Weird Chemist

Planck's Quantum Theory — DPP-3 Solutions

Chapter: Structure of Atom

"Teen formulae, sab questions. Sahi unit mein dalo, answer aa jaayega."

Key Formulae for This DPP

- Energy of one photon: $E = h\nu = \dfrac{hc}{\lambda}$
- Frequency from energy: $\nu = \dfrac{E}{h}$
- Wavelength from energy: $\lambda = \dfrac{hc}{E}$
- Number of photons: $n = \dfrac{\text{Total Energy}}{E_{\text{one photon}}} = \dfrac{P \cdot t}{hc/\lambda} = \dfrac{P \cdot t \cdot \lambda}{hc}$
- Molar bond energy per photon: $E_{\text{one photon}} = \dfrac{\Delta E_{\text{molar}}}{N_A}$
- Energy conservation (emission): $E_{\text{absorbed}} = E_{\text{emitted}_1} + E_{\text{emitted}_2}$, i.e. $\dfrac{hc}{\lambda_{\text{abs}}} = \dfrac{hc}{\lambda_1} + \dfrac{hc}{\lambda_2}$
- Useful: $hc = 6.626 \times 10^{-34} \times 3 \times 10^8 = 1.988 \times 10^{-25}$ J·m
- Unit conversions: 1 nm $= 10^{-9}$ m; 1 Å $= 10^{-10}$ m; 1 eV $= 1.602 \times 10^{-19}$ J

TYPE 1 : Energy and Wavelength of Photon

Q.1 Energy of a photon of sodium light, $\lambda = 5.862 \times 10^{-16}$ m.
($h = 6.6 \times 10^{-34}$ J s, $c = 3 \times 10^8$ m s^{-1})

Explanation

$$E = \frac{hc}{\lambda} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8}{5.862 \times 10^{-16}} = \frac{19.8 \times 10^{-26}}{5.862 \times 10^{-16}} = \frac{19.8}{5.862} \times 10^{-10} = 3.38 \times 10^{-10}\ \text{J}$$

Concept

$E = hc/\lambda$ is the only formula needed here. Plug in SI values directly (λ is already in metres). Numerator: $h \times c = 6.6 \times 10^{-34} \times 3 \times 10^8 = 19.8 \times 10^{-26}$. Then divide by λ, subtract exponents: $10^{-26}/10^{-16} = 10^{-10}$.

Answer

Option (1): 3.38×10^{-10} **J**

Common Student Mistake

Note: $\lambda = 5.862 \times 10^{-16}$ m is unrealistically short for sodium light (real sodium D-line is ~ 589 nm). This appears to be a printing error in the PDF — the exponent should likely be 10^{-7} m. Regardless, solve with the given value. Students lose marks by "correcting" the question instead of using the given data.

Q.2 Energy of light with $\lambda = 45$ nm.

($h = 6.63 \times 10^{-34}$ J s, $c = 3 \times 10^{8}$ m s^{-1})

Explanation

Convert: $\lambda = 45$ nm $= 45 \times 10^{-9}$ m $= 4.5 \times 10^{-8}$ m

$$E = \frac{hc}{\lambda} = \frac{6.63 \times 10^{-34} \times 3 \times 10^{8}}{4.5 \times 10^{-8}} = \frac{19.89 \times 10^{-26}}{4.5 \times 10^{-8}} = 4.42 \times 10^{-18} \text{ J}$$

Concept

The only task beyond the formula is the unit conversion: nm $\rightarrow$ m. 45 nm $= 45 \times 10^{-9}$ m. After that, divide hc by λ as usual.

Answer

Option (4): 4.42×10^{-18} **J**

Common Student Mistake

Options (1) and (2) correspond to accidentally computing $\nu = c/\lambda$ and reporting that as energy — giving 6.67×10^{15} Hz (which has units of s^{-1}, not joules). Always check whether the answer has energy units (J). Option (3) arises from a power-of-10 error in the nm conversion.

Q.3 Wavelength (in Å) of a photon with $E = 4.38 \times 10^{-18}$ J.

Explanation

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{4.38 \times 10^{-18}} = \frac{1.988 \times 10^{-25}}{4.38 \times 10^{-18}} = 0.454 \times 10^{-7} \text{ m} = 4.54 \times 10^{-8} \text{ m}$$

Convert to angstroms: $4.54 \times 10^{-8} \text{ m} \times \dfrac{1 \text{ Å}}{10^{-10} \text{ m}} = 454$ Å

Concept

Two steps: (1) find λ in metres using $\lambda = hc/E$, then (2) convert to Å using 1 Å $= 10^{-10}$ m, so multiply by 10^{10}.

Answer

Option (1): 454 Å

Common Student Mistake

Students often forget to convert metres to Å and report the answer as 4.54×10^{-8} (in metres), which matches none of the options. Always read what unit the answer is required in before finishing.

Q.4 $\Delta E = 3 \times 10^{-8}$ J, $h = 6.64 \times 10^{-34}$ J s, $c = 3 \times 10^{8}$ m/s. Find λ.

Explanation

$$\lambda = \frac{hc}{\Delta E} = \frac{6.64 \times 10^{-34} \times 3 \times 10^{8}}{3 \times 10^{-8}} = \frac{6.64 \times 10^{-26}}{3 \times 10^{-8}/3}$$

Simplify: $\dfrac{6.64 \times 3 \times 10^{-26}}{3 \times 10^{-8}} = 6.64 \times 10^{-26+8} = 6.64 \times 10^{-18}$ m
Convert to Å: 6.64×10^{-18} m $\times 10^{10}$ Å/m $= 6.64 \times 10^{-8}$ Å

Concept

$\lambda = hc/\Delta E$, then convert m $\rightarrow$ Å by multiplying by 10^{10}. The c and ΔE both have $3\times$ in the numerator/denominator, so they cancel cleanly, leaving only h and the power of 10.

Answer

Option (3): 6.64×10^{-8} **Å**

Common Student Mistake

This is an extremely small wavelength (smaller than a nucleus), which makes physical sense here only as a mathematical exercise. Students often pick option (1) or (2) by making a power-of-10 error when converting m to Å. The conversion is: multiply by 10^{10} (since 1 m $= 10^{10}$ Å).

Q.5 Violet light ($\lambda = 4000$ Å) vs red light ($\lambda = 7000$ Å): which photon has higher energy?

Explanation

From $E = hc/\lambda$: energy is **inversely proportional** to wavelength.
Shorter wavelength $\Rightarrow$ higher energy.
$\lambda_{\text{violet}} = 4000$ Å $< \lambda_{\text{red}} = 7000$ Å
Therefore: $E_{\text{violet}} > E_{\text{red}}$

Concept

Key principle: $E \propto 1/\lambda$. No calculation needed. Shorter wavelength = higher frequency = higher energy per photon. Intensity (brightness) determines the *number* of photons, not the energy per photon.

Answer

Option (1): Violet light photon has higher energy

Common Student Mistake

Option (4) is a trap — "cannot be determined without intensity." Intensity tells you how many photons per second, not the energy of each photon. Energy per photon depends only on frequency/wavelength, not on intensity.

Q.6 Ratio of energy of photon at 2000 Å to that at 4000 Å.

Explanation

$E = hc/\lambda$, so for fixed h and c:

$$\frac{E_{2000}}{E_{4000}} = \frac{hc/\lambda_1}{hc/\lambda_2} = \frac{\lambda_2}{\lambda_1} = \frac{4000\ \text{Å}}{2000\ \text{Å}} = 2$$

Concept

Since $E \propto 1/\lambda$, the ratio of energies equals the inverse ratio of wavelengths. When λ halves, E doubles. No need to calculate actual energies.

Answer

Option (4): 2

Common Student Mistake

Students write $E_{2000}/E_{4000} = \lambda_1/\lambda_2 = 2000/4000 = 1/2$ — inverting the relationship. Remember: $E \propto 1/\lambda$, so $E_1/E_2 = \lambda_2/\lambda_1$ (flip the wavelengths when taking the ratio).

TYPE 2 : Energy, Frequency and Wavelength of Photon

Q.7 Frequency and energy of a photon with $\lambda = 4000$ Å.
($h = 6.626 \times 10^{-34}$ J s, $c = 3 \times 10^8$ m s^{-1})

Explanation

Convert: $\lambda = 4000\ \text{Å} = 4000 \times 10^{-10}\ \text{m} = 4 \times 10^{-7}\ \text{m}$
Step 1 — Frequency:

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^8}{4 \times 10^{-7}} = 0.75 \times 10^{15} = 7.5 \times 10^{14}\ \text{s}^{-1}$$

Step 2 — Energy:

$$E = h\nu = 6.626 \times 10^{-34} \times 7.5 \times 10^{14} = 49.695 \times 10^{-20} = 4.97 \times 10^{-19}\ \text{J} \approx 4.96 \times 10^{-19}\ \text{J}$$

Concept

Two-step process: (1) $\nu = c/\lambda$, then (2) $E = h\nu$. Alternatively, do it in one step: $E = hc/\lambda$. Both approaches give the same answer; the two-step method makes it easy to verify frequency and energy separately.

Answer

Option (1): 7.5×10^{14} **s**$^{-1}$ **and** 4.96×10^{-19} **J**

Common Student Mistake

Option (2) has frequency 3.75×10^{14} — this arises from forgetting to convert Å to m and using $\lambda = 4000$ without any conversion, then computing $c/4000$. Always convert Å to m first: $1\ \text{Å} = 10^{-10}$ m.

Q.8 Frequency and wavelength of photon with $E = 3.98 \times 10^{-15}$ J.
$(h = 6.626 \times 10^{-34}\ \text{J s},\ c = 3 \times 10^{8}\ \text{m s}^{-1})$

Explanation

Step 1 — Frequency:

$$\nu = \frac{E}{h} = \frac{3.98 \times 10^{-15}}{6.626 \times 10^{-34}} = \frac{3.98}{6.626} \times 10^{19} = 0.6006 \times 10^{19} = 6.0 \times 10^{18}\ \text{Hz}$$

Step 2 — Wavelength:

$$\lambda = \frac{c}{\nu} = \frac{3 \times 10^{8}}{6.0 \times 10^{18}} = 0.5 \times 10^{-10}\ \text{m} = 0.5\ \text{Å}$$

Concept

From energy: find $\nu = E/h$ first. Then $\lambda = c/\nu$. For the wavelength step: $3 \times 10^{8}/6 \times 10^{18} = 0.5 \times 10^{-10}$ m. Recall 1 Å $= 10^{-10}$ m, so 0.5×10^{-10} m $= 0.5$ Å.

Answer

Option (1): 6.0×10^{18} **Hz and 0.5 Å**

Q.9 Wavelength and frequency of photon with $E = 2$ eV.
$(h = 6.626 \times 10^{-34}\ \text{J s},\ c = 3 \times 10^{8}\ \text{m s}^{-1},\ 1\ \text{eV} = 1.602 \times 10^{-19}\ \text{J})$

Explanation

Step 1 — Convert eV to J:

$$E = 2 \times 1.602 \times 10^{-19} = 3.204 \times 10^{-19}\ \text{J}$$

Step 2 — Wavelength:

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{3.204 \times 10^{-19}} = \frac{1.988 \times 10^{-25}}{3.204 \times 10^{-19}} = 0.6204 \times 10^{-6}\ \text{m} = 6.204 \times 10^{-7}\ \text{m}$$

Step 3 — Frequency:

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^{8}}{6.204 \times 10^{-7}} = 0.483 \times 10^{15} \approx 4.9 \times 10^{14}\ \text{s}^{-1}$$

Concept

The critical first step is converting eV to joules before applying any formula. All formulae require SI units (joules). Skipping the eV→J conversion produces an answer off by a factor of 1.602×10^{-19}.

Answer

Option (1): 6.204×10^{-7} **m and** 4.9×10^{14} $\mathbf{s}^{-1}$

Common Student Mistake

Students plug $E = 2$ (without converting from eV to J) and get a wavelength of $\sim 10^{-7}$ m — off by a factor of $\sim 10^{19}$. Energy in formulas must always be in joules.

TYPE 3 : Number of Photons

Q.10 A 1 kW transmitter at 800 Hz. How many photons per second?

Explanation

Power = energy per second = 1 kW = 1000 J/s.
Energy of one photon: $E = h\nu = 6.626 \times 10^{-34} \times 800 = 5.30 \times 10^{-31}$ J

$$n = \frac{\text{Power}}{E_{\text{photon}}} = \frac{1000}{5.30 \times 10^{-31}} = \frac{10^3}{5.30 \times 10^{-31}} = 1.887 \times 10^{33} \approx 1.88 \times 10^{33} \text{ photons/s}$$

Concept

General formula for number of photons per second from a source of power P:

$$n = \frac{P}{h\nu} = \frac{P\lambda}{hc}$$

When frequency is given directly, use $n = P/(h\nu)$. When wavelength is given, use $n = P\lambda/(hc)$.

Answer

Option (2): 1.88×10^{33}

Common Student Mistake

Option (1) = 1.71×10^{21} arises from confusing the frequency 800 Hz with a wavelength and using c/λ instead of ν directly. When frequency is given in Hz, use $E = h\nu$ directly — do not compute $\lambda = c/\nu$ unless the question specifically asks for it.

Q.11 100 W bulb emits at 400 nm. Photons emitted per second?
($h = 6.626 \times 10^{-34}$ J s, $c = 3 \times 10^8$ m s^{-1})

Explanation

$\lambda = 400$ nm $= 4 \times 10^{-7}$ m. Power = 100 W = 100 J/s.

$$E_{\text{photon}} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^8}{4 \times 10^{-7}} = \frac{1.988 \times 10^{-25}}{4 \times 10^{-7}} = 4.97 \times 10^{-19} \text{ J}$$

$$n = \frac{P}{E_{\text{photon}}} = \frac{100}{4.97 \times 10^{-19}} = 20.12 \times 10^{19} = 2.012 \times 10^{20} \text{ photons/s}$$

Answer

Option (1): 2.012×10^{20}

Q.12 40 W bulb, $\lambda = 620$ nm, 80% efficiency, $t = 20$ s. Number of photons emitted?
$(hc = 12400 \text{ eV} \cdot \text{Å})$

Explanation

Step 1 — Effective power (accounting for efficiency):

$$P_{\text{eff}} = 40 \times 0.80 = 32 \text{ W} = 32 \text{ J/s}$$

Step 2 — Total energy emitted in 20 s:

$$E_{\text{total}} = 32 \times 20 = 640 \text{ J}$$

Step 3 — Energy of one photon using $hc = 12400$ eV · Å:
$\lambda = 620 \text{ nm} = 6200 \text{ Å}$

$$E_{\text{photon}} = \frac{12400 \text{ eV} \cdot \text{Å}}{6200 \text{ Å}} = 2 \text{ eV} = 2 \times 1.6 \times 10^{-19} = 3.2 \times 10^{-19} \text{ J}$$

Step 4 — Number of photons:

$$n = \frac{640}{3.2 \times 10^{-19}} = 200 \times 10^{19} = 2 \times 10^{21}$$

Concept

The shortcut $hc = 12400$ eV·Å is extremely useful: $E\,(\text{eV}) = 12400/\lambda\,(\text{Å})$. This saves computing hc every time. Memorise it.

Answer

Option (4): 2×10^{21}

Common Student Mistake

Two common errors: (1) Ignoring efficiency and using 40 W directly — gives $n = 2.5 \times 10^{21}$, which is not an option. (2) Forgetting to multiply by time $t = 20$ s and reporting photons per second instead of total photons emitted.

Q.13 Number of photons of $\lambda = 7000$ Å equivalent to 1 J.

Explanation

$\lambda = 7000 \text{ Å} = 7 \times 10^{-7}$ m

$$E_{\text{photon}} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{7 \times 10^{-7}} = \frac{1.988 \times 10^{-25}}{7 \times 10^{-7}} = 0.2840 \times 10^{-18} = 2.84 \times 10^{-19} \text{ J}$$

$$n = \frac{1 \text{ J}}{2.84 \times 10^{-19} \text{ J}} = 0.352 \times 10^{19} = 3.52 \times 10^{18} \text{ photons}$$

Answer

Option (2): 3.52×10^{18}

Common Student Mistake

Option (1) = 3.52×10^{-18} is simply the energy of one photon, not the number of photons. The number of photons = Total energy / Energy per photon = $1/(3.52 \times 10^{-19}) = 3.52 \times 10^{18}$. Do not confuse E_{photon} with n.

TYPE 4 : Miscellaneous

Q.14 Iodine molecules dissociate at $\lambda = 4995$ Å. Energy to dissociate 1 mole?

Explanation

$\lambda = 4995$ Å $= 4.995 \times 10^{-7}$ m

Energy of one photon:

$$E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{4.995 \times 10^{-7}} = \frac{1.988 \times 10^{-25}}{4.995 \times 10^{-7}} = 3.979 \times 10^{-19}\ \text{J}$$

Energy per mole (one photon breaks one bond, one mole of bonds needs N_A photons):

$$E_{\text{molar}} = 3.979 \times 10^{-19} \times 6.022 \times 10^{23} = 23.97 \times 10^{4}\ \text{J/mol} = 239.7\ \text{kJ/mol}$$

Convert to kcal: 239.7 kJ/mol $\div$ 4.184 kJ/kcal ≈ 57.3 kcal/mol

Concept

The logic: one photon breaks one bond. To break one mole of bonds, you need N_A photons. So molar dissociation energy $= E_{\text{photon}} \times N_A$. Then convert J to kcal if required: 1 kcal $= 4.184$ kJ.

Answer

Option (4): 57.3 kcal/mol

Common Student Mistake

Students often stop after calculating E_{photon} and forget to multiply by N_A to get molar energy. Also, forgetting to convert kJ to kcal (when the options are in kcal) leads to selecting a wrong option.

Q.15 Energy to break Cl–Cl bond = 242 kJ mol^{-1}. Longest λ to break one bond.
($c = 3 \times 10^{8}$ m s^{-1}, $N_A = 6.02 \times 10^{23}$ mol^{-1}) *[AIEEE 2010]*

Explanation

Step 1 — Energy per bond (per photon needed):

$$E_{\text{bond}} = \frac{242 \times 10^{3}\ \text{J/mol}}{6.02 \times 10^{23}\ \text{mol}^{-1}} = \frac{242 \times 10^{3}}{6.02 \times 10^{23}} = 40.2 \times 10^{-20} = 4.02 \times 10^{-19}\ \text{J}$$

Step 2 — Longest wavelength:
Longest λ = minimum energy photon that can *just* break the bond, i.e. $E_{\text{photon}} = E_{\text{bond}}$.

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \times 3 \times 10^{8}}{4.02 \times 10^{-19}} = \frac{1.988 \times 10^{-25}}{4.02 \times 10^{-19}} = 4.945 \times 10^{-7}\ \text{m} \approx 494\ \text{nm}$$

Concept

Key principle: "longest wavelength" means the photon with minimum energy that can just do the job. At exactly this wavelength, $E_{\text{photon}} = E_{\text{bond}}$. Any longer wavelength photon would have insufficient energy.

Step order: molar energy $\to$ per-bond energy (divide by N_A) $\to \lambda = hc/E$.

Answer

Option (1): 494 nm

Common Student Mistake

Students forget to divide by N_A and use $E = 242 \times 10^3$ J directly in $\lambda = hc/E$, getting a wavelength of the order of 10^{-28} m — physically meaningless. **Always divide molar energy by** N_A to get energy per bond/particle before using photon formulae.

Q.16 Ionisation energy of Na = 495.5 kJ mol^{-1}. Lowest frequency of light to ionise Na.
$(h = 6.626 \times 10^{-34}$ J s, $N_A = 6.022 \times 10^{23}$ mol$^{-1})$ *[JEE Main 2014]*

Explanation

Step 1 — IE per atom:

$$E = \frac{495.5 \times 10^3}{6.022 \times 10^{23}} = \frac{4.955 \times 10^5}{6.022 \times 10^{23}} = 0.8228 \times 10^{-18} = 8.228 \times 10^{-19} \text{ J}$$

Step 2 — Lowest frequency (minimum ν to ionise = photon energy just equals IE):

$$\nu = \frac{E}{h} = \frac{8.228 \times 10^{-19}}{6.626 \times 10^{-34}} = 1.242 \times 10^{15} \approx 1.24 \times 10^{15} \text{ s}^{-1}$$

Concept

"Lowest frequency" = threshold frequency = minimum energy photon that can just ionise the atom. At this frequency, $h\nu_{\text{min}} = IE_{\text{per atom}}$. Frequency lower than this cannot ionise regardless of intensity (photoelectric effect principle).

Answer

Option (3): 1.24×10^{15} **s**$^{-1}$

Common Student Mistake

Same error as Q15: using 495.5 kJ directly without dividing by N_A. The ionisation energy is given *per mole*; you need it *per atom* to use $E = h\nu$. Dividing by N_A is mandatory.

Q.17 Gas absorbs at 355 nm, emits at 680 nm and one unknown λ. Find the unknown.

Explanation

By conservation of energy, the absorbed photon's energy equals the sum of the two emitted photons' energies:

$$E_{abs} = E_1 + E_2 \quad \Rightarrow \quad \frac{hc}{\lambda_{abs}} = \frac{hc}{\lambda_1} + \frac{hc}{\lambda_2}$$

Divide through by hc:

$$\frac{1}{\lambda_{abs}} = \frac{1}{\lambda_1} + \frac{1}{\lambda_2}$$

$$\frac{1}{\lambda_2} = \frac{1}{355} - \frac{1}{680} = \frac{680 - 355}{355 \times 680} = \frac{325}{241400}$$

$$\lambda_2 = \frac{241400}{325} = 742.8 \text{ nm} \approx 743 \text{ nm}$$

Concept

This is a direct application of energy conservation. Since $E = hc/\lambda$, energy conservation $\Rightarrow$ $1/\lambda_{abs} = 1/\lambda_1 + 1/\lambda_2$. The hc cancels. Solve for the unknown λ_2 by simple arithmetic.

Answer

Option (1): 743 nm

Common Student Mistake

Students incorrectly apply conservation as $\lambda_{abs} = \lambda_1 + \lambda_2$, giving $\lambda_2 = 355 - 680 < 0$ (impossible). **Wavelengths do not add — energies add.** Since $E \propto 1/\lambda$, the correct conservation equation is $1/\lambda_{abs} = 1/\lambda_1 + 1/\lambda_2$.

Answer Key — DPP-3

Q	Ans	Q	Ans	Q	Ans	Q	Ans	Q	Ans
1	1	2	4	3	1	4	3	5	1
6	4	7	1	8	1	9	1	10	2
11	1	12	4	13	2	14	4	15	1
16	3	17	1						

"N_A se divide karna bhool gaye? 14 aur 15 dono miss ho jaate. Ek baar check karo unit match ho raha hai ya nahi."

— Weird Chemist