. **Use the mole ratio.** Use the half-reaction to change moles of electrons into moles of the substance.
5. **Find the mass.** $m = n \times M$, with M from the periodic table in the Data Booklet. Round once, at the end.
6. **Working backwards?** Start from the mass. Do the same steps in the opposite order: $n = m \div M$, mole ratio, $q = n(e^-) \times F$, then $t = q \div I$ or $I = q \div t$.

Worked example

A current of 2.50 A runs for 20.0 min through a solution of $Cu^{2+}(aq)$. What mass of copper forms on the cathode?

> **Givens**
> I = 2.50 A and t = 20.0 min (given). $F = 9.65 \times 10^4$ C/mol e^- and M(Cu) = 63.55 g/mol (Data Booklet).
>
> **Step 1**
> $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. 2 mol of electrons make 1 mol of copper.
>
> **Step 2**
> t = 20.0 min × 60 s/min = 1200 s
>
> **Step 3**
> q = 2.50 A × 1200 s = 3000 C. $n(e^-) = 3000\ C \div 9.65 \times 10^4$ C/mol = 0.03109 mol
>
> **Step 4**
> n(Cu) = 0.03109 mol ÷ 2 = 0.01554 mol
>
> **Step 5**
> m = 0.01554 mol × 63.55 g/mol = 0.988 g

The common mistake

Leaving the time in minutes, or skipping the mole ratio. Cu^{2+} needs 2 electrons for each atom and Al^{3+} needs 3. Only ions with a 1+ charge, such as Ag^{+}, are 1 to 1.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · NUMERICAL RESPONSE

A current of 1.50 A runs for 600 s through a solution of $Ag^{+}(aq)$. The half-reaction is $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$. What mass of silver forms, in grams? Record your answer to three significant digits.

Answer: ☐

QUESTION 2 · EASY · MULTIPLE CHOICE

A current of 0.500 A runs for 30.0 min. How much charge passes?

A. 15.0 C

B. 60.0 C

C. 900 C

D. 3600 C

QUESTION 3 · MEDIUM · NUMERICAL RESPONSE

Aluminium is made by the electrolysis of molten aluminium oxide. The half-reaction is $Al^{3+}(l) + 3e^{-} \rightarrow Al(l)$. A current of 15.0 A runs for 2.00 h. What mass of aluminium forms, in grams? Record your answer to three significant digits.

Answer: ☐

QUESTION 4 · MEDIUM · NUMERICAL RESPONSE

A key is plated with 2.00 g of nickel from a solution of $Ni^{2+}(aq)$, using a current of 3.00 A. The half-reaction is $Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s)$. How long does it take, in minutes? Record your answer to three significant digits.

Answer: ☐

QUESTION 5 · EXAM LEVEL · NUMERICAL RESPONSE

A part is plated with 1.25 g of chromium in 45.0 min. The half-reaction is $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$. What current is needed, in amperes? Record your answer to three significant digits.

Answer: ____

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

The same current runs for the same time through two cells. One plates silver from $Ag^+(aq)$. The other plates copper from $Cu^{2+}(aq)$. How do the amounts of metal compare?

A. The moles of silver equal the moles of copper.

B. The moles of silver are twice the moles of copper.

C. The moles of silver are half the moles of copper.

D. The masses of silver and copper are equal.

CHALLENGE

A zinc–copper voltaic cell runs a small motor. It delivers a steady current of 0.150 A for 8.00 h. (a) What mass does the zinc anode lose? (b) What mass does the copper cathode gain?

▼ **Hint**

Faraday's law works for a voltaic cell too. The same moles of electrons leave the anode and arrive at the cathode. Both half-reactions have 2 electrons.

▼ **Check your answers for Card 10**

Question 1: 1.01. q = 1.50 A × 600 s = 900 C. $n(e^-)$ = 900 C ÷ 9.65 × 10^4 C/mol = 0.009326 mol. The ratio is 1 to 1, so n(Ag) = 0.009326 mol. m = that × 107.87 g/mol = 1.01 g.

Question 2: C. t = 30.0 min × 60 s/min = 1800 s. q = I × t = 0.500 A × 1800 s = 900 C.

Question 3: 10.1. t = 2.00 h × 3600 s/h = 7200 s. q = 15.0 A × 7200 s = 1.08 × 10^5 C. $n(e^-)$ = q ÷ 9.65 × 10^4 C/mol = 1.119 mol. n(Al) = 1.119 mol ÷ 3 = 0.3731 mol. m = that × 26.98 g/mol = 10.1 g.

Question 4: 36.5. n(Ni) = 2.00 g ÷ 58.69 g/mol = 0.03408 mol. $n(e^-)$ = 2 × that = 0.06815 mol. q = 0.06815 mol × 9.65 × 10^4 C/mol = 6577 C. t = q ÷ I = 6577 C ÷ 3.00 A = 2192 s = 36.5 min.

Question 5: 2.58. n(Cr) = 1.25 g ÷ 52.00 g/mol = 0.02404 mol. $n(e^-)$ = 3 × that = 0.07212 mol. q = 0.07212 mol × 9.65 × 10^4 C/mol = 6959 C. t = 45.0 min × 60 s/min = 2700 s. I = q ÷ t = 2.58 A.

Question 6: B. The same charge means the same moles of electrons. Each mole of electrons makes 1 mol of Ag but only ½ mol of Cu. So there is twice as much silver, in moles.

Challenge. t = 8.00 h × 3600 s/h = 28800 s. q = 0.150 A × 28800 s = 4320 C. n(e^-) = 4320 C ÷ 9.65 × 10^4 C/mol = 0.04477 mol. Each metal: 0.04477 mol ÷ 2 = 0.02238 mol. (a) Zinc: 0.02238 mol × 65.41 g/mol = 1.46 g lost. (b) Copper: 0.02238 mol × 63.55 g/mol = 1.42 g gained.

Card 10 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. How many moles of electrons are in 9650 C of charge?
2. How many moles of electrons are needed to make 1 mol of Al(l) from Al^{3+}(l)?
3. Change 15.0 min to the unit you need for q = I × t.

Need more? Go back to the lesson: Lesson B10 · Lesson B11. Videos for each idea: Videos by concept.

Card 11: Uses of redox and cells

I can explain how corrosion is stopped, and how plating, refining, batteries and fuel cells work, using what I know about redox.

How to spot it on the diploma

- A paragraph about a real technology: a pipeline, a ship, a battery, a fuel cell car, a plating shop, a water treatment plant.
- “Which metal could protect the iron?”, “The object to be plated is the …”
- A question about a benefit, a risk or an unintended result of a technology.

The steps

1. **Find the redox.** What is oxidized? What is reduced? Write the half-reactions if you can.
2. **Decide the kind of cell.** It makes electrical energy or happens by itself (battery, fuel cell, corrosion): voltaic, spontaneous. It needs electrical energy (plating, refining, making chlorine): electrolytic.
3. **Use the table to explain.** A metal protects iron if it is a stronger reducing agent than iron: lower on the right side of the table.
4. **Place the object.** In plating and refining, metal forms at the cathode. So the object to be coated is the cathode.
5. **For “should we use it?” questions:** name the intended result, then one unintended result (cost, waste, pollution, safety).

Worked example

A steel ship has blocks of zinc bolted to its hull, under the water. Explain how the zinc protects the steel.

> **Step 1**
> Without zinc, iron is oxidized: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$. Oxygen, with water, is reduced: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$.
>
> **Step 2**
> Corrosion happens by itself, so it is spontaneous: $E° = +0.40\ V - (-0.45\ V) = +0.85\ V$.
>
> **Step 3**
> Zn(s) (−0.76 V) is lower on the right than Fe(s) (−0.45 V). Zinc is the stronger reducing agent, so zinc is oxidized in place of iron.
>
> **Step 4**
> The zinc block is a sacrificial anode. The steel hull is the cathode, where oxygen is reduced. This is cathodic protection. The zinc is used up and must be replaced.

The common mistake

Picking a metal that is **above** iron in the table, such as tin or copper, as a sacrificial anode. It must be a stronger reducing agent than iron: below iron on the right.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

Which metal could be used as a sacrificial anode to protect an iron pipe? Use the table of standard electrode potentials in the Data Booklet.

A. Ag(s)

B. Cu(s)

C. Mg(s)

D. Pb(s)

QUESTION 2 · EASY · MULTIPLE CHOICE

A steel spoon is to be plated with silver. Which row describes the set-up?

A. The spoon is the anode, in a solution of $Ag^{+}(aq)$.

B. The spoon is the cathode, in a solution of $Ag^{+}(aq)$.

C. The spoon is the anode, in a solution of $Fe^{2+}(aq)$.

D. The spoon is the cathode, in a solution of $Fe^{2+}(aq)$.

QUESTION 3 · MEDIUM · NUMERICAL RESPONSE

Iron rusts in damp air. Iron is oxidized: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$. Oxygen is reduced: $O_2(g) + 2H_2O(l) + 4e^{-} \rightarrow 4OH^{-}(aq)$. Use the table of standard electrode potentials in the Data Booklet. What is E° for the reaction, in volts? Record your answer to two decimal places.

Answer: ____

QUESTION 4 · MEDIUM · MULTIPLE CHOICE

In a hydrogen fuel cell, hydrogen gas and oxygen gas are fed in and water comes out: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. Which statement is true?

A. $H_2(g)$ is the reducing agent. It is oxidized at the anode.

B. $H_2(g)$ is the oxidizing agent. It is reduced at the cathode.

C. $O_2(g)$ is the reducing agent. It is oxidized at the anode.

D. The cell needs a power source, because the reaction is not spontaneous.

QUESTION 5 · EXAM LEVEL · MULTIPLE CHOICE

One steel can is coated with tin. Another is coated with zinc (galvanized). Each coating is scratched so the iron shows.
At the scratch in the tin-coated can, the metal that corrodes is *i*. At the scratch in the zinc-coated can, the metal that corrodes is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	iron	zinc
B.	tin	iron
C.	iron	iron
D.	tin	zinc

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

Many small devices use batteries that are thrown away when they run down. Which of the following is an unintended result of this technology?

A. The battery changes chemical energy to electrical energy.

B. The device can be used away from a wall outlet.

C. Metals from the batteries can get into soil and water at a landfill.

D. A spontaneous redox reaction happens inside the battery.

CHALLENGE

A buried steel pipeline can be protected in two ways. Way 1: blocks of magnesium are joined to the pipe by wires. Way 2: a power source is joined to the pipe. (a) For way 1, write the half-reaction at the magnesium and the half-reaction at the pipe in damp soil, and find E°. (b) Why must the magnesium blocks be replaced? (c) For way 2, which terminal of the power source is joined to the pipe? Explain.

▼ **Hint**

To protect the iron, make it the cathode. Think about where electrons must go.

▼ **Check your answers for Card 11**

Question 1: C. Fe(s) is at −0.45 V. Mg(s) (−2.37 V) is below it, so magnesium is the stronger reducing agent. Ag, Cu and Pb (−0.13 V) are above iron.

Question 2: B. Silver forms by reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$. Reduction happens at the cathode, so the spoon is the cathode, joined to the negative terminal of the power source.

Question 3: 0.85. Oxygen is reduced (cathode) and iron is oxidized (anode). E° = +0.40 V − (−0.45 V) = +0.85 V. It is positive, so rusting is spontaneous. Net: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$.

Question 4: A. Hydrogen goes from 0 to +1, so it is oxidized, at the anode. Oxygen goes from 0 to −2, so it is reduced, at the cathode. The reaction is spontaneous: a fuel cell is a voltaic cell.

Question 5: A. The strongest reducing agent is oxidized. Sn(s) (−0.14 V) is above Fe(s) (−0.45 V), so iron is the stronger reducing agent and it corrodes. Zn(s) (−0.76 V) is below Fe(s), so zinc corrodes and the iron is protected.

Question 6: C. A, B and D are what the battery is meant to do. Metals leaking from used batteries is a result nobody intended. Recycling used batteries is one answer to it.

Challenge. (a) At the magnesium (anode): $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$. At the pipe (cathode): $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$. E° = +0.40 V − (−2.37 V) = +2.77 V. Net: $2Mg(s) + O_2(g) + 2H_2O(l) \rightarrow 2Mg^{2+}(aq) + 4OH^-(aq)$. (b) The magnesium is oxidized, so it is slowly used up. (c) The negative terminal. The supply pushes electrons onto the pipe. That makes the pipe the cathode, so the iron cannot be oxidized.

Card 11 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. Rusting needs two things besides iron. What are they?
2. In electroplating, is the object to be coated the anode or the cathode?
3. Is a fuel cell a voltaic cell or an electrolytic cell?

Need more? Go back to the lesson: Lesson B12 · Lesson B13 · Lesson B16. Videos for each idea: Videos by concept.

Card 12: Lab questions: variables, design, observations

I can name the variables in an experiment, tell an observation from an inference, judge a design, and explain why a measured voltage differs from the predicted one.

How to spot it on the diploma

- A description of an experiment, often with a table of results or a diagram of the set-up.
- “The manipulated variable is …”, “A controlled variable is …”
- “Which of the following is an observation?”
- “Which change would improve the design?” or “Why is the measured voltage lower than predicted?”

The steps

1. **Find the manipulated variable.** The one thing the experimenter changes on purpose.
2. **Find the responding variable.** What is measured to see the effect. It must be measured **directly**: a voltmeter reading, a mass, a volume, a colour. A calculated value, such as moles, is not it.
3. **Find the controlled variables.** Things that could change but are kept the same: concentration, temperature, the size of the electrodes, the time.
4. **Sort observations from inferences.** An observation is what you see or measure. An inference is what you work out from it.
5. **Compare with the prediction.** A measured voltage is usually a little lower than E°cell. Reasons: the solutions are not exactly 1.0 mol/L, the temperature is not 25 °C, the wires and salt bridge resist the current, and the cell has been running.
6. **Check the safety.** Read the WHMIS 2015 pictogram on each label before you start. A flame over a circle means an oxidizing material, such as potassium permanganate. Wear eye protection. Put waste where your teacher says, not down the sink.

Worked example

A student builds four voltaic cells. Each has the same copper half-cell and a different second half-cell: zinc, lead, nickel or magnesium. The student reads the voltmeter for each cell. Name the variables.

Step 1
Manipulated variable: the metal (and its ion) in the second half-cell.

Step 2
Responding variable: the voltmeter reading, in volts.

Step 3
Controlled variables: the copper half-cell, the concentration of every solution (1.0 mol/L), the temperature, the salt bridge.

Step 4
Observation: “the reading was 1.05 V”. Inference: “zinc is a stronger reducing agent than copper”.

The common mistake

Calling the name of a gas an observation. “Bubbles formed” is an observation. “Hydrogen gas formed” is an inference.

Your turn

Use the Data Booklet. Work on paper. Then check your answers.

QUESTION 1 · EASY · MULTIPLE CHOICE

A solution of potassium iodide is electrolyzed. Which of the following is an observation?

A. Iodine forms at the anode.

B. A brown colour appears near one electrode.

C. Hydroxide ions form at the cathode.

D. Water is reduced to hydrogen gas.

QUESTION 2 · EASY · MULTIPLE CHOICE

A student changes the current in a copper plating cell and wants to see how this affects the amount of copper plated in 10 min. What is the responding variable?

A. The current

B. The time

C. The mass of the cathode

D. The moles of electrons

QUESTION 3 · MEDIUM · MULTIPLE CHOICE

In the experiment in question 2, which of the following is a controlled variable?

A. The concentration of the $Cu^{2+}(aq)$ solution

B. The mass of copper plated

C. The current

D. The Faraday constant

QUESTION 4 · MEDIUM · MULTIPLE CHOICE

A student builds a zinc–copper voltaic cell and measures 1.02 V. The predicted value is 1.10 V. Which is the best explanation?

A. The table values are wrong.

B. The zinc electrode was the cathode.

C. The solutions were not at exactly 1.0 mol/L and 25 °C.

D. Electrons moved through the salt bridge.

QUESTION 5 · EXAM LEVEL · NUMERICAL RESPONSE

A student plates copper from $Cu^{2+}(aq)$ at a steady current and finds the mass of the cathode every 10 min.

Time (min)	0.0	10.0	20.0	30.0
Mass of copper plated (g)	0.000	0.200	0.400	0.600

Use the 20.0 min reading. What was the current, in amperes? Record your answer to three significant digits.

Answer: ____

QUESTION 6 · EXAM LEVEL · MULTIPLE CHOICE

A student wants to rank four metals as reducing agents. The student puts a strip of each metal into each of four ion solutions and looks for a reaction. Which change would most improve the design?

A. Use a different volume of each solution.

B. Use solutions of the same concentration, and clean each strip with sandpaper first.

C. Warm some of the test tubes.

D. Use a larger strip for the metals that react slowly.

CHALLENGE

Metal X is unknown. Design an experiment with voltaic cells to place X in a list of reducing agents with Cu(s), Zn(s) and Mg(s). Give (a) the manipulated variable, (b) the responding variable, (c) two controlled variables, (d) how you would use the results, and (e) one limit of the experiment.

▼ **Hint**

Keep one half-cell the same in every cell, and compare each metal with it.

▼ **Check your answers for Card 12**

Question 1: B. A colour is something you see. Iodine, hydroxide ions and hydrogen gas are worked out from what is seen, so they are inferences.

Question 2: C. The mass of the cathode is measured directly, before and after. The current is the manipulated variable. The time is controlled. Moles of electrons are calculated, not measured.

Question 3: A. The concentration could be changed, but it is kept the same. The mass of copper is the responding variable. The current is the manipulated variable. The Faraday constant cannot be changed, so it is not a variable.

Question 4: C. Table values are for standard conditions. A real cell is rarely at exactly 1.0 mol/L and 25 °C, and its wires and salt bridge resist the current. So the reading is a little different, usually lower.

Question 5: 1.01. n(Cu) = 0.400 g ÷ 63.55 g/mol = 0.006294 mol. $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$, so $n(e^-) = 2 \times$ that = 0.01259 mol. q = 0.01259 mol × 9.65 × 10^4 C/mol = 1215 C. t = 1200 s. I = q ÷ t = 1.01 A. The table also shows that mass is directly proportional to time.

Question 6: B. A fair test changes one thing only. Concentration is a controlled variable, so it should be the same in every test. Cleaning the strips removes any coating that could stop a reaction. A, C and D each add something that changes from test to test.

Challenge. One good answer: build a cell from a standard copper half-cell and a half-cell of each metal in turn (Zn, Mg, X) with a solution of its own ion. (a) The metal in the second half-cell. (b) The voltmeter reading, and which electrode is negative. (c) Any two of: the copper half-cell, the concentration of each solution (1.0 mol/L), the temperature, the salt bridge. (d) The metal that gives the largest voltage against copper, as the negative electrode, is the strongest reducing agent. Put X where its voltage fits. (e) Readings are a little off the table values because conditions are not exactly standard, so two metals with close values may be hard to tell apart.

Card 12 check

Answer these three on paper and show Mr F. They tell you both if you are ready to move on.

1. “The solution turned red litmus blue.” Observation or inference?
2. “The solution became basic.” Observation or inference?

3. A student tests how time affects the mass of silver plated. Name the manipulated variable and the responding variable.

Need more? Go back to the lesson: Lesson B14 · Lesson B15. Videos for each idea: Videos by concept.

AFTER CARD 12

Practice unit check

Twelve mixed questions, in sets, as on the diploma. Give yourself 25 minutes. Use only your Data Booklet and a calculator. If you get 9 or more, you are ready for the unit final.

> **Use this information for questions 1 to 4.** One kind of breath tester uses acidified potassium dichromate. Ethanol in the breath is changed to ethanoic acid, and the orange solution turns green. The skeleton equation is $Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + CH_3COOH(aq)$.

1. What is the change in the oxidation number of chromium?

A. From +6 to +3

B. From +7 to +3

C. From +12 to +6

D. From +3 to +6

2. What is the reducing agent?

A. $Cr_2O_7^{2-}(aq)$

B. $H^+(aq)$

C. $C_2H_5OH(aq)$

D. $Cr^{3+}(aq)$

3. When the equation is balanced for an acidic solution with the smallest whole numbers, the coefficient of $H^+(aq)$ is *ij* and the coefficient of $H_2O(l)$ is *kl*. Record the four digits.

Answer: ____

4. A 10.00 mL sample that contains ethanol is titrated with 0.0500 mol/L acidified $K_2Cr_2O_7(aq)$. The endpoint is reached at 12.60 mL. Use the balanced equation from question 3. What is the concentration of ethanol, in mmol/L? Record your answer to three significant digits.

Answer: ____

> **Use this information for questions 5 to 8.** A student builds this voltaic cell at standard conditions: $Cd(s) \mid Cd^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$.

5. Which row gives the anode and the electrode that electrons move toward?

A. Cd(s); Ag(s)

B. Cd(s); Cd(s)

C. Ag(s); Cd(s)

D. Ag(s); Ag(s)

---

6. What is the standard cell potential, in volts? Record your answer to two decimal places.

Answer: ☐

---

7. Which row describes what happens as the cell runs?

A. The cadmium electrode loses mass, and anions move toward the cadmium half-cell.

B. The cadmium electrode loses mass, and anions move toward the silver half-cell.

C. The cadmium electrode gains mass, and anions move toward the cadmium half-cell.

D. The silver electrode loses mass, and cations move toward the cadmium half-cell.

---

8. Suppose $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$ were chosen as the reference half-cell, at 0.00 V. Which row gives the new value for Ag^+/Ag and the cell potential?

A. +1.20 V; +1.20 V

B. +1.20 V; +0.80 V

C. +0.40 V; +0.40 V

D. +0.80 V; +1.20 V

> **Use this information for questions 9 to 12.** A steel car part is plated with chromium. The part and an inert electrode are put in a solution that contains $Cr^{3+}(aq)$, and they are joined to a power source. The half-reaction at the part is $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$.

---

9. Which row describes the steel part?

A. It is the cathode, joined to the negative terminal.

B. It is the cathode, joined to the positive terminal.

C. It is the anode, joined to the negative terminal.

D. It is the anode, joined to the positive terminal.

10. A current of 25.0 A runs for 30.0 min. What mass of chromium is plated, in grams? Record your answer to three significant digits.

Answer: [ ]

11. Water is oxidized at the inert electrode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$. Which row gives two observations near that electrode?

A. Bubbles form, and blue litmus turns red.

B. Bubbles form, and red litmus turns blue.

C. A metal coating forms, and blue litmus turns red.

D. Oxygen forms, and hydrogen ions form.

12. A student tests how the plating time affects the mass of chromium plated. Which of the following is a controlled variable?

A. The plating time

B. The mass of chromium plated

C. The current

D. The moles of electrons

▼ **Check your answers for the practice unit check**

Question 1: A. In $Cr_2O_7^{2-}$: $2x + 7(-2) = -2$, so $x = +6$. In Cr^{3+} it is +3.

Question 2: C. Chromium is reduced, so $Cr_2O_7^{2-}$ is the oxidizing agent. Ethanol is oxidized, so it is the reducing agent.

Question 3: 1611. Halves: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$, and $C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+ + 4e^-$. Multiply by 2 and by 3 (12 electrons). After cancelling: $2Cr_2O_7^{2-}(aq) + 3C_2H_5OH(aq) + 16H^+(aq) \rightarrow 4Cr^{3+}(aq) + 3CH_3COOH(aq) + 11H_2O(l)$.

Question 4: 94.5. $n(Cr_2O_7^{2-}) = 0.0500\ mol/L \times 0.01260\ L = 6.300 \times 10^{-4}$ mol. The ratio is 3 mol of ethanol to 2 mol of $Cr_2O_7^{2-}$. $n(C_2H_5OH)$ = that $\times 3/2 = 9.450 \times 10^{-4}$ mol. c = that ÷ 0.01000 L = 0.0945 mol/L = 94.5 mmol/L.

Question 5: A. The anode is on the far left: Cd(s). Electrons move through the wire to the cathode, Ag(s).

Question 6: 1.20. $E°cell = +0.80\ V - (-0.40\ V) = +1.20\ V$.

Question 7: A. Cd(s) is oxidized to $Cd^{2+}(aq)$, so the electrode loses mass. Anions move toward the anode, where positive ions are forming.

Question 8: A. Every value goes up by 0.40 V. Silver becomes +1.20 V. The gap between the two half-cells does not change, so E°cell is the same.

Question 9: A. Chromium forms by reduction, and reduction happens at the cathode. In an electrolytic cell the cathode is joined to the negative terminal.

Question 10: 8.08. t = 1800 s. q = 25.0 A × 1800 s = 4.50×10^4 C. n(e^-) = q ÷ 9.65×10^4 C/mol = 0.4663 mol. n(Cr) = that ÷ 3 = 0.1554 mol. m = that × 52.00 g/mol = 8.08 g.

Question 11: A. A gas shows as bubbles. H^+(aq) makes the solution acidic there, so blue litmus turns red. “Oxygen forms” and “hydrogen ions form” are inferences.

Question 12: C. The current could be changed but is kept the same. Time is manipulated. Mass is the responding variable. Moles of electrons are calculated.

AFTER THE PRACTICE UNIT CHECK

Exam-style question sets

8 sets, 28 questions, asked the way the diploma asks them. Each set starts with a box of information. Read the box first. Use only your Data Booklet and a calculator. Do one or two sets at a time. Each answer names the card to go back to.

Set 1

Use the following information to answer questions 1 to 3.
Household bleach is made by bubbling chlorine gas into a cold solution of sodium hydroxide. The reaction can be represented by this equation:
$Cl_2(g) + 2OH^-(aq) \rightarrow Cl^-(aq) + OCl^-(aq) + H_2O(l)$

Four statements	
I	Some chlorine atoms gain electrons.
II	Some chlorine atoms lose electrons.
III	The oxidation number of oxygen changes.
IV	The reaction is a disproportionation.

---

1. The statements above that are true are

A. I and III

B. II and III

C. I, II and IV

D. I, II, III and IV

---

2. In $OCl^-(aq)$, the oxidation number of chlorine is *i*. In this reaction, $OH^-(aq)$ is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	+1	neither oxidized nor reduced
B.	−1	neither oxidized nor reduced
C.	+1	the reducing agent
D.	−1	the oxidizing agent

3. The oxidation number of chlorine in
$Cl_2(g)$ is +/− ___ (record in the first column)
$Cl^-(aq)$ is +/− ___ (record in the second column)
$OCl^-(aq)$ is +/− ___ (record in the third column)
$ClO_3^-(aq)$ is +/− ___ (record in the fourth column)

Answer:

Set 2

Use the following information to answer questions 4 to 6.
A student compared some metals that are used in electronics. The student put a strip of a metal into a solution of a metal ion and recorded these observations.

Mixed	Observation
$Tl^+(aq)$ and Ga(s)	evidence of a reaction
$Pd^{2+}(aq)$ and Tl(s)	evidence of a reaction
$Ga^{3+}(aq)$ and Sc(s)	evidence of a reaction
$Sc^{3+}(aq)$ and Tl(s)	no evidence of a reaction

Chemical species			
1 $Ga^{3+}(aq)$	2 $Pd^{2+}(aq)$	3 $Sc^{3+}(aq)$	4 $Tl^+(aq)$

4. The oxidizing agents, listed from strongest to weakest, are numbered ___, ___, ___ and ___. Record all four digits.

Answer:

5. Reaction types: 1 is spontaneous, 2 is nonspontaneous. Match a number to each pair of reactants.
$Pd^{2+}(aq)$ and Sc(s): ___ (first column)
$Sc^{3+}(aq)$ and Ga(s): ___ (second column)
$Tl^+(aq)$ and Sc(s): ___ (third column)
$Ga^{3+}(aq)$ and Pd(s): ___ (fourth column)

Answer:

6. The strongest reducing agent in the student's experiment is

A. Pd(s)

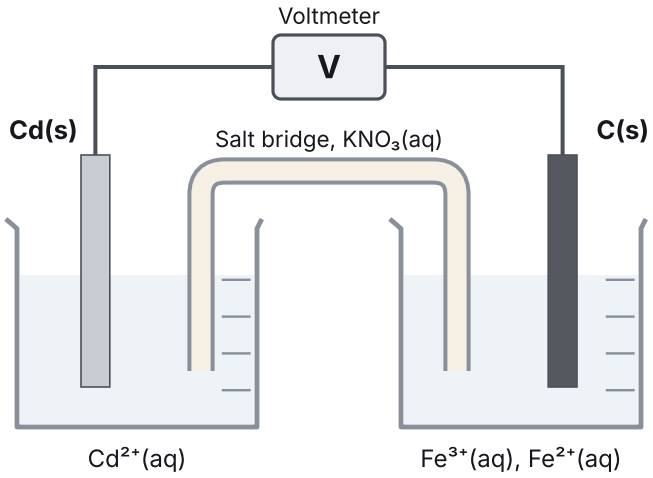
B. Tl(s)

C. Ga(s)

D. Sc(s)

Set 3

Use the following information to answer questions 7 to 10.
A student builds the electrochemical cell in the diagram at standard conditions.



7. In this cell, electrons move through the wire *i*, and anions in the salt bridge move *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	from the Cd(s) electrode to the C(s) electrode	toward the cadmium half-cell
B.	from the Cd(s) electrode to the C(s) electrode	toward the iron half-cell
C.	from the C(s) electrode to the Cd(s) electrode	toward the cadmium half-cell
D.	from the C(s) electrode to the Cd(s) electrode	toward the iron half-cell

8. The cell potential for the electrochemical cell in the diagram is +/− ___ V. Record your three-digit answer.

Answer: [ ]

9. Eight statements about the cathode:

1 It is the carbon rod.
2 It is the cadmium strip.
3 Electrons move toward it through the wire.
4 Electrons move away from it through the wire.
5 Cations move toward it.
6 Anions move toward it.
7 $Fe^{3+}(aq)$ is reduced at it.
8 Its mass goes down.

The statements that describe the cathode in this cell are numbered ___, ___, ___ and ___. Record all four digits in any order.

Answer: ☐

---

10. The net ionic equation for the reaction in the cell in the diagram is

A. $Cd(s) + Fe^{3+}(aq) \rightarrow Cd^{2+}(aq) + Fe^{2+}(aq)$

B. $Cd^{2+}(aq) + 2Fe^{2+}(aq) \rightarrow Cd(s) + 2Fe^{3+}(aq)$

C. $Cd(s) + 2Fe^{3+}(aq) \rightarrow Cd^{2+}(aq) + 2Fe^{2+}(aq)$

D. $3Cd(s) + 2Fe^{3+}(aq) \rightarrow 3Cd^{2+}(aq) + 2Fe(s)$

Set 4

> **Use the following information to answer questions 11 to 14.**
> Hydrogen gas for fuel cells can be made by the electrolysis of a solution of potassium hydroxide, KOH(aq). Nickel electrodes are used. In this cell they act as inert electrodes.

---

11. The half-reaction at the anode is

A. $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

B. $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

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

D. $K^+(aq) + e^- \rightarrow K(s)$

---

12. The minimum voltage needed for this cell is +/− ___ V. Record your three-digit answer.

Answer: ☐

13. A current of 2.00×10^3 A runs for 1.00 h. The mass of hydrogen gas produced, in scientific notation, is $a.bc \times 10^d$ g. The values of a, b, c and d are ___, ___, ___ and ___. Record all four digits.

Answer: [ ]

14. This cell is *i* cell, and its E°cell is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	a voltaic	positive
B.	a voltaic	negative
C.	an electrolytic	positive
D.	an electrolytic	negative

Set 5

Use the following information to answer questions 15 to 17.
The vitamin C in fruit juice is ascorbic acid, $C_6H_8O_6(aq)$. Its concentration can be found by titrating the juice with a solution of iodine. The reaction can be represented by this equation:
$C_6H_8O_6(aq) + I_2(aq) \rightarrow C_6H_6O_6(aq) + 2H^+(aq) + 2I^-(aq)$
A student titrated 25.00 mL samples of a juice with 0.00500 mol/L $I_2(aq)$.

Trial	I	II	III
Final burette reading (mL)	13.10	25.25	37.50
Initial burette reading (mL)	0.50	13.10	25.25

15. The volume of $I_2(aq)$ that the student should use in the calculation is

A. 12.15 mL
B. 12.20 mL
C. 12.33 mL
D. 12.60 mL

16. The concentration of ascorbic acid in the juice, in scientific notation, is $a.bc \times 10^{-d}$ mol/L. The values of a, b, c and d are ___, ___, ___ and ___. Record all four digits.

Answer: [ ]

17. In this titration, ascorbic acid is the *i* agent, and iodine is *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	reducing	reduced
B.	reducing	oxidized
C.	oxidizing	reduced
D.	oxidizing	oxidized

Set 6

> **Use the following information to answer questions 18 to 21.**
> A watch is run by a small silver oxide "button" cell. Inside it, zinc powder and silver oxide are separated by a moist paste of potassium hydroxide. The half-reactions in the cell are given below.
> $Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$
> $Ag_2O(s) + H_2O(l) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$ $E° = +0.34$ V
> The cell potential, E°cell, is +1.59 V.

18. In this cell, the substance reduced at the cathode is *i*, and electrons move through the watch's circuit *ii*.

The statement above is completed by the information in row

Row	*i*	*ii*
A.	$Ag_2O(s)$	from the zinc to the silver oxide
B.	$Ag_2O(s)$	from the silver oxide to the zinc
C.	$Zn(s)$	from the zinc to the silver oxide
D.	$Zn(s)$	from the silver oxide to the zinc

19. The electrical potential, E°, for the half-reaction $ZnO(s) + H_2O(l) + 2e^- \rightarrow Zn(s) + 2OH^-(aq)$ is +/− ___ V. Record your three-digit answer.

Answer: [ ]

20. The watch draws a steady current of 1.50×10^{-5} A for 30.0 days. The mass of zinc that reacts, in scientific notation, is $a.bc \times 10^{-d}$ g. The values of a, b, c and d are ___, ___, ___ and ___. Record all four digits.

Answer: [ ]

21. Eight statements about electrochemical cells:

1 Reduction happens at the cathode.
2 The reaction is spontaneous.
3 A power source is needed.
4 Electrons move through the wire from the anode to the cathode.
5 The cathode is the negative electrode.
6 Anions move toward the anode.
7 Oxidation happens at the anode.
8 E°cell is positive.

The statements that are true for **both** a voltaic cell and an electrolytic cell are numbered ___, ___, ___ and ___. Record all four digits in any order.

Answer: [ ]

Set 7

Use the following information to answer questions 22 to 25.
A student is given these reduction half-reactions. They are not in the Data Booklet.

Reduction half-reaction	E° (V)
$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$	+1.61
$In^{3+}(aq) + 3e^- \rightarrow In(s)$	−0.34
$Ti^{2+}(aq) + 2e^- \rightarrow Ti(s)$	−1.63

The student also has these materials:

1 Cu(s)	2 Mg(s)	3 Ni(s)	4 Ag(s)
5 $Cu(NO_3)_2(aq)$	6 $Mg(NO_3)_2(aq)$	7 $Ni(NO_3)_2(aq)$	8 $AgNO_3(aq)$

22. In the student's data, the strongest reducing agent is

A. $Ce^{4+}(aq)$
B. $Ce^{3+}(aq)$
C. In(s)
D. Ti(s)

23. Use the given half-reactions with the Data Booklet table. Which metal will change $Ce^{4+}(aq)$ to $Ce^{3+}(aq)$ but will **not** change $In^{3+}(aq)$ to In(s)?

A. Zn(s)

B. Al(s)

C. Cu(s)

D. Mg(s)

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24. The Data Booklet table has lines for two solid compounds: $AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$ at +0.07 V, and $PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$ at −0.36 V. Which metals will react spontaneously with AgBr(s) but not with $PbSO_4(s)$?

A. Cu(s) and Hg(l)

B. Pb(s) and Ni(s)

C. Zn(s) and Fe(s)

D. Cd(s) and Al(s)

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25. The student must build a voltaic cell that uses a nickel electrode and gives at least 2.00 V. Electrons must move from electrode A to electrode B. Use the numbered materials.
Electrode A is numbered ___ (first column)
Electrode B is numbered ___ (second column)
The solution around electrode A is numbered ___ (third column)
The solution around electrode B is numbered ___ (fourth column)

Answer: [ ]

Set 8

> **Use the following information to answer questions 26 to 28.**
> Some contact lens cleaners contain hydrogen peroxide, $H_2O_2(aq)$. A technician checks its concentration by titrating an acidified sample with potassium permanganate solution. The reaction can be represented by this unbalanced equation:
> $MnO_4^-(aq) + H_2O_2(aq) + H^+(aq) \rightarrow Mn^{2+}(aq) + O_2(g) + H_2O(l)$
> A 10.0 mL sample needs 17.6 mL of 0.0200 mol/L $KMnO_4(aq)$ to reach the endpoint.

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26. When the equation is balanced with the lowest whole numbers, the coefficient of
$MnO_4^-(aq)$ is ___ (first column)
$H_2O_2(aq)$ is ___ (second column)
$H^+(aq)$ is ___ (third column)
$H_2O(l)$ is ___ (fourth column)

Answer: [ ]

27. During the titration, which of the following is an observation the technician could make?

A. Oxygen gas is produced.

B. Bubbles form in the flask.

C. Permanganate ions are reduced.

D. The solution becomes more acidic.

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28. The concentration of $H_2O_2(aq)$ in the sample is

A. 14.1 mmol/L

B. 35.2 mmol/L

C. 88.0 mmol/L

D. 176 mmol/L

▼ **Check your answers for the exam-style question sets**

Set 1

Question 1: C. Chlorine is 0 in Cl_2, −1 in Cl^- and +1 in OCl^-. Some atoms gain electrons (I) and some lose electrons (II), so the same particle is both reduced and oxidized (IV). Oxygen stays −2, so III is false. *Go back to Card 1.*

Question 2: A. In OCl^-: (−2) + x = −1, so x = +1. In OH^-, oxygen is −2 and hydrogen is +1 before and after, so OH^- is neither oxidized nor reduced. *Go back to Card 1.*

Question 3: 0115. Cl_2: 0. Cl^-: −1, so record 1. OCl^-: +1, so record 1. ClO_3^-: x + 3(−2) = −1, so x = +5, record 5. You record the sizes only. *Go back to Card 1.*

Set 2

Question 4: 2413. Each reaction puts the ion above the metal. Tl^+ is above Ga. Pd^{2+} is above Tl. Ga^{3+} is above Sc. So from the top: Pd^{2+} (2), Tl^+ (4), Ga^{3+} (1), Sc^{3+} (3). The last observation agrees: Sc^{3+} is below Tl. *Go back to Card 4.*

Question 5: 1212. An ion reacts with a metal below it. Order from the top: Pd^{2+}, Tl^+, Ga^{3+}, Sc^{3+}. Pd^{2+} is above Sc: 1. Sc^{3+} is below Ga: 2. Tl^+ is above Sc: 1. Ga^{3+} is below Pd: 2. *Go back to Card 5.*

Question 6: D. The strongest reducing agent is the metal of the weakest oxidizing agent. Sc^{3+} is at the bottom, so Sc(s) is the strongest reducing agent. *Go back to Card 4.*

Set 3

Question 7: A. Fe^{3+}/Fe^{2+} (+0.77 V) is higher than Cd^{2+}/Cd (−0.40 V). The carbon electrode is the cathode and cadmium is the anode. Electrons move from the anode to the cathode. Anions move toward the anode. *Go back to Card 7.*

Question 8: 1.17. E°cell = +0.77 V − (−0.40 V) = +1.17 V. Net: $Cd(s) + 2Fe^{3+}(aq) \rightarrow Cd^{2+}(aq) + 2Fe^{2+}(aq)$. *Go back to Card 8.*

Question 9: 1357. The cathode is the carbon rod (1). Electrons arrive at it (3). Cations move toward it (5). $Fe^{3+}(aq)$ is reduced there (7). Its mass does not change, because carbon is inert and both iron ions stay in the solution. *Go back to Card 7.*

Question 10: C. Anode: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$. Cathode: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$, used twice so the electrons are equal. A is not balanced for charge. B is the reverse, which is nonspontaneous. In D, Fe^{3+} would have to become Fe(s), but the half-cell holds Fe^{3+} and Fe^{2+}. *Go back to Card 7.*

Set 4

Question 11: A. In the cell: K^+, OH^-, H_2O. Reducing agents: OH^- (+0.40 V) and H_2O (+1.23 V). OH^- is lower on the right, so it is the strongest reducing agent. It is oxidized at the anode. C and D are reductions. *Go back to Card 9.*

Question 12: 1.23. Cathode: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$, −0.83 V. E°cell = −0.83 V − (+0.40 V) = −1.23 V. The minimum voltage is 1.23 V. *Go back to Card 9.*

Question 13: 7541. t = 3600 s. q = 2.00×10^3 A × 3600 s = 7.20×10^6 C. n(e^-) = q ÷ 9.65×10^4 C/mol = 74.61 mol. 2 mol of electrons make 1 mol of H_2, so n(H_2) = 37.31 mol. m = that × 2.02 g/mol = 75.4 g = 7.54×10^1 g. *Go back to Card 10.*

Question 14: D. The cell needs electrical energy, so it is electrolytic. Its reaction is nonspontaneous, so E°cell is negative. *Go back to Card 9.*

Set 5

Question 15: B. Volume added = final − initial: 12.60 mL, 12.15 mL, 12.25 mL. Trial I does not agree with the others, so leave it out. Average of trials II and III: 12.20 mL. *Go back to Card 6.*

Question 16: 2443. n(I_2) = 0.00500 mol/L × 0.01220 L = 6.100×10^{-5} mol. The ratio is 1 to 1, so n($C_6H_8O_6$) = 6.100×10^{-5} mol. c = that ÷ 0.02500 L = 2.44×10^{-3} mol/L. *Go back to Card 6.*

Question 17: A. Iodine goes from 0 in I_2 to −1 in I^-. It gains electrons, so it is reduced. Ascorbic acid gives the electrons, so it is the reducing agent. *Go back to Card 1.*

Set 6

Question 18: A. The half-reaction with electrons on the left is the reduction: $Ag_2O(s)$ is reduced, at the cathode. Zn(s) is oxidized, at the anode. Electrons move from the anode to the cathode. *Go back to Card 7.*

Question 19: 1.25. E°cell = E°cathode − E°anode, so +1.59 V = +0.34 V − E°(zinc). E°(zinc) = +0.34 V − 1.59 V = −1.25 V. Record 1.25. *Go back to Card 8.*

Question 20: 1322. t = 30.0 days × 24 h/day × 3600 s/h = 2.59×10^6 s. q = 1.50×10^{-5} A × 2.59×10^6 s = 38.88 C. n(e^-) = q ÷ 9.65×10^4 C/mol = 4.029×10^{-4} mol. 2 mol of electrons come from 1 mol of zinc, so n(Zn) = 2.015×10^{-4} mol. m = that × 65.41 g/mol = 1.32×10^{-2} g. *Go back to Card 10.*

Question 21: 1467. Statements 1, 4, 6 and 7 are true in every cell. Statements 2 and 8 are true for a voltaic cell only. Statements 3 and 5 are true for an electrolytic cell only. *Go back to Card 9.*

Set 7

Question 22: D. Reducing agents are on the right side of each half-reaction: $Ce^{3+}(aq)$, In(s), Ti(s). The strongest is on the lowest line, the one with the lowest E°: Ti(s). *Go back to Card 5.*

Question 23: C. The metal must be below Ce^{4+} (+1.61 V) and above In^{3+} (−0.34 V). Cu(s) is at +0.34 V, between them. Zn(s) (−0.76 V), Al(s) and Mg(s) are below In^{3+}, so they would change both. *Go back to Card 5*.

Question 24: B. The metal must be below AgBr (+0.07 V) and above $PbSO_4$ (−0.36 V). Pb(s) (−0.13 V) and Ni(s) (−0.26 V) are between them. Cu and Hg are above AgBr, so they react with neither. Zn, Fe (−0.45 V), Cd (−0.40 V) and Al are below $PbSO_4$, so they react with both. *Go back to Card 5*.

Question 25: 2367. Electrons leave the anode, so electrode A is the anode. With nickel (−0.26 V): magnesium and nickel give +2.11 V; nickel and copper give +0.60 V; nickel and silver give +1.06 V. Only magnesium with nickel reaches 2.00 V. Magnesium is the anode (A, number 2) in $Mg(NO_3)_2$(aq) (6). Nickel is the cathode (B, number 3) in $Ni(NO_3)_2$(aq) (7). *Go back to Card 8*.

Set 8

Question 26: 2568. Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Oxidation: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$. Multiply by 2 and by 5 (10 electrons), add, and cancel $10H^+$: $2MnO_4^-(aq) + 5H_2O_2(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(l)$. *Go back to Card 3*.

Question 27: B. Bubbles are something you see. “Oxygen gas is produced” and “permanganate ions are reduced” are inferences. D is not true: H^+(aq) is used up, so the solution becomes less acidic. *Go back to Card 12*.

Question 28: C. $n(MnO_4^-) = 0.0200\ mol/L \times 0.0176\ L = 3.52 \times 10^{-4}$ mol. The balanced equation gives 5 mol of H_2O_2 for 2 mol of MnO_4^-. $n(H_2O_2)$ = that × 5/2 = 8.80×10^{-4} mol. c = that ÷ 0.0100 L = 0.0880 mol/L = 88.0 mmol/L. A uses the ratio upside down. B leaves the ratio out. *Go back to Card 6*.