## Chapter Outline: The d- and f-Block Elements

The chapter covers positions, electronic configurations, and properties of d-block (transition) and f-block (inner transition) elements, with emphasis on 3d series trends, key compounds like $K_2Cr_2O_7$ and $KMnO_4$, and comparisons between lanthanoids and actinoids. [1]

- **Introduction and Position**: d-block (groups 3-12, 3d/4d/5d/6d series); f-block (4f lanthanoids Ce-Lu, 5f actinoids Th-Lr). Transition metals have incomplete d subshells in atoms or ions. [1]
- **Electronic Configurations**: General $(n-1)d^{1-10}\ ns^{1-2}$; exceptions like Cr ($3d^5 4s^1$), Cu ($3d^{10} 4s^1$) due to stability of half/full d shells. Zn/Cd/Hg excluded as $d^{10}$. [1]
- **General Properties (3d focus)**: High melting points/enthalpies from d-electron bonding; decreasing atomic/ionic radii; variable oxidation states (max near middle, e.g., Mn +2 to +7); electrode potentials; paramagnetism ($\mu = \sqrt{[n(n+2)]}$ BM); colored ions (d-d transitions); complexes, catalysis, interstitial compounds, alloys. [1]
- **Key Compounds**: Preparation/properties of $K_2Cr_2O_7$ (from chromite, acidic oxidant) and $KMnO_4$ (from $MnO_2$, strong oxidant in acid). [1]
- **f-Block**: Lanthanoids (mostly +3, contraction); actinoids (variable states, more reactive). [1]

## First Topic: Position in the Periodic Table

d-block elements occupy the middle of the periodic table (groups 3-12), where d-orbitals of the penultimate shell fill progressively across four series: 3d (Sc-Zn), 4d (Y-Cd), 5d (Hf-Hg, La separate), 6d (Rf-Cn, Ac separate). f-block elements sit at the bottom, with 4f (lanthanoids) and 5f (actinoids) filling. [1]

**Visual Analogy**: Picture the periodic table as a city—s-block on the left (residential outskirts), p-block on the right (commercial district), d-block as the bustling central highway (groups 3-12) with traffic (electrons) flowing into d-orbital "lanes," and f-block as underground metro lines (hidden deeper shells). Transition metals bridge s/p properties, like a highway connecting suburbs to downtown. [1]

**First Principles**: Electrons fill orbitals by energy order; d-orbitals (higher energy) fill after s, creating a "transitional" zone with partly filled d subshells, enabling unique bonding. Scandium ($3d^1$) qualifies as transition; zinc ($3d^{10} 4s^2$) does not, as full d shell lacks variability. [1]

**Exam Writing Tip**: Start with definition: "d-block: groups 3-12, $(n-1)d^{1-10}\ ns^{1-2}$." List series (e.g., 3d: Sc to Zn). Draw table excerpt. Mention IUPAC: incomplete d in atom/ion. Use 1 mark for exceptions (Zn not transition). Keep neat, underline key terms. [1]

Ready for the next topic (electronic configurations)? Confirm understanding: Why is Cr's config $3d^5 4s^1$, not $3d^4 4s^2$? Reply to proceed step-by-step.

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# I'm so sorry but use first principle mode more. I don't understand. This is complex for me. I think my basics are not clear.

It feels complex because the book is starting from “middle level” ideas, not from zero. Let’s rebuild from the ground, very slowly, using first principles and simple pictures.[2]

### Step 0: What do you already know?

Before d- and f-block, three tiny ideas are needed:

1. Atoms have energy levels (shells): $n = 1, 2, 3, \ldots$
2. Each shell has orbitals: s, p, d, f
   - s holds 2 electrons
   - p holds 6
   - d holds 10
   - f holds 14
3. Electrons fill lower energy first (1s, 2s, 2p, 3s, 3p, 4s, then 3d, etc.).

If any of these three is shaky, say which one, and that will be the next mini-lesson.

### Step 1: First-principle picture of “blocks”

Think from scratch, not from definitions.

- The periodic table is nothing but a big seating plan for electrons.
- Every **column/region** is decided by **which orbital is being filled last**.

So:

- If the last electrons go into an **s** orbital → **s-block** (left side).[2]
- If the last electrons go into a **p** orbital → **p-block** (right side).[2]
- If the last electrons go into a **d** orbital → **d-block** (middle).[2]
- If the last electrons go into an **f** orbital → **f-block** (bottom strip).[2]

That’s all “block” means:
**Block = ‘which orbital is being filled in this region of the table.’**[2]

### Step 2: Visual/story analogy

Imagine a 4-floor parking building:

- Floor 1 = n = 1
- Floor 2 = n = 2
- Floor 3 = n = 3
- Floor 4 = n = 4

On each floor, there are different types of parking spots:

- s-spots: 2 slots
- p-spots: 6 slots
- d-spots: 10 slots
- f-spots: 14 slots

Cars (electrons) always choose **lower floors and easier spots first**, then move to more complicated spots.

For 3rd and 4th floor:

- 4s spots get filled **before** 3d spots (4s is slightly lower in energy than 3d, even though 4 > 3).[2]
- So for elements from **Sc to Zn**, the "new" electrons are going into **3d spots** while 4s is already mostly filled.[2]

So the **3d parking row** filling corresponds to **the 3d series (Sc to Zn)** in the **d-block**.[2]

### Step 3: Now define d-block from first principles

Start from what is actually happening with electrons:

- In certain elements, the **(n−1)d orbitals** are being filled (3d, 4d, 5d, 6d) while the ns is already occupied.[2]

Therefore:

- **d-block elements** are those where **d orbitals of the previous shell are getting filled**.
- In the periodic table, that region is **groups 3–12**, the big middle block.[2]

**Transition metal (IUPAC idea)**:
A transition element is a metal whose atom or common ion has an **incomplete d subshell** (i.e., not $d^{10}$).[2]

- Sc: [Ar] $3d^1 4s^2$ → d is incomplete → transition.[2]
- Zn: [Ar] $3d^{10} 4s^2$ → d is full in atom and in $Zn^{2+}$ ($3d^{10}$) → not a transition element (by definition).[2]

### Step 4: First-principle picture of f-block

Same game, one level deeper:

- After d orbitals, at higher atomic numbers, **f orbitals** begin to fill: 4f (lanthanoids), 5f (actinoids).[2]
- These electrons are going into **(n−2)f** orbitals (for 6s, 4f is filling; for 7s, 5f is filling).[2]

So:

- **f-block elements** are those where **4f or 5f orbitals are being filled**.[2]
- They are written in the **separate bottom two rows** (lanthanoids and actinoids).[2]

### Step 5: How to write this in the exam (very simple form)

For a 2–3 mark "definition + position" question:

- Start with 1 clear sentence.
- Add 1–2 support lines.
- Optionally, add one example or a small sketch.

Example answer (you can copy the structure, not the exact wording):

- "d-block elements are those in which the (n−1)d orbitals are being filled and they occupy groups 3–12 in the middle of the periodic table."[2]
- "Their general outer configuration is $(n-1)d^{1-10}ns^{1-2}$ and many of them are transition metals with incomplete d subshells."[2]

For f-block:

- "f-block elements are those in which 4f or 5f orbitals are progressively filled; they are shown as two separate rows at the bottom and are called lanthanoids (4f) and actinoids (5f)."[2]

### What we should do next

To keep it really clear, next step should be:

1. Very slowly build **electronic configurations** of **Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn**, from first principles (parking-lot picture).[2]
2. Then show how this leads to the idea of "transition."[2]

Reply with which of these feels confusing right now:

- A) Shells and orbitals (s, p, d, f)
- B) Why 4s fills before 3d
- C) How to write configs like Sc: [Ar] $3d^1 4s^2$

and the teaching will continue from that exact point, with more stories, no rush.

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# Tell me more about electrons feeling lower energy. What is n-1 d orbital? What is the meaning of an S? I don't have much knowledge of the basics. Sc: [Ar] $3d^1 4s^2$ What is the meaning of this? I don't understand man, it is complex for me.

This is normal to find confusing; the textbook is jumping too fast. Let's build everything from the very beginning, in tiny steps, with pictures in mind.[3]

### 1. What does "lower energy" mean for electrons?

Think of an atom like a building and electrons like students choosing rooms:

- Ground floor rooms are safest and most comfortable → **lowest energy**.
- Higher floors are less comfortable → **higher energy**.

Electrons "prefer" to sit in the **lowest energy rooms available**.
In atoms, these "floors" are called **shells**:

- Shell 1 → $n = 1$
- Shell 2 → $n = 2$
- Shell 3 → $n = 3$
- Shell 4 → $n = 4$, etc.

Each shell has different types of rooms: **s, p, d, f**. These are called **subshells/orbitals**.[3]

Energy rule (rough idea):

- 1s is lowest
- Then 2s, 2p
- Then 3s, 3p
- Then 4s
- Then 3d
- Then 4p
  ... and so on (this is the order you see in your earlier classes).

So when we say "electrons fill lower energy first", it just means:

> They go to 1s, then 2s, then 2p, etc., following that order.

### 2. What does "s", "p", "d", "f" mean?

They are just **types of rooms** (shapes) where electrons can stay.

- **s**: 1 s-orbital per shell; can hold **2 electrons**.
- **p**: 3 p-orbitals per shell; total **6 electrons**.
- **d**: 5 d-orbitals; total **10 electrons**.
- **f**: 7 f-orbitals; total **14 electrons**.[3]

You do not need the exact shapes now. For d-block and f-block, only these facts matter:

- d-sublevel can take **10** electrons.
- f-sublevel can take **14** electrons.

### 3. What is "$n-1$ d orbital"?

Take an example: 4s and 3d.

- For 4s, $n = 4$.
- For 3d, the principal quantum number is 3, which is $4 - 1$.
  So 3d is written as $(n-1)d$ **when we are talking about 4s and 3d together**.[3]

So:

- In period 4 (K to Kr), the **s** part is **4s**, and the **d** part is **3d**.
- That's why for 3d series elements the general valence configuration is $(n-1)d^{1-10}ns^{1-2}$.[3]

Example: for period 4 d-block (Sc to Zn):

- $n = 4$, so $(n-1)d = 3d$, and $ns = 4s$.[3]

### 4. What does "Sc: [Ar] $3d^1 4s^2$" actually mean?

Scandium (Sc) has atomic number 21 → 21 electrons.

Writing all electrons from 1s onwards is long, so chemists use a shortcut:

- **[Ar]** means "the exact electronic configuration of Argon"
  (which already has 18 electrons).[3]

So:

- Sc has 21 electrons.
- Put the first 18 electrons exactly like Argon → write **[Ar]**.
- Now 3 electrons are left (21 − 18 = 3).
- These 3 go into the next available orbitals: first 4s, then 3d.[3]

Filling order around there:

- After [Ar], next is 4s, then 3d. So:
  - 2 electrons go into 4s → $4s^2$.
  - 1 electron goes into 3d → $3d^1$.

So:

- **Sc: [Ar] $3d^1$ $4s^2$** means:
  - “Start with Argon’s configuration”
  - “Then add 1 electron in 3d and 2 electrons in 4s.”

This is just a compact way to say where the 21 electrons sit.

## 5. Putting this together for d-block

Now that “$3d^1 4s^2$” makes sense:

- In Sc, the **d** subshell (3d) has **1 electron** and can go up to 10.
- As we go from Sc to Zn, we keep adding electrons into **3d**. [3]

For the 3d series (Sc to Zn):

- Sc: [Ar] $3d^1 4s^2$
- Ti: [Ar] $3d^2 4s^2$
- V: [Ar] $3d^3 4s^2$
- Cr: [Ar] $3d^5 4s^1$ (special, half-filled d is extra stable) [3]
- ... and so on, until Zn: [Ar] $3d^{10} 4s^2$. [3]

Because here the **(n−1)d = 3d** orbitals are being filled step by step, these elements lie in the **d-block**. [3]

## 6. Quick exam-style sentences (very basic level)

When you have to write in the exam:

- “s, p, d, f are subshells (types of orbitals) within shells; s holds 2 electrons, p 6, d 10, f 14.” [3]
- “In Sc: [Ar] $3d^1 4s^2$, [Ar] represents the electron configuration of Argon (18 electrons), and the remaining 3 electrons occupy the 4s and 3d orbitals as $4s^2 3d^1$.” [3]
- “For period 4 d-block elements, the general outer configuration is $3d^{1-10}$ $4s^{1-2}$, which can be written as $(n-1)d^{1-10}ns^{1-2}$ with $n = 4$.” [3]

**What you want next?**

Choose one, and the next reply will only focus on that with drawings-in-words:

- A) Step-by-step building of electron configuration from H (1) up to Ar (18).
- B) Only practice on reading/writing configurations like Sc, Ti, V, Cr, Mn.
- C) A very slow explanation of why 4s fills before 3d (the “energy” order).

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# What is this valency configuration thing? These 4s and 3s thing is very confusing to me. I think you should give me few more examples. Write a prompt to make a simulation or animation to help me understand this.

Valence configuration just means: “Where are the outer (reacting) electrons sitting?” Those are usually the electrons in the **last s and p (and for d-block, also d)** subshells.[4]

- Example: Sc: [Ar] $3d^1\ 4s^2$
  Valence/outer configuration = $3d^1\ 4s^2$ (these are the electrons that can take part in bonding).[4]

A few more simple examples (outer/valence part only):

- Na (Z = 11): $1s^2\ 2s^2\ 2p^6\ 3s^1$ → valence = $3s^1$
- Mg (Z = 12): … $3s^2$ → valence = $3s^2$
- Al (Z = 13): … $3s^2\ 3p^1$ → valence = $3s^2\ 3p^1$
- Ca (Z = 20): … $4s^2$ → valence = $4s^2$
- Sc (Z = 21): … $3d^1\ 4s^2$ → valence = $3d^1\ 4s^2$[4]

These outer electrons (valence) decide the **valency** and many chemical properties.

### Prompt for a simulation/animation (you can give this to an animator or coding tool)

Use this exact text (or paste into an AI video/animation tool):

Create a 3D animation to explain shells, subshells (s, p, d, f), and valence configuration for a beginner Class 11–12 student.

1. Visual metaphor:
    - Show an atom as a small glowing nucleus at the center.
    - Around it, show “floors” or circular layers labeled n = 1, 2, 3, 4 (shells).
    - On each floor, show different types of “rooms”:
        - s room: a single spherical room that can hold 2 electrons.
        - p rooms: three dumbbell-shaped rooms that together can hold 6 electrons.
        - d rooms: five petal-like rooms that together can hold 10 electrons.
    - Represent each electron as a small colored ball.
2. Explain filling order (energy):
    - Animate electrons entering orbitals in this order:
      1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p.
    - Show a simple energy ladder on one side, with 1s at the bottom and 3d slightly above 4s, to explain why 4s fills before 3d.
3. Build a few atoms step by step:
    - Hydrogen (Z=1): place 1 electron in 1s.
    - Helium (Z=2): place 2 electrons in 1s.
    - Sodium (Z=11): fill up to Ne, then show 1 electron in 3s (highlight valence electron).
    - Calcium (Z=20): fill up to Ar, then show 2 electrons in 4s (valence = $4s^2$).
    - Scandium (Z=21): start from Argon core [Ar], then add 2 electrons to 4s and 1 electron to 3d ($3d^1$ $4s^2$). Clearly label “valence configuration = $3d^1$ $4s^2$”.
4. Highlight valence configuration:
    - After building each atom, fade the inner electrons and brightly highlight only the outermost subshell(s) (e.g., $3s^1$ for Na, $3s^2 3p^1$ for Al, $3d^1 4s^2$ for Sc).
    - Add text: “Valence electrons = electrons in the outermost s, p (and for transition metals, also d) orbitals.”
5. Interactivity (if possible):
    - Allow the user to change atomic number with a slider from 1 to 30.
    - For each step, automatically:
        - Place electrons into orbitals following the correct order.
        - Show the full configuration and separately highlight the valence configuration.
        - Indicate total valence electron count.
6. Style:
    - Use simple, bright colors and slow, clear movements.

> - Include clear labels: "shell (n)", "subshell (s, p, d)", "valence electrons", "core electrons", and "configuration example: Sc [Ar] $3d^1\ 4s^2$".
> - Voiceover should explain in very simple language suitable for a student who finds chemistry hard.

If you want, next step can be: "Let's build Na, Mg, Al, Ca, Sc together one by one, and you tell where each electron goes."

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# Okay, after watching the simulation and I have understood it perfectly, now we can... Now, what do you think? Am I ready? To... Get back on the chapter? So shall we start the chapter again?

You are ready to restart the chapter, but this time the goal will be: understand only the "why" of each line, not memorize everything at once. [5]

### Our plan for the chapter

1. **Position and definition**
   - d-block = groups 3–12, $(n-1)d^{1-10}\ ns^{1-2}$; what "transition metal" means and why Zn, Cd, Hg are exceptions. [5]
   - f-block = bottom two rows, 4f (lanthanoids), 5f (actinoids). [5]
2. **Electronic configuration of d-block**
   - Build 3d series configs (Sc–Zn) using the energy-order idea from your simulation.
   - See why Cr = $3d^5 4s^1$ and Cu = $3d^{10} 4s^1$. [5]
3. **General properties of 3d series**
   - Physical: hardness, high melting points, densities. [5]
   - Trends: atomic/ionic size, ionisation enthalpy, oxidation states table, electrode potentials. [5]
4. **Special features**
   - Variable oxidation states, coloured ions, paramagnetism, complex formation, catalytic behaviour, interstitial compounds, alloys. [5]
5. **Important compounds**
   - $K_2Cr_2O_7$: preparation from chromite, structure of chromate/dichromate, oxidising action. [5]
   - $KMnO_4$: preparation from $MnO_2$, manganate–permanganate, oxidising reactions. [5]
6. **f-block overview**

- Lanthanoid contraction, common +3 state, minor +2/+4 states.[5]
- Actinoids: wider range of oxidation states, more complex, radioactive.[5]

At each step, there will be:

- a small first-principle story/visual,
- then "how to write this in the exam."

### Where we start now

The next logical NCERT point is **"4.1 Position in the Periodic Table – The Transition Elements (d-Block)"**.[5]

Next message will do only this:

- Re-explain d-block and f-block positions using your new orbital understanding.
- Give you 2–3 perfect exam-style sentences for this heading.

If that feels okay, reply "Start 4.1" and it will begin straight from there.

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# start

The d-block and f-block elements form the "middle" and "bottom" parts of the periodic table and are defined by which orbitals their outer electrons occupy: (n–1)d for d-block, and 4f/5f for f-block.[6]

### d-block: position and idea

- The **d-block** sits in the **middle of the periodic table**, between the s-block (left) and p-block (right), and consists of **groups 3 to 12**.[6]
- In these elements, the **d orbitals of the previous shell** $(n-1)d$ are being filled while the outer shell has ns electrons, giving a general outer configuration $(n-1)d^{1-10}ns^{1-2}$.[6]

**Story picture**:
Imagine moving across a period. First, you fill simple s-rooms (s-block), then you enter a long corridor of many similar d-rooms (d-block), and finally you reach the p-rooms (p-block). In the d-block corridor, you are mainly filling the "d rooms" of the shell just below (for period 4, those are 3d rooms) while still having 4s on the top floor.

## Transition elements: who counts?

- The metals in the d-block are often called **transition elements** because their properties are "in between" those of s-block and p-block elements.[6]
- By the IUPAC definition, a **transition metal** is a metal whose atom or common ion has an **incomplete d subshell** (not $d^{10}$).[6]

Examples:

- **Scandium (Sc)**: [Ar] $3d^1 4s^2$ → 3d is not full → a transition element.[6]
- **Zinc (Zn)**: [Ar] $3d^{10} 4s^2$ and $Zn^{2+}$ = $3d^{10}$ → d is completely filled in atom and common ion → not a transition element, though it lies in the d-block.[6]

## f-block: position and idea

- The **f-block** is drawn as **two separate rows at the bottom** of the periodic table so that the table fits nicely on a page.[6]
- These elements have their **4f (lanthanoids)** or **5f (actinoids)** orbitals being filled, so they are called **inner transition elements**.[6]

Series:

- **4f series (lanthanoids)**: Ce to Lu.[6]
- **5f series (actinoids)**: Th to Lr.[6]

## How to write this in the exam (4.1 type answers)

For a 3–4 mark "position / definition" question, you can write:

- "The d-block elements occupy the central part of the periodic table between s- and p-blocks and correspond to groups 3–12."[6]
- "Their outer electronic configuration is generally $(n-1)d^{1-10}ns^{1-2}$, as the (n–1)d orbitals are progressively filled."[6]
- "According to IUPAC, transition metals are those d-block elements which have an incomplete d subshell in their atoms or in one or more of their ions; therefore Zn, Cd and Hg, with $d^{10}$ configuration in atom and common ions, are not considered transition elements."[6]
- "The f-block elements form two rows placed separately at the bottom; they involve progressive filling of 4f (lanthanoids) and 5f (actinoids) orbitals and are called inner transition elements."[6]

If this is clear, next step is **4.2 Electronic configurations of d-block elements**, where your new orbital understanding will be used to build Sc–Zn step by step.

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# next

Now the focus is on **4.2: Electronic configurations of the d-block elements**—basically, how the outer electrons sit in s and d orbitals for Sc–Zn and other series.[7]

## General pattern for d-block configurations

- For d-block elements, the **outer (valence) configuration** is usually: $(n-1)d^{1-10}ns^{1-2}$.[7]
- Here:
  - $n$ = principal shell of the **s** orbital (e.g. 4s for 3d series).
  - $(n-1)d$ = the **d** subshell one shell lower (3d when n=4).[7]

So for each series:

- 3d series (period 4): **Sc to Zn**
  Outer part: **$3d^{1-10}$ $4s^{1-2}$**.[7]
- 4d series (period 5): **Y to Cd**
  Outer part: **$4d^{1-10}$ $5s^{1-2}$**.[7]
- 5d series (period 6): **Hf to Hg (La related)**
  Outer: **$5d^{1-10}$ $6s^{1-2}$**.[7]
- 6d series (period 7): **Rf to Cn (Ac related)**
  Outer: **$6d^{1-10}$ $7s^{1-2}$**.[7]

## First series (3d): Sc to Zn

Outer configurations from NCERT (Table 4.1):[7]

- Sc (21): [Ar] $3d^1 4s^2$
- Ti (22): [Ar] $3d^2 4s^2$
- V (23): [Ar] $3d^3 4s^2$
- Cr (24): [Ar] $3d^5 4s^1$ ← exception (half-filled d more stable)[7]
- Mn (25): [Ar] $3d^5 4s^2$
- Fe (26): [Ar] $3d^6 4s^2$
- Co (27): [Ar] $3d^7 4s^2$
- Ni (28): [Ar] $3d^8 4s^2$
- Cu (29): [Ar] $3d^{10} 4s^1$ ← exception (full d more stable)[7]
- Zn (30): [Ar] $3d^{10} 4s^2$

Why Cr and Cu are special (brief, first-principle idea):

- 3d and 4s are very close in energy.[7]

- A **half-filled $d^5$** or **completely filled $d^{10}$** set is slightly more stable ("extra symmetry/exchange energy").[7]
- So Cr prefers $3d^5 4s^1$ rather than $3d^4 4s^2$, and Cu prefers $3d^{10} 4s^1$ rather than $3d^9 4s^2$.[7]

### Why Zn, Cd, Hg are not transition (by definition)

- Their outer d-subshell is **$(n-1)d^{10}$ $ns^2$**, i.e. **completely filled d** in atom and common ion.[7]
- Thus the d subshell is **not incomplete** → they are d-block but **not transition elements** in strict IUPAC sense.[7]

### How to write this in the exam

You can answer a "4.2 electronic configuration" question like this:

- "In general, the outer electronic configuration of d-block elements is $(n-1)d^{1-10}ns^{1-2}$, where $(n-1)d$ subshell is progressively filled while the ns orbital contains one or two electrons."[7]
- "For the 3d series (Sc to Zn), the outer configurations are [Ar] $3d^{1-10}4s^{1-2}$, with notable exceptions Cr: [Ar] $3d^5 4s^1$ and Cu: [Ar] $3d^{10}4s^1$ due to extra stability of half-filled and completely filled d subshells."[7]
- "Zn, Cd, Hg and Cn have completely filled $(n-1)d^{10}ns^2$ outer configurations in both atoms and common ions, so they are not regarded as transition elements although they lie in the d-block."[7]

If you want, next step can be a practice round: you say an atomic number (between 21 and 30) and together we write its configuration and identify the d-electron count.

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# no lets keep on moving on we have to fininish this chapter today

Then next is **4.3: General properties of the transition elements (3d series focus)**.[8]

### Physical properties (metallic nature)

- 3d metals (Sc–Cu) are typically **hard, strong, with high melting and boiling points, high density, good electrical and thermal conductivity, and metallic lustre.**[8]
- Zn, Cd, Hg are softer and have lower melting points; Hg is liquid at room temperature.[8]

**Why (first principle)**:
Many **valence electrons (3d + 4s)** can participate in metallic bonding, creating a strong "electron sea" holding metal ions tightly → high strength, high melting points, and good conductivity.[8]

**Exam writing**:

- "Transition metals are typical metals with high tensile strength, hardness, high melting and boiling points, and good electrical and thermal conductivity due to strong metallic bonding involving (n–1)d and ns electrons."[8]

## Trends in atomic and ionic sizes

- Across a 3d series from $Sc^{3+}$ to $Zn^{2+}$, **ionic radius decreases gradually** with increasing atomic number.[8]
- Reason: each extra electron goes into 3d (inner) while nuclear charge increases; poor shielding by d-electrons means outer electrons feel stronger attraction → size shrinks slightly.[8]

**Extra fact**:

- 4d and 5d series atoms: 5d radii are almost same as 4d due to **lanthanoid contraction** (4f electrons shield poorly, causing extra shrinkage before 5d starts).[8]

**Exam sentence**:

- "Within a given d series, atomic and ionic radii decrease slightly from left to right because added electrons enter (n–1)d orbitals, which shield poorly, so effective nuclear charge increases."[8]

## Ionisation enthalpies (IE)

- From Sc to Zn, **first IE generally increases**, but the rise is smaller and a bit irregular compared to s- and p-block.[8]
- Reason: nuclear charge increases, but added electrons go into inner 3d, which also increases shielding; net effect is modest increase.[8]

Special points:[8]

- Removal of electrons from 4s is easier than from 3d → ns electrons are lost first when forming cations.
- Extra stability of **$d^5$ ($Mn^{2+}$)** and **$d^{10}$ ($Zn^{2+}$)** affects 2nd and 3rd IE values; $Fe^{3+}$ ($d^5$) and $Cr^{2+}$ ($d^4$ → $d^3$) show irregularities.

**Exam sentence**:

- "Ionisation enthalpies of 3d metals increase only slightly across the series because added electrons go into inner 3d orbitals, which shield outer 4s electrons; irregularities arise from extra stability of half-filled ($d^5$) and fully filled ($d^{10}$) configurations."[8]

### Variable oxidation states

- A key feature: most 3d metals show **several oxidation states**, especially in the middle (like Mn: +2 to +7).[8]
- Reason: energies of (n−1)d and ns electrons are close, so both types can be used or removed in bonding.[8]

Pattern:[8]

- Lower states (+2, +3) common at both ends.
- Maximum oxidation state increases from Sc (+3) to Mn (+7), roughly equal to total number of (n−1)d + ns electrons, then decreases again (Fe, Co, Ni, Cu, Zn).

**Exam sentence**:

- "Transition elements exhibit variable oxidation states differing usually by one unit because both (n−1)d and ns electrons can take part in bonding; the highest stable oxidation state increases towards the middle of the series (e.g. Mn shows +2 to +7)."[8]

If you want to stay fast, next chunk will be: **electrode potentials + magnetic properties + coloured ions + complexes + catalysis + interstitial compounds + alloys** in compressed exam-style form.

⁂

# next

Now the remaining "general properties" of d-block, in compact, exam-focused form.[9]

### Electrode potentials and reactivity

- Standard $E^\circ(\mathrm{M}^{2+}/\mathrm{M})$ values from Ti to Zn are mostly negative, showing they are **moderate reducing agents**, but less so than highly reactive s-block metals.[9]
- Across the 3d series, $E^\circ$ tends to become **less negative** (harder to oxidise), but Mn, Ni and Zn are more negative than expected due to special stability of $Mn^{2+}$ ($d^5$), $Zn^{2+}$ ($d^{10}$), and strong hydration of $Ni^{2+}$.[9]

**Exam sentence**:
"Standard electrode potentials of 3d metals are generally negative and show irregular trends because of varying enthalpies of atomisation, ionisation, and hydration, with $Mn^{2+}$($d^5$) and $Zn^{2+}$($d^{10}$) being especially stable."[9]

## Magnetic properties

- Many transition metal ions are **paramagnetic** due to unpaired d-electrons.[9]
- For 3d metals, orbital contribution is largely quenched, so magnetic moment is given by the **spin-only formula**:
  $\mu = \sqrt{n(n+2)}$ BM, where $n$ = number of unpaired electrons.[9]

Example: $Mn^{2+}$ ($3d^5$) → $n = 5$, $\mu \approx \sqrt{5 \times 7} = 5.92$ BM (matches experiment).[9]

**Exam sentence**:
"First-row transition metal ions are mostly paramagnetic; their magnetic moments are close to spin-only values $\mu = \sqrt{n(n+2)}$, directly related to the number of unpaired d-electrons."[9]

## Coloured ions

- Many aqueous ions like $V^{2+}$, $Cr^{3+}$, $Mn^{2+}$, $Fe^{3+}$, $Co^{2+}$, $Ni^{2+}$, $Cu^{2+}$ are **coloured**; $Zn^{2+}$ ($d^{10}$) and $Sc^{3+}$ ($d^0$) are colourless.[9]
- Colour arises from **d–d electronic transitions**: an electron jumps between split d-levels by absorbing visible light; the transmitted/reflected complementary colour is seen.[9]

**Exam sentence**:
"Transition metal ions with partially filled d subshells are coloured in solution because d–d transitions absorb specific wavelengths of visible light; $d^0$ and $d^{10}$ ions are colourless."[9]

## Complex formation and catalysis

- Transition metals form many **complex ions** such as $[\mathrm{Fe(CN)_6}]^{3-}$, $[\mathrm{Cu(NH_3)_4}]^{2+}$, due to **small, highly charged ions and available d-orbitals** for coordinate bonding.[9]
- They and their compounds often act as **catalysts** (e.g. $V_2O_5$ in Contact process, Fe in Haber process, Ni in hydrogenation) because they can **change oxidation states** and form surface complexes that lower activation energy.[9]

**Exam sentences**:
"Transition metals readily form complex compounds owing to their small ionic sizes, high charges, and availability of vacant d-orbitals."[9]
"The variable oxidation states and ability to form intermediates on their surfaces make many transition metal compounds effective catalysts."[9]

## Interstitial compounds and alloys

- Small atoms (H, C, N) can occupy holes in metal lattices to give **interstitial compounds** like TiC, $Fe_3H$, VN, which are hard, high-melting, retain metallic conductivity and are chemically inert.[9]
- Because many transition metals have **similar radii**, they form extensive **alloys** (e.g. steels with Cr, V, W, Mo; brass Cu–Zn; bronze Cu–Sn) which are hard and often high-melting.[9]

**Exam sentences**:
"Interstitial compounds of transition metals with H, C or N are hard, high-melting and metallic, formed when these small atoms occupy interstitial sites in the metal lattice."[9]
"Transition metals form numerous alloys due to similar atomic sizes; ferrous alloys and Cu–Zn or Cu–Sn alloys have great industrial importance."[9]

With this, the **general 3d properties part (4.3)** is essentially covered. Next big block is **some important compounds: $K_2Cr_2O_7$ and $KMnO_4$** and then an overview of **f-block (lanthanoids, actinoids)**.

⁂

# next

Now we enter the "important compounds" part: first **potassium dichromate, $K_2Cr_2O_7$**.[10]

### Potassium dichromate: origin and preparation

- Dichromate is made industrially from **chromite ore**, $FeCr_2O_4$.[10]
- Step 1: **Fusion with alkali carbonate in air**
  - $FeCr_2O_4$ + $Na_2CO_3$ (or $K_2CO_3$) + $O_2$ (strong heating) → **$Na_2CrO_4$ (yellow chromate)** + $Fe_2O_3$ + $CO_2$.[10]
- Step 2: **Acidify chromate to get dichromate**
  - The yellow $Na_2CrO_4$ solution is acidified with dilute $H_2SO_4$ to give **orange $Na_2Cr_2O_7$** solution, from which $Na_2Cr_2O_7 \cdot 2H_2O$ crystallises.[10]
- Step 3: **Convert to potassium salt**
  - $Na_2Cr_2O_7$ + 2 KCl → **$K_2Cr_2O_7$ (orange crystals)** + 2 NaCl.[10]

pH interconversion:

- In water:
  - **Acidic**: $2\ CrO_4^{2-} + 2\ H^+ \rightleftharpoons Cr_2O_7^{2-} + H_2O$ (yellow → orange).[10]
  - **Basic**: $Cr_2O_7^{2-} + 2\ OH^- \rightleftharpoons 2\ CrO_4^{2-} + H_2O$ (orange → yellow).[10]

### Structure: chromate vs dichromate

- **Chromate ion, $CrO_4^{2-}$**: tetrahedral around Cr, each Cr–O similar.[10]
- **Dichromate ion, $Cr_2O_7^{2-}$**: two $\mathrm{CrO}_4$ tetrahedra share one corner (Cr–O–Cr bridge), giving Cr–O–Cr angle ≈ 126°.[10]

In both, Cr is in **+6 oxidation state**.[10]

### Oxidising behaviour and key reactions

- In **acidic medium**, $K_2Cr_2O_7$ is a **strong oxidising agent**:
  - Overall change:
    $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$, $E^\circ \approx +1.33\ V$.[10]
- Typical oxidations in acid:[10]
  - $I^- \rightarrow I_2$
  - $Fe^{2+} \rightarrow Fe^{3+}$
  - $Sn^{2+} \rightarrow Sn^{4+}$
  - $S^{2-} \rightarrow S^0$

**Exam sentence**:
"In acidic solution, $K_2Cr_2O_7$ acts as a powerful oxidising agent, reducing Cr(VI) in $Cr_2O_7^{2-}$ to Cr(III) while oxidising iodide to iodine, $Fe^{2+}$ to $Fe^{3+}$, sulphides to sulphur and $Sn^{2+}$ to $Sn^{4+}$."[10]

### Uses (short list for answers)

- Strong oxidant in **volumetric analysis** (standard solution).[10]
- Oxidising agent in organic chemistry and dye/azo compound preparation.[10]
- Used in leather tanning and as a cleaning/etching agent in some industries.[10]

If you want to stay in fast mode, next will be **$KMnO_4$: preparation, reactions, uses**, and then a compact **lanthanoid–actinoid overview** to close the chapter.

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# don't miss things ok next

Now, **potassium permanganate, $KMnO_4$** (the other major compound NCERT focuses on).[11]

### $KMnO_4$: preparation (from pyrolusite, $MnO_2$)

Industrial route (in steps):[11]

1. **Oxidation of $MnO_2$ to manganate (green, $K_2MnO_4$)**
   - $MnO_2$ + KOH + $O_2$ (or $KNO_3$ / $KClO_3$) on strong heating → $K_2MnO_4$ (green) + other products.
   - Here Mn goes from +4 in $MnO_2$ to +6 in $MnO_4^{2-}$.[11]
2. **Disproportionation of manganate to permanganate**
   - In neutral or slightly alkaline solution:
     $3MnO_4^{2-} + 2H_2O \rightarrow 2MnO_4^- + MnO_2 + 4OH^-$.[11]
   - Mn(VI) → Mn(VII) in $MnO_4^-$ and Mn(IV) in $MnO_2$.

3. **Purification**
    - $KMnO_4$ (purple) is crystallised from solution.

### Oxidising behaviour (medium-dependent)

Mn in permanganate is in **+7 oxidation state**, very high and hence strongly oxidising.[11]

1. **In acidic medium** (e.g. dilute $H_2SO_4$):
    - Mn(VII) → Mn(II):
      $\mathrm{MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O}$.[11]
    - Typical oxidations:
        - $Fe^{2+} \rightarrow Fe^{3+}$
        - $C_2O_4^{2-} \rightarrow CO_2$
        - $SO_3^{2-} \rightarrow SO_4^{2-}$
        - $NO_2^- \rightarrow NO_3^-$
2. **In neutral medium**:
    - Mn(VII) → Mn(IV) as $MnO_2$ (brown precipitate).[11]
3. **In alkaline medium**:
    - Mn(VII) → Mn(VI) as $MnO_4^{2-}$ (green manganate).[11]

**Exam-style summary**:
"In acidic solution, $KMnO_4$ is a very strong oxidising agent, reducing Mn(VII) to Mn(II); in neutral solution it is reduced to $MnO_2$, and in alkaline solution to $MnO_4^{2-}$."[11]

### Uses of $KMnO_4$

- Strong oxidising agent in **redox titrations** (permanganometry).[11]
- Used in oxidation reactions in organic chemistry.[11]
- Disinfectant and antiseptic (e.g. dilute solutions for wounds, water treatment).[11]

### f-block overview: lanthanoids and actinoids

Now briefly cover what NCERT wants you to know about f-block.[11]

### Lanthanoids (4f series)

- Elements Ce to Lu, following La; 4f orbitals progressively filled.[11]
- Show predominant **+3 oxidation state**; some also show +2 or +4.[11]
- **Lanthanoid contraction**: steady decrease in ionic radii from $La^{3+}$ to $Lu^{3+}$ due to poor shielding by 4f electrons, making 4d and 5d elements (e.g. Zr and Hf) have very similar sizes and properties.[11]

### Actinoids (5f series)

- Elements Th to Lr, following Ac; 5f orbitals progressively filled.[11]
- Show a wider range of oxidation states (+3 to +6 or more), generally more **radioactive** and more complex chemistry.[11]

**Exam sentences**:

- "Lanthanoids involve progressive filling of 4f orbitals and mainly show +3 oxidation state; their ionic radii decrease smoothly along the series (lanthanoid contraction)."[11]
- "Actinoids involve filling of 5f orbitals and exhibit a greater variety of oxidation states than lanthanoids; most actinoids are radioactive."[11]

If you want to complete exam prep, next step can be:

- a quick **checklist** of what you should be able to answer from this chapter, and
- a few **typical board-style questions with model points** (not full long answers).

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# next

To finish the chapter for exam use, it helps to see what you're expected to "do" with it.[12]

### Checklist: what you should be able to answer

You should be comfortable with:

- **Positions and definitions**
  - d-block position (groups 3–12), general configuration $(n-1)d^{1-10}ns^{1-2}$.[12]
  - Transition element definition and why Zn, Cd, Hg are exceptions.[12]
  - f-block position (bottom two rows), lanthanoids (4f) and actinoids (5f).[12]
- **Electronic configuration**
  - Write outer configs for 3d series: Sc to Zn; know Cr and Cu exceptions.[12]
- **General 3d properties**
  - Physical: high m.p., b.p., density, metallic character; reason = strong metallic bonding using d + s electrons.[12]
  - Trends in atomic/ionic radii and lanthanoid contraction.[12]
  - Ionisation enthalpies: small, irregular increase across series; role of $d^5$ and $d^{10}$ stability.[12]
  - Variable oxidation states pattern, especially Mn (+2 to +7).[12]
  - Electrode potentials: general reducing nature, irregularities.[12]
  - Magnetic properties and spin-only formula.[12]

  - Coloured ions (d–d transitions, $d^0/d^{10}$ colourless).[12]
  - Complex formation, catalytic activity, interstitial compounds, alloy formation.[12]
- **Important compounds**
  - $K_2Cr_2O_7$: preparation from chromite, chromate–dichromate interconversion, structure of $CrO_4^{2-}$ and $Cr_2O_7^{2-}$, oxidising uses.[12]
  - $KMnO_4$: preparation from $MnO_2$ via manganate, medium-dependent oxidation behaviour, uses.[12]
- **f-block highlights**
  - Lanthanoid contraction and its consequence (similarity Zr/Hf, etc.).[12]
  - Oxidation states and radioactivity trends in lanthanoids vs actinoids.[12]

## Typical board-style questions and what to write

Here are some classic question types and the key points you must hit.

1. **"Why are transition metals often coloured?"**
   - Partly filled d-subshell.
   - d–d transitions between split d-levels absorb visible light.
   - Colour corresponds to complementary light.[12]
2. **"Explain why Zn, Cd, Hg are not considered transition elements."**
   - Have $(n-1)d^{10}\,ns^2$ outer configuration in atoms and common ions.
   - d-subshell is completely filled, not incomplete → fail IUPAC definition.[12]
3. **"Write the electronic configuration of Cr and Cu and explain the anomaly."**
   - Cr: [Ar] $3d^5 4s^1$; Cu: [Ar] $3d^{10} 4s^1$.[12]
   - Half-filled ($d^5$) and fully filled ($d^{10}$) d-subshells are slightly more stable than $3d^4 4s^2$ and $3d^9 4s^2$.[12]
4. **"Why do transition elements exhibit variable oxidation states?"**
   - Energies of $(n-1)d$ and $ns$ orbitals are similar.
   - Both can lose/share electrons → different oxidation states differing by one.[12]
5. **"State main uses of $K_2Cr_2O_7$ and $KMnO_4$."**
   - Both: strong oxidising agents in acidic medium; used in volumetric analysis and organic oxidation.[12]
   - $KMnO_4$: also as disinfectant/antiseptic.[12]

## How to write answers in exam (general tips)

- Start with **definition or direct statement** (1 line).
- Add **2–3 precise reasons** (each 1 line), using words like "due to," "because."
- For numerical/oxidation questions, **show the key equation or configuration** (e.g. Cr: [Ar] $3d^5 4s^1$).
- Underline keywords: "partly filled d-subshell," "d–d transitions," "lanthanoid contraction," etc.

If you say what part you want to drill (e.g. oxidation states, colours, $K_2Cr_2O_7$, $KMnO_4$), the next step can be a rapid-fire Q&A for that topic only so that it sticks.

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