مراجعة نهاية الفصل في الحرارية ومعدلات التفاعل – ملاحظات وأسئلة اختبارات الوزارة



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المزيد من الملفات بحسب الصف الحادي عشر المتقدم والمادة كيمياء في الفصل الأول	
1	مراجعة نهاية الفصل في الحرارية ومعدلات التفاعل – ملاحظات وأسئلة اختبارات الوزارة
2	اختبار تجريبي 1 لنهاية الفصل – نسخة المعلم
3	هيكل امتحان نهاية الفصل مع 3 نماذج وإجابات المتقدمة (بريدج (محلول)
4	مراجعة هيكل اختبار نهاية الفصل
5	مراجعة هيكل وحدة الطاقة والتغير الكيميائي

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Grade 11 Advanced Chemistry

EOT Term 1 revision · thermochemistry and reaction rates

Notes, self-test questions and Ministry exam questions · click to show each answer

Quick notes

All the key ideas and formulas on two slides

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Quick notes (1/2)

Quick notes 1 · Thermochemistry

Heat and specific heat



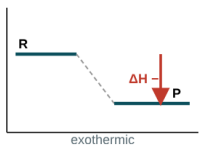
Conversion route: follow the green arrows to the right, the red arrows back to the left · K = °C + 273

- **K = °C + 273**
- **q = m c ΔT** · $\Delta T = T_{final} - T_{initial}$
- Calorimetry equation: **$-q_{lost} = q_{gained}$**
- c of H_2O = **4.184 J/g·°C**
- Specific heat: **c ↑ → ΔT ↓** · **c ↓ → ΔT ↑**

Enthalpy

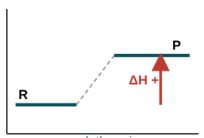
- **$\Delta H = H_p - H_r$**

- **Exo** = gives out heat · **ΔH = −**
- Heat flows from the **system → surroundings**
- Hot pack: heat flows from the pack (system) to the skin/hand (surroundings)



exothermic

- **Endo** → absorbs heat · **ΔH = +**
- Heat flows from the **surroundings → system**
- Cold pack: heat flows from the hand/skin (surroundings) to the pack (system)



endothermic

Changes of state

- **$\Delta H_{vap} = -\Delta H_{cond}$**
- **$\Delta H_{fusion} = -\Delta H_{solid}$**
- **q = n × ΔH** · n = m ÷ M

Change	State
Vaporization	l → g
Condensation	g → l
Fusion	s → l
Solidification	l → s

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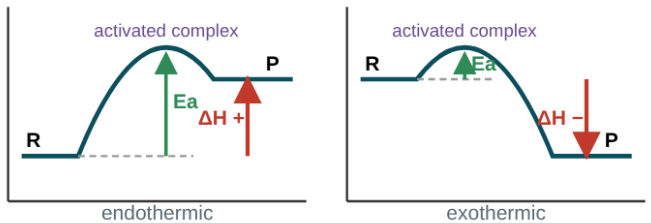
Quick notes 2 · Hess's law, ΔH°f and reaction rates

Hess's law and ΔH°f

- **$\Delta H_{total} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \ldots$**
- **ΔH°f** = forming 1 mol of compound from its elements (standard state)
- ΔH°f of an element = **0** (O_2, H_2, N_2, F_2, Hg(l), C(graphite))
- **$\Delta H_{rxn} = \Sigma \Delta H°f_{products} - \Sigma \Delta H°f_{reactants}$**

Reaction rate

- **Average rate = − Δ[Reactant] ÷ Δt** or **Δ[Product] ÷ Δt**
- **Effective collision** → collision + correct orientation + enough energy (≥ Ea)
- **Activated complex** = top of the curve, unstable transition state
- **Low Ea → faster reaction** · Ea ↑ → speed ↓



Rate law

- **Rate = k[A]**
- **Rate = $k[A]^m[B]^n$**
- **Overall order = m + n**
- Unit of k = **$M^{1 - \text{overall order}}$ s^{-1}** (M = mol L^{-1})
- **k = Rate ÷ ($[A]^m \cdot [B]^n$)**
- **[A] = $\sqrt[m]{\text{Rate} \div k[B]^n}$**
- **[B] = $\sqrt[n]{\text{Rate} \div k[A]^m}$**

When concentration doubles	Order
rate has no effect (constant)	zero order
rate doubles (× 2)	first order
rate quadruples (× 4)	second order
rate × 8	third order

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Module 14 · Lesson 1 · Heat and specific heat

Learning outcomes Q1–Q5 · notes, then 5 self-test questions per outcome

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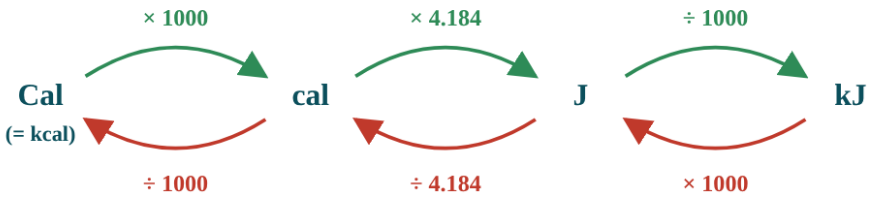
Notes · Q1 and Q2

Q1

Textbook + Table 1 + Example problem 1 + Practice problems · p. 5, 6

Perform interconversion between units of temperature and heat

- **Heat (q)** is energy transferred because of a temperature difference. The SI unit is the **joule (J)**; other units are cal, kJ and kcal.
- The nutritional **Calorie** (capital C) = 1 kcal = 1000 cal = 4184 J — the value printed on food labels.
- **Kelvin:** K = °C + 273. A temperature change ΔT is the same in K and in °C.



Conversion route: follow the green arrows to the right, the red arrows back to the left · K = °C + 273

Worked example

Q. How many joules does a snack containing 240 Cal supply? And how many kJ?

Answer: 240 Cal × 4184 J/Cal = **1.00×10^6 J ≈ 1004 kJ** (1 Cal = 1 kcal = 1000 cal).

Common mistakes: Confusing Cal (nutritional, = 1000 cal) with cal gives a 1000× error. · To convert °C to K add 273 — never multiply.

Q2

Textbook + Table 2 + Practice problems · p. 7, 8, 9

Describe the relationship between the specific heat capacity of a substance and the resistance to change in temperature

- **Specific heat (c)** is the heat needed to raise the temperature of 1 g of a substance by 1 °C; unit J/(g·°C).
- Low c → the temperature changes quickly (metals such as Pb and Au = 0.129).
- Same heat, equal masses: the smaller c gives the larger ΔT (ΔT ∝ 1/c).

c ↑ → ΔT ↓ (heats and cools slowly) · **c ↓ → ΔT ↑** (heats and cools quickly)

Substance	c in J/(g·°C)
Water (l)	4.184
Beryllium	1.825
Magnesium	1.023
Aluminum	0.897
Calcium	0.647
Iron	0.449
Copper	0.385
Silver	0.235
Lead / Gold	0.129

Worked example

Q. Equal masses of Au (0.129), Cu (0.385), Fe (0.449) and Al (0.897) receive the same heat. Which warms up the most?

Answer: Gold: it has the lowest specific heat, so it has the largest ΔT.

Common mistakes: High specific heat does not mean high temperature — it means a large heat capacity with only a small change in T.

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Self-test · Q1 and Q2

Q1 · Perform interconversion between units of temperature and heat

1. A breakfast contains 170 Cal. How many joules of energy does it supply?

   A 7.11×10^5 J B 1.70×10^5 J C 711 J D 170 J

Answer: A · 7.11×10^5 J

2. Y contains 150 Cal and X contains 5.0×10^5 J. Which statement is true?

   A The energy in X is greater than in Y B The energy in Y is 3.6×10^5 J
   C The energy in X is 355 Cal D The energy in Y is greater than in X

Answer: D · The energy in Y is greater than in X

3. Which of the following contains the greatest energy?

   A 10 Cal B 9600 J C 1000 cal D 86.5 kJ

Answer: D · 86.5 kJ

4. An exothermic reaction released 75.0 kJ of energy. How much is this in kcal?

   A 17.9 kcal B 29.7 kcal C 20.6 kcal D 21.5 kcal

Answer: A · 17.9 kcal

5. A water sample is at 25 °C. What is its temperature in kelvin?

   A 248 K B −248 K C 298 K D 25 K

Answer: C · 298 K

Q2 · Describe the relationship between the specific heat capacity of a substance and the resistance to change in temperature

1. Equal masses of water (4.184), calcium (0.647), concrete (0.84) and lead (0.129) J/(g·°C) absorb equal heat. Which shows the greatest temperature increase?

   A Water B Lead C Calcium D Concrete

Answer: B · Lead

2. Equal masses of X and Y receive the same heat. The temperature of X rises by 20 °C and that of Y by 70 °C. Which is correct?

   A Y has a higher specific heat than X B X and Y have the same specific heat
   C X has a higher specific heat than Y D Nothing can be concluded

Answer: C · X has a higher specific heat than Y

3. Equal masses of Ca (0.647), Sr (0.301), Pb (0.129) and Mg (1.023) J/(g·°C) sit in the Sun for the same time. What is the order of temperature increase from smallest to largest?

   A Ca, Mg, Pb, Sr B Mg, Ca, Sr, Pb C Pb, Sr, Ca, Mg D Ca, Sr, Mg, Pb

Answer: B · Mg, Ca, Sr, Pb

4. 1.00 g each of Fe (0.449), Pb (0.129) and Zn (0.388) at 100 °C are dropped into three beakers with 10.0 g of water at 25 °C. What is the order of the final water temperatures, lowest to highest?

   A Fe < Zn < Pb B Pb < Zn < Fe C Zn < Pb < Fe D Pb < Fe < Zn

Answer: B · Pb < Zn < Fe

5. Why do coastal cities have milder temperatures than desert cities?

   A Water has a high specific heat, so it resists temperature change
   B Water has a low specific heat
   C Sand has a high specific heat
   D Water absorbs no heat

Answer: A

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Notes · Q3 and Q4

Q3

Textbook + Example problem 2 + Practice problems · p. 7, 8, 9

Calculate the amount of heat absorbed by a substance as its temperature changes

- **$q = m \times c \times \Delta T$** with m in grams, c the specific heat, and $\Delta T = T(\text{final}) - T(\text{initial})$.
- On heating ΔT is positive, so q is positive (heat absorbed).
- Units: m must be in grams (kg × 1000) and q is in joules unless told otherwise.

Worked example

Q. How much heat is absorbed by 5.50 g of iron (c = 0.449 J/g·°C) heated from 25.0 to 95.0 °C?

Answer: $\Delta T = 95.0 - 25.0 = 70.0$ °C · $q = 5.50 \times 0.449 \times 70.0 =$ **173 J**

Common mistakes: Forgetting to convert kg to g, or using the final T instead of ΔT.

Q4

Textbook + Example problem 2 + Practice problems · p. 7, 8, 9

Calculate the amount of heat released by a substance as its temperature changes

- On cooling ΔT is negative, so q is negative: the substance releases heat.
- The same law $q = m \cdot c \cdot \Delta T$ applies; the amount of heat released is often quoted as a positive magnitude.
- Conservation of energy: heat lost by the hot body = heat gained by the cold body: **$-q(\text{hot}) = q(\text{cold})$** (calorimetry).
- In a calorimeter: if the final temperature **increases** → the reaction is **exothermic** and q(rxn) is **negative (−)**; if the final temperature **decreases** → the reaction is **endothermic** and q(rxn) is **positive (+)**.
- The sign of q is used to find the temperature change: $q = m \cdot c \cdot \Delta T$, so $\Delta T = q \div (m \cdot c)$ and $T(\text{final}) = T(\text{initial}) + \Delta T$.

Worked example

Q. How much heat is released when 200 g of copper (c = 0.385 J/g·°C) cools from 150 °C to 25 °C?

Answer: $\Delta T = 25 - 150 = -125$ °C · $q = 200 \times 0.385 \times (-125) = -9625$ J → **9625 J released**

Common mistakes: Ignoring the negative sign when q (not 'heat released') is asked.

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Self-test · Q3 and Q4

Q3 · Calculate the amount of heat absorbed by a substance as its temperature changes

1. How much heat is absorbed when a 3.00 kg granite boulder warms from 10.0 °C to 45.0 °C? (c = 0.803 J/g·°C)

   A 84 J B 1.1×10^5 J C 11 J D 8.4×10^4 J

   Answer: D · 8.4×10^4 J

2. A 15.5 g piece of aluminum at 30.5 °C absorbs 345 J. What is its final temperature? (c = 0.897 J/g·°C)

   A 5.69 °C B 60.5 °C C 55.3 °C D 8.45 °C

   Answer: C · 55.3 °C

3. The same heat is supplied to 47.0 g of Al (c = 0.897) and to a sample of Fe (c = 0.449), and both warm by the same number of degrees. What is the mass of the iron?

   A 93.9 g B 45.6 g C 35.5 g D 53.5 g

   Answer: A · 93.9 g

4. How much heat is needed to warm 50.0 g of water from 20.0 °C to 30.0 °C? (c = 4.184 J/g·°C)

   A 209 J B 4184 J C 6276 J D 2092 J

   Answer: D · 2092 J

5. A sample is heated from 25 °C to 60 °C. Which statement about q is correct?

   A ΔT is positive, so q is positive (heat absorbed)
   B ΔT is positive, so q is negative
   C ΔT is negative, so q is positive
   D q is zero because the mass is constant

   Answer: A

Q4 · Calculate the amount of heat released by a substance as its temperature changes

1. 250 g of ethanol at 75.0 °C loses 4655 J. What is its final temperature? (c = 2.44 J/g·°C)

   A 82.6 °C B 59.5 °C C 45.8 °C D 67.4 °C

   Answer: D · 67.4 °C

2. How much heat is released when 150 g of aluminum cools from 90.0 °C to 25.0 °C? (c = 0.897 J/g·°C)

   A 1.21×10^4 J B 3.36×10^3 J C 8.75×10^3 J D 5.83×10^3 J

   Answer: C · 8.75×10^3 J

3. 100 g of water at 80.0 °C releases 8368 J. What is its final temperature? (c = 4.184 J/g·°C)

   A 40.0 °C B 20.0 °C C 100 °C D 60.0 °C

   Answer: D · 60.0 °C

4. A hot metal loses 540 J of heat to cold water in an insulated cup. How much heat does the water gain?

   A −540 J B 270 J C 1080 J D 540 J

   Answer: D · 540 J

5. When a substance cools, which statement is correct?

   A ΔT is negative and q is positive B ΔT is positive and q is negative
   C ΔT is positive and q is positive D ΔT is negative and q is negative

   Answer: D · ΔT is negative and q is negative

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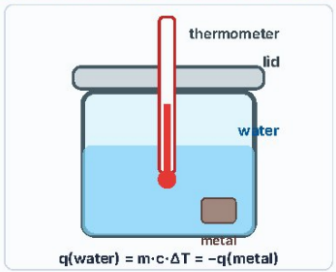
Textbook + Table 2 + Example problem 2 + Practice problems · p. 8, 9

Q5 Calculate the specific heat capacity of a sample given its mass and temperature change

- Rearrange: **c = q / (m × ΔT)**, then compare with the table of specific heats to identify the substance.
- Calorimetry equation: **– q(metal) = q(water)**, i.e. – m(metal)·c(metal)·ΔT(metal) = m(water)·4.184·ΔT(water).
- ΔT of the metal is negative (it cools) and ΔT of the water is positive (it warms), so both sides come out positive.



Worked example

Q. An 11.8 g metal at 108 °C is dropped into 125 g of water at 22.3 °C; the final temperature is 24.0 °C. Find c of the metal.

Answer: – q(metal) = q(water) → – 11.8 × c × (24.0 – 108) = 125 × 4.184 × (24.0 – 22.3) → 11.8 × c × 84.0 = 889 J → c = 889 ÷ 991 = **0.897 J/g·°C (Al)**

Common mistakes: Mixing up ΔT of the metal and of the water, or using the wrong mass with 4.184.

Self-test · Q5

Q5 · Calculate the specific heat capacity of a sample given its mass and temperature change

1. A 360 g hot alloy is placed in 425 g of cold water. The alloy temperature falls by 205 °C and the water rises by 18.7 °C. What is c of the alloy? (c water = 4.184)

A 0.129 J/g·°C B 0.235 J/g·°C C 0.380 J/g·°C D 0.450 J/g·°C

Answer: D · 0.450 J/g·°C

2. A 355 g sample warms from 22.4 °C to 43.6 °C and absorbs 6.75 kJ. Which metal is it? (Au 0.129, Ag 0.235, Fe 0.449, Al 0.897)

A Aluminum B Iron C Silver D Gold

Answer: A · Aluminum

3. A 25.0 g alloy nail absorbs 250 J and warms from 25.0 °C to 78.0 °C. What is its specific heat?

A 0.19 J/g·°C B 5.30 J/g·°C C 1.19 J/g·°C D 0.89 J/g·°C

Answer: A · 0.19 J/g·°C

4. 20.0 g of a metal absorbs 180 J and its temperature rises by 25.0 °C. What is c?

A 0.36 J/g·°C B 2.8 J/g·°C C 0.036 J/g·°C D 9.0 J/g·°C

Answer: A · 0.36 J/g·°C

5. A 50.0 g metal at 100.0 °C is dropped into 50.0 g of water at 20.0 °C. The final temperature is 27.8 °C. What is c of the metal?

A 0.90 J/g·°C B 0.45 J/g·°C C 0.23 J/g·°C D 1.8 J/g·°C

Answer: B · 0.45 J/g·°C

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Module 14 · Lesson 2 · Enthalpy and changes of state

Learning outcomes Q6–Q15 · notes, then 5 self-test questions per outcome

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Notes · Q6 and Q7

Q6

Textbook · p. 14

Define thermochemistry

- **Thermochemistry** is the study of the heat (energy) changes that accompany chemical reactions and physical changes such as changes of state.
- **Energy** is the capacity to do work or produce heat. Chemical potential energy is the energy stored in chemical bonds.
- **Law of conservation of energy:** energy is neither created nor destroyed; it changes form or moves between systems.
- Breaking bonds absorbs energy and forming bonds releases energy; the difference is the heat released or absorbed.

Remember: Thermochemistry = heat changes in reactions and physical changes

Worked example

Q. Burning methane releases heat and melting ice absorbs heat. Which branch of chemistry studies these effects?

Answer: Thermochemistry.

Common mistakes: Thermochemistry does not study reaction speed (that is kinetics) — it studies the amount of energy transferred.

Q7

Textbook + Figure 7 · p. 14

Differentiate among the system, the surroundings, and the universe

- **System:** the part being studied (reactants and products). **Surroundings:** everything else around it (beaker, water, air, your hand).
- **Universe** = system + surroundings.
- **Exothermic:** heat flows from the system to the surroundings, so the surroundings get warmer (hot pack).
- **Endothermic:** heat flows from the surroundings to the system, so the surroundings get colder (cold pack; the $Ba(OH)_2$ beaker freezes to a wet board).

Remember: Universe = System + Surroundings · Exo: system → surroundings · Endo: surroundings → system

Worked example

Q. In a cold pack NH_4NO_3 dissolves in water and the pack feels cold. Identify the system, the surroundings and the direction of heat flow.

Answer: System = NH_4NO_3 and water inside the pack · surroundings = the pack wall, your hand and the air · heat flows **from the surroundings into the system** (endothermic).

Common mistakes: Do not mix up what the hand feels (the surroundings) with what happens to the system.

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Q6 · Define thermochemistry

1 **Thermochemistry is the study of...**

A the speed of chemical reactions
B the shapes of molecules
C heat changes in chemical reactions and changes of state
D the masses of atoms

Answer: C

2 **Where is chemical potential energy stored?**

A Only in the nucleus of atoms
B In the chemical bonds of substances
C In the temperature of the surroundings
D In the container walls

Answer: B · In the chemical bonds of substances

3 **Which statement expresses the law of conservation of energy?**

A Energy cannot be created or destroyed, only transformed or transferred
B Energy is destroyed in exothermic reactions
C Energy is created in endothermic reactions
D Heat can be converted to mass only

Answer: A

4 **Which of the following is NOT studied in thermochemistry?**

A Heat released when fuel burns
B How fast a reaction proceeds
C Heat absorbed when ice melts
D Heat of dissolving a salt

Answer: B · How fast a reaction proceeds

5 **During a chemical reaction, energy is absorbed when...**

A bonds in the products are formed
B the reaction is cooled
C bonds in the reactants are broken
D the container is closed

Answer: C · bonds in the reactants are broken

Q7 · Differentiate among the system, the surroundings, and the universe

1 **Which relationship between system, surroundings and universe is correct?**

A Universe = system − surroundings
B System = universe + surroundings
C Surroundings = system
D Universe = system + surroundings

Answer: D · Universe = system + surroundings

2 **In a hot pack (iron reacting with oxygen) heat is released. Which is correct?**

A Heat flows from the surroundings to the system
B No heat flows
C The pack is the surroundings and the hand is the system
D Heat flows from the system to the surroundings

Answer: D

3 **Barium hydroxide and ammonium thiocyanate are mixed in a beaker placed on a wet board; the beaker sticks to the board. What explains this?**

A An endothermic reaction occurs; heat flows from the wet board and water into the beaker
B An exothermic reaction occurs; heat flows from the beaker to the board
C The mixture releases enough energy to freeze the water
D The mixture is a catalyst for evaporation

Answer: A

4 **In an endothermic reaction in a beaker, the temperature of the surroundings...**

A increases
B stays constant
C first increases then decreases
D decreases

Answer: D · decreases

5 **Which of the following is part of the surroundings when NaOH dissolves in water in a beaker?**

A The beaker and the air around it
B The NaOH only
C The dissolved ions
D The water–NaOH solution

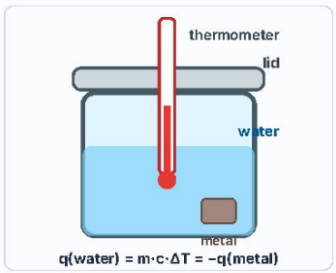
Answer: A · The beaker and the air around it

Q8

Textbook + Figure 7 · p. 14

Perform and conduct an experiment to determine the enthalpy of a chemical reaction

- A **calorimeter** is an insulated vessel (foam cup) with a known mass of water/solution and a thermometer; the reaction is run inside and ΔT is measured.
- Heat gained by the solution: q(soln) = m·c·ΔT, and **q(rxn) = − q(soln)**.
- Temperature rises → exothermic (ΔH negative); temperature falls → endothermic (ΔH positive).
- ΔH = q(rxn) / n in kJ per mole; convert J to kJ.



Worked example

Q. 50.0 mL of 1.0 M HCl is mixed with 50.0 mL of 1.0 M NaOH in a calorimeter; T rises by 6.5 °C (total mass 100 g, c = 4.184). Calculate ΔH in kJ/mol.

Answer: q(soln) = 100 × 4.184 × 6.5 = 2720 J = 2.72 kJ · n = 0.0500 L × 1.0 M = 0.050 mol · ΔH = −2.72 ÷ 0.050 = **−54 kJ/mol**

Common mistakes: Forgetting that q(rxn) has the opposite sign of q(soln), or forgetting to divide by moles.

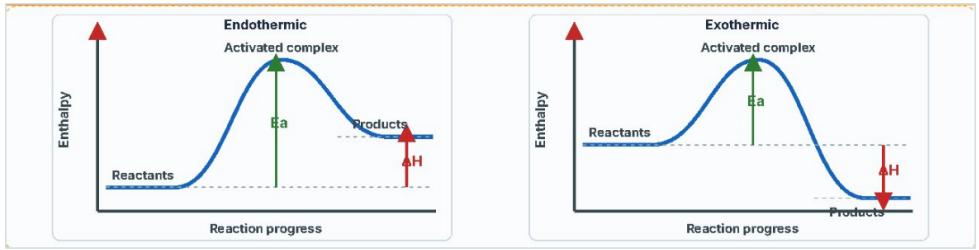
Q9

Textbook + Figures 8, 9 · p. 15, 16, 17

Compare and contrast potential energy diagrams of exothermic and endothermic reactions in terms of general shape, enthalpy of reactants and products, activation energy of forward and backward reactions, and enthalpy of reaction and its sign

- **Exothermic:** products lower than reactants, ΔH < 0, heat released. **Endothermic:** products higher, ΔH > 0, heat absorbed.
- Both have a peak (activated complex). Ea(forward) = peak − reactants; Ea(reverse) = peak − products.
- ΔH = H(products) − H(reactants) = Ea(fwd) − Ea(rev). Exo: Ea(rev) > Ea(fwd); Endo: Ea(fwd) > Ea(rev).
- Heat pack: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ (exothermic, ΔH = −1625 kJ); cold pack: $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$ (endothermic, ΔH = +27 kJ).

Remember: ΔH = H(products) − H(reactants) = Ea(fwd) − Ea(rev)



Worked example

Q. A reaction has Ea(forward) = 50 kJ and Ea(reverse) = 80 kJ. Find ΔH and the type of reaction.

Answer: ΔH = 50 − 80 = **−30 kJ → exothermic**

Common mistakes: ΔH is measured between reactants and products, not to the peak — the peak belongs to Ea.

Self-test · Q8

Q8 · Perform and conduct an experiment to determine the enthalpy of a chemical reaction

1 **In a coffee-cup calorimeter the temperature of the solution increases. What is concluded?**

- **A** The reaction is endothermic and ΔH is negative
- **B** The reaction is exothermic and ΔH is positive
- **C** The reaction is endothermic and ΔH is positive
- **D** The reaction is exothermic and ΔH is negative

Answer: D

2 **A reaction in 100 g of water raises its temperature by 6.5 °C (c = 4.184 J/g·°C). How much heat did the water gain?**

A 2.72 J **B** 27.2 J **C** 6.5×10^2 J **D** 2.72×10^3 J

Answer: D · 2.72×10^3 J

3 **2.72 kJ of heat is released by a reaction of 0.050 mol of acid. What is ΔH per mole?**

A −54.4 kJ/mol **B** +54.4 kJ/mol **C** −0.136 kJ/mol **D** −136 kJ/mol

Answer: A · −54.4 kJ/mol

4 **Why is a calorimeter insulated?**

- **A** To minimize heat exchange with the surroundings
- **B** To speed up the reaction
- **C** To increase the mass of water
- **D** To make c of water larger

Answer: A

5 **A solid dissolves in water in a calorimeter and the temperature falls from 25.0 °C to 21.0 °C. Which statement is correct?**

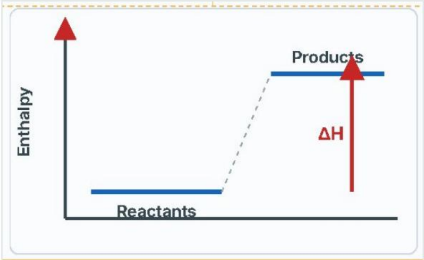
- **A** The process is exothermic and q(rxn) is positive
- **B** The process is endothermic and q(rxn) is negative
- **C** The process is endothermic and q(rxn) is positive
- **D** The process is exothermic and q(rxn) is negative

Answer: C

Self-test · Q9

Q9 · Compare and contrast potential energy diagrams of exothermic and endothermic reactions in terms of general shape, enthalpy of reactants and products, activation energy of forward and backward reactions, and enthalpy of

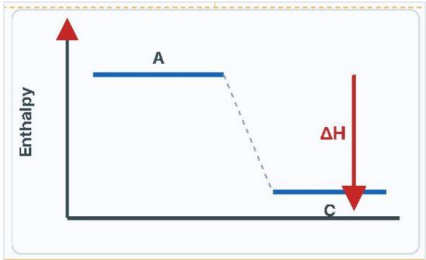
1 Reactants lie at a low level and products at a higher level. Which is correct for this reaction?



A The system absorbs heat from the surroundings
B The system releases heat to the surroundings
C The enthalpy of the system always decreases
D No energy change occurs

Answer: A

2 In the diagram A → C is shown (A at the top, C at the bottom). Which statements are correct?
I. ΔH > 0
II. H(products) < H(reactants)
III. Heat flows from the system to the surroundings



A I only B I and II only C II and III only D I and III only

Answer: C · II and III only

3 Which reaction does NOT agree with its enthalpy diagram?

A $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, $\Delta H = -1625$ kJ ; products drawn lower
B $NO + O_3 \rightarrow NO_2 + O_2 + 199$ kJ ; products drawn higher than reactants
C $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$, $\Delta H = +27$ kJ ; products drawn higher
D $A + B \rightarrow C + 394$ kJ ; products drawn lower

Answer: B

4 Which is true regarding the diagram of $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ (products lower than reactants)?

A The sign of the enthalpy change is positive
B The reaction is used in the heat pack
C The reaction is endothermic
D The reaction is used in the cold pack

Answer: B · The reaction is used in the heat pack

5 A reaction has Ea(forward) = 120 kJ and Ea(reverse) = 85 kJ. What is ΔH and the type of reaction?

A −35 kJ ; exothermic B +35 kJ ; endothermic
C +205 kJ ; endothermic D −205 kJ ; exothermic

Answer: B · +35 kJ ; endothermic

Q10

Textbook + Figures 8, 9 · p. 15, 16, 17

Define enthalpy change of reaction (ΔHrxn) while writing the equation used

- **Enthalpy (H)** is the heat content of a system at constant pressure; we can only measure the change ΔH.
- **ΔH(rxn) = H(products) − H(reactants)** = heat absorbed or released at constant pressure.
- ΔH < 0 → exothermic; ΔH > 0 → endothermic. Unit: kJ (or kJ/mol).

Remember: ΔH(rxn) = H(products) − H(reactants) · ΔH < 0 exo · ΔH > 0 endo



Worked example

Q. H(reactants) = 450 kJ and H(products) = 380 kJ. Find ΔH and the type.

Answer: ΔH = 380 − 450 = **−70 kJ → exothermic**

Common mistakes: Subtracting in the wrong order flips the sign: always products − reactants.

Q11

Textbook + Example problem 4 + Practice problems · p. 19, 21

Perform calculations using the enthalpy of combustion

- **Enthalpy of combustion** is the heat released when 1 mol of a substance burns completely in oxygen (always negative).
- Steps: grams → moles (n = m/M) → q = n × |ΔHcomb|; to find mass: m = (q/|ΔH|) × M.
- Heat released is quoted as a positive kJ value; ΔH carries the negative sign.

Remember: n = m ÷ M · q = n × |ΔH°comb| · m = (q ÷ |ΔH°comb|) × M

Substance	ΔH°comb (kJ/mol)
Methane $CH_4(g)$	−891
Propane $C_3H_8(g)$	−2219
Octane $C_8H_{18}(l)$	−5471
Glucose $C_6H_{12}O_6(s)$	−2808
Sucrose $C_{12}H_{22}O_{11}(s)$	−5644

Worked example

Q. What mass of CH_4 (M = 16.0 g/mol) must burn to release 3564 kJ? (ΔHcomb = −891 kJ/mol)

Answer: n = 3564 ÷ 891 = 4.00 mol · m = 4.00 × 16.0 = **64 g**

Common mistakes: Skipping the grams-to-moles step, or using the equation coefficient instead of 'per 1 mol'.

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Self-test · Q10 and Q11

Q10 · Define enthalpy change of reaction (ΔHrxn) while writing the equation used

1 Enthalpy (H) is best defined as...

A the total energy stored in the bonds of reactants and products
B the energy required to break all bonds in one mole
C the energy change at constant volume
D the heat content of a system at constant pressure

Answer: D

2 Which equation gives the enthalpy change of a reaction?

A ΔHrxn = Hreactants − Hproducts
B ΔHrxn = Hproducts − Hreactants
C ΔHrxn = Hreactants + Hproducts
D ΔHrxn = Hproducts ÷ Hreactants

Answer: B · ΔHrxn = Hproducts – Hreactants

3 For an endothermic reaction (products above reactants in the diagram) which is correct?

A ΔHrxn > 0
B ΔHrxn < 0
C ΔHrxn = 0
D ΔHrxn = Hreactants + Hproducts

Answer: A · ΔHrxn > 0

4 H(reactants) = 450 kJ and H(products) = 380 kJ. What is ΔHrxn?

A +70 kJ
B −70 kJ
C +830 kJ
D −830 kJ

Answer: B · −70 kJ

5 A reaction has ΔH = +120 kJ. Which statement is correct?

A It is endothermic and the products have more enthalpy than the reactants
B It is exothermic and the products have less enthalpy
C It is exothermic and the products have more enthalpy
D It is endothermic and the products have less enthalpy

Answer: A

Q11 · Perform calculations using the enthalpy of combustion

1 What mass of sucrose $C_{12}H_{22}O_{11}$ (342.3 g/mol, ΔHcomb = −5644 kJ/mol) must be burned to release 8466 kJ?

A 228.2 g
B 513.5 g
C 684.6 g
D 171.1 g

Answer: B · 513.5 g

2 How much heat is released by burning 154 g of CH_4 (16.04 g/mol)?

A 4.50×10^3 kJ
B 5.20×10^3 kJ
C 6.70×10^3 kJ
D 8.55×10^3 kJ

Answer: D · 8.55×10^3 kJ

3 How much heat is released by the combustion of 50 g of hydrogen (2.0 g/mol)? $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 286$ kJ

A 29458 kJ
B 7150 kJ
C 14300 kJ
D 572 kJ

Answer: B · 7150 kJ

4 What mass of butane C_4H_{10} (58.12 g/mol, ΔHcomb = −2657 kJ/mol) must burn to release 18599 kJ?

A 145.3 g
B 87.20 g
C 43.60 g
D 406.8 g

Answer: D · 406.8 g

5 Burning 0.5 mol of a compound releases 2822 kJ. Which compound is it? (Sucrose −5644, Propane −2219, Glucose −2808, Octane −5471 kJ/mol)

A Propane
B Glucose
C Sucrose
D Octane

Answer: C · Sucrose

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Notes · Q12 and Q13

Textbook + Table 3 · p. 18, 19

Q12 Write a thermochemical equation for the combustion of a compound, while including the sign of ΔH and identifying whether the reaction is exothermic or endothermic

- Combustion is a reaction with O_2; hydrocarbons give CO_2 and H_2O. It is always exothermic (ΔH < 0).
- A **thermochemical equation** is balanced, shows states, and gives ΔH — either ΔH = −x kJ, or + x kJ on the products side (exothermic).
- In a molar combustion equation the fuel has coefficient 1 (fractions are allowed for O_2).

Remember: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ ΔH = −891 kJ · or $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ + 891 kJ

Substance	ΔH°comb (kJ/mol)
Methane $CH_4(g)$	−891
Propane $C_3H_8(g)$	−2219
Octane $C_8H_{18}(l)$	−5471
Glucose $C_6H_{12}O_6(s)$	−2808
Sucrose $C_{12}H_{22}O_{11}(s)$	−5644

Worked example

Q. Write the thermochemical equation for the combustion of 1 mol of octane (ΔHcomb = −5471 kJ/mol).

Answer: $C_8H_{18}(l) + 25/2\ O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$ ΔH = −5471 kJ

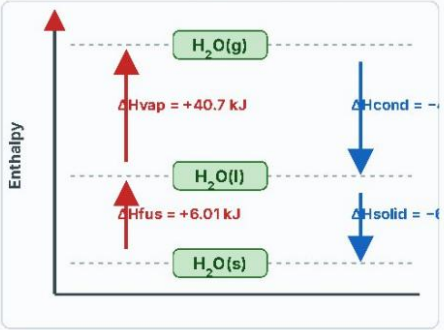
Common mistakes: Writing '+ kJ' on the reactant side for a combustion is wrong; released heat belongs with the products.

Textbook + Table 4 + Figure 10 · p. 19, 20

Q13 Define molar enthalpy (heat) of vaporization, ΔHvap · Define molar enthalpy (heat) of condensation, ΔHcond

- **ΔHvap:** heat absorbed to vaporize 1 mol of liquid at its boiling point — positive (endothermic). Water: +40.7 kJ/mol.
- **ΔHcond:** heat released when 1 mol of vapor condenses — negative (exothermic). Water: −40.7 kJ/mol.
- The processes are reverses: **ΔHcond = − ΔHvap**.

Remember: $H_2O(l) \rightarrow H_2O(g)$ ΔHvap = +40.7 kJ · $H_2O(g) \rightarrow H_2O(l)$ ΔHcond = −40.7 kJ



Worked example

Q. ΔHvap of ethanol = +38.6 kJ/mol. What are ΔHcond and the type of process?

Answer: ΔHcond = **−38.6 kJ/mol** → condensation is **exothermic**.

Common mistakes: Mixing up vaporization (endothermic, +) and condensation (exothermic, −).

Self-test · Q12

Q12 · Write a thermochemical equation for the combustion of a compound, while including the sign of ΔH and identifying whether the reaction is exothermic or endothermic.

1 **Which equation is a molar combustion equation?**

A $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

B $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

C $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

D $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Answer: B · $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

2 **Which equation does NOT represent a combustion reaction?**

A $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + 891\ kJ$

B $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 286\ kJ$

C $NH_4NO_3(s) + 27\ kJ \rightarrow NH_4^+(aq) + NO_3^-(aq)$

D $C_8H_{18}(l) + 25/2\ O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l) + 5471\ kJ$

Answer: C · $NH_4NO_3(s) + 27\ kJ \rightarrow NH_4^+(aq) + NO_3^-(aq)$

3 **Which of the following processes is exothermic? A: $NH_3(g) \rightarrow NH_3(l)$ B: $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$ C: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O(l)$ D: $H_2O(s) \rightarrow H_2O(l)$**

A A only B A and C C B and D D B only

Answer: B · A and C

4 **The combustion of propane has $\Delta H°comb = -2219$ kJ/mol. Which thermochemical equation is correct?**

A $C_3H_8(g) + 5O_2(g) + 2219\ kJ \rightarrow 3CO_2(g) + 4H_2O(l)$

B $3CO_2(g) + 4H_2O(l) \rightarrow C_3H_8(g) + 5O_2(g) + 2219\ kJ$

C $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l) + 2219\ kJ$

D $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)\ \Delta H = +2219\ kJ$

Answer: C

5 **The reaction $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ has $\Delta H = -891$ kJ. Which description is correct?**

A Exothermic; 891 kJ is released to the surroundings

B Endothermic; 891 kJ is absorbed from the surroundings

C Exothermic; 891 kJ is absorbed

D Endothermic; 891 kJ is released

Answer: A

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Self-test · Q13

Q13 · Define molar enthalpy (heat) of vaporization, ΔHvap · Define molar enthalpy (heat) of condensation, ΔHcond

1 **The molar enthalpy of vaporization (ΔHvap) is the heat…**

A released when 1 mol of a liquid vaporizes
B absorbed when 1 mol of a liquid vaporizes
C absorbed when 1 mol of a solid melts
D released when 1 g of vapor condenses

Answer: B · absorbed when 1 mol of a liquid vaporizes

2 **ΔHvap of ethanol is +38.6 kJ/mol. What is ΔHcond of ethanol?**

A +38.6 kJ/mol B −19.3 kJ/mol C −38.6 kJ/mol D 0 kJ/mol

Answer: C · −38.6 kJ/mol

3 **Which thermochemical equation represents the vaporization of water (ΔHvap = +40.7 kJ/mol)?**

A $H_2O(g) \rightarrow H_2O(l)$ ΔH = +40.7 kJ
B $H_2O(l) \rightarrow H_2O(g)$ ΔH = −40.7 kJ
C $H_2O(s) \rightarrow H_2O(g)$ ΔH = +40.7 kJ
D $H_2O(l) \rightarrow H_2O(g)$ ΔH = +40.7 kJ

Answer: D · $H_2O(l) \rightarrow H_2O(g)$ ΔH = +40.7 kJ

4 **Which process has a negative enthalpy change?**

A Boiling of water
B Evaporation of ethanol
C Condensation of steam
D Sublimation of dry ice

Answer: C · Condensation of steam

5 **Which statement about ΔHvap and ΔHcond of water is correct?**

A They have the same value and the same sign
B They have the same numerical value but opposite signs
C Both are positive
D Both are negative

Answer: B

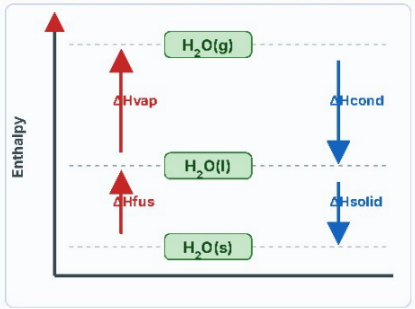
Q14

Textbook + Table 4 + Figure 10 · p. 19, 20

Define molar enthalpy (heat) of fusion, ΔHfus · Define molar enthalpy (heat) of solidification, ΔHsolid

- **ΔHfus:** heat absorbed to melt 1 mol of solid at its melting point — positive (endothermic). Water: +6.01 kJ/mol.
- **ΔHsolid:** heat released when 1 mol of liquid freezes — negative (exothermic). Water: −6.01 kJ/mol.
- **ΔHsolid = − ΔHfus.** Example: farmers spray orchards with water; as it freezes it releases heat that protects the fruit.

Remember: $H_2O(s) \rightarrow H_2O(l)$ ΔHfus = +6.01 kJ · $H_2O(l) \rightarrow H_2O(s)$ ΔHsolid = −6.01 kJ



Worked example

Q. ΔHfus of ethanol = +4.93 kJ/mol. Write the thermochemical equation for freezing ethanol.

Answer: $C_2H_5OH(l) \rightarrow C_2H_5OH(s)$ ΔHsolid = −4.93 kJ

Common mistakes: Melting and freezing occur at constant temperature — the heat changes the state, not the temperature.

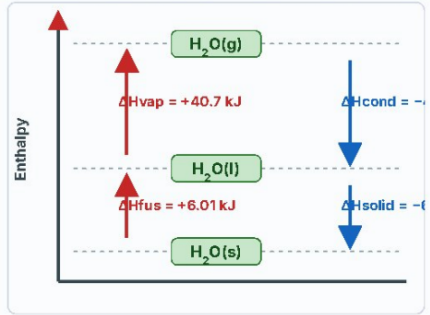
Q15

Textbook + Figure 10 + Practice problems · p. 20, 21

Perform phase-change calculations using enthalpy of fusion, vaporization, condensation, and solidification

- **q = n × ΔH(phase)** where n = m / M; choose the right ΔH: vap / cond / fus / solid.
- Sign: vaporization and melting absorb heat (+); condensation and freezing release heat (−).
- The temperature stays constant during a phase change, so q = m·c·ΔT is not used here.

Remember: n = m ÷ M · q = n × ΔH(phase) · H_2O: ΔHvap = 40.7 · ΔHfus = 6.01 kJ/mol



Worked example

Q. How much heat is released when 4.50 g of water freezes (M = 18.02 g/mol)?

Answer: n = 4.50 ÷ 18.02 = 0.250 mol · q = 0.250 × 6.01 = **1.50 kJ released** (ΔH = −1.50 kJ)

Common mistakes: Using ΔHvap instead of ΔHfus, or forgetting grams → moles.

Self-test · Q14 and Q15

Q14 · Define molar enthalpy (heat) of fusion, ΔHfus · Define molar enthalpy (heat) of solidification, ΔHsolid

1 Which of the following processes are exothermic? I: vaporization of water II: melting of ice III: condensation of water vapor IV: freezing of water

A I and II B I and III C III and IV D II and IV

Answer: C · III and IV

2 Which processes are endothermic?

1: $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$

2: $NH_3(g) \rightarrow NH_3(l)$

3: $NaCl(s) \rightarrow NaCl(l)$

4: $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

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

Answer: D · 1 and 3

3 ΔHfus of ethanol is +4.93 kJ/mol. What is ΔHsolid?

A +4.93 kJ/mol B −4.93 kJ/mol C −9.86 kJ/mol D +2.47 kJ/mol

Answer: B · −4.93 kJ/mol

4 Which thermochemical equation represents the freezing of water?

A $H_2O(s) \rightarrow H_2O(l)$ ΔH = −6.01 kJ B $H_2O(l) \rightarrow H_2O(s)$ ΔH = −6.01 kJ

C $H_2O(l) \rightarrow H_2O(s)$ ΔH = +6.01 kJ D $H_2O(l) \rightarrow H_2O(g)$ ΔH = −6.01 kJ

Answer: B · $H_2O(l) \rightarrow H_2O(s)$ ΔH = −6.01 kJ

5 Which of the following is endothermic?

A $Br_2(l) \rightarrow Br_2(s)$ B $NH_3(g) \rightarrow NH_3(l)$

C $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + 286$ kJ D $C_2H_5OH(s) \rightarrow C_2H_5OH(l)$

Answer: D · $C_2H_5OH(s) \rightarrow C_2H_5OH(l)$

Q15 · Perform phase-change calculations using enthalpy of fusion, vaporization, condensation, and solidification

1 How much heat is released when 29.5 g of water freezes on an orange? (18.02 g/mol, ΔHsolid = −6.01 kJ/mol)

A 66.6 kJ B 9.84 kJ C 3195 kJ D 21636 kJ

Answer: B · 9.84 kJ

2 How much heat is absorbed when 9.01 g of water vaporizes? (18.02 g/mol, ΔHvap = +40.7 kJ/mol)

A 20.35 kJ B 3.00 kJ C 366.7 kJ D 81.40 kJ

Answer: A · 20.35 kJ

3 A certain amount of water needs 173 kJ to evaporate. What is its mass? (18.02 g/mol, ΔHvap = 40.7 kJ/mol)

A 76.6 g B 58.6 g C 45.1 g D 27.1 g

Answer: A · 76.6 g

4 How much heat is needed to melt 36.0 g of ice? (18.02 g/mol, ΔHfus = 6.01 kJ/mol)

A 12.0 kJ B 6.01 kJ C 81.3 kJ D 216 kJ

Answer: A · 12.0 kJ

5 How much heat is absorbed when 2.30 g of ethanol (46.07 g/mol) vaporizes? (ΔHvap = 38.6 kJ/mol)

A 1.93 kJ B 88.8 kJ C 0.519 kJ D 19.3 kJ

Answer: A · 1.93 kJ

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Module 14 · Lesson 3 · Hess's law and enthalpy of formation

Learning outcomes Q16–Q20 · notes, then 5 self-test questions per outcome

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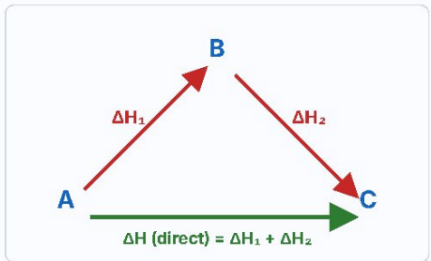
Q16

Textbook + Figure 13 + Example problem 5 + Practice problems · p. 23, 24, 25, 26

Calculate, using Hess's law, the ΔH of a reaction

- **Hess's law:** the enthalpy change of a reaction is the same whether it happens in one step or several (ΔH is a state function).
- Method: (1) reverse an equation → flip the sign of ΔH; (2) multiply an equation → multiply ΔH by the same factor; (3) add the equations and their ΔH values.
- Use Hess's law when ΔH is hard to measure directly (e.g. C → CO, or diamond → graphite because it is too slow).
- If the question asks for the enthalpy change **per mole** of a substance, divide the enthalpy change by the **coefficient** of that substance (e.g. ΔH = −1625 kJ for 4Fe → per mol Fe = −1625 ÷ 4 = −406 kJ/mol).



Worked example

Q. Find ΔH for $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$ given (1) $C + O_2 \rightarrow CO_2$ ΔH = −393.5 kJ; (2) $CO + \frac{1}{2}O_2 \rightarrow CO_2$ ΔH = −283.0 kJ.

Answer: Keep (1); reverse (2): $CO_2 \rightarrow CO + \frac{1}{2}O_2$ ΔH = +283.0 kJ · add: ΔH = −393.5 + 283.0 = **−110.5 kJ**

Common mistakes: Forgetting to flip the sign when reversing, or to multiply ΔH when scaling.

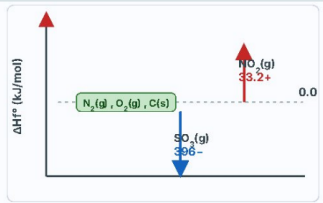
Q17

Textbook · p. 26, 27, 28

Define the standard enthalpy of formation of a compound, ΔH°f, while determining on what basis it is defined

- **ΔH°f** is the enthalpy change when 1 mol of a compound forms from its elements in their standard states (25 °C, 1 atm).
- A formation equation has only elements as reactants and exactly 1 mol of compound as product (fractions allowed).
- Basis: per mole of compound formed. ΔH°f tables (Table 5) let us calculate ΔH for any reaction.
- When writing equations: for the **standard enthalpy of formation** the **product** must be **1 mol** ($S + 3/2\ O_2 \rightarrow$ **1** SO_3); for the **standard enthalpy of combustion** the substance burned (the **reactant**) must be **1 mol** (**1** $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$). Fractions are allowed for the other substances.

Compound	ΔH°f (kJ/mol)
$H_2S(g)$	−21
$HF(g)$	−273
$SO_3(g)$	−396
$SF_6(g)$	−1220
$NO_2(g)$	+33.2



Worked example

Q. Which equation represents ΔH°f of SO_3? (a) $2S + 3O_2 \rightarrow 2SO_3$ ΔH = −792 kJ; (b) $S + 3/2\ O_2 \rightarrow SO_3$

Answer: (b): it forms exactly 1 mol of SO_3 from its elements. (a) forms 2 mol.

Common mistakes: An equation producing 2 mol, or starting from a compound (e.g. $NO + \frac{1}{2}O_2 \rightarrow NO_2$), is not a standard formation equation.

Self-test · Q16

Q16 · Calculate, using Hess's law, the ΔH of a reaction

1. $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ $\Delta H = +66.4$ kJ ; $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$ $\Delta H = -56.6$ kJ. What is ΔH for $N_2(g) + O_2(g) \rightarrow 2NO(g)$?

   A −150.5 kJ **B** +350.0 kJ **C** −179.6 kJ **D** +179.6 kJ

Answer: D · +179.6 kJ

2. Using (a) $S + O_2 \rightarrow SO_2$ $\Delta H = -296.8$; (b) $H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l)$ $\Delta H = -285.8$; (c) $H_2 + S \rightarrow H_2S$ $\Delta H = -20.6$ kJ, find ΔH for $2H_2S + 3O_2 \rightarrow 2H_2O(l) + 2SO_2$

   A −603.2 kJ **B** −1206 kJ **C** −562.0 kJ **D** −1124 kJ

Answer: D · −1124 kJ

3. Find ΔH for $B_2H_6(g) + 3O_2(g) \rightarrow B_2O_3(s) + 3H_2O(l)$ given: $2B + 3H_2 \rightarrow B_2H_6$ $\Delta H = +36$; $H_2 + \frac{1}{2}O_2 \rightarrow H_2O(l)$ $\Delta H = -286$; $2B + 1.5O_2 \rightarrow B_2O_3$ $\Delta H = -1274$ kJ

   A +2168 kJ **B** −1235 kJ **C** −2168 kJ **D** +1235 kJ

Answer: C · −2168 kJ

4. Using: (I) $2OF_2 \rightarrow O_2 + 2F_2$ $\Delta H° = -49.9$; (II) $2ClF + O_2 \rightarrow Cl_2O + OF_2$ $\Delta H° = +205.6$; (III) $ClF_3 + O_2 \rightarrow \frac{1}{2}Cl_2O + 3/2\ OF_2$ $\Delta H° = +266.7$ kJ, find ΔH° for $ClF + F_2 \rightarrow ClF_3$

   A −139 kJ **B** +422 kJ **C** +394 kJ **D** −188 kJ

Answer: A · −139 kJ

5. Use $C + O_2 \rightarrow CO_2$ ($\Delta H = -393.5$ kJ) and $CO + \frac{1}{2}O_2 \rightarrow CO_2$ ($\Delta H = -283.0$ kJ) to find ΔH of $C + \frac{1}{2}O_2 \rightarrow CO$.

   A +110.5 kJ **B** −676.5 kJ **C** −110.5 kJ **D** +676.5 kJ

Answer: C · −110.5 kJ

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Self-test · Q17

Q17 · Define the standard enthalpy of formation of a compound, ΔH°f, while determining on what basis it is defined

1 **Which of the following is NOT a standard enthalpy of formation (ΔH°f)?**

A $\frac{1}{2}H_2(g) + \frac{1}{2}F_2(g) \rightarrow HF(g)$ ΔH = −273 kJ

B $S(s) + 3/2\ O_2(g) \rightarrow SO_3(g)$ ΔH = −396 kJ

C $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$ ΔH = −58.1 kJ

D $S(s) + 3F_2(g) \rightarrow SF_6(g)$ ΔH = −1220 kJ

Answer: C · $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$ ΔH = −58.1 kJ

2 **Which ΔH value represents the standard enthalpy of formation of NO_2?**

A $NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$ ΔH = −199 kJ

B $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ ΔH = +66.4 kJ

C $NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$ ΔH = −58.1 kJ

D $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g)$ ΔH = +33.2 kJ

Answer: D · $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g)$ ΔH = +33.2 kJ

3 **Which equation is the thermochemical equation for the formation of 1 mol of SO_3 from its elements in their standard states?**

A $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ ΔH° = −198 kJ

B $S(s) + 3/2\ O_2(g) \rightarrow SO_3(g)$ ΔH° = −396 kJ

C $3S(s) + 9/2\ O_2(g) \rightarrow 3SO_3(g)$ ΔH° = −1188 kJ

D $2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ ΔH° = −792 kJ

Answer: B · $S(s) + 3/2\ O_2(g) \rightarrow SO_3(g)$ ΔH° = −396 kJ

4 **$2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ has ΔH = −792 kJ. What is the standard enthalpy of formation of SO_3?**

A −396 kJ/mol **B** −792 kJ/mol **C** +1584 kJ/mol **D** +792 kJ/mol

Answer: A · −396 kJ/mol

5 **Which reaction has a ΔH that represents BOTH the standard enthalpy of combustion and the standard enthalpy of formation?**

A $2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ ΔH = −792 kJ

B $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ ΔH = −198 kJ

C $C(s) + O_2(g) \rightarrow CO_2(g)$ ΔH = −394 kJ

D $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ ΔH = −891 kJ

Answer: C · $C(s) + O_2(g) \rightarrow CO_2(g)$ ΔH = −394 kJ

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Notes · Q18 and Q19

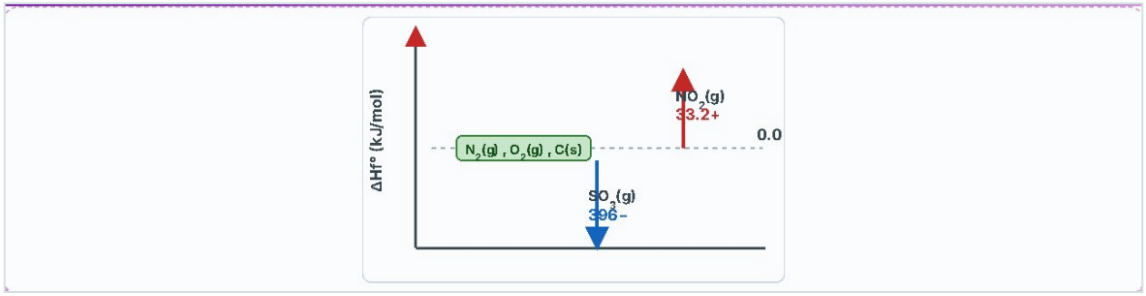
Q18

Textbook + Figure 15 · p. 28

Identify the standard enthalpy (heat) of formation of a free element in its standard state

- ΔH°f of any free element in its standard state is **zero** ($H_2(g)$, $O_2(g)$, $N_2(g)$, $Br_2(l)$, Na(s), C(graphite)).
- Forming an element from itself involves no change; zero is the reference point.
- Careful: non-standard forms are not zero — $O_3(g)$, $Br_2(g)$, C(diamond) and all compounds (H_2O, CO_2).

Remember: ΔH°f [element in its standard state] = 0 · $Br_2(l)$ = 0 · $Br_2(g)$ ≠ 0 · C(graphite) = 0



Worked example

Q. Which has ΔH°f = 0: $Br_2(l)$ or $Br_2(g)$? Why?

Answer: $Br_2(l)$: liquid is the standard state of bromine at 25 °C and 1 atm.

Common mistakes: Not every gaseous element is zero — it must be in its standard state (O_2 yes, O_3 no).

Q19

Textbook + Example problem 6 + Practice problems · p. 28, 29, 30

Calculate, using the standard enthalpies of formation, the enthalpy of a reaction, ΔHrxn

- **$\Delta H°_{rxn} = \Sigma\, n \cdot \Delta H°_f(\text{products}) - \Sigma\, n \cdot \Delta H°_f(\text{reactants})$**, n = coefficient in the balanced equation.
- Free elements (e.g. O_2) are zero and do not affect the sum.
- Multiply each ΔH°f by its coefficient and keep careful track of negative signs when subtracting.

Remember: ΔH°rxn = ΣnΔH°f(products) − ΣnΔH°f(reactants)

Worked example

Q. $2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$; ΔH°f: ZnS −206.0, ZnO −348.3, SO_2 −296.8 kJ/mol. Find ΔH°rxn.

Answer: Products: 2(−348.3) + 2(−296.8) = −1290.2 · reactants: 2(−206.0) + 0 = −412.0 · ΔH° = −1290.2 − (−412.0) = **−878.2 kJ**

Common mistakes: Forgetting the coefficients, or subtracting in the wrong order.

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Self-test · Q18 and Q19

Q18 · Identify the standard enthalpy (heat) of formation of a free element in its standard state

1. **Which of the following has $\Delta H°f = 0$ at standard conditions?**

   A $CO_2(g)$ B $N_2(g)$ C $H_2O(l)$ D $O_3(g)$

Answer: B · $N_2(g)$

2. **Which form of bromine has $\Delta H°f = 0$?**

   A $Br_2(l)$ B $Br_2(g)$ C $Br(g)$ D $Br^-(aq)$

Answer: A · $Br_2(l)$

3. **Which allotrope of carbon is assigned $\Delta H°f = 0$?**

   A Graphite B Diamond
   C Both graphite and diamond D Neither

Answer: A · Graphite

4. **Why is $\Delta H°f$ of $H_2(g)$ equal to zero?**

   A It has no energy
   B It is an element in its standard state, so no change is needed to form it
   C It is a gas
   D It is a compound of hydrogen

Answer: B

5. **Which pair contains two substances that BOTH have $\Delta H°f = 0$?**

   A $O_2(g)$ and $O_3(g)$ B $H_2(g)$ and $H_2O(l)$
   C Na(s) and $O_2(g)$ D C(diamond) and C(graphite)

Answer: C · Na(s) and $O_2(g)$

Q19 · Calculate, using the standard enthalpies of formation, the enthalpy of a reaction, ΔHrxn

1. **Propanoic acid burns: $2C_2H_5COOH(l) + 7O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$. $\Delta H°f$: CO_2 = −394, $H_2O(l)$ = −286, C_2H_5COOH = −510.8 kJ/mol. What is ΔHcomb per mole of acid?**

   A +1529.2 kJ/mol B −3058.4 kJ/mol C +3058.4 kJ/mol D −1529.2 kJ/mol

Answer: D · −1529.2 kJ/mol

2. **Using $\Delta H°f$: $CH_4(g)$ = −74.8, $CO_2(g)$ = −393.5, $H_2O(l)$ = −285.8 kJ/mol, find $\Delta H°$ for $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O(l)$.**

   A −1040 kJ B −890.3 kJ C +890.3 kJ D −540.5 kJ

Answer: B · −890.3 kJ

3. **Using $\Delta H°f$: NO(g) = +91.3 and $NO_2(g)$ = +33.2 kJ/mol, find $\Delta H°$ for $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$.**

   A +116.2 kJ B −58.1 kJ C +248.8 kJ D −116.2 kJ

Answer: D · −116.2 kJ

4. **Using $\Delta H°f$: $C_2H_4(g)$ = +52.3, $CO_2(g)$ = −393.5, $H_2O(l)$ = −285.8 kJ/mol, find $\Delta H°$ for $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O(l)$.**

   A −1306 kJ B +1411 kJ C −1411 kJ D −706 kJ

Answer: C · −1411 kJ

5. **Thermite: $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$. $\Delta H°f$: Al_2O_3 = −1675.7, Fe_2O_3 = −824.2 kJ/mol. What is $\Delta H°$?**

   A +851.5 kJ B −851.5 kJ C −2499.9 kJ D +2499.9 kJ

Answer: B · −851.5 kJ

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Textbook + Example problem 6 + Practice problems · p. 29, 30

Q20 Calculate the enthalpy of reaction (ΔHrxn) using standard enthalpies of formation of products and reactants while using the sign of ΔH for determining if the reaction is exothermic or endothermic

- First compute ΔH°rxn = Σ n ΔH°f(products) − Σ n ΔH°f(reactants), then look at the sign.
- ΔH negative → exothermic (heat released); ΔH positive → endothermic (heat absorbed).
- The sign also tells the direction of heat flow and whether a pack is hot or cold.

Remember: ΔH < 0 → exothermic · ΔH > 0 → endothermic

Worked example

Q. Calculate ΔH° for $2SO_2 + O_2 \rightarrow 2SO_3$ and state the type ($\Delta H°f$: SO_2 −296.8, SO_3 −396 kJ/mol).

Answer: ΔH° = 2(−396) − [2(−296.8) + 0] = −792 + 593.6 = **−198.4 kJ → exothermic**

Common mistakes: A negative ΔH does not mean 'a smaller amount of heat' — it means heat is released.

Self-test · Q20

Q20 · Calculate the enthalpy of reaction (ΔHrxn) using standard enthalpies of formation of products and reactants while using the sign of ΔH for determining if the reaction is exothermic or endothermic

1 **Using ΔH°f: $SO_2(g)$ = −296.8 and $SO_3(g)$ = −396 kJ/mol, what is ΔH° of $2SO_2 + O_2 \rightarrow 2SO_3$ and is it exo- or endothermic?**

A −198.4 kJ ; exothermic
B +198.4 kJ ; endothermic
C −198.4 kJ ; endothermic
D +198.4 kJ ; exothermic

Answer: A · −198.4 kJ ; exothermic

2 **Using ΔH°f: NO(g) = +90.3 kJ/mol, what is ΔH° of $N_2 + O_2 \rightarrow 2NO$ and what type of reaction is it?**

A +180.6 kJ ; endothermic
B −180.6 kJ ; exothermic
C +90.3 kJ ; endothermic
D +180.6 kJ ; exothermic

Answer: A · +180.6 kJ ; endothermic

3 **$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$. ΔH°f: $H_2O_2(l)$ = −187.8, $H_2O(l)$ = −285.8 kJ/mol. Find ΔH°.**

A +196.0 kJ ; endothermic
B −98.0 kJ ; exothermic
C +98.0 kJ ; endothermic
D −196.0 kJ ; exothermic

Answer: D · −196.0 kJ ; exothermic

4 **$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$. ΔH°f: $CaCO_3$ = −1206.9, CaO = −635.1, CO_2 = −393.5 kJ/mol. Which is correct?**

A ΔH° = −178.3 kJ ; exothermic
B ΔH° = +1206.9 kJ ; endothermic
C ΔH° = +178.3 kJ ; endothermic
D ΔH° = −1028.6 kJ ; exothermic

Answer: C · ΔH° = +178.3 kJ ; endothermic

5 **A reaction has ΔH°rxn = −1124 kJ. Which statement is correct?**

A The reaction is exothermic: the products have lower enthalpy than the reactants
B The reaction is endothermic: the products have lower enthalpy
C The reaction is exothermic: the products have higher enthalpy
D The reaction absorbs 1124 kJ from the surroundings

Answer: A

Module 15 · Reaction rates

Learning outcomes Q21–Q30 · notes, then 5 self-test questions per outcome

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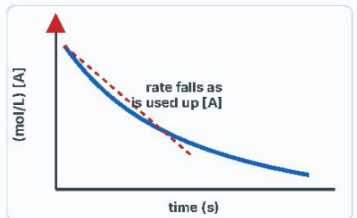
Notes · Q21 and Q22

Textbook + Example problem 1 + Practice problems · p. 46, 47, 48, 49

Q21 Calculate the average reaction rate using the rate of consumption of reactants or the rate of production of products

- **Reaction rate** is the change in concentration of a reactant or product per unit time; unit mol/(L·s) (or mol/(L·min)).
- Average rate = − Δ[reactant]/Δt (the minus makes the rate positive) or + Δ[product]/Δt.
- Δ[ ] = final − initial. The rate falls with time as reactants are used up, hence 'average'.
- Rates come from experimental data, not the balanced equation; rates of different species are proportional to their coefficients.

Remember: Average rate = − Δ[reactant] ÷ Δt = Δ[product] ÷ Δt · Δ[ ] = final − initial



Worked example

Q. [HI] falls from 0.045 M to 0.030 M in 3.00 s. What is the average rate of consumption of HI?

Answer: Δ[HI] = 0.030 − 0.045 = −0.015 M · rate = −(−0.015) ÷ 3.00 = **0.0050 mol/(L·s)**

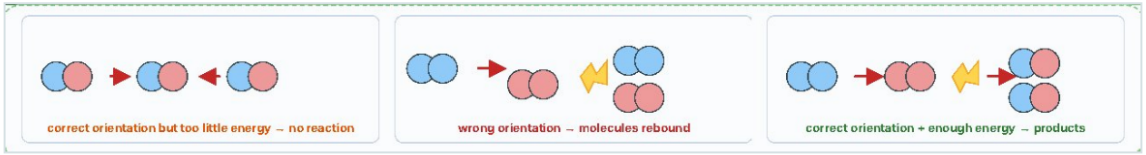
Common mistakes: Dropping the minus sign for a reactant gives a negative rate; rates are always positive.

Textbook + Figures 3 and 4 + Table 1 · p. 49, 50, 51

Q22 Explain why a collision between two particles is necessary for a reaction to occur

- **Collision theory:** particles must collide before they can react, so bonds can break and new bonds form.
- Not every collision gives a reaction: it needs (1) **correct orientation** and (2) **sufficient energy** to form the activated complex.
- Effective collisions are few compared with all collisions; with wrong orientation or too little energy the particles simply rebound.

Remember: Effective collision = collision + correct orientation + sufficient energy



Worked example

Q. $H_2 + Cl_2 \rightarrow 2HCl$: why do many collisions between H_2 and Cl_2 not produce HCl?

Answer: Those collisions have the **wrong orientation** or **not enough energy** (less than Ea), so the molecules rebound.

Common mistakes: 'Every collision gives products' and 'orientation alone is enough' are both wrong.

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Self-test · Q21

Q21 · Calculate the average reaction rate using the rate of consumption of reactants or the rate of production of products

1 **The rate of a chemical reaction is defined as...**

- **A** the ratio of collisions to time
- **B** the change in concentration of a reactant or product per unit time
- **C** the ratio of moles of reactants to products at equilibrium
- **D** the minimum energy needed to start the reaction

Answer: B

2 **$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$. $[Cu^{2+}]$ falls from 0.50 M to 0.25 M in 10.0 s. What is the average rate?**

A 0.015 mol/(L·s) **B** 0.025 mol/(L·s) **C** 0.005 mol/(L·s) **D** 0.002 mol/(L·s)

Answer: B · 0.025 mol/(L·s)

3 **$2NO_2(g) \rightarrow 2NO(g) + O_2(g)$. $[NO_2]$ falls from 0.080 M to 0.040 M in 20 s. What is the average rate with respect to NO_2?**

- **A** 2.0×10^{-3} mol/(L·s)
- **B** 1.0×10^{-3} mol/(L·s)
- **C** 4.0×10^{-3} mol/(L·s)
- **D** 8.0×10^{-3} mol/(L·s)

Answer: A · 2.0×10^{-3} mol/(L·s)

4 **For a hypothetical reaction, [A] changes from 3.00 M to 0.75 M in 3.00 s. What is the average rate of the reaction?**

A 0.85 mol/(L·s) **B** 0.75 mol/(L·s) **C** 1.25 mol/(L·s) **D** 1.50 mol/(L·s)

Answer: B · 0.75 mol/(L·s)

5 **$I_2 + Cl_2 \rightarrow 2ICl$. $[I_2]$ changes from 0.400 M at 0.00 min to 0.3840 M at 4.00 min. What is the average rate of consumption of I_2?**

- **A** 0.010 mol/(L·min)
- **B** 0.016 mol/(L·min)
- **C** 0.004 mol/(L·min)
- **D** 0.096 mol/(L·min)

Answer: C · 0.004 mol/(L·min)

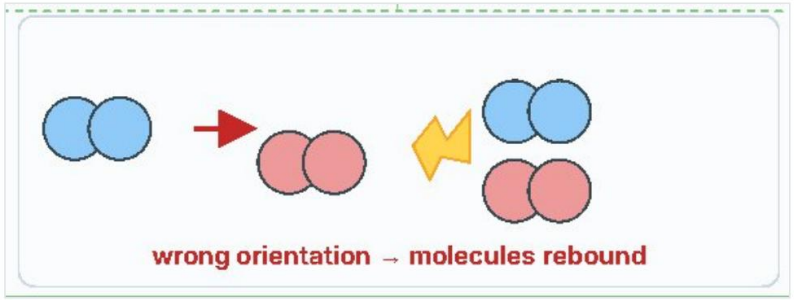
Q22 · Explain why a collision between two particles is necessary for a reaction to occur

1 Why is it important that colliding molecules have the correct orientation?

A Only the correct orientation allows bonds to break and new bonds to form
B Incorrect orientation always increases the rate
C Orientation only matters in endothermic reactions
D Orientation changes the products formed

Answer: A

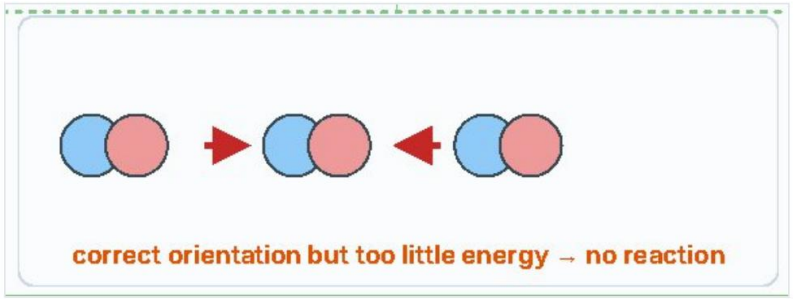
2 In the figure the two molecules collide and rebound. What is the most likely reason?



A The collision had the wrong orientation
B The molecules had too much energy
C The molecules did not collide
D A catalyst was present

Answer: A · The collision had the wrong orientation

3 The molecules have the correct orientation but do not react. Why?



A The collision was too gentle to be counted as a collision
B Correct orientation always prevents a reaction
C Bonds cannot break in the correct orientation
D They collide with insufficient energy to form the activated complex

Answer: D

4 According to collision theory, which is NOT required for a reaction between two particles?

A The particles must collide
B A catalyst must be present
C Correct orientation
D Sufficient energy

Answer: B · A catalyst must be present

5 A reaction between two gases is carried out. Which statement about collisions is correct?

A Only a small fraction of all collisions are effective
B All collisions are effective
C No collisions are needed
D Effective collisions need no energy

Answer: A

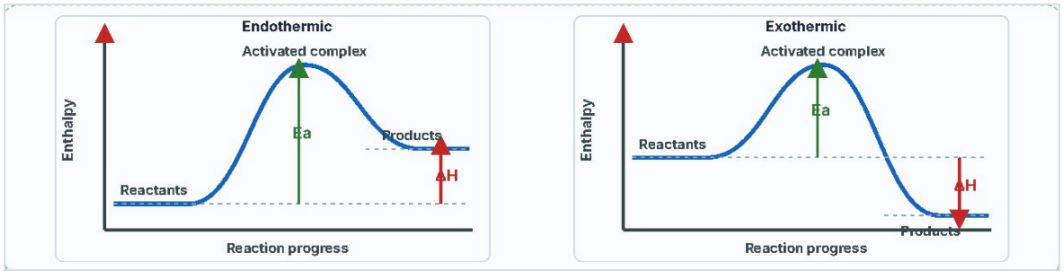
Q23

Textbook + Figures 5 and 6 · p. 51, 52

Define activated complex

- The **activated complex** (transition state) is a temporary, unstable arrangement of atoms in which old bonds break and new bonds form.
- It sits at the top of the energy curve and forms when colliding particles have energy ≥ the activation energy Ea.
- It may form the products or fall back to the reactants; it cannot be isolated. Ea = energy of the peak − energy of the reactants.
- Example: $CO + NO_2 \rightarrow [O–C–O–N–O]$ (complex) $\rightarrow CO_2 + NO$.

Remember: Ea = E(activated complex) − E(reactants)



Worked example

Q. In an energy diagram, what does the top of the curve represent and what does the arrow from reactants to the peak represent?

Answer: Top of the curve = **activated complex** · arrow from reactants to the peak = **activation energy (Ea)**.

Common mistakes: The activated complex is neither a reactant nor a product — it is a temporary state at the peak.

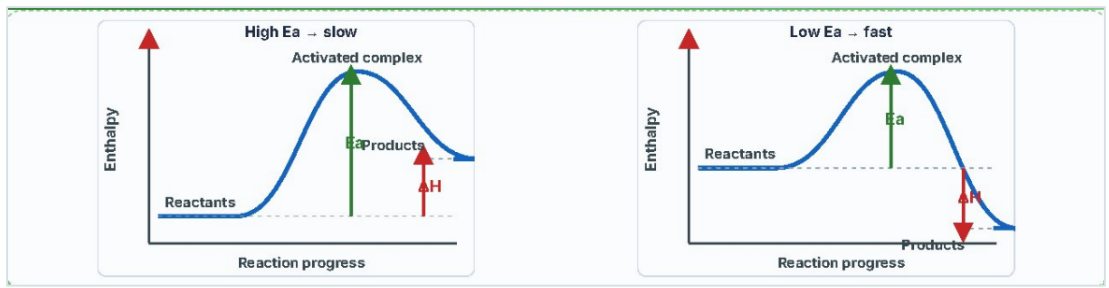
Q24

Textbook + Figures 5 and 6 + Table 1 · p. 51, 52

Describe the relationship between activation energy and the rate of the reaction

- **Activation energy** is the minimum energy needed to form the activated complex (the energy barrier).
- The higher the Ea, the slower the reaction (fewer particles have enough energy); the lower the Ea, the faster the reaction.
- Under identical conditions the reaction with the smaller Ea is faster; Ea does not depend on whether the reaction is exo- or endothermic.

Remember: Ea ↓ ⇒ rate ↑ · Ea ↑ ⇒ rate ↓



Worked example

Q. Reactions X (Ea = 25 kJ/mol) and Y (Ea = 75 kJ/mol) run under identical conditions. Which is faster and why?

Answer: X: its barrier is lower, so a larger fraction of collisions has enough energy to react.

Common mistakes: A large Ea does not mean an endothermic reaction — it is only the barrier.

Q23 · Define activated complex

1 **Which statement best defines an activated complex?**

A The stable product of a reaction
B The lowest energy state of the reactants
C A temporary, unstable arrangement of atoms in which old bonds are breaking and new bonds are forming
D A catalyst that speeds a reaction

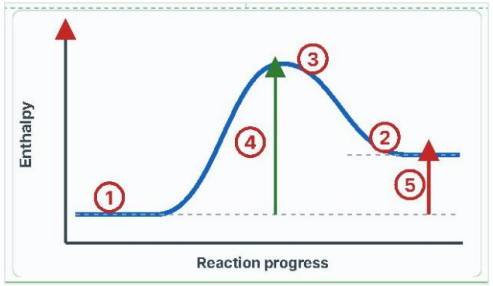
Answer: C

2 **Where does the activated complex appear on a potential-energy diagram?**

A At the starting level of the reactants
B At the level of the products
C At the bottom of the curve
D At the highest point (peak) of the curve

Answer: D · At the highest point (peak) of the curve

3 **In the endothermic diagram, arrows and points are numbered (1 reactants, 2 products, 3 peak, 4 arrow to the peak, 5 arrow between the levels). Which statement is correct?**



A 4 = activation energy, 5 = energy absorbed by the reaction
B 1 = products, 2 = reactants
C 4 = activation energy, 5 = energy released by the reaction
D 1 = activated complex, 2 = reactants

Answer: A

4 **What can the activated complex do?**

A Only form the products
B Only re-form the reactants
C Remain unchanged forever
D Form the products or break apart to re-form the reactants

Answer: D

5 **In the reaction $CO + NO_2 \rightarrow CO_2 + NO$ the species OCONO forms briefly. What is it called?**

A The product
B The catalyst
C The reactant
D The activated complex

Answer: D · The activated complex

Self-test · Q24

Q24 · Describe the relationship between activation energy and the rate of the reaction

1 **Which statement correctly relates activation energy and reaction rate?**

A The higher the activation energy, the faster the reaction
B Activation energy has no effect on rate
C The rate depends only on ΔH
D The higher the activation energy, the slower the reaction

Answer: D

2 **Four reactions have Ea = 40, 55, 90 and 120 kJ/mol under the same conditions. Which is the fastest?**

A The one with Ea = 55 kJ/mol
B The one with Ea = 40 kJ/mol
C The one with Ea = 90 kJ/mol
D The one with Ea = 120 kJ/mol

Answer: B · The one with Ea = 40 kJ/mol

3 **Why does a reaction with a lower activation energy proceed faster?**

A The molecules collide less often
B The reaction is always exothermic
C A larger fraction of collisions has enough energy to react
D The products have higher energy

Answer: C

4 **Reaction X has Ea = 25 kJ/mol and reaction Y has Ea = 75 kJ/mol under identical conditions. Which is correct?**

A Y is faster because it needs more energy
B Both have the same rate
C Ea does not affect the rate
D X is faster because more molecules can overcome the lower Ea

Answer: D

5 **Which statement is true about the activation energy of a reaction?**

A It equals ΔH of the reaction
B It is always negative
C It is the energy barrier that colliding particles must overcome
D It is the energy of the products

Answer: C

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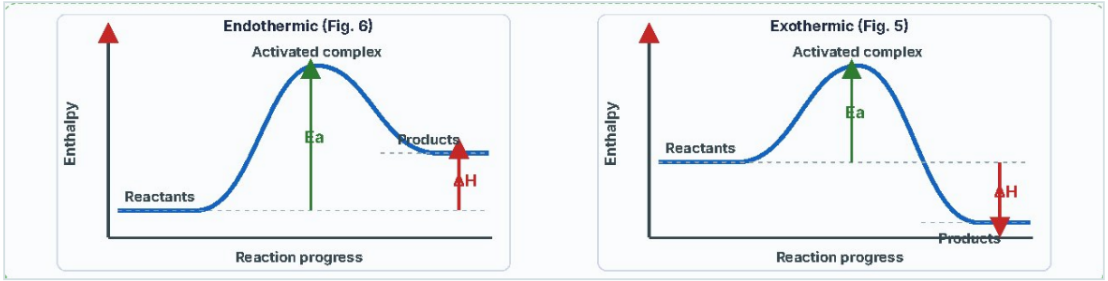
Q25

Textbook + Figures 5, 6 · p. 51, 52

Interpret exothermic and endothermic reactions using the collision theory

- **Exothermic (Fig. 5):** molecules collide with enough energy to pass the barrier, form the complex, then release energy and give products of lower energy (ΔH < 0).
- **Endothermic (Fig. 6):** the reactants absorb enough energy to overcome the barrier and form higher-energy products (ΔH > 0); the surroundings cool.
- In an exothermic reaction Ea(forward) < Ea(reverse); in an endothermic reaction Ea(forward) > Ea(reverse).

Remember: Exo: products lower · energy released · ΔH < 0 · Endo: products higher · energy absorbed · ΔH > 0



Worked example

Q. Figure 1 is exothermic and Figure 2 is endothermic. Which statement is correct?

Answer: In Figure 1 the products have **lower** energy and energy is **released**; in Figure 2 the products have **higher** energy and energy is **absorbed**.

Common mistakes: Do not confuse released energy with Ea — Ea is put in to pass the barrier.

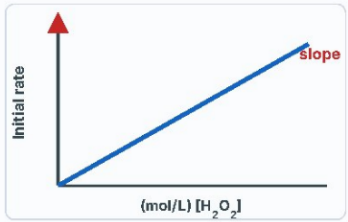
Q26

Textbook + Practice problems · p. 62, 63, 64, 65

Write the rate law for a chemical reaction at a given temperature

- **Rate law:** rate = k[A]^m[B]^n at a given temperature, with k the specific rate constant.
- Orders (m, n) are found experimentally and need not equal the coefficients; overall order = m + n.
- First order: doubling [A] doubles the rate. Second order: doubling [A] quadruples the rate (tripling → ×9). k changes only with temperature.
- Example: $2NO + 2H_2 \rightarrow N_2 + 2H_2O$ has rate = k[NO]^2[H_2]: second order in NO, first in H_2, third overall.

Remember: rate = k [A]^m [B]^n · overall order = m + n · new rate = old rate × (factor)^order



Worked example

Q. rate = k[A]^2[B]. What happens if [A] is doubled with [B] constant? And if [B] is tripled?

Answer: [A] × 2 → rate × 2^2 = **× 4** · [B] × 3 → rate × 3^1 = **× 3**

Common mistakes: Taking orders directly from balanced-equation coefficients without data.

Q25 · Interpret exothermic and endothermic reactions using the collision theory

1 **Which statement is correct regarding the two figures? (Figure 2: products higher than reactants; Figure 1: products lower)**

A In figure 1 the reactants collide and form higher-energy products

B The activation energy of figure 1 is larger than that of figure 2

C In figure 2 the reactant molecules absorb enough energy to overcome the barrier and form higher-energy products

D Figure 2 is exothermic and figure 1 is endothermic

Answer: C

2 **For $CO_2 + NO \rightarrow CO + NO_2$ the products are higher than the reactants (arrow 1 = activation energy, arrow 2 = energy change). Which description is correct?**

A 1: activation energy · 2: energy released by reaction · exothermic

B 1: energy absorbed · 2: activation energy · endothermic

C 1: activation energy · 2: energy absorbed by reaction · endothermic

D 1: energy released · 2: activation energy · exothermic

Answer: C

3 **In an exothermic reaction, which statement fits collision theory?**

A Molecules absorb energy to form higher-energy products

B No energy barrier exists

C Molecules with enough energy form the complex and then release energy to give lower-energy products

D The products always have more energy

Answer: C

4 **Which pair is correct for an endothermic reaction?**

A Ea(forward) < Ea(reverse) ; ΔH > 0

B Ea(forward) > Ea(reverse) ; ΔH > 0

C Ea(forward) > Ea(reverse) ; ΔH < 0

D Ea(forward) = Ea(reverse) ; ΔH > 0

Answer: B · Ea(forward) > Ea(reverse) ; ΔH > 0

5 **In an endothermic reaction in a beaker, why does the surroundings' temperature fall?**

A The reactants release energy to the surroundings

B The reactants absorb energy from the surroundings to pass the barrier

C The collisions stop

D The products have lower energy

Answer: B

Self-test · Q26

Q26 · Write the rate law for a chemical reaction at a given temperature

1 **For a reaction that is second order in A, how does the rate change if [A] is doubled and all other conditions remain constant?**

A It decreases to a quarter of its value

B It increases four times

C It decreases to half its value

D It doubles

Answer: B · It increases four times

2 **The rate law of a reaction is rate = $k[A]^2[B]$. What happens to the rate if [A] is doubled?**

A It is halved

B It triples

C It increases four times

D It doubles

Answer: C · It increases four times

3 **The rate law for $2NO + 2H_2 \rightarrow N_2 + 2H_2O$ is rate = $k[NO]^2[H_2]$. What is the overall order?**

A Second order

B Fourth order

C Third order

D First order

Answer: C · Third order

4 **Reaction orders in a rate law are determined from...**

A the coefficients of the balanced equation

B the sign of ΔH

C the states of matter

D experimental data

Answer: D · experimental data

5 **A reaction is first order in A. If [A] is tripled, the rate...**

A increases 9 times

B stays the same

C triples

D increases 6 times

Answer: C · triples

Notes · Q27 and Q28

Q27

Textbook + Table 2 + Practice problems · p. 64, 65, 66

Use the method of initial rates to write the rate law of the reaction

- Method: pick two trials in which only one concentration changes and compare the rate ratio with the concentration ratio.
- Order x from $(rate_2/rate_1) = ([A]_2/[A]_1)^x$: doubling [A] gives ×2 (order 1), ×4 (order 2) or no change (order 0).
- After all orders are found write rate = $k[A]^m[B]^n$.

Remember: $rate_2 / rate_1 = ([A]_2 / [A]_1)^m$

Trial	[A] (M)	[B] (M)	Initial rate (mol/(L·s))
1	0.10	0.10	2.0×10^{-3}
2	0.20	0.10	8.0×10^{-3}
3	0.10	0.20	4.0×10^{-3}

Worked example

Q. From the table, what is the rate law?

Answer: Trials 1→2: [A] × 2 ([B] constant), rate × 4 → second order in A · trials 1→3: [B] × 2, rate × 2 → first order in B · **rate = $k[A]^2[B]$**

Common mistakes: Comparing trials where more than one concentration changes.

Q28

Textbook · p. 63, 64

Use the method of initial rates to calculate the specific rate constant and identify its unit

- Once the rate law is known, substitute the data of any trial: k = rate / ($[A]^m[B]^n$). k is the same in every trial at one temperature.
- The unit of k depends on the overall order: zero → mol/(L·s); first → s^{-1}; second → L/(mol·s); third → $L^2/(mol^2 \cdot s)$.
- Derive the unit by dividing mol/(L·s) by the concentration units in the rate law.

Remember: k = rate ÷ ($[A]^m$ $[B]^n$) · 1st: s^{-1} · 2nd: L/(mol·s) · 3rd: $L^2/(mol^2 \cdot s)$

Overall order	Rate law example	Unit of k
0	rate = k	mol/(L·s)
1	rate = k[A]	s^{-1}
2	rate = k[A][B]	L/(mol·s)
3	rate = $k[A]^2[B]$	$L^2/(mol^2 \cdot s)$

Worked example

Q. $2NO + O_2 \rightarrow 2NO_2$, rate = $k[NO]^2[O_2]$; trial 1: [NO] = 0.030 M, $[O_2]$ = 0.020 M, rate = 0.0041 M/s. Calculate k and its unit.

Answer: $k = 0.0041 \div (0.030^2 \times 0.020) = 0.0041 \div 1.8 \times 10^{-5} =$ **2.3×10^2 $L^2/(mol^2 \cdot s)$**

Common mistakes: Forgetting to square a second-order concentration, or using the same unit of k for every reaction.

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Self-test · Q27 and Q28

Q27 · Use the method of initial rates to write the rate law of the reaction

1 Using the table in the Q27 notes (trials 1–3), which rate law is correct?

A rate = k[A][B] B rate = k[A]²[B] C rate = k[A]²[B]² D rate = k[A][B]²

Answer: B · rate = k[A]²[B]

2 Trial 1: [A] = 0.10 M, rate = 1.5×10^{-3} M/s. Trial 2: [A] = 0.20 M, rate = 3.0×10^{-3} M/s. What is the order in A?

A Zero B Second C Third D First

Answer: D · First

3 Tripling [A] raises the rate by a factor of 9. What is the order in A?

A Second B First C Third D Zero

Answer: A · Second

4 $2ICl + H_2 \rightarrow I_2 + 2HCl$. Doubling [ICl] doubles the rate; halving $[H_2]$ halves the rate. What is the rate law?

A rate = k[ICl]²[H_2] B rate = k[ICl]²[H_2]²

C rate = k[ICl] D rate = k[ICl][H_2]

Answer: D · rate = k[ICl][H_2]

5 Doubling [A] makes the rate 8 times larger. What is the order in A?

A Second B Third C First D Fourth

Answer: B · Third

Q28 · Use the method of initial rates to calculate the specific rate constant and identify its unit

1 From rate = k[NO]²[O_2], trial: [NO] = 0.030 M, [O_2] = 0.020 M, rate = 0.0041 M/s. What is k?

A 2.3×10^2 L/(mol·s) B 6.8 L²/(mol²·s)

C 2.3×10^2 L²/(mol²·s) D 1.4×10^{-2} L²/(mol²·s)

Answer: C · 2.3×10^2 L²/(mol²·s)

2 What is the unit of k for a reaction that is second order overall?

A s^{-1} B L²/(mol²·s) C L/(mol·s) D mol/(L·s)

Answer: C · L/(mol·s)

3 Rate = k[A]²; when [A] = 0.20 M the rate is 0.0040 M/s. What is k?

A 0.020 L/(mol·s) B 0.80 L/(mol·s) C 0.10 s^{-1} D 0.10 L/(mol·s)

Answer: D · 0.10 L/(mol·s)

4 Which rate law corresponds to a unit of k of L²/(mol²·s)?

A rate = k[A]²[B] B rate = k[A] C rate = k[A][B] D rate = k

Answer: A · rate = k[A]²[B]

5 The decomposition of H_2O_2 is first order. What is the unit of k?

A L/(mol·s) B L²/(mol²·s) C s^{-1} D mol/(L·s)

Answer: C · s^{-1}

Notes · Q29 and Q30

Q29

Textbook + Table 2 + Practice problems · p. 63, 64, 65, 66

Use the method of initial rates to calculate the concentration of reactants

- Knowing the rate law, k and the rate, rearrange the law to find the unknown concentration.
- Second order: [A] = √(rate/k); first order: [A] = rate/k; with two reactants evaluate the known part first, then divide.
- Check that the unit of k fits the overall order and that the answer is in mol/L (M).

Remember: [A] = rate ÷ (k [B]^n) → then take the m-th root

Worked example

Q. $2SO_2 + O_2 \rightarrow 2SO_3$: first order in O_2, second in SO_2, k = 3.40×10^3, rate = 6.5×10^{-4} mol/(L·s), $[O_2]$ = 0.00150 M. Find $[SO_2]$.

Answer: rate = $k[O_2][SO_2]^2 \rightarrow [SO_2]^2 = 6.5 \times 10^{-4} \div (3.40 \times 10^3 \times 0.00150) = 1.27 \times 10^{-4} \rightarrow [SO_2] =$ **0.011 M**

Common mistakes: Forgetting the square root for a second-order term, or multiplying instead of dividing.

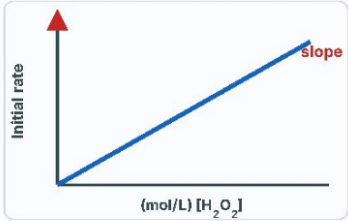
Q30

Textbook + Table 2 + Practice problems · p. 63, 64, 65, 66

Use the method of initial rates to calculate the initial rate of the reaction

- With the rate law and k known, substitute the initial concentrations raised to their orders: rate = k[A]^m[B]^n.
- The initial rate is the rate at t = 0, using the initial concentrations; unit mol/(L·s).
- On a graph of initial rate vs [A] for a first-order reaction the slope is k.

Remember: rate = k [A]^m[B]^n · slope (rate vs [A]) = k (first order)



Worked example

Q. rate = k[A]^2[C], k = 7.62×10^{-6} L^2/(mol^2·s), [A] = 0.175 M, [C] = 0.315 M. Find the initial rate.

Answer: rate = $7.62 \times 10^{-6} \times (0.175)^2 \times 0.315 = 7.62 \times 10^{-6} \times 0.0306 \times 0.315 =$ **7.35×10^{-8} mol/(L·s)**

Common mistakes: Forgetting to raise a concentration to its order, or including a zero-order species.

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Self-test · Q29

Q29 · Use the method of initial rates to calculate the concentration of reactants

1. rate = $k[A]^2$ with k = 0.50 L/(mol·s). The rate is 0.020 mol/(L·s). What is [A]?

A 0.040 M B 0.010 M C 0.50 M D 0.20 M

Answer: D · 0.20 M

2. rate = k[A][B], k = 2.0 L/(mol·s), rate = 0.040 mol/(L·s), [B] = 0.10 M. What is [A]?

A 0.80 M B 0.20 M C 0.080 M D 0.40 M

Answer: B · 0.20 M

3. rate = k[A] with k = 0.025 s^{-1}. The rate is 1.0×10^{-3} mol/(L·s). What is [A]?

A 0.40 M B 2.5×10^{-5} M C 0.040 M D 25 M

Answer: C · 0.040 M

4. rate = $k[NO]^2[O_2]$, k = 2.3×10^2 $L^2/(mol^2 \cdot s)$, $[O_2]$ = 0.020 M, rate = 0.0164 mol/(L·s). What is [NO]?

A 0.0036 M B 0.060 M C 0.60 M D 0.12 M

Answer: B · 0.060 M

5. rate = $k[A]^2[B]$, k = 15 $L^2/(mol^2 \cdot s)$, [A] = 0.10 M, rate = 0.030 mol/(L·s). What is [B]?

A 0.20 M B 0.020 M C 2.0 M D 0.45 M

Answer: A · 0.20 M

Self-test · Q30

Q30 · Use the method of initial rates to calculate the initial rate of the reaction

1 rate = $k[NO]^2[O_2]$ with k = 228 $L^2/(mol^2 \cdot s)$. [NO] = 0.050 M and $[O_2]$ = 0.020 M. What is the initial rate?

A 1.14×10^{-2} mol/(L·s) B 5.7×10^{-1} mol/(L·s)
C 2.28×10^{-1} mol/(L·s) D 4.6×10^{-3} mol/(L·s)

Answer: A · 1.14×10^{-2} mol/(L·s)

2 rate = k[A] with k = 0.025 s^{-1} and [A] = 0.40 M. What is the initial rate?

A 0.010 mol/(L·s) B 0.016 mol/(L·s) C 0.063 mol/(L·s) D 0.40 mol/(L·s)

Answer: A · 0.010 mol/(L·s)

3 rate = $k[H_2][NO]^2$ with k = 1.23 $L^2/(mol^2 \cdot s)$; $[H_2]$ = 0.30 M, [NO] = 0.20 M. What is the initial rate?

A 1.48×10^{-2} mol/(L·s) B 7.4×10^{-2} mol/(L·s)
C 7.4×10^{-3} mol/(L·s) D 2.2×10^{-1} mol/(L·s)

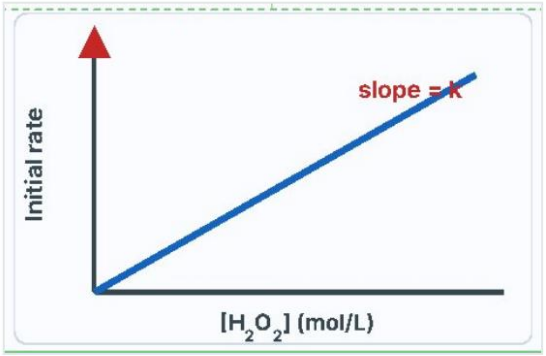
Answer: A · 1.48×10^{-2} mol/(L·s)

4 rate = $k[A]^2$ with k = 0.50 L/(mol·s) and [A] = 0.30 M. What is the initial rate?

A 0.15 mol/(L·s) B 0.075 mol/(L·s) C 0.045 mol/(L·s) D 0.30 mol/(L·s)

Answer: C · 0.045 mol/(L·s)

5 The graph of initial rate vs $[H_2O_2]$ is a straight line with slope 0.40 s^{-1}. What is the initial rate when $[H_2O_2]$ = 1.50 mol/L?



A 0.27 mol/(L·s) B 3.75 mol/(L·s) C 0.60 mol/(L·s) D 0.40 mol/(L·s)

Answer: C · 0.60 mol/(L·s)

Ministry exam 2025–2026

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Ministry exam 2025–2026 · Questions 1–7

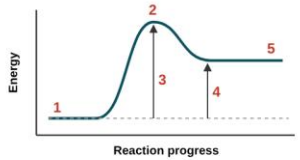
1 For the reaction $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$, why might many collisions between H_2 and Cl_2 molecules fail to produce products?

1	Because the molecules collide without the proper orientation of H–H and Cl–Cl bonds
2	Because many collisions lack sufficient energy
3	Because the activation energy decreases with time

A 1 and 2 B 2 only C 2 and 3 D 1 only

Answer: A · 1 and 2

2 Regarding the energy curve below, which of the following is correct?



A Number 1 indicates the reactants while number 4 indicates the products

B Number 4 indicates the activated complex while number 2 indicates the energy released by the reaction

C Number 3 indicates the reactants while number 4 indicates the activated complex

D Number 2 indicates the activated complex while number 3 indicates the activation energy

Answer: D

3 Which of the following definitions correctly describes the rate of reaction?

A The ratio of moles of reactants to products at equilibrium

B Ratio of the number of collisions per unit time

C The change in concentrations of reactants or products per unit time

D The minimum amount of energy required for a reaction to occur

Answer: C

4 For the reaction $2NO_2(g) \rightarrow 2NO(g) + O_2(g)$, which occurs in a closed system, which of the following is correct regarding the concentration changes?

A The concentration of NO_2 decreases continuously, while the concentrations of NO and O_2 increase

B The concentration of NO_2 increases continuously, while the concentrations of NO and O_2 decrease

C The concentrations of NO_2, NO and O_2 all remain constant

D The concentration of NO_2 decreases, but the concentration of NO remains constant

Answer: A

5 For the reaction $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$, the concentration of Cu^{2+} is 0.50 M at the beginning of the reaction. After 10.0 s the concentration has become 0.25 M. What is the average reaction rate over the given time period?

A 0.025 mol/(L·s) B 0.002 mol/(L·s) C 0.015 mol/(L·s) D 0.005 mol/(L·s)

Answer: A · 0.025 mol/(L·s)

6 Consider the combustion of ethanol: $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$, $\Delta H = -1367$ kJ/mol. Which of the following correctly describes this reaction?

A Endothermic, energy absorbed from the surroundings

B Endothermic, energy released to the surroundings

C Exothermic, energy released to the surroundings

D Exothermic, energy absorbed from the surroundings

Answer: C

7 Which of the following processes is exothermic?

A	$NH_3(g) \rightarrow NH_3(l)$
B	$C_2H_5OH(l) \rightarrow C_2H_5OH(g)$
C	$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$
D	$H_2O(s) \rightarrow H_2O(l)$

A A only B B and D C A and C D B only

Answer: C · A and C

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Ministry exam 2025–2026 · Questions 8–13

8 **Which of the following correctly expresses the concept of enthalpy (heat) of combustion (ΔH_{comb})?**

A One mole of the substance produced from its primary elements
B One mole of reactant oxygen
C One mole of a substance incompletely burned in oxygen
D One mole of a substance burned completely in oxygen

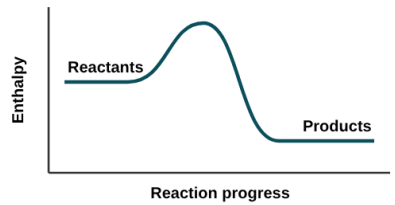
Answer: D

9 **Which of the following statements correctly defines enthalpy (H)?**

A The energy required to break all bonds in one mole of substance
B The heat content of a system at constant pressure
C The total energy stored in the bonds of reactants and products
D The energy change that occurs at constant volume

Answer: B

10 **Regarding the potential energy diagram shown below, which of the following is correct?**



A The reactants and the products have the same energy level
B No difference in enthalpy is observed
C The energy of the products is higher than the energy of the reactants
D The energy of the products is lower than the energy of the reactants

Answer: D

11 **When hydrochloric acid is mixed with sodium hydroxide in a calorimeter, the temperature of the solution increases. Which of the following correctly describes the heat flow and type of reaction?**

A Heat flows out of the system; exothermic reaction
B No change in enthalpy occurs; ΔH = 0
C Heat circulates only within the system
D Heat flows into the system; endothermic reaction

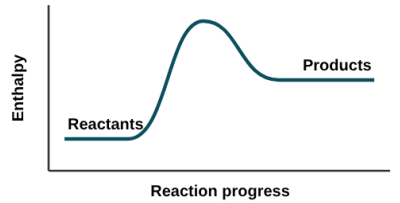
Answer: A

12 **A 200 g piece of copper cools from 150 °C to 25 °C. What is the amount of heat released? (c of copper = 0.385 J/g·°C)**

A 8250 J B 4800 J C 5780 J D 9625 J

Answer: D · 9625 J

13 **For the chemical reaction represented by the diagram below, which of the following is correct?**



A The system absorbs heat from the surroundings
B The system heat always decreases
C The system releases heat to the surroundings
D No energy change occurs

Answer: A

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Ministry exam 2025–2026 · Questions 14–18

14 **Which of the following substances has the greatest increase in temperature when the same amount of heat is added to equal masses of it?**

Substance	Specific heat
Water (l)	4.184 J/(g·°C)
Calcium (s)	0.647 J/(g·°C)
Concrete (s)	0.84 J/(g·°C)
Lead (s)	0.129 J/(g·°C)

A Calcium B Concrete C Water D Lead

Answer: D · Lead

15 **In an exothermic reaction, 75.0 kJ of energy was released. How much energy is released in kcal?**

A 17.9 kcal B 29.7 kcal C 20.6 kcal D 21.5 kcal

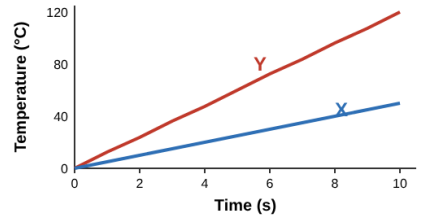
Answer: A · 17.9 kcal

16 **A nail of mass 25.0 g made of an alloy absorbs 250 J of heat when its temperature changes from 25.0 °C to 78.0 °C. What is the specific heat of the nail?**

A 0.19 J/(g·°C) B 1.19 J/(g·°C) C 0.89 J/(g·°C) D 5.30 J/(g·°C)

Answer: A · 0.19 J/(g·°C)

17 **The graph below shows the heating curves of two different substances with the same mass. Which of the following is correct?**



A They both have the same specific heat

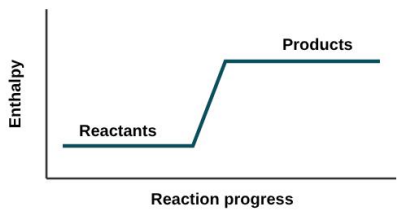
B It is not possible from the diagram to determine which of the two substances has the higher specific heat

C Substance X has a higher specific heat than substance Y

D Substance Y has a higher specific heat than substance X

Answer: C

18 **For the endothermic reaction shown below, which of the following is correct?**



A $\Delta H_{rxn} = H_{reactants} + H_{products}$

B $\Delta H_{rxn} > 0$

C $\Delta H_{rxn} = 0$

D $\Delta H_{rxn} < 0$

Answer: B · ΔHrxn > 0

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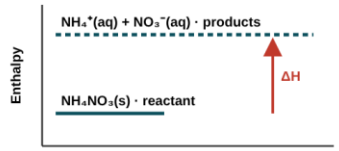
Ministry exam 2025–2026 · Questions 19–25

19 Which of the following is correct regarding the enthalpy change in the cold-pack process shown in the figure below?



1	$H_{products} > H_{reactants}$
2	The sign of ΔH is negative
3	The reaction is exothermic

A 1 and 3 B 2 only C 2 and 3 D 1 only

Answer: D · 1 only

20 In a chemical reaction, 2822 kJ of energy is produced when 0.5 mol of a certain compound undergoes combustion. Which of the substances in the table represents this compound?

Substance	Formula	ΔH°comb (kJ/mol)
Sucrose	$C_{12}H_{22}O_{11}(s)$	−5644
Propane	$C_3H_8(g)$	−2219
Glucose	$C_6H_{12}O_6(s)$	−2808
Octane	$C_8H_{18}(l)$	−5471

A Sucrose B Glucose C Propane D Octane

Answer: A · Sucrose

21 The complete combustion of 1 mol of methane (CH_4) releases 890 kJ of energy. How much energy is released when 3 mol of methane are combusted?

A 297 kJ B 1780 kJ C 2670 kJ D 445 kJ

Answer: C · 2670 kJ

22 At 100 °C and 1 atm, the molar enthalpy of vaporization ΔH_{vap} of water equals +40.7 kJ/mol. Which thermochemical equation represents the condensation of water vapor?

A $H_2O(l) \rightarrow H_2O(g)$ ΔH = +40.7 kJ B $H_2O(g) \rightarrow H_2O(l)$ ΔH = +40.7 kJ
C $H_2O(g) \rightarrow H_2O(l)$ ΔH = −40.7 kJ D $H_2O(l) \rightarrow H_2O(g)$ ΔH = −40.7 kJ

Answer: C · $H_2O(g) \rightarrow H_2O(l)$ ΔH = −40.7 kJ

23 How much heat is evolved when 90.6 g of propane (C_3H_8, 44.09 g/mol) is burned according to the equation $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$, ΔH_{comb} = −2219 kJ/mol?

A 4560 kJ B 2455 kJ C 8942 kJ D 3574 kJ

Answer: A · 4560 kJ

24 Using Hess's law and the enthalpy changes in the table, what is ΔH for the reaction $N_2(g) + O_2(g) \rightarrow 2NO(g)$?

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$	ΔH = +66.4 kJ
$NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$	ΔH = −56.6 kJ

A −179.6 kJ B +179.6 kJ C +350.0 kJ D −150.5 kJ

Answer: B · +179.6 kJ

25 Using the standard enthalpies of formation in the table, what is $\Delta H°_{rxn}$ for $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$?

Substance	$CO_2(g)$	$H_2O(g)$	$C_2H_6(g)$
ΔH°f (kJ/mol)	−393.5	−242	−84

A +2858 kJ B −2858 kJ C +1890 kJ D −1560 kJ

Answer: B · −2858 kJ

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Ministry exam 2025–2026 · Questions 26–30

26 **Which of the following has ΔH°f = 0 at standard conditions (298 K, 1 atm)?**

1	$N_2(g)$
2	$H_2O(l)$
3	$CO_2(g)$

A 3 only **B** 2 and 3 **C** 1 and 2 **D** 1 only

Answer: D · 1 only

27 **For the reaction $2NO_2(g) \rightarrow 2NO(g) + O_2(g)$, the concentration of NO_2 decreases from 0.080 M to 0.040 M in 20 seconds. What is the average rate of reaction with respect to NO_2?**

A 8.0×10^{-3} mol/(L·s) **B** 2.0×10^{-3} mol/(L·s)

C 4.0×10^{-3} mol/(L·s) **D** 1.0×10^{-3} mol/(L·s)

Answer: B · 2.0×10^{-3} mol/(L·s)

28 **Which of the following explains the importance of orientation in particle collisions for a reaction to occur?**

1	Only collisions with the correct orientation allow bonds to break and new bonds to form
2	Incorrect orientation collisions always increase reaction rate
3	Orientation only affects endothermic reactions

A 2 and 3 **B** 1 only **C** 1 and 3 **D** 2 only

Answer: B · 1 only

29 **Two reactions occur under identical conditions. Reaction X has an activation energy of 25 kJ/mol, while reaction Y has an activation energy of 75 kJ/mol. Therefore, which of the following is correct?**

A Reaction Y is faster because it requires more activation energy to proceed

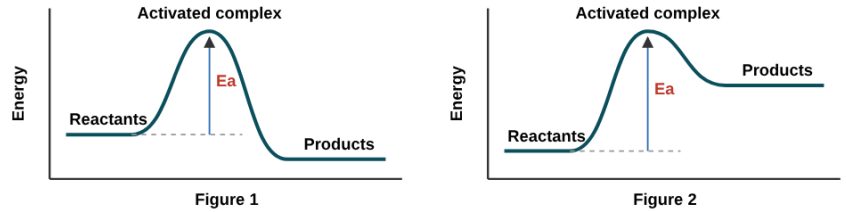
B Reaction X is faster because a larger number of molecules can overcome the lower activation energy

C Both reactions occur at the same rate since temperature is constant

D The activation energy does not affect reaction rate in this case

Answer: B

30 **Figures 1 and 2 below represent the energy diagrams of two reactions. Which of the following is correct?**



A In figure 1, collisions result in products at a higher energy level than the reactants and the sign of ΔH is positive

B In figure 1, collisions result in products at a lower energy level than the reactants and the sign of ΔH is negative

C In figure 2, collisions result in products at a higher energy level than the reactants and the sign of ΔH is negative

D In figure 2, collisions result in products at a lower energy level than the reactants and the sign of ΔH is positive

Answer: B

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More Ministry exam questions

12 more Ministry questions on the EOT outcomes

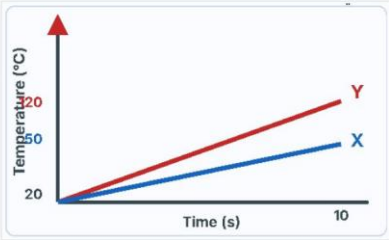
More Ministry questions · Questions 1–6

1 Q2 **Equal masses of silver (0.235), iron (0.449), gold (0.129) and aluminum (0.897 J/g·°C) are left in the Sun for the same time. Which metal shows the highest increase in temperature?**

A Aluminum B Gold C Iron D Silver

Answer: B · Gold

2 Q2 **The graph shows the heating curves of equal masses of substances X and Y heated at the same rate. Which statement is correct?**



A Y has a higher specific heat than X
B X and Y have the same specific heat
C It cannot be determined
D X has a higher specific heat than Y

Answer: D · X has a higher specific heat than Y

3 Q3 **15.0 g of cadmium absorbs 32.0 cal and its final temperature is 62.7 °C. What was its initial temperature? (c = 0.231 J/g·°C)**

A 24.1 °C B 38.7 °C C 9.3 °C D 53.4 °C

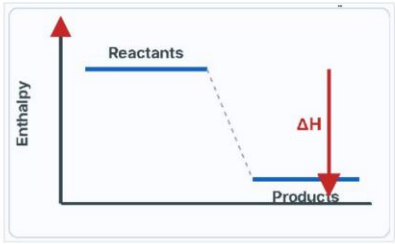
Answer: A · 24.1 °C

4 Q8 **Why can a calorimeter NOT be used to measure ΔH for the change of diamond to graphite?**

A The reaction occurs too quickly to measure
B The reaction occurs too slowly
C Calorimeters only measure endothermic reactions
D Diamond and graphite have the same enthalpy

Answer: B · The reaction occurs too slowly

5 Q9 **The enthalpy diagram of a reaction shows the reactants at a high level and the products at a lower level. Which statement is correct?**



A The energy of the products is lower than that of the reactants
B The energy of the products is higher than that of the reactants
C The products and reactants are at the same level
D There is no difference in energy

Answer: A

6 Q9 **The enthalpy diagram shows the products at a higher level than the reactants. Which statement is correct?**



A It represents the reaction used in a cold pack
B It is an exothermic reaction
C Heat is released to the surroundings
D H(products) < H(reactants)

Answer: A

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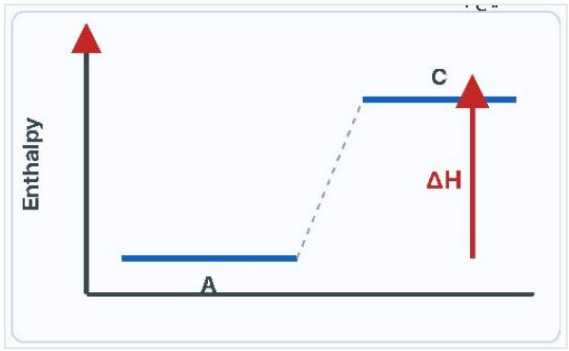
More Ministry questions · Questions 7–12

7 Q9 **In the diagram A → C (A at the bottom, C at the top), which type of pack does the reaction represent and what is true about it?**



A Hot pack; H(products) > H(reactants)

B Hot pack; ΔH < 0

C Cold pack; ΔH > 0

D Cold pack; H(products) < H(reactants)

Answer: C · Cold pack; ΔH > 0

8 Q10 **Which statement correctly describes the cold-pack process $NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$ with ΔH = +27 kJ?**

A $NH_4NO_3(s) + 27\ kJ \rightarrow NH_4^+(aq) + NO_3^-(aq)$

B ΔH(rxn) < 0

C ΔH(rxn) = −27 kJ

D 27 kJ appears on both sides of the equation

Answer: A · $NH_4NO_3(s) + 27\ kJ \rightarrow NH_4^+(aq) + NO_3^-(aq)$

9 Q11 **Complete combustion of 1 mol of CH_4 releases 890 kJ. How much energy is released when 3 mol of CH_4 burn completely?**

A 2670 kJ **B** 1780 kJ **C** 445 kJ **D** 297 kJ

Answer: A · 2670 kJ

10 Q11 **What mass of propane C_3H_8 (44.097 g/mol, ΔHcomb = −2219 kJ/mol) must burn to release 9985.5 kJ?**

A 154.4 g **B** 110.3 g **C** 66.20 g **D** 198.4 g

Answer: D · 198.4 g

11 Q12 **The enthalpy of combustion is the heat released when...**

A one mole of oxygen is used as a reactant

B one mole of a substance is burned completely in oxygen

C one mole of a substance is formed from its elements

D one mole of a substance is burned incompletely

Answer: B

12 Q21 **In a closed system, $2NO_2(g) \rightarrow 2NO(g) + O_2(g)$ takes place. What happens to the concentrations as the reaction proceeds?**

A $[NO_2]$ increases while [NO] and $[O_2]$ decrease

B All concentrations remain constant

C $[NO_2]$ decreases but [NO] stays constant

D $[NO_2]$ decreases continuously while [NO] and $[O_2]$ increase

Answer: D